UTSC FIRST-YEAR TEXTBOOK (WINTER 2025)



Marco Zimmer-De Iuliis & Anna Galang
University of Toronto Scarbo

LibreTexts™

UTSC First-Year Textbook

Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar

University of Toronto Scarborough

This text is disseminated via the Open Education Resource (OER) LibreTexts Project (https://LibreTexts.org) and like the thousands of other texts available within this powerful platform, it is freely available for reading, printing, and "consuming."

The LibreTexts mission is to bring together students, faculty, and scholars in a collaborative effort to provide and accessible, and comprehensive platform that empowers our community to develop, curate, adapt, and adopt openly licensed resources and technologies; through these efforts we can reduce the financial burden born from traditional educational resource costs, ensuring education is more accessible for students and communities worldwide.

Most, but not all, pages in the library have licenses that may allow individuals to make changes, save, and print this book. Carefully consult the applicable license(s) before pursuing such effects. Instructors can adopt existing LibreTexts texts or Remix them to quickly build course-specific resources to meet the needs of their students. Unlike traditional textbooks, LibreTexts' web based origins allow powerful integration of advanced features and new technologies to support learning.



LibreTexts is the adaptable, user-friendly non-profit open education resource platform that educators trust for creating, customizing, and sharing accessible, interactive textbooks, adaptive homework, and ancillary materials. We collaborate with individuals and organizations to champion open education initiatives, support institutional publishing programs, drive curriculum development projects, and more.

The LibreTexts libraries are Powered by NICE CXone Expert and was supported by the Department of Education Open Textbook Pilot Project, the California Education Learning Lab, the UC Davis Office of the Provost, the UC Davis Library, the California State University Affordable Learning Solutions Program, and Merlot. This material is based upon work supported by the National Science Foundation under Grant No. 1246120, 1525057, and 1413739.

Any opinions, findings, and conclusions or recommendations expressed in this material are those of the author(s) and do not necessarily reflect the views of the National Science Foundation nor the US Department of Education.

Have questions or comments? For information about adoptions or adaptions contact info@LibreTexts.org or visit our main website at https://LibreTexts.org.

This text was compiled on 09/01/2025

TABLE OF CONTENTS

Licensing

1: Introduction to Chemistry - Fundamental Concepts and Skills

- 1.1: Introduction
- 1.2: Phases and Classification of Matter
- 1.3: Physical and Chemical Properties
- 1.4: Measurements
- 1.5: Measurement Uncertainty, Accuracy, and Precision
- 1.6: Mathematical Treatment of Measurement Results
- 1.7: Key Terms
- 1.8: Key Equations
- 1.9: Summary
- 1.10: Exercises

2: Atoms, Molecules, and Ions

- 2.1: Introduction
- 2.2: Early Ideas in Atomic Theory
- 2.3: Evolution of Atomic Theory
- 2.4: Atomic Structure and Symbolism
- 2.5: Chemical Formulas
- 2.6: Key Terms
- 2.7: Key Equations
- 2.8: Summary
- 2.9: Exercises

3: The Quantum Model of the Atom

- 3.1: Introduction
- 3.2: Electromagnetic Radiation
- 3.3: Atomic Spectra
- 3.4: Quantum Theory
- 3.5: The Bohr Atom
- 3.6: Two Ideas Leading to a New Quantum Mechanics
- 3.7: Wave Mechanics
- 3.8: Electron Spin- A Fourth Quantum Number
- 3.9: The Shape of Atomic Orbitals
- 3.10: Multielectron Atoms
- 3.11: Key Terms
- 3.12: Key Equations
- 3.13: Summary
- 3.14: Exercises

4: The Periodic Table and Some Atomic Properties

- 4.1: Classifying the Elements- The Periodic Law and the Periodic Table
- 4.2: Metals and Nonmetals and their Ions
- 4.3: Sizes of Atoms and Ions
- 4.4: Ionization Energy

- 4.5: Electron Affinity
- 4.6: Magnetic Properties
- 4.7: Periodic Properties of the Elements
- 4.8: Key Terms
- 4.9: Key Equations
- 4.10: Summary
- 4.11: Exercises

5: Chemical Bonding and Molecular Geometry

- 5.1: Introduction
- 5.2: Ionic Bonding
- 5.3: Covalent Bonding
- 5.4: Chemical Nomenclature
- 5.5: Lewis Symbols and Structures
- 5.6: Formal Charges and Resonance
- 5.7: Molecular Structure and Polarity
- 5.8: Key Terms
- 5.9: Key Equations
- 5.10: Summary
- 5.11: Exercises

6: Advanced Theories of Covalent Bonding

- 6.1: Introduction
- 6.2: Valence Bond Theory
- 6.3: Hybrid Atomic Orbitals
- 6.4: Multiple Bonds
- 6.5: Molecular Orbital Theory
- 6.6: Key Terms
- 6.7: Key Equations
- 6.8: Summary
- 6.9: Exercises

7: Thermochemistry

- 7.1: Introduction
- 7.2: Energy Basics
- 7.3: Calorimetry
- 7.4: Enthalpy
- 7.5: Strengths of Ionic and Covalent Bonds
- 7.6: Key Terms
- 7.7: Key Equations
- 7.8: Summary
- 7.9: Exercises

8: Liquids and Solids and Phase Changes

- 8.1: Introduction
- 8.2: Intermolecular Forces
- 8.3: Properties of Liquids
- 8.4: Phase Transitions
- 8.5: Phase Diagrams
- 8.6: The Solid State of Matter

- 8.7: Key Terms
- 8.8: Key Equations
- 8.9: Summary
- 8.10: Exercises

9: Gases

- 9.1: Introduction
- 9.2: Gas Pressure
- 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law
- 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions
- 9.5: Effusion and Diffusion of Gases
- 9.6: The Kinetic-Molecular Theory
- 9.7: Non-Ideal Gas Behavior
- 9.8: Key Terms
- 9.9: Key Equations
- 9.10: Summary
- 9.11: Exercises

10: Three Major Classes of Chemical Reactions

- 10.1: Formula Mass
- 10.2: Writing and Balancing Chemical Equations
- 10.3: Reaction Stoichiometry
- 10.4: Reaction Yields
- 10.5: Molarity
- 10.6: Precipitation Reactions
- 10.7: Acid-Base Reactions
- 10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations
- 10.9: Oxidation-Reduction- Some General Principles
- 10.10: Balancing Oxidation-Reduction Equations
- 10.11: Oxidizing and Reducing Agents
- 10.12: Key Terms
- 10.13: Key Equations
- 10.E: Reactions in Aqueous Solution (Exercises)
- 10.S: Summary

11: Nuclear Chemistry

- 11.1: Introduction
- 11.2: Nuclear Structure and Stability
- 11.3: Nuclear Equations
- 11.4: Radioactive Decay
- 11.5: Transmutation and Nuclear Energy
- 11.6: Uses of Radioisotopes
- 11.7: Biological Effects of Radiation
- 11.8: Key Terms
- 11.9: Key Equations
- 11.10: Summary
- 11.11: Exercises

12: Kinetics

- 12.1: Introduction
- 12.2: Chemical Reaction Rates
- 12.3: Factors Affecting Reaction Rates
- 12.4: Rate Laws
- 12.5: Integrated Rate Laws
- 12.6: Collision Theory
- 12.7: Reaction Mechanisms
- 12.8: Catalysis
- 12.9: Key Terms
- 12.10: Key Equations
- 12.11: Summary
- 12.12: Exercises

13: Solutions and their Physical Properties

- 13.1: Types of Solutions - Some Terminology
- 13.2: Solution Concentration
- 13.3: Intermolecular Forces and the Solution Process
- 13.4: Solution Formation and Equilibrium
- 13.5: Solubilities of Gases
- 13.6: Vapor Pressures of Solutions
- 13.7: Osmotic Pressure
- 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions
- 13.9: Solutions of Electrolytes
- 13.10: Colloidal Mixtures
- 13.11: Exercises

14: Fundamental Equilibrium Concepts

- 14.1: Introduction
- 14.2: Chemical Equilibria
- 14.3: Equilibrium Constants
- 14.4: Shifting Equilibria - Le Chatelier’s Principle
- 14.5: Equilibrium Calculations
- 14.6: Key Terms
- 14.7: Key Equations
- 14.8: Summary
- 14.9: Exercises

15: Acid-Base Equilibria - A More Detailed Look

- 15.1: Introduction
- 15.2: Arrhenius Theory- A Brief Review
- 15.3: Brønsted-Lowry Theory of Acids and Bases
- 15.4: Self-Ionization of Water and the pH Scale
- 15.5: Strong Acids and Strong Bases
- 15.6: Weak Acids and Weak Bases
- 15.7: Polyprotic Acids
- 15.8: Ions as Acids and Bases
- 15.9: Molecular Structure and Acid-Base Behavior
- 15.10: Lewis Acids and Bases
- 15.11: Common-Ion Effect in Acid-Base Equilibria

- 15.12: Buffer Solutions
- 15.13: Acid-Base Indicators
- 15.14: Neutralization Reactions and Titration Curves
- 15.15: Solutions of Salts of Polyprotic Acids
- 15.16: Acid-Base Equilibrium Calculations- A Summary
- 15.17: Key Terms
- 15.18: Key Equations
- 15.19: Summary
- 15.20: Exercises

16: Solubility and Complex-Ion Equilibria

- 16.1: Solubility Product Constant, Ksp
- 16.2: Relationship Between Solubility and Ksp
- 16.3: Common-Ion Effect in Solubility Equilibria
- 16.4: Limitations of the K_{sp} Concept
- 16.5: Criteria for Precipitation and its Completeness
- 16.6: Fractional Precipitation
- 16.7: Solubility and pH
- 16.8: Equilibria Involving Complex Ions
- 16.9: Qualitative Cation Analysis
- 16.10: Exercises

17: Thermodynamics

- 17.1: Introduction
- 17.2: Enthalpy
- 17.3: Spontaneity
- 17.4: Entropy
- 17.5: The Second and Third Laws of Thermodynamics
- 17.6: Free Energy
- 17.7: Key Terms
- 17.8: Key Equations
- 17.9: Summary
- 17.10: Exercises

18: Electrochemistry

- 18.1: Introduction
- 18.2: Review of Redox Chemistry
- 18.3: Galvanic Cells
- 18.4: Electrode and Cell Potentials
- 18.5: Potential, Free Energy, and Equilibrium
- 18.6: Batteries and Fuel Cells
- 18.7: Electrolysis
- 18.8: Key Terms
- 18.9: Key Equations
- 18.10: Summary
- 18.11: Exercises

19: Transition Metals and Coordination Chemistry

- 19.1: Introduction
- 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds

- 19.3: Coordination Chemistry of Transition Metals
- 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds
- 19.5: Key Terms
- 19.6: Summary
- 19.7: Exercises

Index

Glossary

Detailed Licensing

Detailed Licensing

Detailed Licensing

Detailed Licensing

Licensing

*A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.*

CHAPTER OVERVIEW

1: Introduction to Chemistry - Fundamental Concepts and Skills

1.1: Introduction
1.2: Phases and Classification of Matter
1.3: Physical and Chemical Properties
1.4: Measurements
1.5: Measurement Uncertainty, Accuracy, and Precision
1.6: Mathematical Treatment of Measurement Results
1.7: Key Terms
1.8: Key Equations
1.9: Summary
1.10: Exercises

This page titled 1: Introduction to Chemistry - Fundamental Concepts and Skills is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

1.1: Introduction

Your alarm goes off and, after hitting “snooze” once or twice, you pry yourself out of bed. You make a cup of coffee to help you get going, and then you shower, get dressed, eat breakfast, and check your phone for messages. On your way to school, you stop to fill your car’s gas tank, almost making you late for the first day of chemistry class. As you find a seat in the classroom, you read the question projected on the screen: “Welcome to class! Why should we study chemistry?”



Figure 1.1.1: Chemical substances and processes are essential for our existence, providing sustenance, keeping us clean and healthy, fabricating electronic devices, enabling transportation, and much more. (credit “left”: modification of work by “vxla”/Flickr; credit “left middle”: modification of work by “the Italian voice”/Flickr; credit “right middle”: modification of work by Jason Trim; credit “right”: modification of work by “gosheshe”/Flickr)

Do you have an answer? You may be studying chemistry because it fulfills an academic requirement, but if you consider your daily activities, you might find chemistry interesting for other reasons. Most everything you do and encounter during your day involves chemistry. Making coffee, cooking eggs, and toasting bread involve chemistry. The products you use—like soap and shampoo, the fabrics you wear, the electronics that keep you connected to your world, the gasoline that propels your car—all of these and more involve chemical substances and processes. Whether you are aware or not, chemistry is part of your everyday world. In this course, you will learn many of the essential principles underlying the chemistry of modern-day life.

This page titled 1.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.0: Introduction** is licensed CC BY 4.0.

1.2: Phases and Classification of Matter

Learning Objectives

By the end of this section, you will be able to:

- Describe the basic properties of each physical state of matter: solid, liquid, and gas
- Distinguish between mass and weight
- Apply the law of conservation of matter
- Classify matter as an element, compound, homogeneous mixture, or heterogeneous mixture with regard to its physical state and composition
- Define and give examples of atoms and molecules

Matter is defined as anything that occupies space and has mass, and it is all around us. Solids and liquids are more obviously matter: We can see that they take up space, and their weight tells us that they have mass. Gases are also matter; if gases did not take up space, a balloon would not inflate (increase its volume) when filled with gas.

Solids, liquids, and gases are the three states of matter commonly found on earth (Figure 1.2.1). A solid is rigid and possesses a definite shape. A liquid flows and takes the shape of its container, except that it forms a flat or slightly curved upper surface when acted upon by gravity. (In zero gravity, liquids assume a spherical shape.) Both liquid and solid samples have volumes that are very nearly independent of pressure. A gas takes both the shape and volume of its container.

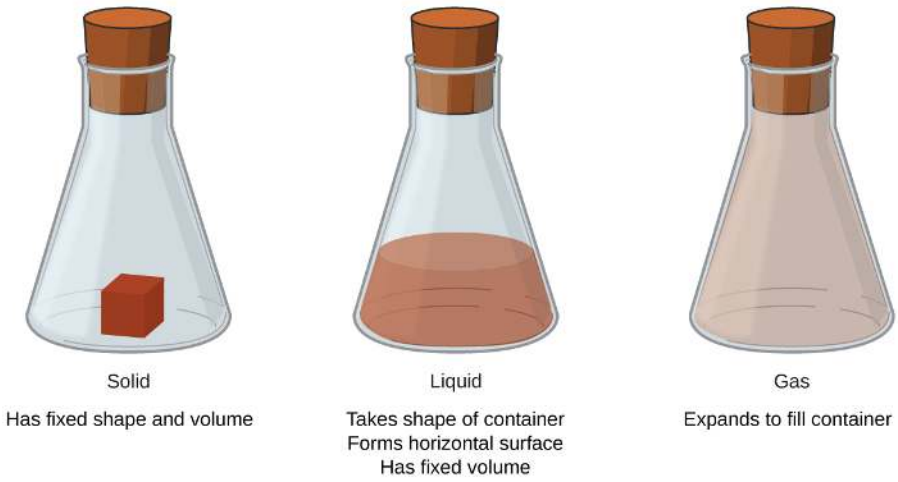


Figure 1.2.1: The three most common states or phases of matter are solid, liquid, and gas.

A fourth state of matter, plasma, occurs naturally in the interiors of stars. A plasma is a gaseous state of matter that contains appreciable numbers of electrically charged particles (Figure 1.2.2). The presence of these charged particles imparts unique properties to plasmas that justify their classification as a state of matter distinct from gases. In addition to stars, plasmas are found in some other high-temperature environments (both natural and man-made), such as lightning strikes, certain television screens, and specialized analytical instruments used to detect trace amounts of metals.



Figure 1.2.2: A plasma torch can be used to cut metal. (credit: “Hypertherm”/Wikimedia Commons)

Link to Learning

In a tiny cell in a plasma television, the plasma emits ultraviolet light, which in turn causes the display at that location to appear a specific color. The composite of these tiny dots of color makes up the image that you see. Watch this video to learn more about plasma and the places you encounter it.

Some samples of matter appear to have properties of solids, liquids, and/or gases at the same time. This can occur when the sample is composed of many small pieces. For example, we can pour sand as if it were a liquid because it is composed of many small grains of solid sand. Matter can also have properties of more than one state when it is a mixture, such as with clouds. Clouds appear to behave somewhat like gases, but they are actually mixtures of air (gas) and tiny particles of water (liquid or solid).

The mass of an object is a measure of the amount of matter in it. One way to measure an object’s mass is to measure the force it takes to accelerate the object. It takes much more force to accelerate a car than a bicycle because the car has much more mass. A more common way to determine the mass of an object is to use a balance to compare its mass with a standard mass.

Although weight is related to mass, it is not the same thing. Weight refers to the force that gravity exerts on an object. This force is directly proportional to the mass of the object. The weight of an object changes as the force of gravity changes, but its mass does not. An astronaut’s mass does not change just because she goes to the moon. But her weight on the moon is only one-sixth her earth-bound weight because the moon’s gravity is only one-sixth that of the earth’s. She may feel “weightless” during her trip when she experiences negligible external forces (gravitational or any other), although she is, of course, never “massless.”

The law of conservation of matter summarizes many scientific observations about matter: It states that *there is no detectable change in the total quantity of matter present when matter converts from one type to another (a chemical change) or changes among solid, liquid, or gaseous states (a physical change)*. Brewing beer and the operation of batteries provide examples of the conservation of matter (Figure 1.2.3). During the brewing of beer, the ingredients (water, yeast, grains, malt, hops, and sugar) are converted into beer (water, alcohol, carbonation, and flavoring substances) with no actual loss of substance. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. This can also be seen in a lead-acid car battery: The original substances (lead, lead oxide, and sulfuric acid), which are capable of producing electricity, are changed into other substances (lead sulfate and water) that do not produce electricity, with no change in the actual amount of matter.

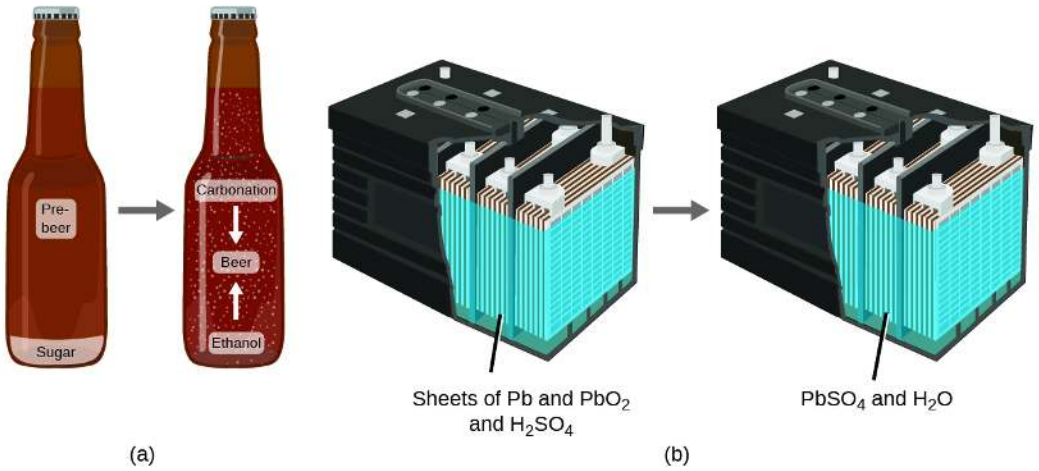


Figure 1.2.3: (a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbon dioxide. (b) The mass of the lead, lead oxide, and sulfuric acid consumed by the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Although this conservation law holds true for all conversions of matter, convincing examples are few and far between because, outside of the controlled conditions in a laboratory, we seldom collect all of the material that is produced during a particular conversion. For example, when you eat, digest, and assimilate food, all of the matter in the original food is preserved. But because some of the matter is incorporated into your body, and much is excreted as various types of waste, it is challenging to verify by measurement.

Classifying Matter

Matter can be classified into several categories. Two broad categories are mixtures and pure substances. A pure substance has a constant composition. All specimens of a pure substance have exactly the same makeup and properties. Any sample of sucrose (table sugar) consists of 42.1% carbon, 6.5% hydrogen, and 51.4% oxygen by mass. Any sample of sucrose also has the same physical properties, such as melting point, color, and sweetness, regardless of the source from which it is isolated.

Pure substances may be divided into two classes: elements and compounds. Pure substances that cannot be broken down into simpler substances by chemical changes are called elements. Iron, silver, gold, aluminum, sulfur, oxygen, and copper are familiar examples of the more than 100 known elements, of which about 90 occur naturally on the earth, and two dozen or so have been created in laboratories.

Pure substances that are comprised of two or more elements are called compounds. Compounds may be broken down by chemical changes to yield either elements or other compounds, or both. Mercury(II) oxide, an orange, crystalline solid, can be broken down by heat into the elements mercury and oxygen (Figure 1.2.4). When heated in the absence of air, the compound sucrose is broken down into the element carbon and the compound water. (The initial stage of this process, when the sugar is turning brown, is known as caramelization—this is what imparts the characteristic sweet and nutty flavor to caramel apples, caramelized onions, and caramel). Silver(I) chloride is a white solid that can be broken down into its elements, silver and chlorine, by absorption of light. This property is the basis for the use of this compound in photographic films and photochromic eyeglasses (those with lenses that darken when exposed to light).

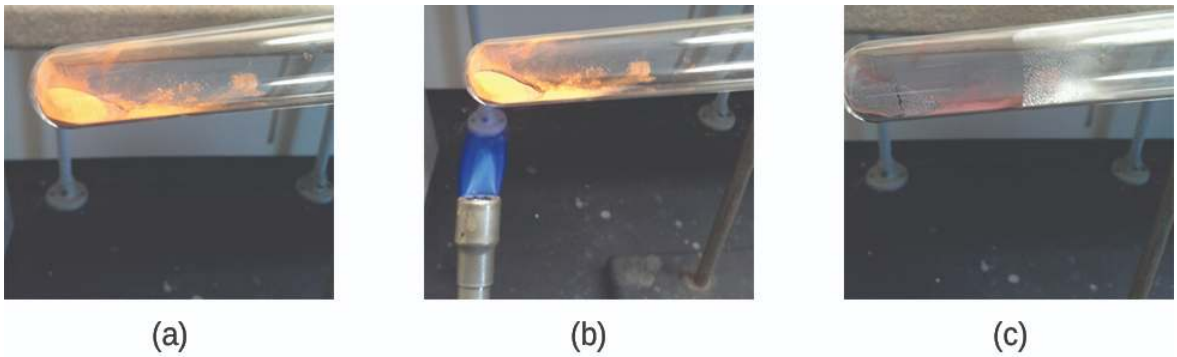


Figure 1.2.4: (a) The compound mercury(II) oxide, (b) when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)

Link to Learning

Many compounds break down when heated. This site shows the breakdown of mercury oxide, HgO. You can also view an example of the photochemical decomposition of silver chloride (AgCl), the basis of early photography.

The properties of combined elements are different from those in the free, or uncombined, state. For example, white crystalline sugar (sucrose) is a compound resulting from the chemical combination of the element carbon, which is a black solid in one of its uncombined forms, and the two elements hydrogen and oxygen, which are colorless gases when uncombined. Free sodium, an element that is a soft, shiny, metallic solid, and free chlorine, an element that is a yellow-green gas, combine to form sodium chloride (table salt), a compound that is a white, crystalline solid.

A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes, such as evaporation (you will learn more about this later). A mixture with a composition that varies from point to point is called a heterogeneous mixture. Italian dressing is an example of a heterogeneous mixture (Figure 1.2.5). Its composition can vary because it may be prepared from varying amounts of oil, vinegar, and herbs. It is not the same from point to point throughout the mixture—one drop may be mostly vinegar, whereas a different drop may be mostly oil or herbs because the oil and vinegar separate and the herbs settle. Other examples of heterogeneous mixtures are chocolate chip cookies (we can see the separate bits of chocolate, nuts, and cookie dough) and granite (we can see the quartz, mica, feldspar, and more).

A homogeneous mixture, also called a solution, exhibits a uniform composition and appears visually the same throughout. An example of a solution is a sports drink, consisting of water, sugar, coloring, flavoring, and electrolytes mixed together uniformly (Figure 1.2.5). Each drop of a sports drink tastes the same because each drop contains the same amounts of water, sugar, and other components. Note that the composition of a sports drink can vary—it could be made with somewhat more or less sugar, flavoring, or other components, and still be a sports drink. Other examples of homogeneous mixtures include air, maple syrup, gasoline, and a solution of salt in water.



Figure 1.2.5: (a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout. (b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; credit b “left: modification of work by Jeff Bedford)

Although there are just over 100 elements, tens of millions of chemical compounds result from different combinations of these elements. Each compound has a specific composition and possesses definite chemical and physical properties that distinguish it

from all other compounds. And, of course, there are innumerable ways to combine elements and compounds to form different mixtures. A summary of how to distinguish between the various major classifications of matter is shown in (Figure 1.2.6).

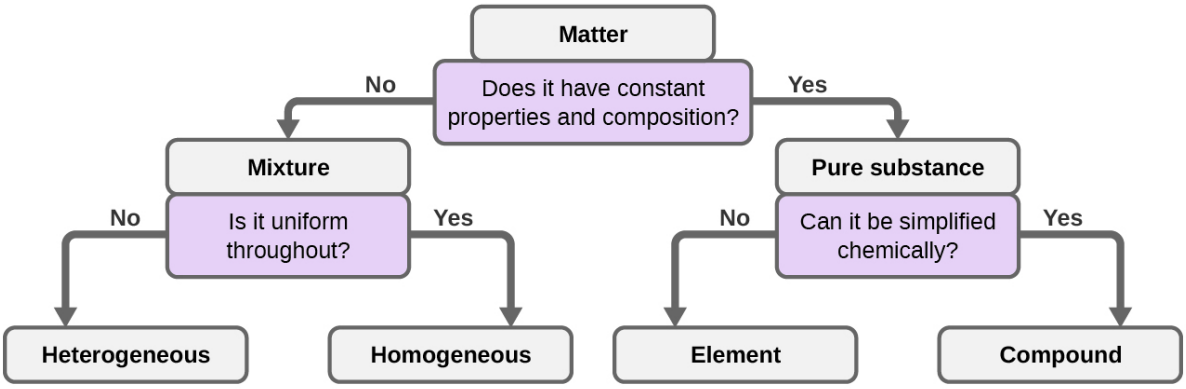


Figure 1.2.6: Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound, or an element.

Eleven elements make up about 99% of the earth's crust and atmosphere (Table 1.2.1). Oxygen constitutes nearly one-half and silicon about one-quarter of the total quantity of these elements. A majority of elements on earth are found in chemical combinations with other elements; about one-quarter of the elements are also found in the free state.

Table 1.2.1: Elemental Composition of Earth

Element	Symbol	Percent Mass	Element	Symbol	Percent Mass
oxygen	O	49.20	chlorine	Cl	0.19
silicon	Si	25.67	phosphorus	P	0.11
aluminum	Al	7.50	manganese	Mn	0.09
iron	Fe	4.71	carbon	C	0.08
calcium	Ca	3.39	sulfur	S	0.06
sodium	Na	2.63	barium	Ba	0.04
potassium	K	2.40	nitrogen	N	0.03
magnesium	Mg	1.93	fluorine	F	0.03
hydrogen	H	0.87	strontium	Sr	0.02
titanium	Ti	0.58	all others	-	0.47

Atoms and Molecules

An atom is the smallest particle of an element that has the properties of that element and can enter into a chemical combination. Consider the element gold, for example. Imagine cutting a gold nugget in half, then cutting one of the halves in half, and repeating this process until a piece of gold remained that was so small that it could not be cut in half (regardless of how tiny your knife may be). This minimally sized piece of gold is an atom (from the Greek *atomos*, meaning "indivisible") (Figure 1.2.7). This atom would no longer be gold if it were divided any further.

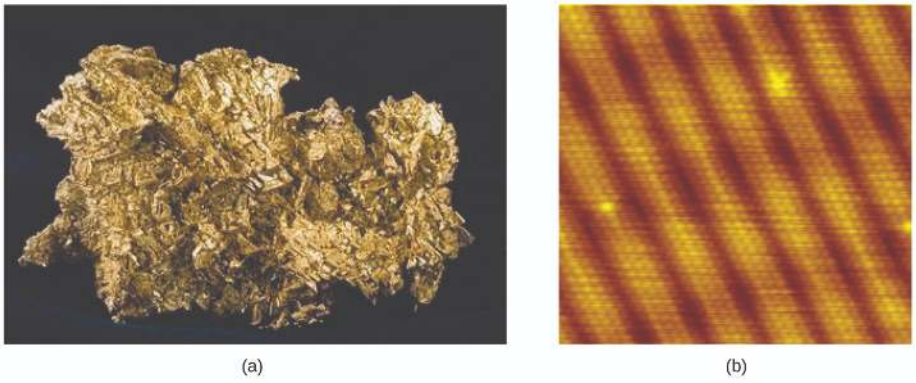


Figure 1.2.7: (a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom. (credit a: modification of work by United States Geological Survey; credit b: modification of work by "Erwinrossen"/Wikimedia Commons)

The first suggestion that matter is composed of atoms is attributed to the Greek philosophers Leucippus and Democritus, who developed their ideas in the 5th century BCE. However, it was not until the early nineteenth century that John Dalton (1766–1844), a British schoolteacher with a keen interest in science, supported this hypothesis with quantitative measurements. Since that time, repeated experiments have confirmed many aspects of this hypothesis, and it has become one of the central theories of chemistry. Other aspects of Dalton's atomic theory are still used but with minor revisions (details of Dalton's theory are provided in the chapter on atoms and molecules).

An atom is so small that its size is difficult to imagine. One of the smallest things we can see with our unaided eye is a single thread of a spider web: These strands are about 1/10,000 of a centimeter (0.0001 cm) in diameter. Although the cross-section of one strand is almost impossible to see without a microscope, it is huge on an atomic scale. A single carbon atom in the web has a diameter of about 0.000000015 centimeter, and it would take about 7000 carbon atoms to span the diameter of the strand. To put this in perspective, if a carbon atom were the size of a dime, the cross-section of one strand would be larger than a football field, which would require about 150 million carbon atom "dimes" to cover it. (Figure 1.2.8) shows increasingly close microscopic and atomic-level views of ordinary cotton.

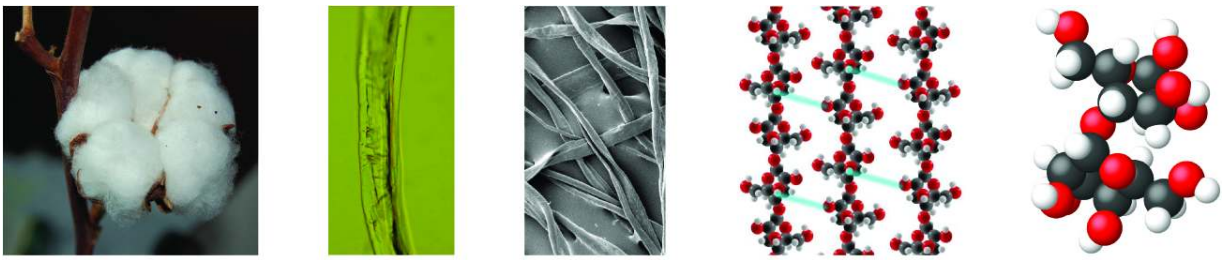


Figure 1.2.8: These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d and e) atomic-level models of the fiber (spheres of different colors represent atoms of different elements). (credit c: modification of work by "Featheredtar"/Wikimedia Commons)

An atom is so light that its mass is also difficult to imagine. A billion lead atoms (1,000,000,000 atoms) weigh about $3\ 10^{-13}$ grams, a mass that is far too light to be weighed on even the world's most sensitive balances. It would require over 300,000,000,000,000 lead atoms (300 trillion, or $3\ 10^{14}$) to be weighed, and they would weigh only 0.0000001 gram.

It is rare to find collections of individual atoms. Only a few elements, such as the gases helium, neon, and argon, consist of a collection of individual atoms that move about independently of one another. Other elements, such as the gases hydrogen, nitrogen, oxygen, and chlorine, are composed of units that consist of pairs of atoms (Figure 1.2.9). One form of the element phosphorus consists of units composed of four phosphorus atoms. The element sulfur exists in various forms, one of which consists of units composed of eight sulfur atoms. These units are called molecules. A molecule consists of two or more atoms joined by strong forces called chemical bonds. The atoms in a molecule move around as a unit, much like the cans of soda in a six-pack or a bunch of keys joined together on a single key ring. A molecule may consist of two or more identical atoms, as in the molecules found in the elements hydrogen, oxygen, and sulfur, or it may consist of two or more different atoms, as in the molecules found in water. Each water molecule is a unit that contains two hydrogen atoms and one oxygen atom. Each glucose molecule is a unit that

contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. Like atoms, molecules are incredibly small and light. If an ordinary glass of water were enlarged to the size of the earth, the water molecules inside it would be about the size of golf balls.

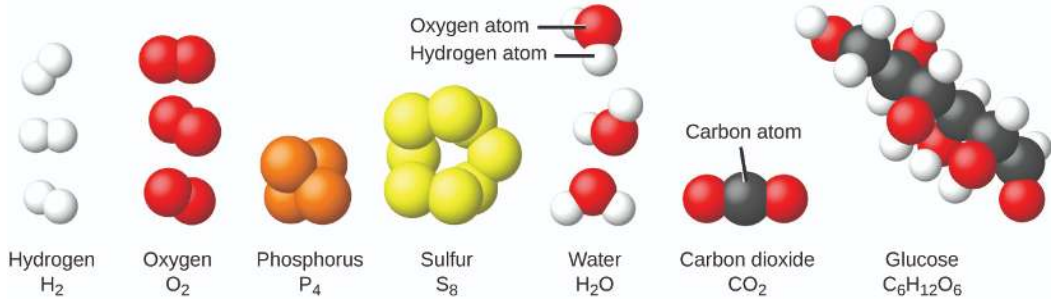


Figure 1.2.9: The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

Chemistry in Everyday Life: Decomposition of Water / Production of Hydrogen

Water consists of the elements hydrogen and oxygen combined in a 2 to 1 ratio. Water can be broken down into hydrogen and oxygen gases by the addition of energy. One way to do this is with a battery or power supply, as shown in (Figure 1.2.10).

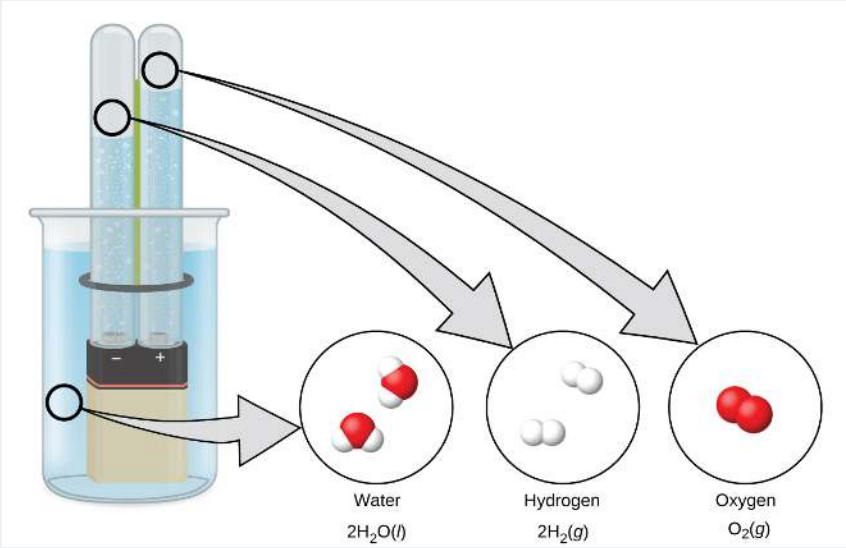


Figure 1.2.10: The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, this change is presented by showing how liquid H_2O separates into H_2 and O_2 gases.

The breakdown of water involves a rearrangement of the atoms in water molecules into different molecules, each composed of two hydrogen atoms and two oxygen atoms, respectively. Two water molecules form one oxygen molecule and two hydrogen molecules. The representation for what occurs, will be explored in more depth in later chapters.

The two gases produced have distinctly different properties. Oxygen is not flammable but is required for combustion of a fuel, and hydrogen is highly flammable and a potent energy source. How might this knowledge be applied in our world? One application involves research into more fuel-efficient transportation. Fuel-cell vehicles (FCV) run on hydrogen instead of gasoline (Figure 1.2.11). They are more efficient than vehicles with internal combustion engines, are nonpolluting, and reduce greenhouse gas emissions, making us less dependent on fossil fuels. FCVs are not yet economically viable, however, and current hydrogen production depends on natural gas. If we can develop a process to economically decompose water, or produce hydrogen in another environmentally sound way, FCVs may be the way of the future.

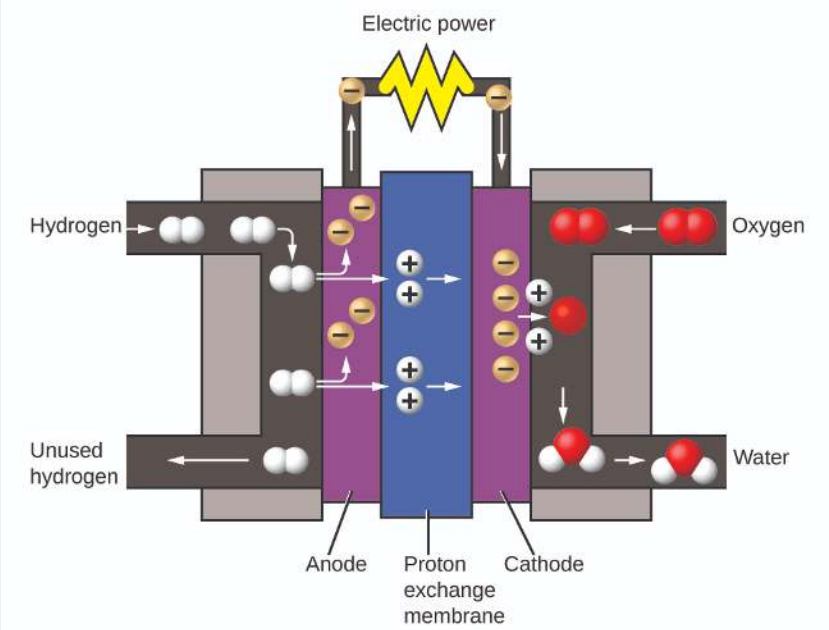


Figure 1.2.11: A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.

Chemistry in Everyday Life: Chemistry of Cell Phones

Imagine how different your life would be without cell phones (Figure 1.2.12) and other smart devices. Cell phones are made from numerous chemical substances, which are extracted, refined, purified, and assembled using an extensive and in-depth understanding of chemical principles. About 30% of the elements that are found in nature are found within a typical smart phone. The case/body/frame consists of a combination of sturdy, durable polymers composed primarily of carbon, hydrogen, oxygen, and nitrogen [acrylonitrile butadiene styrene (ABS) and polycarbonate thermoplastics], and light, strong, structural metals, such as aluminum, magnesium, and iron. The display screen is made from a specially toughened glass (silica glass strengthened by the addition of aluminum, sodium, and potassium) and coated with a material to make it conductive (such as indium tin oxide). The circuit board uses a semiconductor material (usually silicon); commonly used metals like copper, tin, silver, and gold; and more unfamiliar elements such as yttrium, praseodymium, and gadolinium. The battery relies upon lithium ions and a variety of other materials, including iron, cobalt, copper, polyethylene oxide, and polyacrylonitrile.

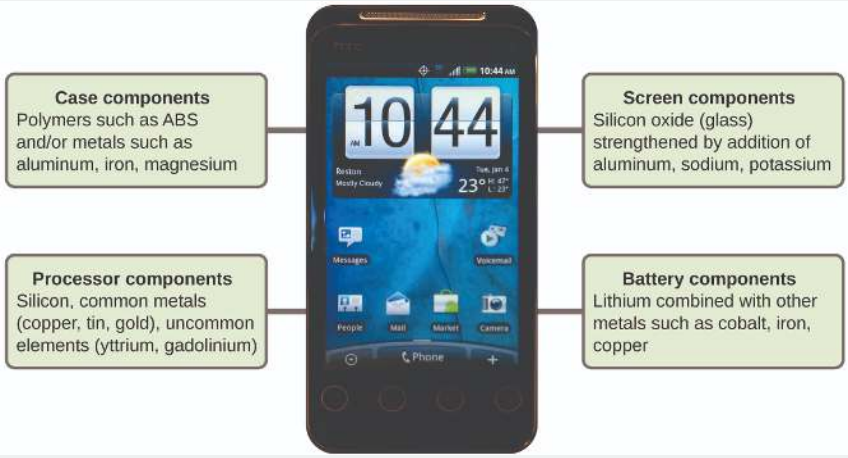


Figure 1.2.12: Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)

This page titled 1.2: Phases and Classification of Matter is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar via source content that was edited to the style and standards of the LibreTexts platform.

- **1.2: Phases and Classification of Matter** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

1.3: Physical and Chemical Properties

Learning Objectives

By the end of this section, you will be able to:

- Identify properties of and changes in matter as physical or chemical
- Identify properties of matter as extensive or intensive

The characteristics that distinguish one substance from another are called properties. A physical property is a characteristic of matter that is not associated with a change in its chemical composition. Familiar examples of physical properties include density, color, hardness, melting and boiling points, and electrical conductivity. Some physical properties, such as density and color, may be observed without changing the physical state of the matter. Other physical properties, such as the melting temperature of iron or the freezing temperature of water, can only be observed as matter undergoes a physical change. A physical change is a change in the state or properties of matter without any accompanying change in the chemical identities of the substances contained in the matter. Physical changes are observed when wax melts, when sugar dissolves in coffee, and when steam condenses into liquid water (Figure 1.3.1). Other examples of physical changes include magnetizing and demagnetizing metals (as is done with common antitheft security tags) and grinding solids into powders (which can sometimes yield noticeable changes in color). In each of these examples, there is a change in the physical state, form, or properties of the substance, but no change in its chemical composition.



(a) (b)

Figure 1.3.1: (a) Wax undergoes a physical change when solid wax is heated and forms liquid wax. (b) Steam condensing inside a cooking pot is a physical change, as water vapor is changed into liquid water. (credit a: modification of work by "95jb14"/Wikimedia Commons; credit b: modification of work by "mjneuby"/Flickr)

The change of one type of matter into another type (or the inability to change) is a chemical property. Examples of chemical properties include flammability, toxicity, acidity, and many other types of reactivity. Iron, for example, combines with oxygen in the presence of water to form rust; chromium does not oxidize (Figure 1.3.2). Nitroglycerin is very dangerous because it explodes easily; neon poses almost no hazard because it is very unreactive.



(a) (b)

Figure 1.3.2: (a) One of the chemical properties of iron is that it rusts; (b) one of the chemical properties of chromium is that it does not. (credit a: modification of work by Tony Hisgett; credit b: modification of work by "Atoma"/Wikimedia Commons)

A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure 1.3.3).

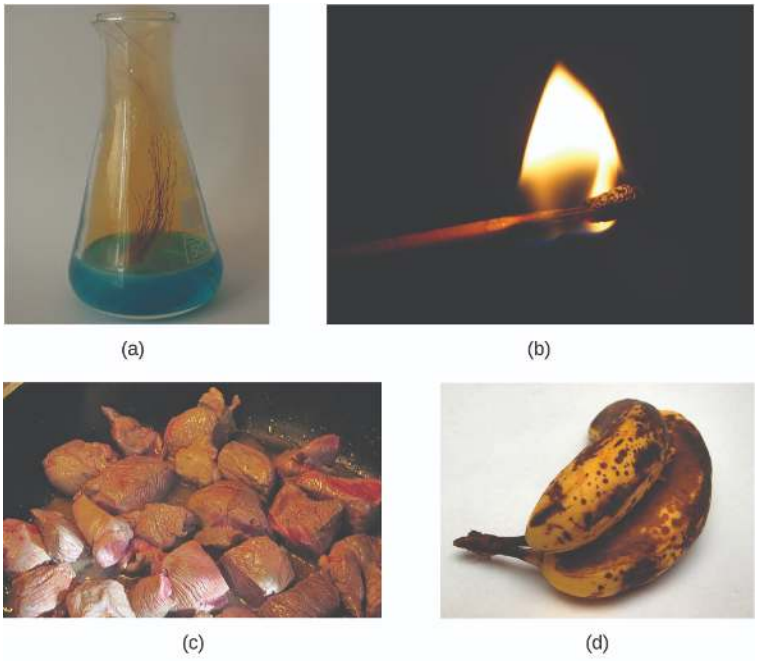


Figure 1.3.3: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo)

Properties of matter fall into one of two categories. If the property depends on the amount of matter present, it is an extensive property. The mass and volume of a substance are examples of extensive properties; for instance, a gallon of milk has a larger mass than a cup of milk. The value of an extensive property is directly proportional to the amount of matter in question. If the property of a sample of matter does not depend on the amount of matter present, it is an intensive property. Temperature is an example of an intensive property. If the gallon and cup of milk are each at 20 °C (room temperature), when they are combined, the temperature remains at 20 °C. As another example, consider the distinct but related properties of heat and temperature. A drop of hot cooking oil spattered on your arm causes brief, minor discomfort, whereas a pot of hot oil yields severe burns. Both the drop and the pot of oil are at the same temperature (an intensive property), but the pot clearly contains much more heat (extensive property).

Chemistry in Everyday Life: Hazard Diamond

You may have seen the symbol shown in Figure 1.3.4: on containers of chemicals in a laboratory or workplace. Sometimes called a “fire diamond” or “hazard diamond,” this chemical hazard diamond provides valuable information that briefly summarizes the various dangers of which to be aware when working with a particular substance.

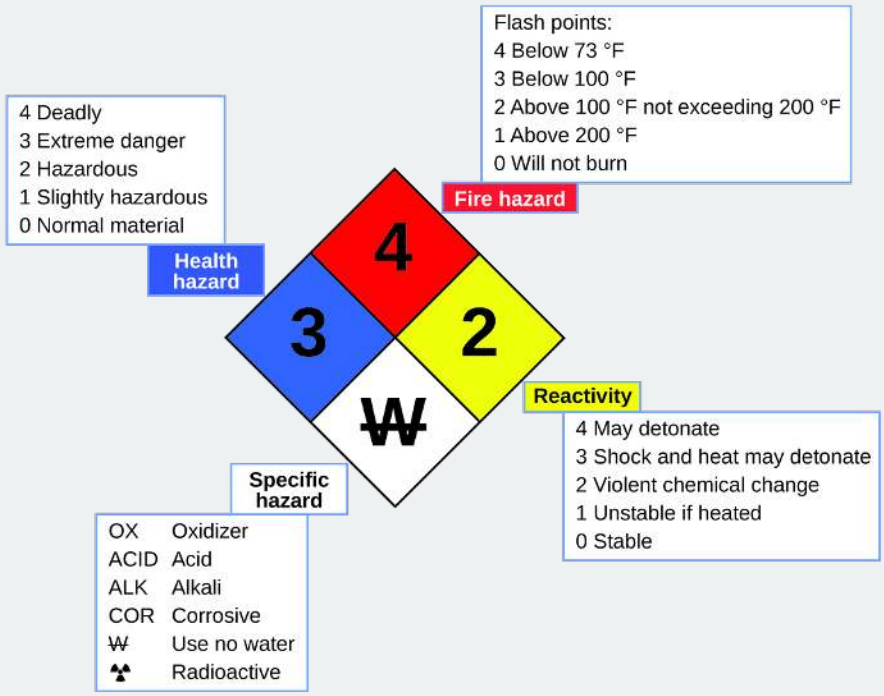


Figure 1.3.4: The National Fire Protection Agency (NFPA) hazard diamond summarizes the major hazards of a chemical substance.

The National Fire Protection Agency (NFPA) 704 Hazard Identification System was developed by NFPA to provide safety information about certain substances. The system details flammability, reactivity, health, and other hazards. Within the overall diamond symbol, the top (red) diamond specifies the level of fire hazard (temperature range for flash point). The blue (left) diamond indicates the level of health hazard. The yellow (right) diamond describes reactivity hazards, such as how readily the substance will undergo detonation or a violent chemical change. The white (bottom) diamond points out special hazards, such as if it is an oxidizer (which allows the substance to burn in the absence of air/oxygen), undergoes an unusual or dangerous reaction with water, is corrosive, acidic, alkaline, a biological hazard, radioactive, and so on. Each hazard is rated on a scale from 0 to 4, with 0 being no hazard and 4 being extremely hazardous.

While many elements differ dramatically in their chemical and physical properties, some elements have similar properties. For example, many elements conduct heat and electricity well, whereas others are poor conductors. These properties can be used to sort the elements into three classes: metals (elements that conduct well), nonmetals (elements that conduct poorly), and metalloids (elements that have intermediate conductivities).

The periodic table is a table of elements that places elements with similar properties close together (Figure 1.3.5). You will learn more about the periodic table as you continue your study of chemistry.

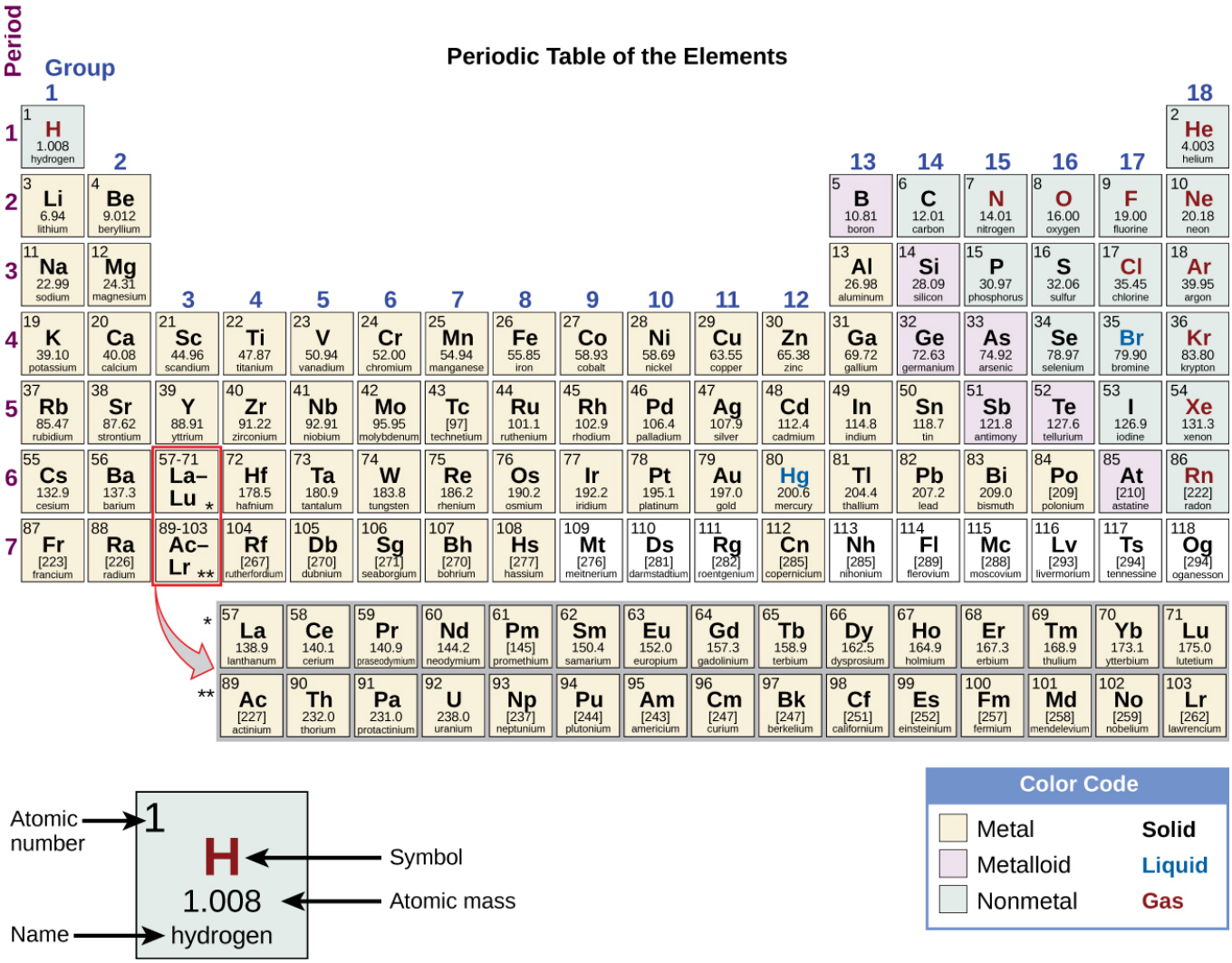


Figure 1.3.5: The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.

This page titled 1.3: Physical and Chemical Properties is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang via source content that was edited to the style and standards of the LibreTexts platform.

- **1.3: Physical and Chemical Properties** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

1.4: Measurements

Learning Objectives

By the end of this section, you will be able to:

- Explain the process of measurement
- Identify the three basic parts of a quantity
- Describe the properties and units of length, mass, volume, density, temperature, and time
- Perform basic unit calculations and conversions in the metric and other unit systems

Measurements provide much of the information that informs the hypotheses, theories, and laws describing the behavior of matter and energy in both the macroscopic and microscopic domains of chemistry. Every measurement provides three kinds of information: the size or magnitude of the measurement (a number); a standard of comparison for the measurement (a unit); and an indication of the uncertainty of the measurement. While the number and unit are explicitly represented when a quantity is written, the uncertainty is an aspect of the measurement result that is more implicitly represented and will be discussed later.

The number in the measurement can be represented in different ways, including decimal form and scientific notation. (Scientific notation is also known as exponential notation; a review of this topic can be found in Appendix B.) For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as $2.98\ 10^5$ kg. The mass of the average mosquito is about 0.0000025 kilograms, which can be written as $2.5\ 10^{-6}$ kg.

Units, such as liters, pounds, and centimeters, are standards of comparison for measurements. A 2-liter bottle of a soft drink contains a volume of beverage that is twice that of the accepted volume of 1 liter. The meat used to prepare a 0.25-pound hamburger weighs one-fourth as much as the accepted weight of 1 pound. Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of "100" without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

The measurement units for seven fundamental properties ("base units") are listed in Table 1.4.1. The standards for these units are fixed by international agreement, and they are called the **International System of Units** or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964. Units for other properties may be derived from these seven base units.

Table 1.4.1: Base Units of the SI System

Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
electric current	ampere	A
amount of substance	mole	mol
luminous intensity	candela	cd

Everyday measurement units are often defined as fractions or multiples of other units. Milk is commonly packaged in containers of 1 gallon (4 quarts), 1 quart (0.25 gallon), and one pint (0.5 quart). This same approach is used with SI units, but these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means "one thousand," which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m). The prefixes used and the powers to which 10 are raised are listed in Table 1.4.2.

Table 1.4.2: Common Unit Prefixes

Prefix	Symbol	Factor	Example
femto	f	10^{-15}	1 femtosecond (fs) = 1 x 10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1 x 10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4 x 10^{-9} g (0.000000004 g)
micro	µ	10^{-6}	1 microliter (µL) = 1 x 10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2 x 10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7 x 10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1 x 10^{-1} L (0.1 L)
kilo	k	10^{3}	1 kilometer (km) = 1 x 10^{3} m (1000 m)
mega	M	10^{6}	3 megahertz (MHz) = 3 x 10^{6} Hz (3,000,000 Hz)

Prefix	Symbol	Factor	Example
giga	G	10^9	8 gigayears (Gyr) = 8 x 10^9 yr (8,000,000,000 yr)
tera	T	10^{12}	5 terawatts (TW) = 5 x 10^{12} W (5,000,000,000,000 W)

Link to Learning

Need a refresher or more practice with scientific notation? Visit this site to go over the basics of scientific notation.

SI Base Units

The initial units of the metric system, which eventually evolved into the SI system, were established in France during the French Revolution. The original standards for the meter and the kilogram were adopted there in 1799 and eventually by other countries. This section introduces four of the SI base units commonly used in chemistry. Other SI units will be introduced in subsequent chapters.

Length

The standard unit of **length** in both the SI and original metric systems is the **meter** (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 3 inches longer than a yard (Figure 1.4.1); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

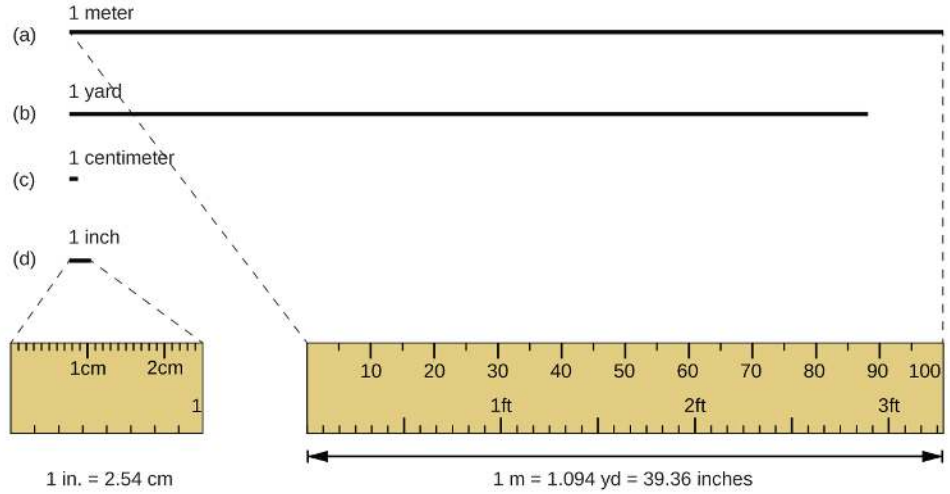


Figure 1.4.1: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Mass

The standard unit of mass in the SI system is the **kilogram** (kg). The kilogram was previously defined by the International Union of Pure and Applied Chemistry (IUPAC) as the mass of a specific reference object. This object was originally one liter of pure water, and more recently it was a metal cylinder made from a platinum-iridium alloy with a height and diameter of 39 mm (Figure 1.4.2). In May 2019, this definition was changed to one that is based instead on precisely measured values of several fundamental physical constants.[1]. One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg).



Figure 1.4.2: This replica prototype kilogram as previously defined is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology)

Temperature

Temperature is an intensive property. The SI unit of temperature is the kelvin (K). The IUPAC convention is to use **kelvin** (all lowercase) for the word, K (uppercase) for the unit symbol, and neither the word “degree” nor the degree symbol (°). The degree **Celsius** (°C) is also allowed in the SI system, with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C) by definition, and normal human body temperature is approximately 310 K (37 °C). The conversion between these two units and the Fahrenheit scale will be discussed later in this chapter.

Time

The SI base unit of time is the **second** (s). Small and large time intervals can be expressed with the appropriate prefixes; for example, 3 microseconds = 0.000003 s = 3 x 10^{-6} and 5 megaseconds = 5,000,000 s = 5 x 10^{6} s. Alternatively, hours, days, and years can be used.

Derived SI Units

We can derive many units from the seven SI base units. For example, we can use the base unit of length to define a unit of volume, and the base units of mass and length to define a unit of density.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure 1.4.3). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m, or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter (dm^3). A liter (L) is the more common name for the cubic decimeter. One **liter** is about 1.06 quarts.

A **cubic centimeter** (cm^3) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for **c**ubic **c**entimeter) is often used by health professionals. A cubic centimeter is equivalent to a milliliter (mL) and is 1/1000 of a liter.

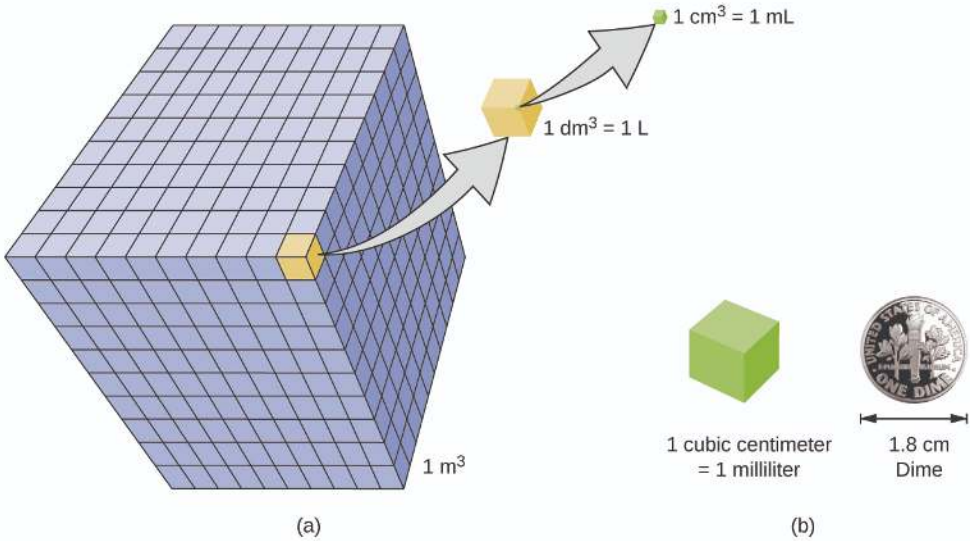


Figure 1.4.3: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

Density

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length.

The **density** of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this is an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L. Table 1.4.3 shows the densities of some common substances.

Table 1.4.3: Densities of Common Substances

Solids	Liquids	Gases (at 25 °C and 1 atm)
ice (at 0 °C) 0.92 g/cm^3	water 1.0 g/cm^3	dry air 1.20 g/L
oak (wood) 0.60–0.90 g/cm^3	ethanol 0.79 g/cm^3	oxygen 1.31 g/L
iron 7.9 g/cm^3	acetone 0.79 g/cm^3	nitrogen 1.14 g/L
copper 9.0 g/cm^3	glycerin 1.26 g/cm^3	carbon dioxide 1.80 g/L
lead 11.3 g/cm^3	olive oil 0.92 g/cm^3	helium 0.16 g/L
silver 10.5 g/cm^3	gasoline 0.70–0.77 g/cm^3	neon 0.83 g/L
gold 19.3 g/cm^3	mercury 13.6 g/cm^3	radon 9.1 g/L

While there are many ways to determine the density of an object, perhaps the most straightforward method involves separately finding the mass and volume of the object, and then dividing the mass of the sample by its volume. In the following example, the mass is found directly by weighing, but the volume is found indirectly through length measurements.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.4.1}$$

Example 1.4.1: Calculation of Density

Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm^3. What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g?

Solution

The density of a substance can be calculated by dividing its mass by its volume. The volume of a cube is calculated by cubing the edge length.

$$\text{volume of lead cube } = 2.00 \text{ cm} \times 2.00 \text{ cm} \times 2.00 \text{ cm} = 8.00 \text{ cm}^3$$

$$\text{density } = \frac{\text{mass}}{\text{volume}} = \frac{90.7 \text{ g}}{8.00 \text{ cm}^3} = 11.3 \text{ g/cm}^3$$

(We will discuss the reason for rounding to the first decimal place in the next section.)

Exercise 1.4.1

a. To three decimal places, what is the volume of a cube (cm^3) with an edge length of 0.843 cm?
b. If the cube in part (a) is copper and has a mass of 5.34 g, what is the density of copper to two decimal places?

Answer

a. 0.599 cm^3
b. 8.91 g/cm^3

Link to Learning

To learn more about the relationship between mass, volume, and density, use this interactive simulator to explore the density of different materials.

Example 1.4.2: Using Displacement of Water to Determine Density

This exercise uses a simulation to illustrate an alternative approach to the determination of density that involves measuring the object's volume via displacement of water. Use the simulator to determine the densities iron and wood.

Solution

Click the "turn fluid into water" button in the simulator to adjust the density of liquid in the beaker to 1.00 g/mL. Remove the red block from the beaker and note the volume of water is 25.5 mL. Select the iron sample by clicking "iron" in the table of materials at the bottom of the screen, place the iron block on the balance pan, and observe its mass is 31.48 g. Transfer the iron block to the beaker and notice that it sinks, displacing a volume of water equal to its own volume and causing the water level to rise to 29.5 mL. The volume of the iron block is therefore:

$$v_{\text{iron}} = 29.5mL - 25.5mL = 4.0mL$$

The density of the iron is then calculated to be:

$$\text{density } = \frac{\text{mass}}{\text{volume}} = \frac{31.48 \text{ g}}{4.0 \text{ mL}} = 7.9 \text{ g/mL}$$

Remove the iron block from the beaker, change the block material to wood, and then repeat the mass and volume measurements. Unlike iron, the wood block does not sink in the water but instead floats on the water's surface. To measure its volume, drag it beneath the water's surface so that it is fully submerged.

$$\text{density } = \frac{\text{mass}}{\text{volume}} = \frac{1.95 \text{ g}}{3.0 \text{ mL}} = 0.65 \text{ g/mL}$$

Note: The sink versus float behavior illustrated in this example demonstrates the property of “buoyancy”.

? Exercise 1.4.2

Following the water displacement approach, use the simulator to measure the density of the foam sample.

Answer

0.230 g/mL

Footnotes

- 1For details see https://www.nist.gov/pml/weights-and.../si-units-mass

---

This page titled 1.4: Measurements is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **1.4: Measurements** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

1.5: Measurement Uncertainty, Accuracy, and Precision

Learning Objectives

By the end of this section, you will be able to:

- Define accuracy and precision
- Distinguish exact and uncertain numbers
- Correctly represent uncertainty in quantities using significant figures
- Apply proper rounding rules to computed quantities

Counting is the only type of measurement that is free from uncertainty, provided the number of objects being counted does not change while the counting process is underway. The result of such a counting measurement is an example of an **exact number**. By counting the eggs in a carton, one can determine *exactly* how many eggs the carton contains. The numbers of defined quantities are also exact. By definition, 1 foot is exactly 12 inches, 1 inch is exactly 2.54 centimeters, and 1 gram is exactly 0.001 kilogram. Quantities derived from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used.

Significant Figures in Measurement

The numbers of measured quantities, unlike defined or directly counted quantities, are not exact. To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid.

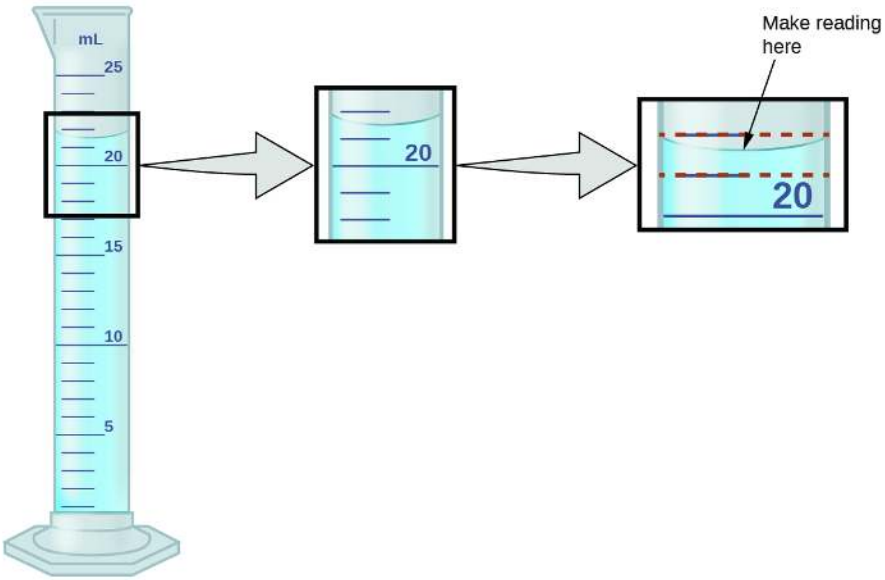
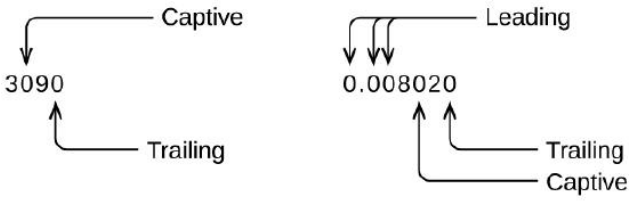


Figure 1.5.1: To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in Figure 1.5.1. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs *about* 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If the coin is weighed on a more sensitive balance, the mass might be 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some **uncertainty**, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called **significant figures** or **significant digits**. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

A measurement result is properly reported when its significant digits accurately represent the certainty of the measurement process. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them.



Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.

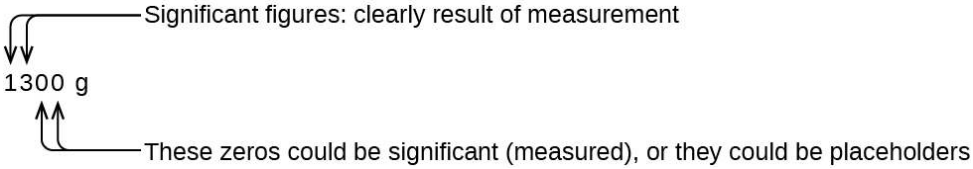


Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.



The leading zeros in this example are not significant. We could use exponential notation (as described in Appendix B) and express the number as 8.32407×10^{-3}; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point.

The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant.



When determining significant figures, be sure to pay attention to reported values and think about the measurement and significant figures in terms of what is reasonable or likely when evaluating whether the value makes sense. For example, the official January 2014 census reported the resident population of the US as 317,297,725. Do you think the US population was correctly determined to the reported nine significant figures, that is, to the exact number of people? People are constantly being born, dying, or moving into or out of the country, and assumptions are made to account for the large number of people who are not actually counted. Because of these uncertainties, it might be more reasonable to expect that we know the population to within perhaps a million or so, in which case the population should be reported as $3.17\ 10^8$ people.

Significant Figures in Calculations

A second important principle of uncertainty is that results calculated from a measurement are at least as uncertain as the measurement itself. Take the uncertainty in measurements into account to avoid misrepresenting the uncertainty in calculated results. One way to do this is to report the result of a calculation with the correct number of significant figures, which is determined by the following three rules for **rounding** numbers:

1. When adding or subtracting numbers, round the result to the same number of decimal places as the number with the least number of decimal places (the least certain value in terms of addition and subtraction).
2. When multiplying or dividing numbers, round the result to the same number of digits as the number with the least number of significant figures (the least certain value in terms of multiplication and division).
3. If the digit to be dropped (the one immediately to the right of the digit to be retained) is less than 5, “round down” and leave the retained digit unchanged; if it is more than 5, “round up” and increase the retained digit by 1. If the dropped digit is 5, and it’s either the last digit in the number or it’s followed only by zeros, round up or down, whichever yields an even value for the retained digit. If any nonzero digits follow the dropped 5, round up. (The last part of this rule may strike you as a bit odd, but it’s based on reliable statistics and is aimed at avoiding any bias when dropping the digit “5,” since it is equally close to both possible values of the retained digit.)

The following examples illustrate the application of this rule in rounding a few different numbers to three significant figures:

- 0.028675 rounds “up” to 0.0287 (the dropped digit, 7, is greater than 5)
- 18.3384 rounds “down” to 18.3 (the dropped digit, 3, is less than 5)
- 6.8752 rounds “up” to 6.88 (the dropped digit is 5, and a nonzero digit follows it)
- 92.85 rounds “down” to 92.8 (the dropped digit is 5, and the retained digit is even)

Let’s work through these rules with a few examples.

✔ Example 1.5.1: Rounding Numbers

Round the following to the indicated number of significant figures:

a. 31.57 (to two significant figures)
b. 8.1649 (to three significant figures)
c. 0.051065 (to four significant figures)
d. 0.90275 (to four significant figures)

Solution

a. 31.57 rounds “up” to 32 (the dropped digit is 5, and the retained digit is even)
b. 8.1649 rounds “down” to 8.16 (the dropped digit, 4, is less than 5)
c. 0.051065 rounds “down” to 0.05106 (the dropped digit is 5, and the retained digit is even)
d. 0.90275 rounds “up” to 0.9028 (the dropped digit is 5, and the retained digit is even)

? Exercise 1.5.1

Round the following to the indicated number of significant figures:

a. 0.424 (to two significant figures)
b. 0.0038661 (to three significant figures)
c. 421.25 (to four significant figures)
d. 28,683.5 (to five significant figures)

Answer

a. 0.42
b. 0.00387
c. 421.2
d. 28,684

Example 1.5.2: Addition and Subtraction with Significant Figures

Rule: When adding or subtracting numbers, round the result to the same number of decimal places as the number with the fewest decimal places (i.e., the least certain value in terms of addition and subtraction).

a. Add 1.0023 g and 4.383 g.
b. Subtract 421.23 g from 486 g.

Solution

(a)

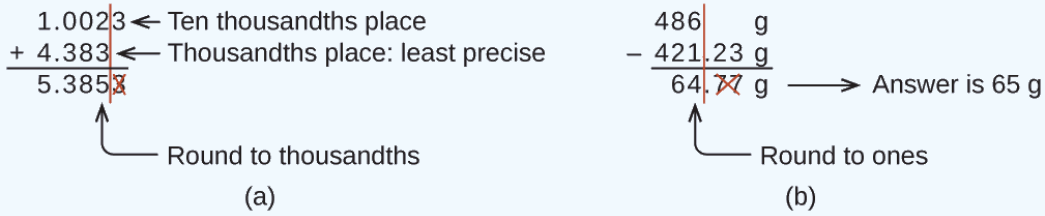
$$\begin{array}{rr} & 1.0023\ \text{g} \\ + & 4.383\ \text{g} \\ \hline & 5.3853\ \text{g} \end{array}$$

Answer is 5.385 g (round to the thousandths place; three decimal places)

(b)

$$\begin{array}{rr} & 486\ \text{g} \\ - & 421.23\ \text{g} \\ \hline & 64.77\ \text{g} \end{array}$$

Answer is 65 g (round to the ones place; no decimal places)



Exercise 1.5.2

a. Add 2.334 mL and 0.31 mL.
b. Subtract 55.8752 m from 56.533 m.

Answer

a. 2.64 mL
b. 0.658 m

Example 1.5.3: Multiplication and Division with Significant Figures

Rule: When multiplying or dividing numbers, round the result to the same number of digits as the number with the fewest significant figures (the least certain value in terms of multiplication and division).

a. Multiply 0.6238 cm by 6.6 cm.
b. Divide 421.23 g by 486 mL.

Solution

(a)

$$\underbrace{0.6238\ \text{cm}}_{\text{four significant figures}} \times \underbrace{6.6\ \text{cm}}_{\text{two significant figures}} = 4.11708\ \text{cm}^2 \rightarrow \underbrace{4.1\ \text{cm}^2}_{\text{round to two significant figures}}$$

(b)

$$\underbrace{421.23\ \text{g}}_{\text{five significant figures}} \div \underbrace{486\ \text{mL}}_{\text{three significant figures}} = 0.866728\ldots\ \text{g/mL} \rightarrow \underbrace{0.867\ \text{g/mL}}_{\text{round to three significant figures}}$$

Exercise 1.5.3

a. Multiply 2.334 cm and 0.320 cm.
b. Divide 55.8752 m by 56.53 s.

Answer

a. 0.747 cm^2
b. 0.9884 m/s

In the midst of all these technicalities, it is important to keep in mind the reason for these rules about significant figures and rounding—to correctly represent the certainty of the values reported and to ensure that a calculated result is not represented as being more certain than the least certain value used in the calculation.

Example 1.5.4: Calculation with Significant Figures

One common bathtub is 13.44 dm long, 5.920 dm wide, and 2.54 dm deep. Assume that the tub is rectangular and calculate its approximate volume in liters.

Solution

$$\begin{aligned} V &= l \times w \times d \\ &= 13.44\ \text{dm} \times 5.920\ \text{dm} \times 2.54\ \text{dm} \\ &= 202.09459\ldots\ \text{dm}^3\ \text{(value from calculator)} \\ &= 202\ \text{dm}^3\ \text{or}\ 202\ \text{L}\ \text{(answer rounded to three significant figures)} \end{aligned}$$

Exercise 1.5.4

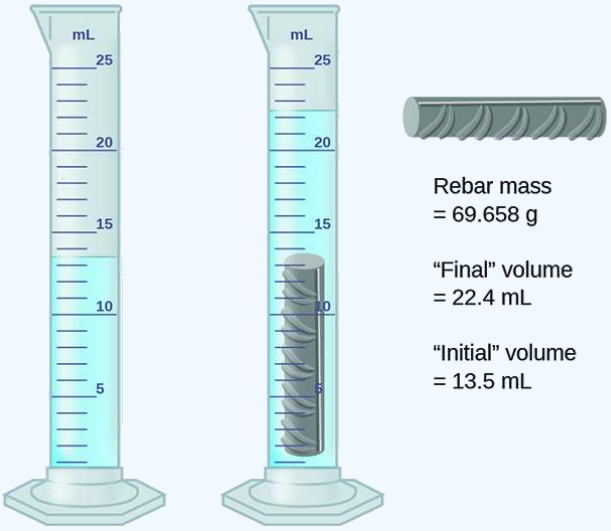
What is the density of a liquid with a mass of 31.1415 g and a volume of 30.13 cm^3?

Answer

1.034 g/mL

Example 1.5.5: Experimental Determination of Density Using Water Displacement

A piece of rebar is weighed and then submerged in a graduated cylinder partially filled with water, with results as shown.



Rebar mass
= 69.658 g

“Final” volume
= 22.4 mL

“Initial” volume
= 13.5 mL

a. Use these values to determine the density of this piece of rebar.
b. Rebar is mostly iron. Does your result in (a) support this statement? How?

Solution

The volume of the piece of rebar is equal to the volume of the water displaced:

$$\text{volume } = 22.4\,\text{mL} - 13.5\,\text{mL} = 8.9\,\text{mL} = 8.9\,\text{cm}^3$$

(rounded to the nearest 0.1 mL, per the rule for addition and subtraction)

The density is the mass-to-volume ratio:

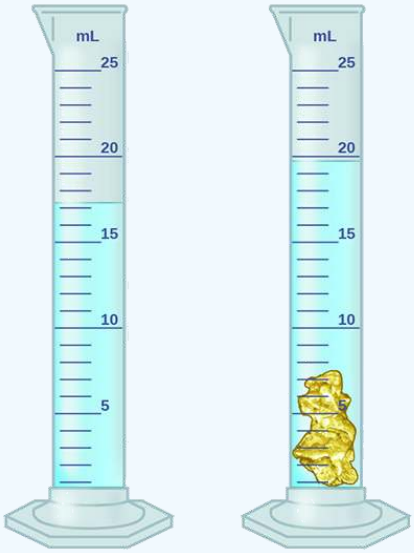
$$\text{density } = \frac{\text{mass}}{\text{volume}} = \frac{69.658\,\text{g}}{8.9\,\text{cm}^3} = 7.8\,\text{g/cm}^3$$

(rounded to two significant figures, per the rule for multiplication and division)

The density of iron is 7.9 g/cm^3, very close to that of rebar, which lends some support to the fact that rebar is mostly iron.

? Exercise 1.5.5

An irregularly shaped piece of a shiny yellowish material is weighed and then submerged in a graduated cylinder, with results as shown.





Mass
= 51.842 g

“Final” volume
= 19.8 mL

“Initial” volume
= 17.1 mL

a. Use these values to determine the density of this material.
b. Do you have any reasonable guesses as to the identity of this material? Explain your reasoning.

Answer

19 g/cm^3

It is likely gold; the right appearance for gold and very close to the known density of gold at $19.3\ \mathrm{g/cm^3}$

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to evaluate both the precision and the accuracy of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (Figure 1.5.2).

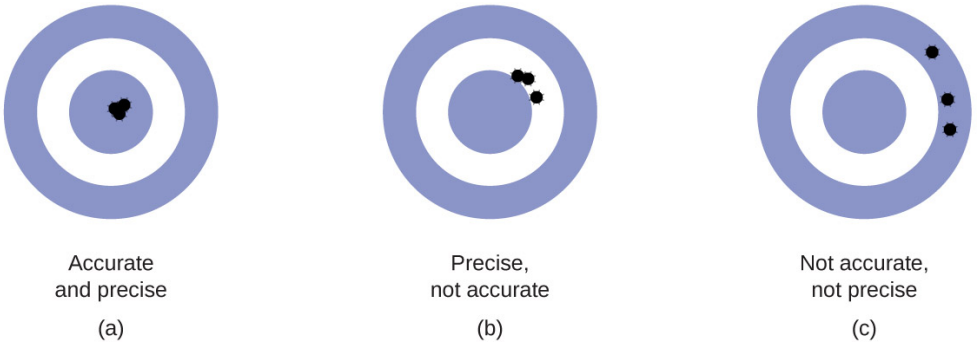


Figure 1.5.2: (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Suppose a quality control chemist at a pharmaceutical company is tasked with checking the accuracy and precision of three different machines that are meant to dispense 10 ounces (296 mL) of cough syrup into storage bottles. She proceeds to use each machine to fill five bottles and then carefully determines the actual volume dispensed, obtaining the results tabulated in Table 1.5.1.

Table 1.5.1: Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers

Dispenser #1	Dispenser #2	Dispenser #3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

Considering these results, she will report that dispenser #1 is precise (values all close to one another, within a few tenths of a milliliter) but not accurate (none of the values are close to the target value of 296 mL, each being more than 10 mL too low). Results for dispenser #2 represent improved accuracy (each volume is less than 3 mL away from 296 mL) but worse precision (volumes vary by more than 4 mL). Finally, she can report that dispenser #3 is working well, dispensing cough syrup both accurately (all volumes within 0.1 mL of the target volume) and precisely (volumes differing from each other by no more than 0.2 mL).

This page titled 1.5: Measurement Uncertainty, Accuracy, and Precision is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.5: Measurement Uncertainty, Accuracy, and Precision** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

1.6: Mathematical Treatment of Measurement Results

Learning Objectives

By the end of this section, you will be able to:

- Explain the dimensional analysis (factor label) approach to mathematical calculations involving quantities
- Use dimensional analysis to carry out unit conversions for a given property and computations involving two or more properties

It is often the case that a quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships. For example, consider measuring the average speed of an athlete running sprints. This is typically accomplished by measuring the *time* required for the athlete to run from the starting line to the finish line, and the *distance* between these two lines, and then computing *speed* from the equation that relates these three properties:

$$\text{speed} = \frac{\text{distance}}{\text{time}}$$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of

$$\frac{100\ \text{m}}{10\ \text{s}} = 10\ \text{m/s}$$

(For this and the next calculation, assume the trailing zeros are significant digits.) Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity (100/10 = 10) *and likewise* dividing the units of each measured quantity to yield the unit of the computed quantity (m/s = m/s). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation among the three properties is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriately:

$$\text{time} = \frac{\text{distance}}{\text{speed}}$$

The time can then be computed as:

$$\frac{25\ \text{m}}{10\ \text{m/s}} = 2.5\ \text{s}$$

Again, arithmetic on the numbers (25/10 = 2.5) was accompanied by the same arithmetic on the units (m/(m/s) = s) to yield the number and unit of the result, 2.5 s. Note that, just as for numbers, when a unit is divided by an identical unit (in this case, m/m), the result is "1"—or, as commonly phrased, the units "cancel."

These calculations are examples of a versatile mathematical approach known as **dimensional analysis** (or the **factor-label method**). Dimensional analysis is based on this premise: *the units of quantities must be subjected to the same mathematical operations as their associated numbers*. This method can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities.

Conversion Factors and Dimensional Analysis

A ratio of two equivalent quantities expressed with different measurement units can be used as a **unit conversion factor**. For example, the lengths of 2.54 cm and 1 in. are equivalent (by definition), and so a unit conversion factor may be derived from the ratio,

$$\frac{2.54\ \text{cm}}{1\ \text{in.}}\ (2.54\ \text{cm} = 1\ \text{in.})\ \text{or}\ 2.54\ \frac{\text{cm}}{\text{in.}}$$

Several other commonly used conversion factors are given in Example 1.6.1.

Example 1.6.1: Common Conversion Factors

Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft^3 = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.787 mL	1 (troy) oz = 31.103 g

When a quantity (such as distance in inches) is multiplied by an appropriate unit conversion factor, the quantity is converted to an equivalent value with different units (such as distance in centimeters). For example, a basketball player's vertical jump of 34 inches can be converted to centimeters by:

$$34 \text{ in.} \times \frac{2.54 \text{ cm}}{1 \text{ in.}} = 86 \text{ cm}$$

Since this simple arithmetic involves *quantities*, the premise of dimensional analysis requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the units are multiplied to yield . Just as for numbers, a ratio of identical units is also numerically equal to one, and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to "cancel.") Dimensional analysis may be used to confirm the proper application of unit conversion factors as demonstrated in the following example.

Example 1.6.1: Using a Unit Conversion Factor

The mass of a competition frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g (Example 1.6.1).

Solution

Given the conversion factor, the mass in ounces may be derived using an equation similar to the one used for converting length from inches to centimeters.

$$x oz = 125 g \times \text{ unit conversion factor}$$

The unit conversion factor may be represented as:

$$\frac{1 \text{ oz}}{28.349 \text{ g}} \text{ and } \frac{28.349 \text{ g}}{1 \text{ oz}}$$

The correct unit conversion factor is the ratio that cancels the units of grams and leaves ounces.

$$\begin{aligned} x \text{ oz} &= 125 \text{ g} \times \frac{1 \text{ oz}}{28.349 \text{ g}} \\ &= \left(\frac{125}{28.349}\right) \text{ oz} \\ &= 4.41 \text{ oz (three significant figures)} \end{aligned}$$

Exercise 1.6.1

Convert a volume of 9.345 qt to liters.

Answer

8.844 L

Beyond simple unit conversions, the factor-label method can be used to solve more complex problems involving computations. Regardless of the details, the basic approach is the same—all the *factors* involved in the calculation must be appropriately oriented

to ensure that their *labels* (units) will appropriately cancel and/or combine to yield the desired unit in the result. As your study of chemistry continues, you will encounter many opportunities to apply this approach.

Example 1.6.2: Computing Quantities from Measurement Results and Known Mathematical Relations

What is the density of common antifreeze in units of g/mL? A 4.00-qt sample of the antifreeze weighs 9.26 lb.

Solution

Since $\text{density} = \frac{\text{mass}}{\text{volume}}$, we need to divide the mass in grams by the volume in milliliters. In general: the number of units of B = the number of units of A × unit conversion factor. The necessary conversion factors are given in Example 1.6.1: 1 lb = 453.59 g; 1 L = 1.0567 qt; 1 L = 1,000 mL. Mass may be converted from pounds to grams as follows:

$$9.26\ \cancel{\text{lb}} \times \left(\frac{453.59\ \text{g}}{1\ \cancel{\text{lb}}}\right) = 4.20 \times 10^3\ \text{g}$$

Volume may be converted from quarts to milliliters via two steps:

Step 1: Convert quarts to liters.

$$4.00\ \cancel{\text{qt}} \times \left(\frac{1\ \text{L}}{1.0567\ \cancel{\text{qt}}}\right) = 3.78\ \text{L}$$

Step 2: Convert liters to milliliters.

$$3.78\ \cancel{\text{L}} \times \left(\frac{1000\ \text{mL}}{1\ \cancel{\text{L}}}\right) = 3.78 \times 10^3\ \text{mL}$$

Then,

$$\text{density} = \frac{4.20 \times 10^3\ \text{g}}{3.78 \times 10^3\ \text{mL}} = 1.11\ \text{g/mL}$$

Alternatively, the calculation could be set up in a way that uses three unit conversion factors sequentially as follows:

$$\frac{9.26\ \cancel{\text{lb}}}{4.00\ \cancel{\text{qt}}} \times \frac{453.59\ \text{g}}{1\ \cancel{\text{lb}}} \times \frac{1.0567\ \cancel{\text{qt}}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{L}}}{1000\ \text{mL}} = 1.11\ \text{g/mL}$$

Exercise 1.6.2

What is the volume in liters of 1.000 oz, given that 1 L = 1.0567 qt and 1 qt = 32 oz (exactly)?

Answer

2.956×10^{-2} L

Example 1.6.3: Computing Quantities from Measurement Results and Known Mathematical Relations

While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L gasoline.

a. What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?
b. If gasoline costs $3.80 per gallon, what was the fuel cost for this trip?

Solution

(a) First convert distance from kilometers to miles:

$$1250\ \cancel{\text{km}} \times \frac{0.62137\ \text{mi}}{1\ \cancel{\text{km}}} = 777\ \text{mi}$$

and then convert volume from liters to gallons:

$$213\ \cancel{\text{L}} \times \frac{1.0567\ \cancel{\text{qt}}}{1\ \cancel{\text{L}}} \times \frac{1\ \text{gal}}{4\ \cancel{\text{qt}}} = 56.3\ \text{gal}$$

Finally,

$$(\text{average})\ \text{mileage} = \frac{777\ \text{mi}}{56.3\ \text{gal}} = 13.8\ \text{miles / gallon} = 13.8, \text{mpg}$$

Alternatively, the calculation could be set up in a way that uses all the conversion factors sequentially, as follows:

$$\frac{1250\ \cancel{\text{km}}}{213\ \cancel{\text{L}}} \times \frac{0.62137\ \text{mi}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{L}}}{1.0567\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = 13.8\ \text{mpg}$$

(b) Using the previously calculated volume in gallons, we find:

$$56.3\ \cancel{\text{gal}} \times \frac{\$3.80}{1\ \cancel{\text{gal}}} = \$214$$

? Exercise 1.6.3

A Toyota Prius Hybrid uses 59.7 L gasoline to drive from San Francisco to Seattle, a distance of 1300 km (two significant digits).

a. What (average) fuel economy, in miles per gallon, did the Prius get during this trip?
b. If gasoline costs $3.90 per gallon, what was the fuel cost for this trip?

Answer

a. 51 mpg
b. $62

Conversion of Temperature Units

We use the word **temperature** to refer to the hotness or coldness of a substance. One way we measure a change in temperature is to use the fact that most substances expand when their temperature increases and contract when their temperature decreases. The liquid in a common glass thermometer changes its volume as the temperature changes, and the position of the trapped liquid's surface along a printed scale may be used as a measure of temperature.

Temperature scales are defined relative to selected reference temperatures: Two of the most commonly used are the freezing and boiling temperatures of water at a specified atmospheric pressure. On the Celsius scale, 0 °C is defined as the freezing temperature of water and 100 °C as the boiling temperature of water. The space between the two temperatures is divided into 100 equal intervals, which we call degrees. On the **Fahrenheit** scale, the freezing point of water is defined as 32 °F and the boiling temperature as 212 °F. The space between these two points on a Fahrenheit thermometer is divided into 180 equal parts (degrees).

Defining the Celsius and Fahrenheit temperature scales as described in the previous paragraph results in a slightly more complex relationship between temperature values on these two scales than for different units of measure for other properties. Most measurement units for a given property are directly proportional to one another (y = mx). Using familiar length units as one example:

$$\text{length in feet} = \left(\frac{1ft}{12in.}\right) \times \text{length in inches}$$

where y = length in feet, x = length in inches, and the proportionality constant, m, is the conversion factor. The Celsius and Fahrenheit temperature scales, however, do not share a common zero point, and so the relationship between these two scales is a

linear one rather than a proportional one (y = mx + b). Consequently, converting a temperature from one of these scales into the other requires more than simple multiplication by a conversion factor, m; it also must take into account differences in the scales' zero points (b).

The linear equation relating Celsius and Fahrenheit temperatures is easily derived from the two temperatures used to define each scale. Representing the Celsius temperature as *x* and the Fahrenheit temperature as *y*, the slope, *m*, is computed to be:

$$m = \frac{\Delta y}{\Delta x} = \frac{212^\circ F - 32^\circ F}{100^\circ C - 0^\circ C} = \frac{180^\circ F}{100^\circ C} = \frac{9^\circ F}{5^\circ C}$$

The *y*-intercept of the equation, *b*, is then calculated using either of the equivalent temperature pairs, (100 °C, 212 °F) or (0 °C, 32 °F), as:

$$b = y - mx = 32^\circ F - \frac{9^\circ F}{5^\circ C} \times 0^\circ C = 32^\circ F$$

The equation relating the temperature (T) scales is then:

$$T^\circ F = \left(\frac{9^\circ F}{5^\circ C} \times T^\circ C\right) + 32^\circ F$$

An abbreviated form of this equation that omits the measurement units is:

$$T^\circ F = \left(\frac{9}{5} \times T^\circ C\right) + 32$$

Rearrangement of this equation yields the form useful for converting from Fahrenheit to Celsius:

$$T^\circ F = \left(\frac{9}{5} \times T^\circ C\right) + 32$$

As mentioned earlier in this chapter, the SI unit of temperature is the kelvin (K). Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale in which 0 (zero) K corresponds to the lowest temperature that can theoretically be achieved. Since the kelvin temperature scale is absolute, a degree symbol is not included in the unit abbreviation, K. The early 19th-century discovery of the relationship between a gas's volume and temperature suggested that the volume of a gas would be zero at −273.15 °C. In 1848, British physicist William Thompson, who later adopted the title of Lord Kelvin, proposed an absolute temperature scale based on this concept (further treatment of this topic is provided in this text's chapter on gases).

The freezing temperature of water on this scale is 273.15 K and its boiling temperature is 373.15 K. Notice the numerical difference in these two reference temperatures is 100, the same as for the Celsius scale, and so the linear relation between these two temperature scales will exhibit a slope of . Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

$$\begin{aligned} T_K &= T^\circ C + 273.15 \\ T^{\circ C} &= T_K - 273.15 \end{aligned}$$

The 273.15 in these equations has been determined experimentally, so it is not exact. Figure 1.6.1: shows the relationship among the three temperature scales.

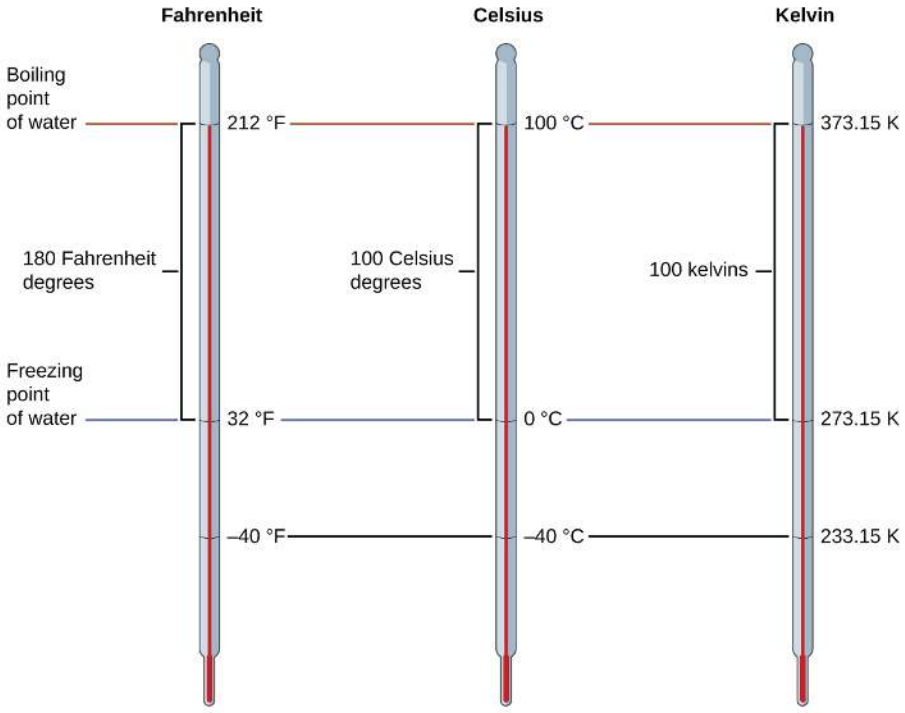


Figure 1.6.1: The Fahrenheit, Celsius, and kelvin temperature scales are compared.

Although the kelvin (absolute) temperature scale is the official SI temperature scale, Celsius is commonly used in many scientific contexts and is the scale of choice for nonscience contexts in almost all areas of the world. Very few countries (the U.S. and its territories, the Bahamas, Belize, Cayman Islands, and Palau) still use Fahrenheit for weather, medicine, and cooking.

✔ Example 1.6.4: Conversion from Celsius

Normal body temperature has been commonly accepted as 37.0 °C (although it varies depending on time of day and method of measurement, as well as among individuals). What is this temperature on the kelvin scale and on the Fahrenheit scale?

Solution

$$K = {}^\circ C + 273.15 = 37.0 + 273.2 = 310.2 K$$

$${}^\circ F = \frac{9}{5}{}^\circ C + 32.0 = \left(\frac{9}{5} \times 37.0\right) + 32.0 = 66.6 + 32.0 = 98.6^\circ F$$

? Exercise 1.6.1

Convert 80.92 °C to K and °F.

Answer

354.07 K, 177.7 °F

✔ Example 1.6.5: Conversion from Fahrenheit

Baking a ready-made pizza calls for an oven temperature of 450 °F. If you are in Europe, and your oven thermometer uses the Celsius scale, what is the setting? What is the kelvin temperature?

Solution

$${}^\circ C = \frac{5}{9}({}^\circ F - 32) = \frac{5}{9}(450 - 32) = \frac{5}{9} \times 418 = 232^\circ C \longrightarrow \text{ set oven to } 230^\circ C \quad (\text{two significant figures}$$

$$K = {}^{\circ}C + 273.15 = \frac{5}{9}({}^{\circ}F - 32) = \frac{5}{9}(450 - 32) + 273.15 = 505.4K \longrightarrow 5.1 \times 10^{2} K$$

Exercise 1.6.1

Convert 50 °F to °C and K.

Answer

10 °C, 280 K

Footnotes

- 2 Strictly speaking, the ounce and pound are units of weight, *W* (a force equal to the product of mass and gravitational acceleration, *W* = *mg*). The conversion relations in this table are commonly used to equate masses and weight assuming a nominal value for *g* at the surface of the earth.

This page titled 1.6: Mathematical Treatment of Measurement Results is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.6: Mathematical Treatment of Measurement Results** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

1.7: Key Terms

accuracy | how closely a measurement aligns with a correct value

atom | smallest particle of an element that can enter into a chemical combination

Celsius (°C) | unit of temperature; water freezes at 0 °C and boils at 100 °C on this scale

chemical change | change producing a different kind of matter from the original kind of matter

chemical property | behavior that is related to the change of one kind of matter into another kind of matter

chemistry | study of the composition, properties, and interactions of matter

compound | pure substance that can be decomposed into two or more elements

cubic centimeter (cm^3 or cc) | volume of a cube with an edge length of exactly 1 cm

cubic meter (m^3) | SI unit of volume

density | ratio of mass to volume for a substance or object

dimensional analysis | (also, factor-label method) versatile mathematical approach that can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities

element | substance that is composed of a single type of atom; a substance that cannot be decomposed by a chemical change

exact number | number derived by counting or by definition

extensive property | property of a substance that depends on the amount of the substance

Fahrenheit | unit of temperature; water freezes at 32 °F and boils at 212 °F on this scale

gas | state in which matter has neither definite volume nor shape

heterogeneous mixture | combination of substances with a composition that varies from point to point

homogeneous mixture | (also, solution) combination of substances with a composition that is uniform throughout

hypothesis | tentative explanation of observations that acts as a guide for gathering and checking information

intensive property | property of a substance that is independent of the amount of the substance

kelvin (K) | SI unit of temperature; 273.15 K = 0 ºC

kilogram (kg) | standard SI unit of mass; 1 kg = approximately 2.2 pounds

law | statement that summarizes a vast number of experimental observations, and describes or predicts some aspect of the natural world

law of conservation of matter | when matter converts from one type to another or changes form, there is no detectable change in the total amount of matter present

length | measure of one dimension of an object

liquid | state of matter that has a definite volume but indefinite shape

liter (L) | (also, cubic decimeter) unit of volume; 1 L = 1,000 cm^3

macroscopic domain | realm of everyday things that are large enough to sense directly by human sight and touch

mass | fundamental property indicating amount of matter

matter | anything that occupies space and has mass

meter (m) | standard metric and SI unit of length; 1 m = approximately 1.094 yards

microscopic domain | realm of things that are much too small to be sensed directly

milliliter (mL) | 1/1,000 of a liter; equal to 1 cm^3

mixture | matter that can be separated into its components by physical means

molecule | bonded collection of two or more atoms of the same or different elements

physical change | change in the state or properties of matter that does not involve a change in its chemical composition

physical property | characteristic of matter that is not associated with any change in its chemical composition

plasma | gaseous state of matter containing a large number of electrically charged atoms and/or molecules

precision | how closely a measurement matches the same measurement when repeated

pure substance | homogeneous substance that has a constant composition

rounding | procedure used to ensure that calculated results properly reflect the uncertainty in the measurements used in the calculation

scientific method | path of discovery that leads from question and observation to law or hypothesis to theory, combined with experimental verification of the hypothesis and any necessary modification of the theory

second (s) | SI unit of time

SI units (International System of Units) | standards fixed by international agreement in the International System of Units (*Le Système International d'Unités*)

significant figures | (also, significant digits) all of the measured digits in a determination, including the uncertain last digit

solid | state of matter that is rigid, has a definite shape, and has a fairly constant volume

symbolic domain | specialized language used to represent components of the macroscopic and microscopic domains, such as chemical symbols, chemical formulas, chemical equations, graphs, drawings, and calculations

temperature | intensive property representing the hotness or coldness of matter

theory | well-substantiated, comprehensive, testable explanation of a particular aspect of nature

uncertainty | estimate of amount by which measurement differs from true value

unit | standard of comparison for measurements

unit conversion factor | ratio of equivalent quantities expressed with different units; used to convert from one unit to a different unit

volume | amount of space occupied by an object

weight | force that gravity exerts on an object

This page titled 1.7: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.7: Key Terms** is licensed CC BY 4.0.

1.8: Key Equations

This page titled 1.8: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **1.8: Key Equations** is licensed CC BY 4.0.

1.9: Summary

Chemistry deals with the composition, structure, and properties of matter, and the ways by which various forms of matter may be interconverted. Thus, it occupies a central place in the study and practice of science and technology. Chemists use the scientific method to perform experiments, pose hypotheses, and formulate laws and develop theories, so that they can better understand the behavior of the natural world. To do so, they operate in the macroscopic, microscopic, and symbolic domains. Chemists measure, analyze, purify, and synthesize a wide variety of substances that are important to our lives.

Matter is anything that occupies space and has mass. The basic building block of matter is the atom, the smallest unit of an element that can enter into combinations with atoms of the same or other elements. In many substances, atoms are combined into molecules. On earth, matter commonly exists in three states: solids, of fixed shape and volume; liquids, of variable shape but fixed volume; and gases, of variable shape and volume. Under high-temperature conditions, matter also can exist as a plasma. Most matter is a mixture: It is composed of two or more types of matter that can be present in varying amounts and can be separated by physical means. Heterogeneous mixtures vary in composition from point to point; homogeneous mixtures have the same composition from point to point. Pure substances consist of only one type of matter. A pure substance can be an element, which consists of only one type of atom and cannot be broken down by a chemical change, or a compound, which consists of two or more types of atoms.

All substances have distinct physical and chemical properties, and may undergo physical or chemical changes. Physical properties, such as hardness and boiling point, and physical changes, such as melting or freezing, do not involve a change in the composition of matter. Chemical properties, such flammability and acidity, and chemical changes, such as rusting, involve production of matter that differs from that present beforehand.

Measurable properties fall into one of two categories. Extensive properties depend on the amount of matter present, for example, the mass of gold. Intensive properties do not depend on the amount of matter present, for example, the density of gold. Heat is an example of an extensive property, and temperature is an example of an intensive property.

Measurements provide quantitative information that is critical in studying and practicing chemistry. Each measurement has an amount, a unit for comparison, and an uncertainty. Measurements can be represented in either decimal or scientific notation. Scientists primarily use SI (International System) units such as meters, seconds, and kilograms, as well as derived units, such as liters (for volume) and g/cm^3 (for density). In many cases, it is convenient to use prefixes that yield fractional and multiple units, such as microseconds (10^{-6} seconds) and megahertz (10^6 hertz), respectively.

Quantities can be defined or measured. Measured quantities have an associated uncertainty that is represented by the number of significant figures in the quantity's number. The uncertainty of a calculated quantity depends on the uncertainties in the quantities used in the calculation and is reflected in how the value is rounded. Quantities are characterized with regard to accuracy (closeness to a true or accepted value) and precision (variation among replicate measurement results).

Measurements are made using a variety of units. It is often useful or necessary to convert a measured quantity from one unit into another. These conversions are accomplished using unit conversion factors, which are derived by simple applications of a mathematical approach called the factor-label method or dimensional analysis. This strategy is also employed to calculate sought quantities using measured quantities and appropriate mathematical relations.

This page titled 1.9: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.9: Summary** is licensed CC BY 4.0.

1.10: Exercises

1. Explain how you could experimentally determine whether the outside temperature is higher or lower than 0 °C (32 °F) without using a thermometer.
2. Identify each of the following statements as being most similar to a hypothesis, a law, or a theory. Explain your reasoning.
    a. Falling barometric pressure precedes the onset of bad weather.
    b. All life on earth has evolved from a common, primitive organism through the process of natural selection.
    c. My truck's gas mileage has dropped significantly, probably because it's due for a tune-up.
3. Identify each of the following statements as being most similar to a hypothesis, a law, or a theory. Explain your reasoning.
    a. The pressure of a sample of gas is directly proportional to the temperature of the gas.
    b. Matter consists of tiny particles that can combine in specific ratios to form substances with specific properties.
    c. At a higher temperature, solids (such as salt or sugar) will dissolve better in water.
4. Identify each of the underlined items as a part of either the macroscopic domain, the microscopic domain, or the symbolic domain of chemistry. For any in the symbolic domain, indicate whether they are symbols for a macroscopic or a microscopic feature.
    a. The mass of a lead pipe is 14 lb.
    b. The mass of a certain chlorine atom is 35 amu.
    c. A bottle with a label that reads Al contains aluminum metal.
    d. Al is the symbol for an aluminum atom.
5. Identify each of the underlined items as a part of either the macroscopic domain, the microscopic domain, or the symbolic domain of chemistry. For those in the symbolic domain, indicate whether they are symbols for a macroscopic or a microscopic feature.
    a. A certain molecule contains one H atom and one Cl atom.
    b. Copper wire has a density of about 8 g/cm^3.
    c. The bottle contains 15 grams of Ni powder.
    d. A sulfur molecule is composed of eight sulfur atoms.
6. According to one theory, the pressure of a gas increases as its volume decreases because the molecules in the gas have to move a shorter distance to hit the walls of the container. Does this theory follow a macroscopic or microscopic description of chemical behavior? Explain your answer.
7. The amount of heat required to melt 2 lbs of ice is twice the amount of heat required to melt 1 lb of ice. Is this observation a macroscopic or microscopic description of chemical behavior? Explain your answer.
8. Why is an object's mass, rather than its weight, used to indicate the amount of matter it contains?
9. What properties distinguish solids from liquids? Liquids from gases? Solids from gases?
10. How does a heterogeneous mixture differ from a homogeneous mixture? How are they similar?
11. How does a homogeneous mixture differ from a pure substance? How are they similar?
12. How does an element differ from a compound? How are they similar?
13. How do molecules of elements and molecules of compounds differ? In what ways are they similar?
14.

How does an atom differ from a molecule? In what ways are they similar?

15.

Many of the items you purchase are mixtures of pure compounds. Select three of these commercial products and prepare a list of the ingredients that are pure compounds.

16.

Classify each of the following as an element, a compound, or a mixture:

a. copper
b. water
c. nitrogen
d. sulfur
e. air
f. sucrose
g. a substance composed of molecules each of which contains two iodine atoms
h. gasoline

17.

Classify each of the following as an element, a compound, or a mixture:

a. iron
b. oxygen
c. mercury oxide
d. pancake syrup
e. carbon dioxide
f. a substance composed of molecules each of which contains one hydrogen atom and one chlorine atom
g. baking soda
h. baking powder

18.

A sulfur atom and a sulfur molecule are not identical. What is the difference?

19.

How are the molecules in oxygen gas, the molecules in hydrogen gas, and water molecules similar? How do they differ?

20.

Why are astronauts in space said to be “weightless,” but not “massless”?

21.

Prepare a list of the principal chemicals consumed and produced during the operation of an automobile.

22.

Matter is everywhere around us. Make a list by name of fifteen different kinds of matter that you encounter every day. Your list should include (and label at least one example of each) the following: a solid, a liquid, a gas, an element, a compound, a homogenous mixture, a heterogeneous mixture, and a pure substance.

23.

When elemental iron corrodes it combines with oxygen in the air to ultimately form red brown iron(III) oxide called rust. (a) If a shiny iron nail with an initial mass of 23.2 g is weighed after being coated in a layer of rust, would you expect the mass to have increased, decreased, or remained the same? Explain. (b) If the mass of the iron nail increases to 24.1 g, what mass of oxygen combined with the iron?

24.

As stated in the text, convincing examples that demonstrate the law of conservation of matter outside of the laboratory are few and far between. Indicate whether the mass would increase, decrease, or stay the same for the following scenarios where chemical reactions take place:

a. Exactly one pound of bread dough is placed in a baking tin. The dough is cooked in an oven at 350 °F releasing a wonderful aroma of freshly baked bread during the cooking process. Is the mass of the baked loaf less than, greater than, or the same as the one pound of original dough? Explain.
b. When magnesium burns in air a white flaky ash of magnesium oxide is produced. Is the mass of magnesium oxide less than, greater than, or the same as the original piece of magnesium? Explain.
c. Antoine Lavoisier, the French scientist credited with first stating the law of conservation of matter, heated a mixture of tin and air in a sealed flask to produce tin oxide. Did the mass of the sealed flask and contents decrease, increase, or remain the same after the heating?

25.

Yeast converts glucose to ethanol and carbon dioxide during anaerobic fermentation as depicted in the simple chemical equation here:

a. If 200.0 g of glucose is fully converted, what will be the total mass of ethanol and carbon dioxide produced?
b. If the fermentation is carried out in an open container, would you expect the mass of the container and contents after fermentation to be less than, greater than, or the same as the mass of the container and contents before fermentation? Explain.
c. If 97.7 g of carbon dioxide is produced, what mass of ethanol is produced?

26.

Classify the six underlined properties in the following paragraph as chemical or physical:

Fluorine is a pale yellow gas that reacts with most substances. The free element melts at −220 °C and boils at −188 °C. Finely divided metals burn in fluorine with a bright flame. Nineteen grams of fluorine will react with 1.0 gram of hydrogen.

27.

Classify each of the following changes as physical or chemical:

a. condensation of steam
b. burning of gasoline
c. souring of milk
d. dissolving of sugar in water
e. melting of gold

28.

Classify each of the following changes as physical or chemical:

a. coal burning
b. ice melting
c. mixing chocolate syrup with milk
d. explosion of a firecracker
e. magnetizing of a screwdriver

29.

The volume of a sample of oxygen gas changed from 10 mL to 11 mL as the temperature changed. Is this a chemical or physical change?

30.

A 2.0-liter volume of hydrogen gas combined with 1.0 liter of oxygen gas to produce 2.0 liters of water vapor. Does oxygen undergo a chemical or physical change?

31.

Explain the difference between extensive properties and intensive properties.

32.

Identify the following properties as either extensive or intensive.

a. volume
b. temperature

c. humidity
d. heat
e. boiling point

33.

The density (d) of a substance is an intensive property that is defined as the ratio of its mass (m) to its volume (V).

Considering that mass and volume are both extensive properties, explain why their ratio, density, is intensive.

34.

Is one liter about an ounce, a pint, a quart, or a gallon?

35.

Is a meter about an inch, a foot, a yard, or a mile?

36.

Indicate the SI base units or derived units that are appropriate for the following measurements:

a. (f) the area of a football field

(g) the maximum temperature at the South Pole on April 1, 1913

37.

Indicate the SI base units or derived units that are appropriate for the following measurements:

a. (f) the area of the state of Delaware

(g) the volume of a flu shot or a measles vaccination

38.

Give the name and symbol of the prefixes used with SI units to indicate multiplication by the following exact quantities.

a. (f) 0.000001

39.

Give the name of the prefix and the quantity indicated by the following symbols that are used with SI base units.

a. c
b. d
c. G
d. k
e. m
f. n
g. p
h. T

40.

A large piece of jewelry has a mass of 132.6 g. A graduated cylinder initially contains 48.6 mL water. When the jewelry is submerged in the graduated cylinder, the total volume increases to 61.2 mL.

a. Determine the density of this piece of jewelry.
b. Assuming that the jewelry is made from only one substance, what substance is it likely to be? Explain.

41.

Visit this density simulation and click the "turn fluid into water" button to adjust the density of liquid in the beaker to 1.00 g/mL.

a. Use the water displacement approach to measure the mass and volume of the unknown material (select the green block with question marks).
b. Use the measured mass and volume data from step (a) to calculate the density of the unknown material.
c. Link out to the link provided.

d. Assuming this material is a copper-containing gemstone, identify its three most likely identities by comparing the measured density to the values tabulated at this gemstone density guide.
e. How are mass and density related for blocks of the same volume?

42.

Visit this density simulation and click the "reset" button to ensure all simulator parameters are at their default values.

a. Use the water displacement approach to measure the mass and volume of the red block.
b. Use the measured mass and volume data from step (a) to calculate the density of the red block.
c. Use the vertical green slide control to adjust the fluid density to values well above, then well below, and finally nearly equal to the density of the red block, reporting your observations.

43.

Visit this density simulation and click the “turn fluid into water” button to adjust the density of liquid in the beaker to 1.00 g/mL. Change the block material to foam, and then wait patiently until the foam block stops bobbing up and down in the water.

(a) The foam block should be floating on the surface of the water (that is, only partially submerged). What is the volume of water displaced?
(b) Use the water volume from part (a) and the density of water (1.00 g/mL) to calculate the mass of water displaced.
(c) Remove and weigh the foam block. How does the block’s mass compare to the mass of displaced water from part (b)?

44.

Express each of the following numbers in scientific notation with correct significant figures:

a. (f) 10000.0

   (g) 0.000000738592

45.

Express each of the following numbers in exponential notation with correct significant figures:

a. (f) 0.0000000651

   (g) 0.007157

46.

Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

a. (f) the time required to drive from San Francisco to Kansas City at an average speed of 53 mi/h

47.

Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

a. (f) the distance from San Francisco to Kansas City

48.

How many significant figures are contained in each of the following measurements?

a. (f) 17.0 kg

   (g) 0.01400 g/mL

49.

How many significant figures are contained in each of the following measurements?

a. (f) 0.17 g/mL

   (g) 0.88400 s

50.

The following quantities were reported on the labels of commercial products. Determine the number of significant figures in each.

a. 0.0055 g active ingredients
b. 12 tablets

c. 3% hydrogen peroxide
d. 5.5 ounces
e. 473 mL
f. 1.75% bismuth
g. 0.001% phosphoric acid
h. 99.80% inert ingredients

51.

Round off each of the following numbers to two significant figures:

a. (f) 0.445

52.

Round off each of the following numbers to two significant figures:

a. (f) 0.885

53.

Perform the following calculations and report each answer with the correct number of significant figures.

a. (f) 42.7 + 0.259

54.

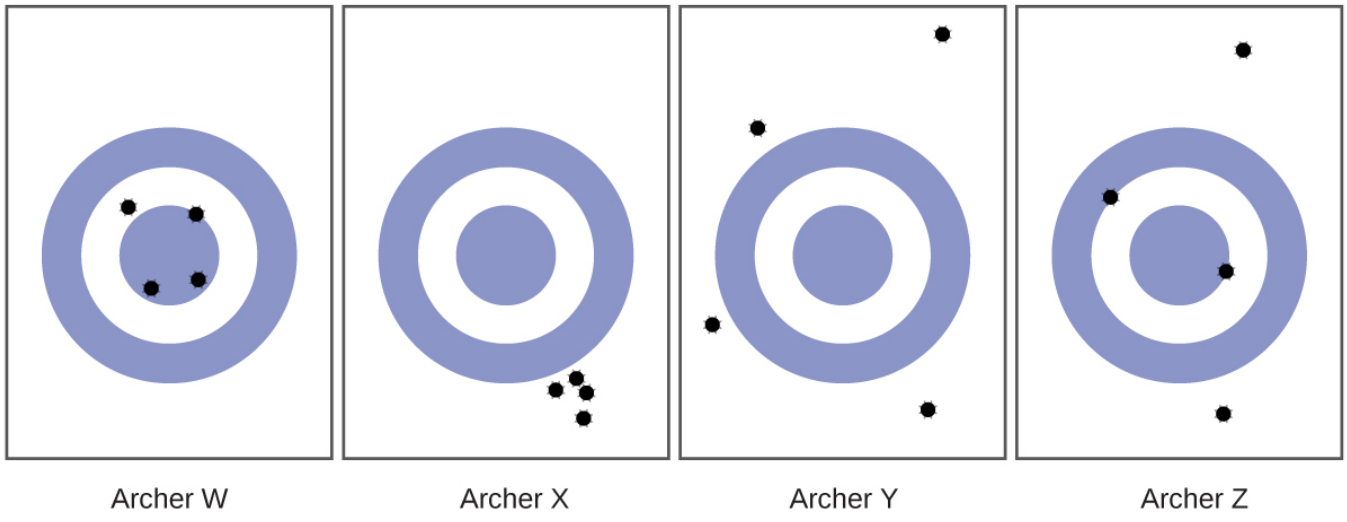
Perform the following calculations and report each answer with the correct number of significant figures.

a. 62.8 34
b. 0.147 + 0.0066 + 0.012
c. 38 95 1.792
d. 15 – 0.15 – 0.6155
e.
f. 140 + 7.68 + 0.014
g. 28.7 – 0.0483
h.

55.

Consider the results of the archery contest shown in this figure.

a. Which archer is most precise?
b. Which archer is most accurate?
c. Who is both least precise and least accurate?



56.

Classify the following sets of measurements as accurate, precise, both, or neither.

a. Checking for consistency in the weight of chocolate chip cookies: 17.27 g, 13.05 g, 19.46 g, 16.92 g
b. Testing the volume of a batch of 25-mL pipettes: 27.02 mL, 26.99 mL, 26.97 mL, 27.01 mL
c. Determining the purity of gold: 99.9999%, 99.9998%, 99.9998%, 99.9999%

57.

Write conversion factors (as ratios) for the number of:

a. yards in 1 meter
b. liters in 1 liquid quart
c. pounds in 1 kilogram

58.

Write conversion factors (as ratios) for the number of:

a. kilometers in 1 mile
b. liters in 1 cubic foot
c. grams in 1 ounce

59.

The label on a soft drink bottle gives the volume in two units: 2.0 L and 67.6 fl oz. Use this information to derive a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

60.

The label on a box of cereal gives the mass of cereal in two units: 978 grams and 34.5 oz. Use this information to find a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

61.

Soccer is played with a round ball having a circumference between 27 and 28 in. and a weight between 14 and 16 oz. What are these specifications in units of centimeters and grams?

62.

A woman's basketball has a circumference between 28.5 and 29.0 inches and a maximum weight of 20 ounces (two significant figures). What are these specifications in units of centimeters and grams?

63.

How many milliliters of a soft drink are contained in a 12.0-oz can?

64.

A barrel of oil is exactly 42 gal. How many liters of oil are in a barrel?

65.

The diameter of a red blood cell is about $3\ 10^{-4}$ in. What is its diameter in centimeters?

66.

The distance between the centers of the two oxygen atoms in an oxygen molecule is $1.21\ 10^{-8}$ cm. What is this distance in inches?

67.

Is a 197-lb weight lifter light enough to compete in a class limited to those weighing 90 kg or less?

68.

A very good 197-lb weight lifter lifted 192 kg in a move called the clean and jerk. What was the mass of the weight lifted in pounds?

69.

Many medical laboratory tests are run using 5.0 μL blood serum. What is this volume in milliliters?

70.

If an aspirin tablet contains 325 mg aspirin, how many grams of aspirin does it contain?

71.

Use scientific (exponential) notation to express the following quantities in terms of the SI base units in Table 1.2:

a. (i) 0.15 mK

72.

Complete the following conversions between SI units.

a. (i) 1.9 Gg =_____ g

73.

Gasoline is sold by the liter in many countries. How many liters are required to fill a 12.0-gal gas tank?

74.

Milk is sold by the liter in many countries. What is the volume of exactly 1/2 gal of milk in liters?

75.

A long ton is defined as exactly 2240 lb. What is this mass in kilograms?

76.

Make the conversion indicated in each of the following:

a. (f) the mass of a 3525-lb car to kilograms

(g) the mass of a 2.3-oz egg to grams

77.

Make the conversion indicated in each of the following:

a. (f) the mass of a bushel of rye, 32.0 lb, to kilograms

(g) the mass of a 5.00-grain aspirin tablet to milligrams (1 grain = 0.00229 oz)

78.

Many chemistry conferences have held a 50-Trillion Angstrom Run (two significant figures). How long is this run in kilometers and in miles? (1 Å = 1 10^{-10} m)

79.

A chemist's 50-Trillion Angstrom Run (see Exercise 1.78) would be an archeologist's 10,900 cubit run. How long is one cubit in meters and in feet? (1 Å = 1 10^{-8} cm)

80.

The gas tank of a certain luxury automobile holds 22.3 gallons according to the owner's manual. If the density of gasoline is 0.8206 g/mL, determine the mass in kilograms and pounds of the fuel in a full tank.

81.

As an instructor is preparing for an experiment, he requires 225 g phosphoric acid. The only container readily available is a 150-mL Erlenmeyer flask. Is it large enough to contain the acid, whose density is 1.83 g/mL?

82.

To prepare for a laboratory period, a student lab assistant needs 125 g of a compound. A bottle containing 1/4 lb is available. Did the student have enough of the compound?

83.

A chemistry student is 159 cm tall and weighs 45.8 kg. What is her height in inches and weight in pounds?

84.

In a recent Grand Prix, the winner completed the race with an average speed of 229.8 km/h. What was his speed in miles per hour, meters per second, and feet per second?

85.

Solve these problems about lumber dimensions.

a. To describe to a European how houses are constructed in the US, the dimensions of “two-by-four” lumber must be converted into metric units. The thickness width length dimensions are 1.50 in. 3.50 in. 8.00 ft in the US. What are the dimensions in cm cm m?
b. This lumber can be used as vertical studs, which are typically placed 16.0 in. apart. What is that distance in centimeters?

86.

The mercury content of a stream was believed to be above the minimum considered safe—1 part per billion (ppb) by weight. An analysis indicated that the concentration was 0.68 parts per billion. What quantity of mercury in grams was present in 15.0 L of the water, the density of which is 0.998 g/ml?

87.

Calculate the density of aluminum if 27.6 cm^3 has a mass of 74.6 g.

88.

Osmium is one of the densest elements known. What is its density if 2.72 g has a volume of 0.121 cm^3?

89.

Calculate these masses.

a. What is the mass of 6.00 cm^3 of mercury, density = 13.5939 g/cm^3?
b. What is the mass of 25.0 mL octane, density = 0.702 g/cm^3?

90.

Calculate these masses.

a. What is the mass of 4.00 cm^3 of sodium, density = 0.97 g/cm^3 ?
b. What is the mass of 125 mL gaseous chlorine, density = 3.16 g/L?

91.

Calculate these volumes.

a. What is the volume of 25 g iodine, density = 4.93 g/cm^3?
b. What is the volume of 3.28 g gaseous hydrogen, density = 0.089 g/L?

92.

Calculate these volumes.

a. What is the volume of 11.3 g graphite, density = 2.25 g/cm^3?
b. What is the volume of 39.657 g bromine, density = 2.928 g/cm^3?

93.

Convert the boiling temperature of gold, 2966 °C, into degrees Fahrenheit and kelvin.

94.

Convert the temperature of scalding water, 54 °C, into degrees Fahrenheit and kelvin.

95.

Convert the temperature of the coldest area in a freezer, −10 °F, to degrees Celsius and kelvin.

96.

Convert the temperature of dry ice, −77 °C, into degrees Fahrenheit and kelvin.

97.

Convert the boiling temperature of liquid ammonia, −28.1 °F, into degrees Celsius and kelvin.

98.

The label on a pressurized can of spray disinfectant warns against heating the can above 130 °F. What are the corresponding temperatures on the Celsius and kelvin temperature scales?

99.

The weather in Europe was unusually warm during the summer of 1995. The TV news reported temperatures as high as 45 °C. What was the temperature on the Fahrenheit scale?

This page titled 1.10: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **1.10: Exercises** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

CHAPTER OVERVIEW

2: Atoms, Molecules, and Ions

This chapter will describe some of the fundamental chemical principles related to the composition of matter, including those central to the concept of molecular identity.

2.1: Introduction
2.2: Early Ideas in Atomic Theory
2.3: Evolution of Atomic Theory
2.4: Atomic Structure and Symbolism
2.5: Chemical Formulas
2.6: Key Terms
2.7: Key Equations
2.8: Summary
2.9: Exercises

This page titled 2: Atoms, Molecules, and Ions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

2.1: Introduction

Lung diseases and lung cancers are among the world's most devastating illnesses partly due to delayed detection and diagnosis. Most noninvasive screening procedures aren't reliable, and patients often resist more accurate methods due to discomfort with the procedures or with the potential danger that the procedures cause. But what if you could be accurately diagnosed through a simple breath test?

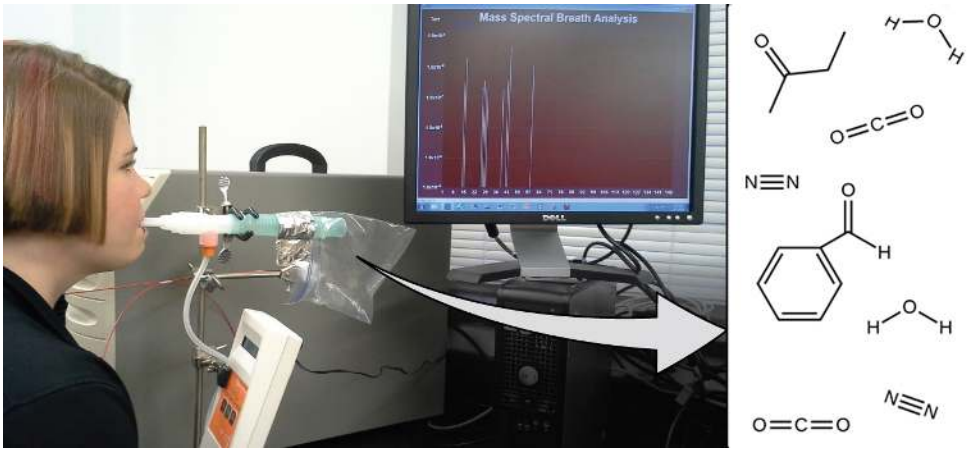


Figure 2.1.1: Analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers)

Early detection of biomarkers, substances that indicate an organism's disease or physiological state, could allow diagnosis and treatment before a condition becomes serious or irreversible. Recent studies have shown that your exhaled breath can contain molecules that may be biomarkers for recent exposure to environmental contaminants or for pathological conditions ranging from asthma to lung cancer. Scientists are working to develop biomarker "fingerprints" that could be used to diagnose a specific disease based on the amounts and identities of certain molecules in a patient's exhaled breath. In Sangeeta Bhatia's lab at MIT, a team used substances that react specifically inside diseased lung tissue; the products of the reactions will be present as biomarkers that can be identified through mass spectrometry (an analytical method discussed later in the chapter). A potential application would allow patients with early symptoms to inhale or ingest a "sensor" substance, and, minutes later, to breathe into a detector for diagnosis. Similar research by scientists such as Laura López-Sánchez has provided similar processes for lung cancer. An essential concept underlying this goal is that of a molecule's identity, which is determined by the numbers and types of atoms it contains, and how they are bonded together. This chapter will describe some of the fundamental chemical principles related to the composition of matter, including those central to the concept of molecular identity.

This page titled 2.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **2.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.2: Early Ideas in Atomic Theory

Learning Objectives

By the end of this section, you will be able to:

- State the postulates of Dalton's atomic theory
- Use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions

The earliest recorded discussion of the basic structure of matter comes from ancient Greek philosophers, the scientists of their day. In the fifth century BC, Leucippus and Democritus argued that all matter was composed of small, finite particles that they called *atomos*, a term derived from the Greek word for "indivisible." They thought of atoms as moving particles that differed in shape and size, and which could join together. Later, Aristotle and others came to the conclusion that matter consisted of various combinations of the four "elements"—fire, earth, air, and water—and could be infinitely divided. Interestingly, these philosophers thought about atoms and "elements" as philosophical concepts, but apparently never considered performing experiments to test their ideas.

The Aristotelian view of the composition of matter held sway for over two thousand years, until English schoolteacher John Dalton helped to revolutionize chemistry with his hypothesis that the behavior of matter could be explained using an atomic theory. First published in 1807, many of Dalton's hypotheses about the microscopic features of matter are still valid in modern atomic theory. Here are the postulates of Dalton's atomic theory.

1. Matter is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.
2. An element consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element (Figure 2.2.1). A macroscopic sample of an element contains an incredibly large number of atoms, all of which have identical chemical properties.

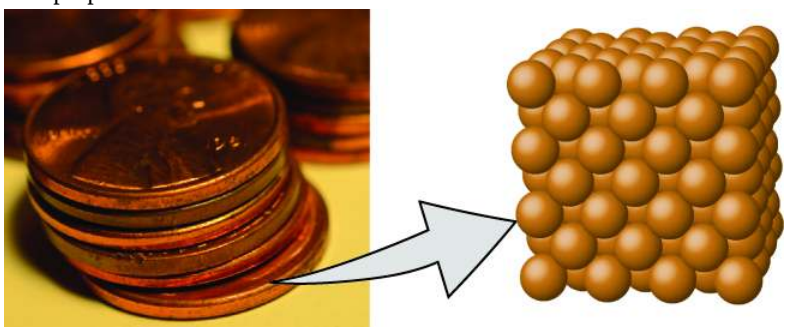


Figure 2.2.1: A pre-1982 copper penny (left) contains approximately 3 10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by "slgckgc"/Flickr)

3. Atoms of one element differ in properties from atoms of all other elements.
4. A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the numbers of atoms of each of its elements are always present in the same ratio (Figure 2.2.2).

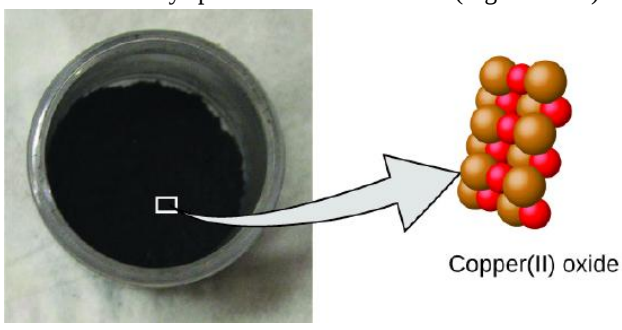


Figure 2.2.2: Copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown spheres) and oxygen (red spheres)—in a 1:1 ratio. (credit: modification of work by "Chemicalinterest"/Wikimedia Commons)

5. Atoms are neither created nor destroyed during a chemical change, but are instead rearranged to yield substances that are different from those present before the change (Figure 2.2.3).

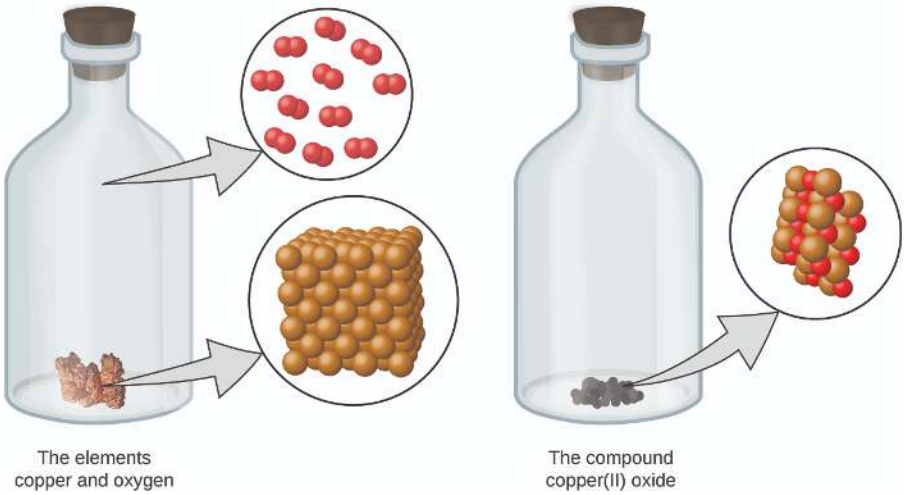


Figure 2.2.3: When the elements copper (a shiny, red-brown solid, shown here as brown spheres) and oxygen (a clear and colorless gas, shown here as red spheres) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work by http://images-of-elements.com/copper.php)

Dalton's atomic theory provides a microscopic explanation of the many macroscopic properties of matter that you've learned about. For example, if an element such as copper consists of only one kind of atom, then it cannot be broken down into simpler substances, that is, into substances composed of fewer types of atoms. And if atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant (the law of conservation of matter).

✔ Example 2.2.1: Testing Dalton's Atomic Theory

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one?



Solution

The starting materials consist of two green spheres and two purple spheres. The products consist of only one green sphere and one purple sphere. This violates Dalton's postulate that atoms are neither created nor destroyed during a chemical change, but are merely redistributed. (In this case, atoms appear to have been destroyed.)

? Exercise 2.2.1

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one?



Answer

The starting materials consist of four green spheres and two purple spheres. The products consist of four green spheres and two purple spheres. This does not violate any of Dalton's postulates: Atoms are neither created nor destroyed, but are redistributed in small, whole-number ratios.

Dalton knew of the experiments of French chemist Joseph Proust, who demonstrated that *all samples of a pure compound contain the same elements in the same proportion by mass*. This statement is known as the law of definite proportions or the law of constant composition. The suggestion that the numbers of atoms of the elements in a given compound always exist in the same ratio is consistent with these observations. For example, when different samples of isooctane (a component of gasoline and one of the standards used in the octane rating system) are analyzed, they are found to have a carbon-to-hydrogen mass ratio of 5.33:1, as shown in Table 2.1.

Table 2.1: Constant Composition of Isooctane

Sample	Carbon	Hydrogen	Mass Ratio
A	14.82 g	2.78 g	
B	22.33 g	4.19 g	
C	19.40 g	3.64 g	

It is worth noting that although all samples of a particular compound have the same mass ratio, the converse is not true in general. That is, samples that have the same mass ratio are not necessarily the same substance. For example, there are many compounds other than isooctane that also have a carbon-to-hydrogen mass ratio of 5.33:1.00.

Dalton also used data from Proust, as well as results from his own experiments, to formulate another interesting law. The law of multiple proportions states that *when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small, whole numbers*. For example, copper and chlorine can form a green, crystalline solid with a mass ratio of 0.558 g chlorine to 1 g copper, as well as a brown crystalline solid with a mass ratio of 1.116 g chlorine to 1 g copper. These ratios by themselves may not seem particularly interesting or informative; however, if we take a ratio of these ratios, we obtain a useful and possibly surprising result: a small, whole-number ratio.

This 2-to-1 ratio means that the brown compound has twice the amount of chlorine per amount of copper as the green compound.

This can be explained by atomic theory if the copper-to-chlorine ratio in the brown compound is 1 copper atom to 2 chlorine atoms, and the ratio in the green compound is 1 copper atom to 1 chlorine atom. The ratio of chlorine atoms (and thus the ratio of their masses) is therefore 2 to 1 (Figure 2.2.4).

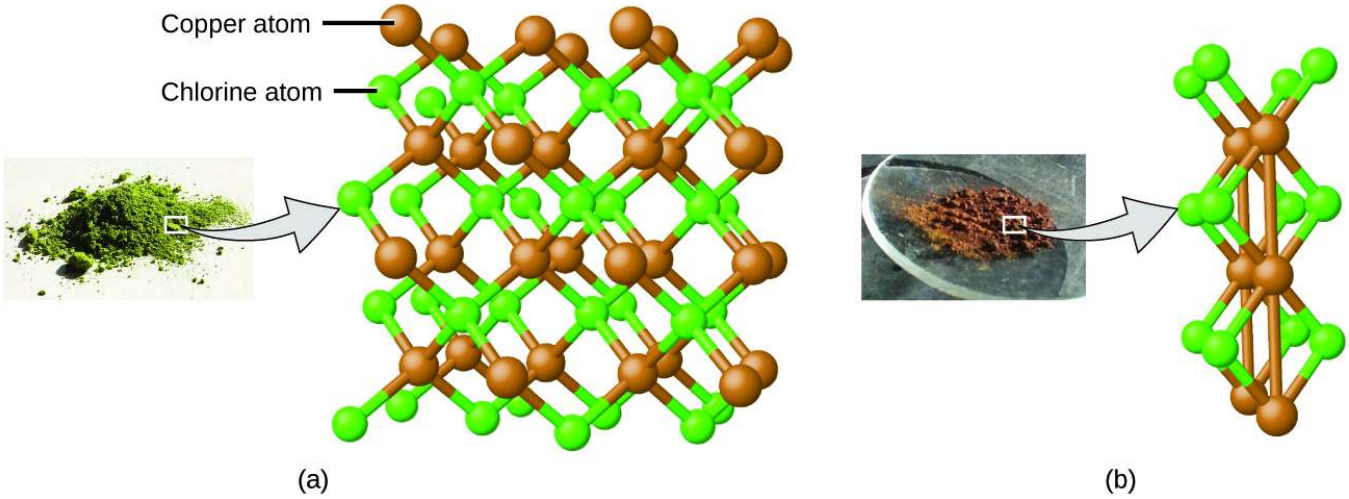


Figure 2.2.4: Compared to the copper chlorine compound in (a), where copper is represented by brown spheres and chlorine by green spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by "Benjah-bmm27"/Wikimedia Commons; credit b: modification of work by "Walkerma"/Wikimedia Commons)

Example 2.2.2: Laws of Definite and Multiple Proportions

A sample of compound A (a clear, colorless gas) is analyzed and found to contain 4.27 g carbon and 5.69 g oxygen. A sample of compound B (also a clear, colorless gas) is analyzed and found to contain 5.19 g carbon and 13.84 g oxygen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances A and B?

Solution

In compound A, the mass ratio of oxygen to carbon is:

In compound B, the mass ratio of oxygen to carbon is:

The ratio of these ratios is:

This supports the law of multiple proportions. This means that A and B are different compounds, with A having one-half as much oxygen per amount of carbon (or twice as much carbon per amount of oxygen) as B. A possible pair of compounds that would fit this relationship would be A = CO and B = CO_2.

Exercise 2.2.1

A sample of compound X (a clear, colorless, combustible liquid with a noticeable odor) is analyzed and found to contain 14.13 g carbon and 2.96 g hydrogen. A sample of compound Y (a clear, colorless, combustible liquid with a noticeable odor that is slightly different from X's odor) is analyzed and found to contain 19.91 g carbon and 3.34 g hydrogen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances X and Y?

Answer

In compound X, the mass ratio of carbon to hydrogen is In compound Y, the mass ratio of carbon to hydrogen is The ratio of these ratios is This small, whole-number ratio supports the law of multiple proportions. This means that X and Y are different compounds.

This page titled 2.2: Early Ideas in Atomic Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **2.1: Early Ideas in Atomic Theory** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.3: Evolution of Atomic Theory

Learning Objectives

By the end of this section, you will be able to:

- Outline milestones in the development of modern atomic theory
- Summarize and interpret the results of the experiments of Thomson, Millikan, and Rutherford
- Describe the three subatomic particles that compose atoms
- Define isotopes and give examples for several elements

If matter is composed of atoms, what are atoms composed of? Are they the smallest particles, or is there something smaller? In the late 1800s, a number of scientists interested in questions like these investigated the electrical discharges that could be produced in low-pressure gases, with the most significant discovery made by English physicist J. J. Thomson using a cathode ray tube. This apparatus consisted of a sealed glass tube from which almost all the air had been removed; the tube contained two metal electrodes. When high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them. This beam was deflected toward the positive charge and away from the negative charge, and was produced in the same way with identical properties when different metals were used for the electrodes. In similar experiments, the ray was simultaneously deflected by an applied magnetic field, and measurements of the extent of deflection and the magnetic field strength allowed Thomson to calculate the charge-to-mass ratio of the cathode ray particles. The results of these measurements indicated that these particles were much lighter than atoms (Figure 2.3.1).

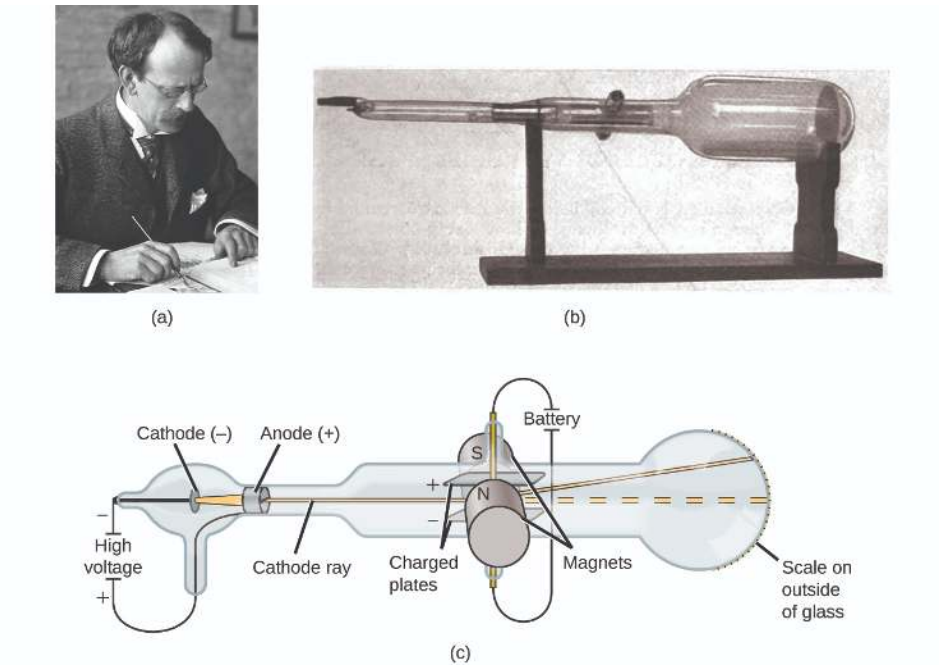


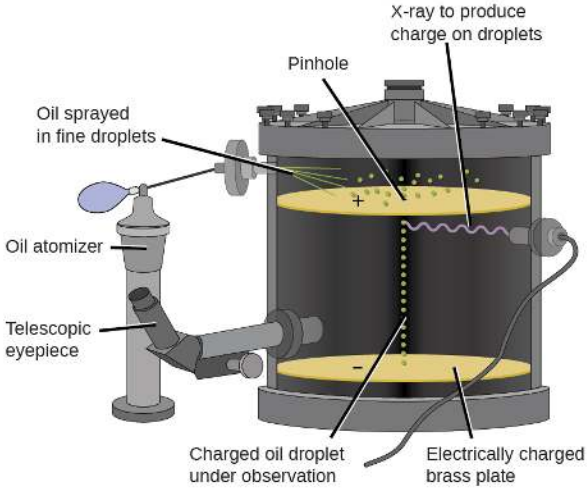
Figure 2.3.1: (a) J. J. Thomson produced a visible beam in a cathode ray tube. (b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun. (c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end of the tube. Simultaneous deflections by applied electric and magnetic fields permitted Thomson to calculate the mass-to-charge ratio of the particles composing the cathode ray. (credit a: modification of work by Nobel Foundation; credit b: modification of work by Eugen Nesper; credit c: modification of work by "Kurzon"/Wikimedia Commons)

Based on his observations, here is what Thomson proposed and why: The particles are attracted by positive (+) charges and repelled by negative (−) charges, so they must be negatively charged (like charges repel and unlike charges attract); they are less massive than atoms and indistinguishable, regardless of the source material, so they must be fundamental, subatomic constituents of all atoms. Although controversial at the time, Thomson's idea was gradually accepted, and his cathode ray particle is what we now call an electron, a negatively charged, subatomic particle with a mass more than one thousand-times less that of an atom. The term "electron" was coined in 1891 by Irish physicist George Stoney, from "*electr*ic i*on*."

Link to Learning

Click here to hear Thomson describe his discovery in his own voice.

In 1909, more information about the electron was uncovered by American physicist Robert A. Millikan via his "oil drop" experiments. Millikan created microscopic oil droplets, which could be electrically charged by friction as they formed or by using X-rays. These droplets initially fell due to gravity, but their downward progress could be slowed or even reversed by an electric field lower in the apparatus. By adjusting the electric field strength and making careful measurements and appropriate calculations, Millikan was able to determine the charge on individual drops (Figure 2.3.2).



Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Figure 2.3.2: Millikan's experiment measured the charge of individual oil drops. The tabulated data are examples of a few possible values.

Looking at the charge data that Millikan gathered, you may have recognized that the charge of an oil droplet is always a multiple of a specific charge, 1.6 10^{-19} C. Millikan concluded that this value must therefore be a fundamental charge—the charge of a single electron—with his measured charges due to an excess of one electron (1 times 1.6 10^{-19} C), two electrons (2 times 1.6 10^{-19} C), three electrons (3 times 1.6 10^{-19} C), and so on, on a given oil droplet. Since the charge of an electron was now known due to Millikan's research, and the charge-to-mass ratio was already known due to Thomson's research (1.759 10^{11} C/kg), it only required a simple calculation to determine the mass of the electron as well.

Scientists had now established that the atom was not indivisible as Dalton had believed, and due to the work of Thomson, Millikan, and others, the charge and mass of the negative, subatomic particles—the electrons—were known. However, the positively charged part of an atom was not yet well understood. In 1904, Thomson proposed the "plum pudding" model of atoms, which described a positively charged mass with an equal amount of negative charge in the form of electrons embedded in it, since all atoms are electrically neutral. A competing model had been proposed in 1903 by Hantaro Nagaoka, who postulated a Saturn-like atom, consisting of a positively charged sphere surrounded by a halo of electrons (Figure 2.3.3).

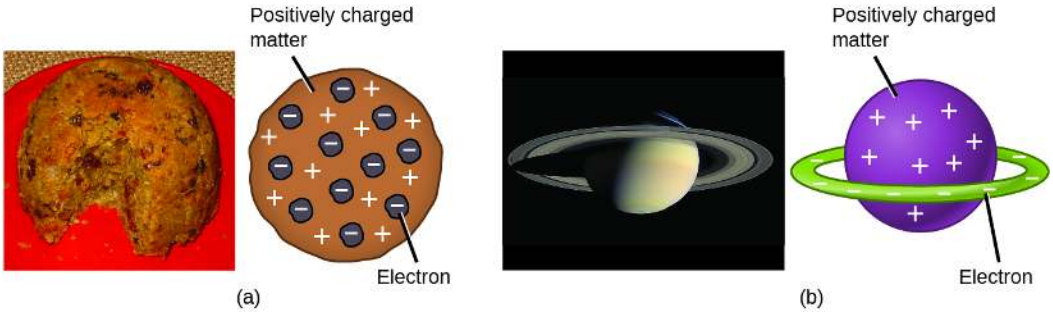


Figure 2.3.3: (a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins ("plums"). (b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive "planet." (credit a: modification of work by "Man vyi"/Wikimedia Commons; credit b: modification of work by "NASA"/Wikimedia Commons)

The next major development in understanding the atom came from Ernest Rutherford, a physicist from New Zealand who largely spent his scientific career in Canada and England. He performed a series of experiments using a beam of high-speed, positively charged alpha particles (α particles) that were produced by the radioactive decay of radium; α particles consist of two protons and two neutrons (you will learn more about radioactive decay in the chapter on nuclear chemistry). Rutherford and his colleagues Hans Geiger (later famous for the Geiger counter) and Ernest Marsden aimed a beam of α particles, the source of which was embedded in a lead block to absorb most of the radiation, at a very thin piece of gold foil and examined the resultant scattering of the α particles using a luminescent screen that glowed briefly where hit by an α particle.

What did they discover? Most particles passed right through the foil without being deflected at all. However, some were diverted slightly, and a very small number were deflected almost straight back toward the source (Figure 2.3.4). Rutherford described finding these results: "It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you."[1]

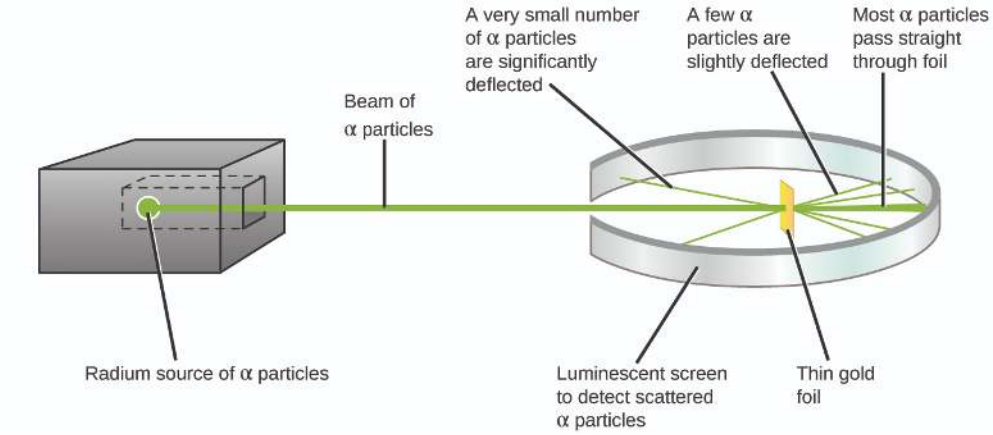


Figure 2.3.4: Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected.

Here is what Rutherford deduced: Because most of the fast-moving α particles passed through the gold atoms undeflected, they must have traveled through essentially empty space inside the atom. Alpha particles are positively charged, so deflections arose when they encountered another positive charge (like charges repel each other). Since like charges repel one another, the few positively charged α particles that changed paths abruptly must have hit, or closely approached, another body that also had a highly concentrated, positive charge. Since the deflections occurred a small fraction of the time, this charge only occupied a small amount of the space in the gold foil. Analyzing a series of such experiments in detail, Rutherford drew two conclusions:

1. The volume occupied by an atom must consist of a large amount of empty space.
2. A small, relatively heavy, positively charged body, the nucleus, must be at the center of each atom.

Link to Learning

View this simulation of the Rutherford gold foil experiment. Adjust the slit width to produce a narrower or broader beam of α particles to see how that affects the scattering pattern.

This analysis led Rutherford to propose a model in which an atom consists of a very small, positively charged nucleus, in which most of the mass of the atom is concentrated, surrounded by the negatively charged electrons, so that the atom is electrically neutral (Figure 2.3.5). After many more experiments, Rutherford also discovered that the nuclei of other elements contain the hydrogen nucleus as a "building block," and he named this more fundamental particle the proton, the positively charged, subatomic particle found in the nucleus. With one addition, which you will learn next, this nuclear model of the atom, proposed over a century ago, is still used today.

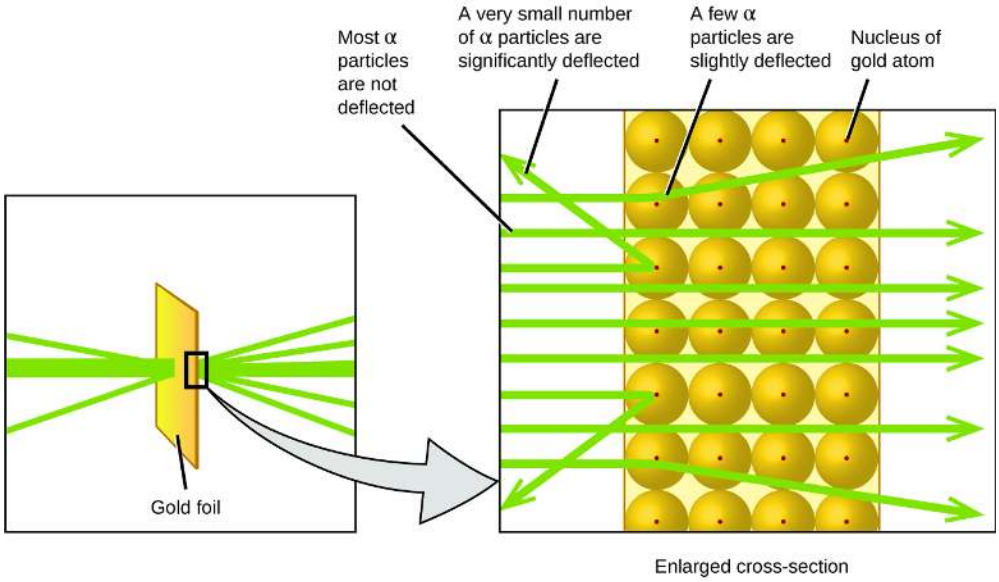


Figure 2.3.5: The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by electrons, which are too light to deflect the rapidly moving particles.

Link to Learning

The Rutherford Scattering simulation allows you to investigate the differences between a "plum pudding" atom and a Rutherford atom by firing α particles at each type of atom.

Another important finding was the discovery of isotopes. During the early 1900s, scientists identified several substances that appeared to be new elements, isolating them from radioactive ores. For example, a "new element" produced by the radioactive decay of thorium was initially given the name mesothorium. However, a more detailed analysis showed that mesothorium was chemically identical to radium (another decay product), despite having a different atomic mass. This result, along with similar findings for other elements, led the English chemist Frederick Soddy to realize that an element could have types of atoms with different masses that were chemically indistinguishable. These different types are called isotopes—atoms of the same element that differ in mass. Soddy was awarded the Nobel Prize in Chemistry in 1921 for this discovery.

One puzzle remained: The nucleus was known to contain almost all of the mass of an atom, with the number of protons only providing half, or less, of that mass. Different proposals were made to explain what constituted the remaining mass, including the existence of neutral particles in the nucleus. As you might expect, detecting uncharged particles is very challenging, and it was not until 1932 that James Chadwick found evidence of neutrons, uncharged, subatomic particles with a mass approximately the same as that of protons. The existence of the neutron also explained isotopes: They differ in mass because they have different numbers of neutrons, but they are chemically identical because they have the same number of protons. This will be explained in more detail later in this chapter.

Footnotes

- 1Ernest Rutherford, "The Development of the Theory of Atomic Structure," ed. J. A. Ratcliffe, in *Background to Modern Science*, eds. Joseph Needham and Walter Pagel, (Cambridge, UK: Cambridge University Press, 1938), 61–74. Accessed September 22, 2014, https://ia600508.us.archive.org/3/it...e032734mbp.pdf.

This page titled 2.3: Evolution of Atomic Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **2.2: Evolution of Atomic Theory** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.4: Atomic Structure and Symbolism

Learning Objectives

By the end of this section, you will be able to:

- Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion
- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure 2.4.1).



Figure 2.4.1: If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by "babyknight"/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

Atoms—and the protons, neutrons, and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g, and an electron has a charge of less than 2×10^{-19} C (coulomb). When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the atomic mass unit (amu) and the fundamental unit of charge (e). The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as "carbon-12" as will be discussed later in this module.) Thus, one amu is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom: 1 amu = 1.6605×10^{-24} g. (The **Dalton** (Da) and the **unified atomic mass** unit (u) are alternative units that are equivalent to the amu.) The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with e = 1.602×10^{-19} C.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1− and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton). The properties of these fundamental particles are summarized in Table 2.4.1. (An observant student might notice that the sum of an atom's subatomic particles does not equal the atom's actual mass: The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This "missing" mass is known as the mass defect, and you will learn about it in the chapter on nuclear chemistry.)

Table 2.4.1: Properties of Subatomic Particles

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	$-1.602\ 10^{-19}$	1−	0.00055	$0.00091\ 10^{-24}$
proton	nucleus	$1.602\ 10^{-19}$	1+	1.00727	$1.67262\ 10^{-24}$

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
neutron	nucleus	0	0	1.00866	$1.67493\ 10^{-24}$

The number of protons in the nucleus of an atom is its **atomic number** (Z). This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number** (A). The number of neutrons is therefore the difference between the mass number and the atomic number: A – Z = number of neutrons.

$$\begin{aligned} \text{atomic number } (Z) &= \text{ number of protons} \\ \text{mass number } (A) &= \text{ number of protons } + \text{ number of neutrons} \\ A - Z &= \text{ number of neutrons} \end{aligned}$$

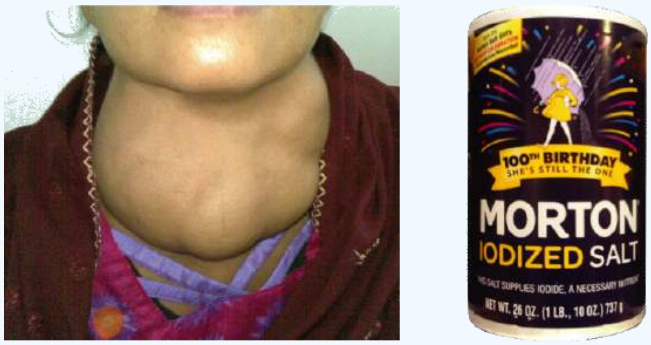
Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an **ion**. The charge of an atom is defined as follows:

$$\text{Atomic charge} = \text{number of protons} - \text{number of electrons}$$

As will be discussed in more detail later in this chapter, atoms (and molecules) typically acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. Positively charged atoms called **cations** are formed when an atom loses one or more electrons. For example, a neutral sodium atom (Z = 11) has 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge (11 − 10 = 1+). A neutral oxygen atom (Z = 8) has eight electrons, and if it gains two electrons it will become an anion with a 2− charge (8 − 10 = 2−).

Example 2.4.1: Composition of an Atom

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland (Figure 2.4.2).



(a) (b)

Figure 2.4.2: (a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

The addition of small amounts of iodine to table salt (iodized salt) has essentially eliminated this health concern in the United States, but as much as 40% of the world's population is still at risk of iodine deficiency. The iodine atoms are added as anions, and each has a 1− charge and a mass number of 127. Determine the numbers of protons, neutrons, and electrons in one of these iodine anions.

Solution

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the numbers of protons and neutrons equals the mass number, 127, the number of neutrons is

74 (127 − 53 = 74). Since the iodine is added as a 1− anion, the number of electrons is 54 [53 – (1–) = 54].

? Exercise 2.4.1

An ion of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

Answer

78 protons; 117 neutrons; charge is 4+

Chemical Symbols

A chemical symbol is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg (Figure 2.4.3). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).



Figure 2.4.3: The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

The symbols for several common elements and their atoms are listed in Table 2.4.2. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All known elements and their symbols are in the periodic table (also found in Appendix A).

Table 2.4.2: Some Common Elements and Their Symbols

Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from *ferrum*)
bromine	Br	lead	Pb (from *plumbum*)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from *hydrargyrum*)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from *kalium*)
copper	Cu (from *cuprum*)	silicon	Si
fluorine	F	silver	Ag (from *argentum*)
gold	Au (from *aurum*)	sodium	Na (from *natrium*)

Element	Symbol	Element	Symbol
helium	He	sulfur	S
hydrogen	H	tin	Sn (from *stannum*)
iodine	I	zinc	Zn

Traditionally, the discoverer (or discoverers) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists (or occasionally locations); for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements. Element 109 was named in honor of Lise Meitner, who discovered nuclear fission, a phenomenon that would have world-changing impacts; Meitner also contributed to the discovery of some major isotopes, discussed immediately below.

Link to Learning

Visit this site to learn more about IUPAC, the International Union of Pure and Applied Chemistry, and explore its periodic table.

Isotopes

The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol (Figure 2.4.4). The atomic number is sometimes written as a subscript preceding the symbol, but since this number defines the element's identity, as does its symbol, it is often omitted. For example, magnesium exists as a mixture of three isotopes, each with an atomic number of 12 and with mass numbers of 24, 25, and 26, respectively. These isotopes can be identified as ^{24}Mg, ^{25}Mg, and ^{26}Mg. These isotope symbols are read as "element, mass number" and can be symbolized consistent with this reading. For instance, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24." ^{25}Mg is read as "magnesium 25," and can be written as "magnesium-25" or "Mg-25." All magnesium atoms have 12 protons in their nucleus. They differ only because a ^{24}Mg atom has 12 neutrons in its nucleus, a ^{25}Mg atom has 13 neutrons, and a ^{26}Mg has 14 neutrons.

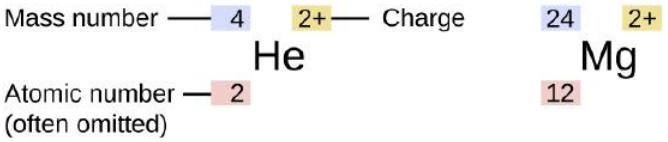


Figure 2.4.4: The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

Information about the naturally occurring isotopes of elements with atomic numbers 1 through 10 is given in Table 2.4.3. Note that in addition to standard names and symbols, the isotopes of hydrogen are often referred to using common names and accompanying symbols. Hydrogen-2, symbolized ^{2}H, is also called deuterium and sometimes symbolized D. Hydrogen-3, symbolized ^{3}H, is also called tritium and sometimes symbolized T.

Table 2.4.3: Nuclear Compositions of Atoms of the Very Light Elements

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	$^{1}_{1}$H (protium)	1	1	0	1.0078	99.989
	$^{2}_{1}$H (deuterium)	1	1	1	2.0141	0.0115
	$^{3}_{1}$H (tritium)	1	1	2	3.01605	— (trace)

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
helium	$^{3}_{2}\mathrm{He}$	2	2	1	3.01603	0.00013
	$^{4}_{2}\mathrm{He}$	2	2	2	4.0026	100
lithium	$^{6}_{3}\mathrm{Li}$	3	3	3	6.0151	7.59
	$^{7}_{3}\mathrm{Li}$	3	3	4	7.0160	92.41
beryllium	$^{9}_{4}\mathrm{Be}$	4	4	5	9.0122	100
boron	$^{10}_{5}\mathrm{B}$	5	5	5	10.0129	19.9
	$^{11}_{5}\mathrm{B}$	5	5	6	11.0093	80.1
carbon	$^{12}_{6}\mathrm{C}$	6	6	6	12.0000	98.89
	$^{13}_{6}\mathrm{C}$	6	6	7	13.0034	1.11
	$^{14}_{6}\mathrm{C}$	6	6	8	14.0032	— (trace)
nitrogen	$^{14}_{7}\mathrm{N}$	7	7	7	14.0031	99.63
	$^{15}_{7}\mathrm{N}$	7	7	8	15.0001	0.37
oxygen	$^{16}_{8}\mathrm{O}$	8	8	8	15.9949	99.757
	$^{17}_{8}\mathrm{O}$	8	8	9	16.9991	0.038
	$^{18}_{8}\mathrm{O}$	8	8	10	17.9992	0.205
fluorine	$^{19}_{9}\mathrm{F}$	9	9	10	18.9984	100
neon	$^{20}_{10}\mathrm{Ne}$	10	10	10	19.9924	90.48
	$^{21}_{10}\mathrm{Ne}$	10	10	11	20.9938	0.27
	$^{22}_{10}\mathrm{Ne}$	10	10	12	21.9914	9.25

Link to Learning

Use this Build an Atom simulator to build atoms of the first 10 elements, see which isotopes exist, check nuclear stability, and gain experience with isotope symbols.

Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the **atomic mass** of a single atom is approximately equal to its mass number (a whole number). However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

$$\text{average mass} = \sum_i (\text{ fractional abundance} \times \text{isotopic mass})_i$$

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are $^{10}\mathrm{B}$ with a mass of 10.0129 amu, and the remaining 80.1% are $^{11}\mathrm{B}$ with a mass of 11.0093 amu. The average atomic mass for boron is calculated to be:

$$\begin{aligned}\text{boron average mass} &= (0.199 \times 10.0129\ \text{amu}) + (0.801 \times 11.0093\ \text{amu}) \\ &= 1.99\ \text{amu} + 8.82\ \text{amu} \\ &= 10.81\ \text{amu}\end{aligned}$$

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

Example 2.4.2: Calculation of Average Atomic Mass

A meteorite found in central Indiana contains traces of the noble gas neon picked up from the solar wind during the meteorite's trip through the solar system. Analysis of a sample of the gas showed that it consisted of 91.84% ^{20}Ne (mass 19.9924 amu), 0.47% ^{21}Ne (mass 20.9940 amu), and 7.69% ^{22}Ne (mass 21.9914 amu). What is the average mass of the neon in the solar wind?

Solution

$$\begin{aligned}\text{average mass} &= (0.9184 \times 19.9924\ \text{amu}) + (0.0047 \times 20.9940\ \text{amu}) + (0.0769 \times 21.9914\ \text{amu} \\ &= (18.36 + 0.099 + 1.69)\ \text{amu} \\ &= 20.15\ \text{amu}\end{aligned}$$

The average mass of a neon atom in the solar wind is 20.15 amu. (The average mass of a terrestrial neon atom is 20.1796 amu. This result demonstrates that we may find slight differences in the natural abundance of isotopes, depending on their origin.)

Exercise 2.4.2

A sample of magnesium is found to contain 78.70% of ^{24}Mg atoms (mass 23.98 amu), 10.13% of ^{25}Mg atoms (mass 24.99 amu), and 11.17% of ^{26}Mg atoms (mass 25.98 amu). Calculate the average mass of a Mg atom.

Answer

24.31 amu

We can also do variations of this type of calculation, as shown in the next example.

Example 2.4.3: Calculation of Percent Abundance

Naturally occurring chlorine consists of ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent composition of Cl in terms of these two isotopes?

Solution

The average mass of chlorine is the fraction that is ^{35}Cl times the mass of ^{35}Cl plus the fraction that is ^{37}Cl times the mass of ^{37}Cl.

$$\text{average mass} = \left(\text{fraction of } {}^{35}Cl \times \text{ mass of } {}^{35}Cl\right) + \left(\text{fraction of } {}^{37}Cl \times \text{ mass of } {}^{37}Cl\right)$$

If we let *x* represent the fraction that is ^{35}Cl, then the fraction that is ^{37}Cl is represented by 1.00 − *x*.

(The fraction that is ^{35}Cl + the fraction that is ^{37}Cl must add up to 1, so the fraction of ^{37}Cl must equal 1.00 − the fraction of ^{35}Cl.)

Substituting this into the average mass equation, we have:

$$\begin{aligned}35.453amu &= (x \times 34.96885amu) + [(1.00 - x) \times 36.96590amu] \\ 35.453 &= 34.96885x + 36.96590 - 36.96590x \\ 1.99705x &= 1.513 \\ x &= \frac{1.513}{1.99705} = 0.7576\end{aligned}$$

So solving yields: $x = 0.7576$, which means that $1.00 - 0.7576 = 0.2424$. Therefore, chlorine consists of 75.76% ^{35}Cl and 24.24% ^{37}Cl.

Exercise 2.4.3

Naturally occurring copper consists of ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu. What is the percent composition of Cu in terms of these two isotopes?

Answer

69.15% Cu-63 and 30.85% Cu-65

Link to Learning

Visit this site to make mixtures of the main isotopes of the first 18 elements, gain experience with average atomic mass, and check naturally occurring isotope ratios using the Isotopes and Atomic Mass simulation.

As you will learn, isotopes are important in nature and especially in human understanding of science and medicine. Let's consider just one natural, stable isotope: Oxygen-18, which is noted in the table above and is referred to as one of the environmental isotopes. It is important in paleoclimatology, for example, because scientists can use the ratio between Oxygen-18 and Oxygen-16 in an ice core to determine the temperature of precipitation over time. Oxygen-18 was also critical to the discovery of metabolic pathways and the mechanisms of enzymes. Mildred Cohn pioneered the usage of these isotopes to act as tracers, so that researchers could follow their path through reactions and gain a better understanding of what is happening. One of her first discoveries provided insight into the phosphorylation of glucose that takes place in mitochondria. And the methods of using isotopes for this research contributed to entire fields of study.

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer. **Mass spectrometry** (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer (Figure 2.4.5), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons. These cations then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge (similar to how the path of a large steel ball rolling past a magnet is deflected to a lesser extent that that of a small steel ball). The ions are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis in a wide range of applications.

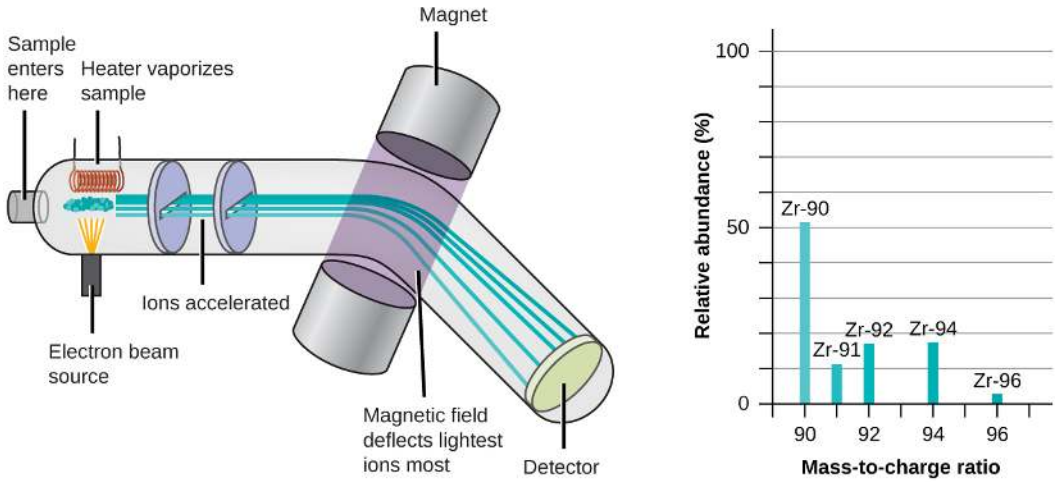


Figure 2.4.5: Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

Link to Learning

See an animation that explains mass spectrometry. Watch this video from the Royal Society for Chemistry for a brief description of the rudiments of mass spectrometry.

This page titled 2.4: Atomic Structure and Symbolism is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **2.3: Atomic Structure and Symbolism** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.5: Chemical Formulas

Learning Objectives

By the end of this section, you will be able to:

- Symbolize the composition of molecules using molecular formulas and empirical formulas
- Represent the bonding arrangement of atoms within molecules using structural formulas

A molecular formula is a representation of a molecule that uses chemical symbols to indicate the types of atoms followed by subscripts to show the number of atoms of each type in the molecule. (A subscript is used only when more than one atom of a given type is present.) Molecular formulas are also used as abbreviations for the names of compounds.

The structural formula for a compound gives the same information as its molecular formula (the types and numbers of atoms in the molecule) but also shows how the atoms are connected in the molecule. The structural formula for methane contains symbols for one C atom and four H atoms, indicating the number of atoms in the molecule (Figure 2.5.1). The lines represent bonds that hold the atoms together. (A chemical bond is an attraction between atoms or ions that holds them together in a molecule or a crystal.) We will discuss chemical bonds and see how to predict the arrangement of atoms in a molecule later. For now, simply know that the lines are an indication of how the atoms are connected in a molecule. A ball-and-stick model shows the geometric arrangement of the atoms with atomic sizes not to scale, and a space-filling model shows the relative sizes of the atoms.

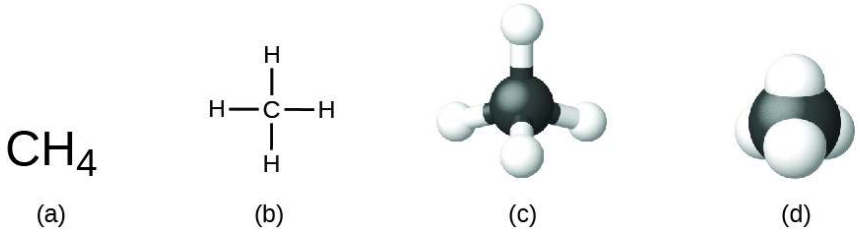


Figure 2.5.1 : A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

Although many elements consist of discrete, individual atoms, some exist as molecules made up of two or more atoms of the element chemically bonded together. For example, most samples of the elements hydrogen, oxygen, and nitrogen are composed of molecules that contain two atoms each (called diatomic molecules) and thus have the molecular formulas H_2, O_2, and N_2, respectively. Other elements commonly found as diatomic molecules are fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S_8 (Figure 2.5.2).

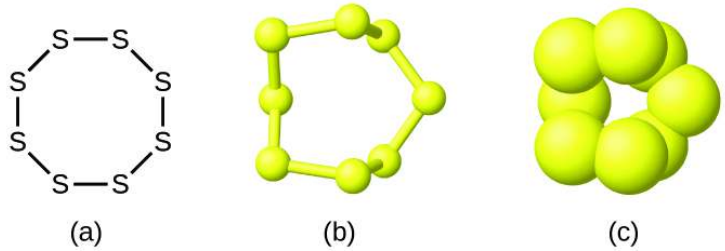


Figure 2.5.2 : A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and 2H represent distinctly different species. H_2 is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression 2H, on the other hand, indicates two separate hydrogen atoms that are not combined as a unit. The expression $2H_2$ represents two molecules of diatomic hydrogen (Figure 2.5.3).

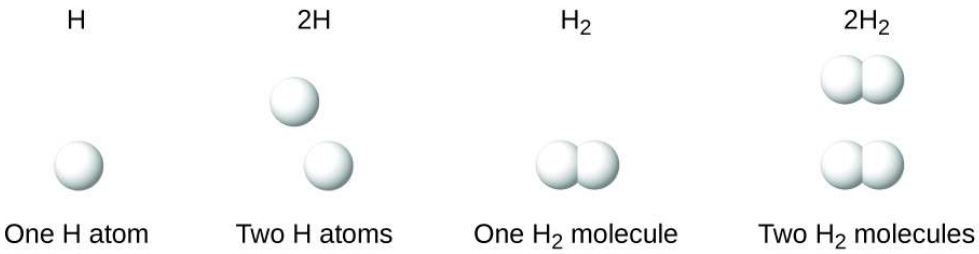


Figure 2.5.3 : The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

Compounds are formed when two or more elements chemically combine, resulting in the formation of bonds. For example, hydrogen and oxygen can react to form water, and sodium and chlorine can react to form table salt. We sometimes describe the composition of these compounds with an empirical formula, which indicates the types of atoms present and *the simplest whole-number ratio of the number of atoms (or ions) in the compound*. For example, titanium dioxide (used as pigment in white paint and in the thick, white, blocking type of sunscreen) has an empirical formula of TiO_2. This identifies the elements titanium (Ti) and oxygen (O) as the constituents of titanium dioxide, and indicates the presence of twice as many atoms of the element oxygen as atoms of the element titanium (Figure 2.5.4).

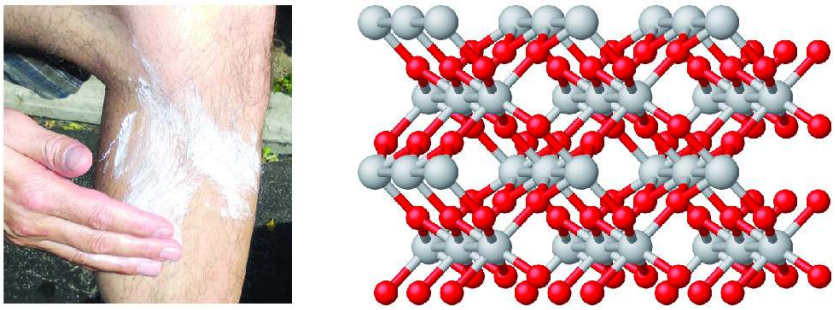


Figure 2.5.4 : (a) The white compound titanium dioxide provides effective protection from the sun. (b) A crystal of titanium dioxide, TiO_2, contains titanium and oxygen in a ratio of 1 to 2. The titanium atoms are gray and the oxygen atoms are red. (credit a: modification of work by “osseous”/Flickr)

As discussed previously, we can describe a compound with a molecular formula, in which the subscripts indicate the *actual numbers of atoms* of each element in a molecule of the compound. In many cases, the molecular formula of a substance is derived from experimental determination of both its empirical formula and its molecular mass (the sum of atomic masses for all atoms composing the molecule). For example, it can be determined experimentally that benzene contains two elements, carbon (C) and hydrogen (H), and that for every carbon atom in benzene, there is one hydrogen atom. Thus, the empirical formula is CH. An experimental determination of the molecular mass reveals that a molecule of benzene contains six carbon atoms and six hydrogen atoms, so the molecular formula for benzene is C_6H_6 (Figure 2.5.5).

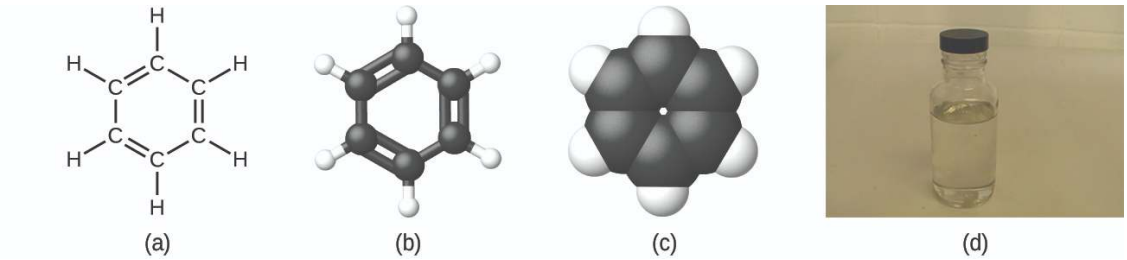
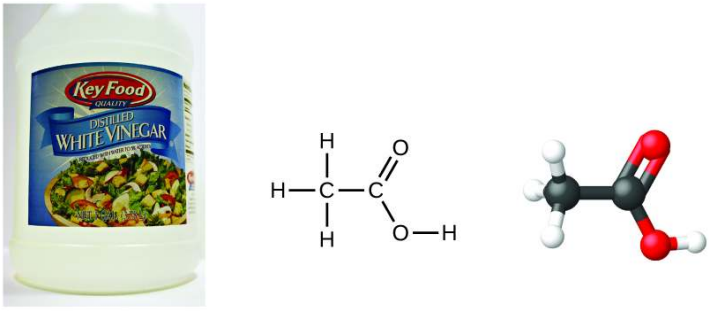


Figure 2.5.5 : Benzene, C_6H_6, is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa)

If we know a compound’s formula, we can easily determine the empirical formula. (This is somewhat of an academic exercise; the reverse chronology is generally followed in actual practice.) For example, the molecular formula for acetic acid, the component that gives vinegar its sharp taste, is $C_2H_4O_2$. This formula indicates that a molecule of acetic acid (Figure 2.5.6) contains two carbon atoms, four hydrogen atoms, and two oxygen atoms. The ratio of atoms is 2:4:2. Dividing by the lowest common

denominator (2) gives the simplest, whole-number ratio of atoms, 1:2:1, so the empirical formula is CH_2O. Note that a molecular formula is always a whole-number multiple of an empirical formula.



(a) (b) (c)

Figure 2.5.6: (a) Vinegar contains acetic acid, $C_2H_4O_2$, which has an empirical formula of CH_2O. It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by "HomeSpot HQ"/Flickr)

Example 2.5.1: Empirical and Molecular Formulas

Molecules of glucose (blood sugar) contain 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. What are the molecular and empirical formulas of glucose?

Solution

The molecular formula is $C_6H_{12}O_6$ because one molecule actually contains 6 C, 12 H, and 6 O atoms. The simplest whole-number ratio of C to H to O atoms in glucose is 1:2:1, so the empirical formula is CH_2O.

Exercise 2.5.1

A molecule of metaldehyde (a pesticide used for snails and slugs) contains 8 carbon atoms, 16 hydrogen atoms, and 4 oxygen atoms. What are the molecular and empirical formulas of metaldehyde?

Answer

Molecular formula, $C_8H_{16}O_4$; empirical formula, C_2H_4O

Link to Learning

You can explore molecule building using an online simulation.

Portrait of a Chemist: Lee Cronin

What is it that chemists do? According to Lee Cronin (Figure 2.5.7), chemists make very complicated molecules by "chopping up" small molecules and "reverse engineering" them. He wonders if we could "make a really cool universal chemistry set" by what he calls "app-ing" chemistry. Could we "app" chemistry?



Figure 2.5.7: Chemist Lee Cronin has been named one of the UK's 10 most inspirational scientists. The youngest chair at the University of Glasgow, Lee runs a large research group, collaborates with many scientists worldwide, has published over 250 papers in top scientific journals, and has given more than 150 invited talks. His research focuses on complex chemical systems and their potential to transform technology, but also branches into nanoscience, solar fuels, synthetic biology, and even artificial life and evolution. (credit: image courtesy of Lee Cronin)

In a 2012 TED talk, Lee describes one fascinating possibility: combining a collection of chemical "inks" with a 3D printer capable of fabricating a reaction apparatus (tiny test tubes, beakers, and the like) to fashion a "universal toolkit of chemistry." This toolkit could be used to create custom-tailored drugs to fight a new superbug or to "print" medicine personally configured to your genetic makeup, environment, and health situation. Says Cronin, "What Apple did for music, I'd like to do for the discovery and distribution of prescription drugs."[2] View his full talk at the TED website.

It is important to be aware that it may be possible for the same atoms to be arranged in different ways: Compounds with the same molecular formula may have different atom-to-atom bonding and therefore different structures. For example, could there be another compound with the same formula as acetic acid, $C_2H_4O_2$? And if so, what would be the structure of its molecules?

If you predict that another compound with the formula $C_2H_4O_2$ could exist, then you demonstrated good chemical insight and are correct. Two C atoms, four H atoms, and two O atoms can also be arranged to form a methyl formate, which is used in manufacturing, as an insecticide, and for quick-drying finishes. Methyl formate molecules have one of the oxygen atoms between the two carbon atoms, differing from the arrangement in acetic acid molecules. Acetic acid and methyl formate are examples of isomers—compounds with the same chemical formula but different molecular structures (Figure 2.5.8). Note that this small difference in the arrangement of the atoms has a major effect on their respective chemical properties. You would certainly not want to use a solution of methyl formate as a substitute for a solution of acetic acid (vinegar) when you make salad dressing.

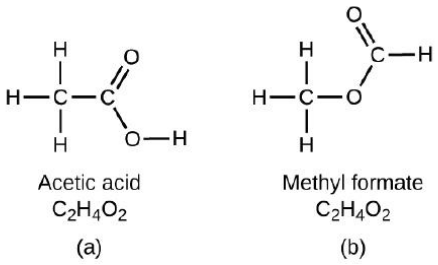


Figure 2.5.8: Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula ($C_2H_4O_2$) but different structures (and therefore different chemical properties).

Many types of isomers exist (Figure 2.5.9). Acetic acid and methyl formate are structural isomers, compounds in which the molecules differ in how the atoms are connected to each other. There are also various types of spatial isomers, in which the relative orientations of the atoms in space can be different. For example, the compound carvone (found in caraway seeds, spearmint, and mandarin orange peels) consists of two isomers that are mirror images of each other. *S*-(+)-carvone smells like caraway, and *R*-(−)-carvone smells like spearmint.

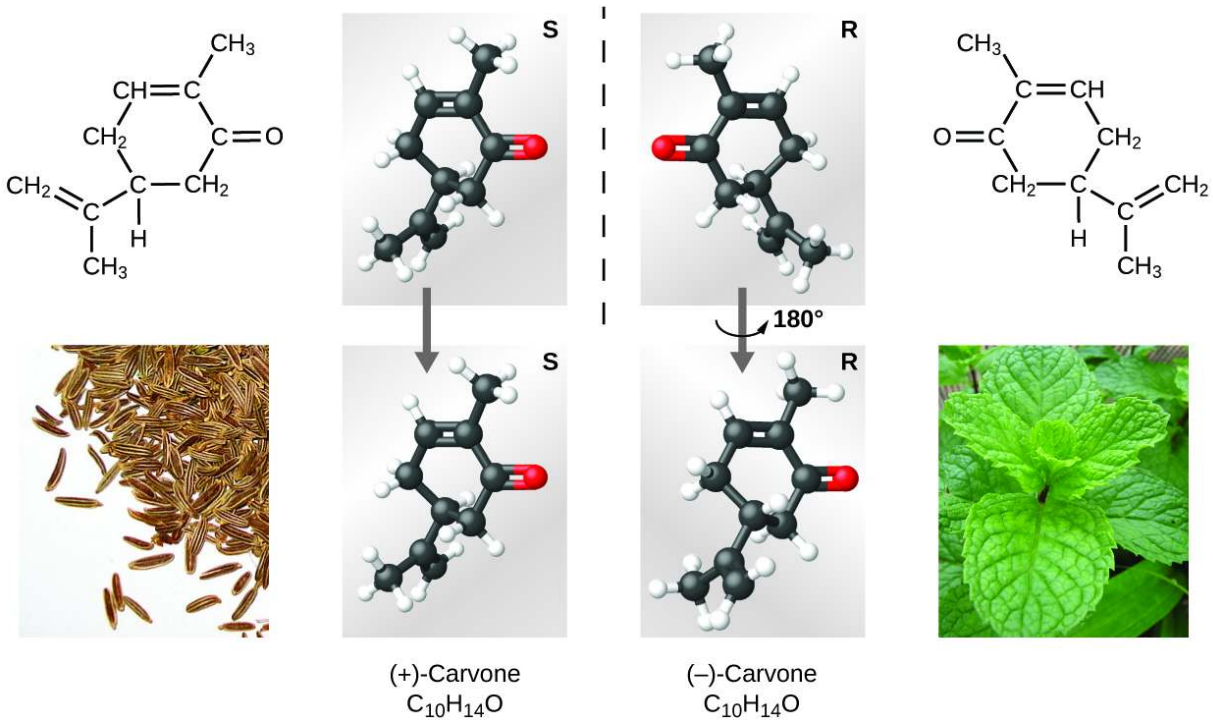


Figure 2.5.0: Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by "Miansari66"/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)

Link to Learning

Select this link to view an explanation of isomers, spatial isomers, and why they have different smells (select the video titled "Mirror Molecule: Carvone").

Footnotes

- 2Lee Cronin, "Print Your Own Medicine," Talk presented at TED Global 2012, Edinburgh, Scotland, June 2012.

This page titled 2.5: Chemical Formulas is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **2.4: Chemical Formulas** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.6: Key Terms

alpha particle (α particle)
positively charged particle consisting of two protons and two neutrons

anion
negatively charged atom or molecule (contains more electrons than protons)

atomic mass
average mass of atoms of an element, expressed in amu

atomic mass unit (amu)
(also, unified atomic mass unit, u, or Dalton, Da) unit of mass equal to $\frac{1}{12}$ of the mass of a ^{12}C atom

atomic number (Z)
number of protons in the nucleus of an atom

cation
positively charged atom or molecule (contains fewer electrons than protons)

chemical symbol
one-, two-, or three-letter abbreviation used to represent an element or its atoms

Dalton (Da)
alternative unit equivalent to the atomic mass unit

Dalton's atomic theory
set of postulates that established the fundamental properties of atoms

electron
negatively charged, subatomic particle of relatively low mass located outside the nucleus

empirical formula
formula showing the composition of a compound given as the simplest whole-number ratio of atoms

fundamental unit of charge
(also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19}$ C

ion
electrically charged atom or molecule (contains unequal numbers of protons and electrons)

isomers
compounds with the same chemical formula but different structures

isotopes
atoms that contain the same number of protons but different numbers of neutrons

law of constant composition
(also, law of definite proportions) all samples of a pure compound contain the same elements in the same proportions by mass

law of definite proportions
(also, law of constant composition) all samples of a pure compound contain the same elements in the same proportions by mass

law of multiple proportions

when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small whole numbers

mass number (A)
sum of the numbers of neutrons and protons in the nucleus of an atom

molecular formula
formula indicating the composition of a molecule of a compound and giving the actual number of atoms of each element in a molecule of the compound.

neutron
uncharged, subatomic particle located in the nucleus

nucleus
massive, positively charged center of an atom made up of protons and neutrons

proton
positively charged, subatomic particle located in the nucleus

spatial isomers
compounds in which the relative orientations of the atoms in space differ

structural formula
shows the atoms in a molecule and how they are connected

structural isomer
one of two substances that have the same molecular formula but different physical and chemical properties because their atoms are bonded differently

unified atomic mass unit (u)
alternative unit equivalent to the atomic mass unit

This page titled 2.6: Key Terms is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

2.7: Key Equations

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})$$

This page titled 2.7: Key Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

2.8: Summary

The ancient Greeks proposed that matter consists of extremely small particles called atoms. Dalton postulated that each element has a characteristic type of atom that differs in properties from atoms of all other elements, and that atoms of different elements can combine in fixed, small, whole-number ratios to form compounds. Samples of a particular compound all have the same elemental proportions by mass. When two elements form different compounds, a given mass of one element will combine with masses of the other element in a small, whole-number ratio. During any chemical change, atoms are neither created nor destroyed.

Although no one has actually seen the inside of an atom, experiments have demonstrated much about atomic structure. Thomson's cathode ray tube showed that atoms contain small, negatively charged particles called electrons. Millikan discovered that there is a fundamental electric charge—the charge of an electron. Rutherford's gold foil experiment showed that atoms have a small, dense, positively charged nucleus; the positively charged particles within the nucleus are called protons. Chadwick discovered that the nucleus also contains neutral particles called neutrons. Soddy demonstrated that atoms of the same element can differ in mass; these are called isotopes.

An atom consists of a small, positively charged nucleus surrounded by electrons. The nucleus contains protons and neutrons; its diameter is about 100,000 times smaller than that of the atom. The mass of one atom is usually expressed in atomic mass units (amu), which is referred to as the atomic mass. An amu is defined as exactly $\frac{1}{12}$ of the mass of a carbon-12 atom and is equal to 1.6605×10^{-24} g.

Protons are relatively heavy particles with a charge of 1+ and a mass of 1.0073 amu. Neutrons are relatively heavy particles with no charge and a mass of 1.0087 amu. Electrons are light particles with a charge of 1− and a mass of 0.00055 amu. The number of protons in the nucleus is called the atomic number (Z) and is the property that defines an atom's elemental identity. The sum of the numbers of protons and neutrons in the nucleus is called the mass number and, expressed in amu, is approximately equal to the mass of the atom. An atom is neutral when it contains equal numbers of electrons and protons.

Isotopes of an element are atoms with the same atomic number but different mass numbers; isotopes of an element, therefore, differ from each other only in the number of neutrons within the nucleus. When a naturally occurring element is composed of several isotopes, the atomic mass of the element represents the average of the masses of the isotopes involved. A chemical symbol identifies the atoms in a substance using symbols, which are one-, two-, or three-letter abbreviations for the atoms.

A molecular formula uses chemical symbols and subscripts to indicate the exact numbers of different atoms in a molecule or compound. An empirical formula gives the simplest, whole-number ratio of atoms in a compound. A structural formula indicates the bonding arrangement of the atoms in the molecule. Ball-and-stick and space-filling models show the geometric arrangement of atoms in a molecule. Isomers are compounds with the same molecular formula but different arrangements of atoms. A convenient amount unit for expressing very large numbers of atoms or molecules is the mole. Experimental measurements have determined the number of entities composing 1 mole of substance to be 6.022×10^{23}, a quantity called Avogadro's number. The mass in grams of 1 mole of substance is its molar mass.

This page titled 2.8: Summary is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

2.9: Exercises

1.

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres of different elements touch, they are part of a single unit of a compound. The following chemical change represented by these spheres may violate one of the ideas of Dalton's atomic theory. Which one?



2.

Which postulate of Dalton's theory is consistent with the following observation concerning the weights of reactants and products? When 100 grams of solid calcium carbonate is heated, 44 grams of carbon dioxide and 56 grams of calcium oxide are produced.

3.

Identify the postulate of Dalton's theory that is violated by the following observations: 59.95% of one sample of titanium dioxide is titanium; 60.10% of a different sample of titanium dioxide is titanium.

4.

Samples of compound X, Y, and Z are analyzed, with results shown here.

Compound	Description	Mass of Carbon	Mass of Hydrogen
X	clear, colorless, liquid with strong odor	1.776 g	0.148 g
Y	clear, colorless, liquid with strong odor	1.974 g	0.329 g
Z	clear, colorless, liquid with strong odor	7.812 g	0.651 g

Do these data provide example(s) of the law of definite proportions, the law of multiple proportions, neither, or both? What do these data tell you about compounds X, Y, and Z?

5.

The existence of isotopes violates one of the original ideas of Dalton's atomic theory. Which one?

6.

How are electrons and protons similar? How are they different?

7.

How are protons and neutrons similar? How are they different?

8.

Predict and test the behavior of α particles fired at a "plum pudding" model atom.

(a) Predict the paths taken by α particles that are fired at atoms with a Thomson's plum pudding model structure. Explain why you expect the α particles to take these paths.

(b) If α particles of higher energy than those in (a) are fired at plum pudding atoms, predict how their paths will differ from the lower-energy α particle paths. Explain your reasoning.

(c) Now test your predictions from (a) and (b). Open the Rutherford Scattering simulation and select the “Plum Pudding Atom” tab. Set “Alpha Particles Energy” to “min,” and select “show traces.” Click on the gun to start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Hit the pause button, or “Reset All.” Set “Alpha Particles Energy” to “max,” and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual paths as shown in the simulation.

9.

Predict and test the behavior of α particles fired at a Rutherford atom model.

(a) Predict the paths taken by α particles that are fired at atoms with a Rutherford atom model structure. Explain why you expect the α particles to take these paths.

(b) If α particles of higher energy than those in (a) are fired at Rutherford atoms, predict how their paths will differ from the lower-energy α particle paths. Explain your reasoning.

(c) Predict how the paths taken by the α particles will differ if they are fired at Rutherford atoms of elements other than gold. What factor do you expect to cause this difference in paths, and why?

(d) Now test your predictions from (a), (b), and (c). Open the Rutherford Scattering simulation and select the “Rutherford Atom” tab. Due to the scale of the simulation, it is best to start with a small nucleus, so select “20” for both protons and neutrons, “min” for energy, show traces, and then start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Pause or reset, set energy to “max,” and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual path as shown in the simulation. Pause or reset, select “40” for both protons and neutrons, “min” for energy, show traces, and fire away. Does this match your prediction from (c)? If not, explain why the actual path would be that shown in the simulation. Repeat this with larger numbers of protons and neutrons. What generalization can you make regarding the type of atom and effect on the path of α particles? Be clear and specific.

10.

In what way are isotopes of a given element always different? In what way(s) are they always the same?

11.

Write the symbol for each of the following ions:

(a) the ion with a 1+ charge, atomic number 55, and mass number 133

(b) the ion with 54 electrons, 53 protons, and 74 neutrons

(c) the ion with atomic number 15, mass number 31, and a 3− charge

(d) the ion with 24 electrons, 30 neutrons, and a 3+ charge

12.

Write the symbol for each of the following ions:

(a) the ion with a 3+ charge, 28 electrons, and a mass number of 71

(b) the ion with 36 electrons, 35 protons, and 45 neutrons

(c) the ion with 86 electrons, 142 neutrons, and a 4+ charge

(d) the ion with a 2+ charge, atomic number 38, and mass number 87

13.

Open the Build an Atom simulation and click on the Atom icon.

(a) Pick any one of the first 10 elements that you would like to build and state its symbol.

(b) Drag protons, neutrons, and electrons onto the atom template to make an atom of your element.
State the numbers of protons, neutrons, and electrons in your atom, as well as the net charge and mass number.

(c) Click on “Net Charge” and “Mass Number,” check your answers to (b), and correct, if needed.

(d) Predict whether your atom will be stable or unstable. State your reasoning.

(e) Check the “Stable/Unstable” box. Was your answer to (d) correct? If not, first predict what you can do to make a stable atom of your element, and then do it and see if it works. Explain your reasoning.

14.

Open the Build an Atom simulation

(a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Oxygen-16 and give the isotope symbol for this atom.

(b) Now add two more electrons to make an ion and give the symbol for the ion you have created.

15.

Open the Build an Atom simulation

(a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Lithium-6 and give the isotope symbol for this atom.

(b) Now remove one electron to make an ion and give the symbol for the ion you have created.

16.

Determine the number of protons, neutrons, and electrons in the following isotopes that are used in medical diagnoses:

(a) atomic number 9, mass number 18, charge of 1−

(b) atomic number 43, mass number 99, charge of 7+

(c) atomic number 53, atomic mass number 131, charge of 1−

(d) atomic number 81, atomic mass number 201, charge of 1+

(e) Name the elements in parts (a), (b), (c), and (d).

17.

The following are properties of isotopes of two elements that are essential in our diet. Determine the number of protons, neutrons and electrons in each and name them.

(a) atomic number 26, mass number 58, charge of 2+

(b) atomic number 53, mass number 127, charge of 1−

18.

Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

(a) $^{10}_{5}\mathrm{B}$

(b) $^{199}_{80}\mathrm{Hg}$

(c) $^{63}_{29}\mathrm{Cu}$

(d) $^{13}_{6}\mathrm{C}$

(e) $^{77}_{34}\mathrm{Se}$

19.

Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

(a) $^{7}_{3}\mathrm{Li}$

(b) $^{125}_{52}\mathrm{Te}$

(c) $^{109}_{47}\mathrm{Ag}$

(d) $^{15}_{7}\mathrm{N}$

(e) $^{31}_{15}\mathrm{P}$

20.

Click on the site and select the “Mix Isotopes” tab, hide the “Percent Composition” and “Average Atomic Mass” boxes, and then select the element boron.

(a) Write the symbols of the isotopes of boron that are shown as naturally occurring in significant amounts.

(b) Predict the relative amounts (percentages) of these boron isotopes found in nature. Explain the reasoning behind your choice.

(c) Add isotopes to the black box to make a mixture that matches your prediction in (b). You may drag isotopes from their bins or click on “More” and then move the sliders to the appropriate amounts.

(d) Reveal the “Percent Composition” and “Average Atomic Mass” boxes. How well does your mixture match with your prediction? If necessary, adjust the isotope amounts to match your prediction.

(e) Select “Nature’s” mix of isotopes and compare it to your prediction. How well does your prediction compare with the naturally occurring mixture? Explain. If necessary, adjust your amounts to make them match “Nature’s” amounts as closely as possible.

21.

Repeat Exercise 2.20 using an element that has three naturally occurring isotopes.

22.

An element has the following natural abundances and isotopic masses: 90.92% abundance with 19.99 amu, 0.26% abundance with 20.99 amu, and 8.82% abundance with 21.99 amu. Calculate the average atomic mass of this element.

23.

Average atomic masses listed by IUPAC are based on a study of experimental results. Bromine has two isotopes, ^{79}Br and ^{81}Br, whose masses (78.9183 and 80.9163 amu, respectively) and abundances (50.69% and 49.31%, respectively) were determined in earlier experiments. Calculate the average atomic mass of bromine based on these experiments.

24.

Variations in average atomic mass may be observed for elements obtained from different sources. Lithium provides an example of this. The isotopic composition of lithium from naturally occurring minerals is 7.5% ^{6}Li and 92.5% ^{7}Li, which have masses of 6.01512 amu and 7.01600 amu, respectively. A commercial source of lithium, recycled from a military source, was 3.75% ^{6}Li (and the rest ^{7}Li). Calculate the average atomic mass values for each of these two sources.

25.

The average atomic masses of some elements may vary, depending upon the sources of their ores. Naturally occurring boron consists of two isotopes with accurately known masses (^{10}B, 10.0129 amu and ^{11}B, 11.00931 amu). The actual atomic mass of boron can vary from 10.807 to 10.819, depending on whether the mineral source is from Turkey or the United States. Calculate the percent abundances leading to the two values of the average atomic masses of boron from these two countries.

26.

The ^{18}O:^{16}O abundance ratio in some meteorites is greater than that used to calculate the average atomic mass of oxygen on earth. Is the average mass of an oxygen atom in these meteorites greater than, less than, or equal to that of a terrestrial oxygen atom?

27.

Explain why the symbol for an atom of the element oxygen and the formula for a molecule of oxygen differ.

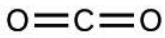
28.

Explain why the symbol for the element sulfur and the formula for a molecule of sulfur differ.

29.

Write the molecular and empirical formulas of the following compounds:

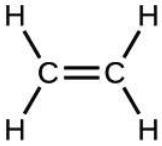
(a)



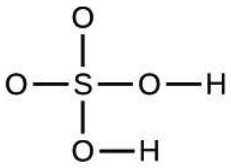
(b)



(c)



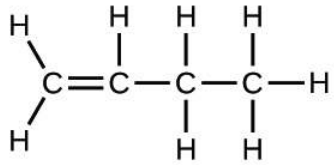
(d)



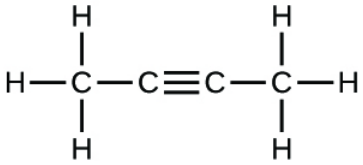
30.

Write the molecular and empirical formulas of the following compounds:

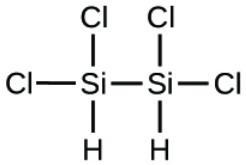
(a)



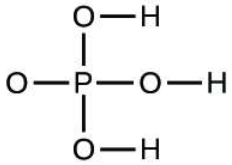
(b)



(c)



(d)



31.

Determine the empirical formulas for the following compounds:

(a) caffeine, $C_8H_{10}N_4O_2$

(b) sucrose, $C_{12}H_{22}O_{11}$

(c) hydrogen peroxide, H_2O_2

(d) glucose, $C_6H_{12}O_6$

(e) ascorbic acid (vitamin C), $C_6H_8O_6$

32.

Determine the empirical formulas for the following compounds:

(a) acetic acid, $C_2H_4O_2$

(b) citric acid, $C_6H_8O_7$

(c) hydrazine, N_2H_4

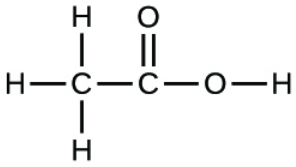
(d) nicotine, $C_{10}H_{14}N_2$

(e) butane, C_4H_{10}

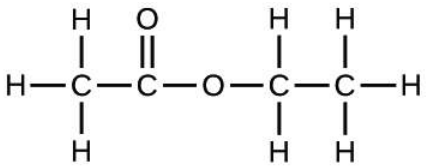
33.

Write the empirical formulas for the following compounds:

(a)



(b)



34.

Open the Build a Molecule simulation and select the “Larger Molecules” tab. Select an appropriate atom's “Kit” to build a molecule with two carbon and six hydrogen atoms. Drag atoms into the space above the “Kit” to make a molecule. A name will appear when you have made an actual molecule that exists (even if it is not the one you want). You can use the scissors tool to separate atoms if you would like to change the connections. Click on “3D” to see the molecule, and look at both the space-filling and ball-and-stick possibilities.

(a) Draw the structural formula of this molecule and state its name.

(b) Can you arrange these atoms in any way to make a different compound?

35.

Use the Build a Molecule simulation to repeat Exercise 2.34, but build a molecule with two carbons, six hydrogens, and one oxygen.

(a) Draw the structural formula of this molecule and state its name.

(b) Can you arrange these atoms to make a different molecule? If so, draw its structural formula and state its name.

(c) How are the molecules drawn in (a) and (b) the same? How do they differ? What are they called (the type of relationship between these molecules, not their names)?

36.

Use the Build a Molecule simulation to repeat Exercise 2.34, but build a molecule with three carbons, seven hydrogens, and one chlorine.

(a) Draw the structural formula of this molecule and state its name.

(b) Can you arrange these atoms to make a different molecule? If so, draw its structural formula and state its name.

(c) How are the molecules drawn in (a) and (b) the same? How do they differ? What are they called (the type of relationship between these molecules, not their names)?

37.

Write a sentence that describes how to determine the number of moles of a compound in a known mass of the compound if we know its molecular formula.

38.

Compare 1 mole of H_2, 1 mole of O_2, and 1 mole of F_2.

(a) Which has the largest number of molecules? Explain why.

(b) Which has the greatest mass? Explain why.

39.

Which contains the greatest mass of oxygen: 0.75 mol of ethanol (C_2H_5OH), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

40.

Which contains the greatest number of moles of oxygen atoms: 1 mol of ethanol (C_2H_5OH), 1 mol of formic acid (HCO_2H), or 1 mol of water (H_2O)? Explain why.

41.

How are the molecular mass and the molar mass of a compound similar and how are they different?

42.

Calculate the molar mass of each of the following compounds:

(a) hydrogen fluoride, HF

(b) ammonia, NH_3

(c) nitric acid, HNO_3

(d) silver sulfate, Ag_2SO_4

(e) boric acid, $B(OH)_3$

43.

Calculate the molar mass of each of the following:

(a) S_8

(b) C_5H_{12}

(c) $Sc_2(SO_4)_3$

(d) CH_3COCH_3 (acetone)

(e) $C_6H_{12}O_6$ (glucose)

44.

Calculate the empirical or molecular formula mass and the molar mass of each of the following minerals:

(a) limestone, $CaCO_3$

(b) halite, NaCl

(c) beryl, $Be_3Al_2Si_6O_{18}$

(d) malachite, $Cu_2(OH)_2CO_3$

(e) turquoise, $CuAl_6(PO_4)_4(OH)_8(H_2O)_4$

45.

Calculate the molar mass of each of the following:

(a) the anesthetic halothane, $C_2HBrClF_3$

(b) the herbicide paraquat, $C_{12}H_{14}N_2Cl_2$

(c) caffeine, $C_8H_{10}N_4O_2$

(d) urea, $CO(NH_2)_2$

(e) a typical soap, $C_{17}H_{35}CO_2Na$

46.

Determine the number of moles of compound and the number of moles of each type of atom in each of the following:

(a) 25.0 g of propylene, C_3H_6

(b) 3.06×10^{-3} g of the amino acid glycine, $C_2H_5NO_2$

(c) 25 lb of the herbicide Treflan, $C_{13}H_{16}N_2O_4F$ (1 lb = 454 g)

(d) 0.125 kg of the insecticide Paris Green, $Cu_4(AsO_3)_2(CH_3CO_2)_2$

(e) 325 mg of aspirin, $C_6H_4(CO_2H)(CO_2CH_3)$

47.

Determine the mass of each of the following:

(a) 0.0146 mol KOH

(b) 10.2 mol ethane, C_2H_6

(c) 1.6×10^{-3} mol $Na_2\ SO_4$

(d) 6.854×10^3 mol glucose, $C_6\ H_{12}\ O_6$

(e) 2.86 mol $Co(NH_3)_6Cl_3$

48.

Determine the number of moles of the compound and determine the number of moles of each type of atom in each of the following:

(a) 2.12 g of potassium bromide, KBr

(b) 0.1488 g of phosphoric acid, H_3PO_4

(c) 23 kg of calcium carbonate, $CaCO_3$

(d) 78.452 g of aluminum sulfate, $Al_2(SO_4)_3$

(e) 0.1250 mg of caffeine, $C_8H_{10}N_4O_2$

49.

Determine the mass of each of the following:

(a) 2.345 mol LiCl

(b) 0.0872 mol acetylene, C_2H_2

(c) 3.3×10^{-2} mol $Na_2\ CO_3$

(d) 1.23×10^3 mol fructose, $C_6\ H_{12}\ O_6$

(e) 0.5758 mol $FeSO_4(H_2O)_7$

50.

The approximate minimum daily dietary requirement of the amino acid leucine, $C_6H_{13}NO_2$, is 1.1 g. What is this requirement in moles?

51.

Determine the mass in grams of each of the following:

(a) 0.600 mol of oxygen atoms

(b) 0.600 mol of oxygen molecules, O_2

(c) 0.600 mol of ozone molecules, O_3

52.

A 55-kg woman has 7.5×10^{-3} mol of hemoglobin (molar mass = 64,456 g/mol) in her blood. How many hemoglobin molecules is this? What is this quantity in grams?

53.

Determine the number of atoms and the mass of zirconium, silicon, and oxygen found in 0.3384 mol of zircon, $ZrSiO_4$, a semiprecious stone.

54.

Determine which of the following contains the greatest mass of hydrogen: 1 mol of CH_4, 0.6 mol of C_6H_6, or 0.4 mol of C_3H_8.

55.

Determine which of the following contains the greatest mass of aluminum: 122 g of $AlPO_4$, 266 g of Al_2Cl_6, or 225 g of Al_2S_3.

56.

Diamond is one form of elemental carbon. An engagement ring contains a diamond weighing 1.25 carats (1 carat = 200 mg). How many atoms are present in the diamond?

57.

The Cullinan diamond was the largest natural diamond ever found (January 25, 1905). It weighed 3104 carats (1 carat = 200 mg). How many carbon atoms were present in the stone?

58.

One 55-gram serving of a particular cereal supplies 270 mg of sodium, 11% of the recommended daily allowance. How many moles and atoms of sodium are in the recommended daily allowance?

59.

A certain nut crunch cereal contains 11.0 grams of sugar (sucrose, $C_{12}H_{22}O_{11}$) per serving size of 60.0 grams. How many servings of this cereal must be eaten to consume 0.0278 moles of sugar?

60.

A tube of toothpaste contains 0.76 g of sodium monofluorophosphate (Na_2PO_3F) in 100 mL.

(a) What mass of fluorine atoms in mg was present?

(b) How many fluorine atoms were present?

61.

Which of the following represents the least number of molecules?

(a) 20.0 g of H_2O (18.02 g/mol)

(b) 77.0 g of CH_4 (16.06 g/mol)

(c) 68.0 g of C_3H_6 (42.08 g/mol)

(d) 100.0 g of N_2O (44.02 g/mol)

(e) 84.0 g of HF (20.01 g/mol)

This page titled 2.9: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

3: The Quantum Model of the Atom

The study of chemistry must at some point extend to the molecular level, for the physical and chemical properties of a substance are ultimately explained in terms of the structure and bonding of molecules. This module introduces some basic facts and principles that are needed for a discussion of organic molecules.

3.1: Introduction
3.2: Electromagnetic Radiation
3.3: Atomic Spectra
3.4: Quantum Theory
3.5: The Bohr Atom
3.6: Two Ideas Leading to a New Quantum Mechanics
3.7: Wave Mechanics
3.8: Electron Spin- A Fourth Quantum Number
3.9: The Shape of Atomic Orbitals
3.10: Multielectron Atoms
3.11: Key Terms
3.12: Key Equations
3.13: Summary
3.14: Exercises

This page titled 3: The Quantum Model of the Atom is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

3.1: Introduction

In 1054, Chinese astronomers recorded the appearance of a “guest star” in the sky, visible even during the day, which then disappeared slowly over the next two years. The sudden appearance was due to a supernova explosion, which was much brighter than the original star. Even though this supernova was observed almost a millennium ago, the remaining Crab Nebula (Figure 3.1.1) continues to release energy today. It emits not only visible light but also infrared light, X-rays, and other forms of electromagnetic radiation. The nebula emits both continuous spectra (the blue-white glow) and atomic emission spectra (the colored filaments). In this chapter, we will discuss light and other forms of electromagnetic radiation and how they are related to the electronic structure of atoms. We will also see how this radiation can be used to identify elements, even from thousands of light years away.

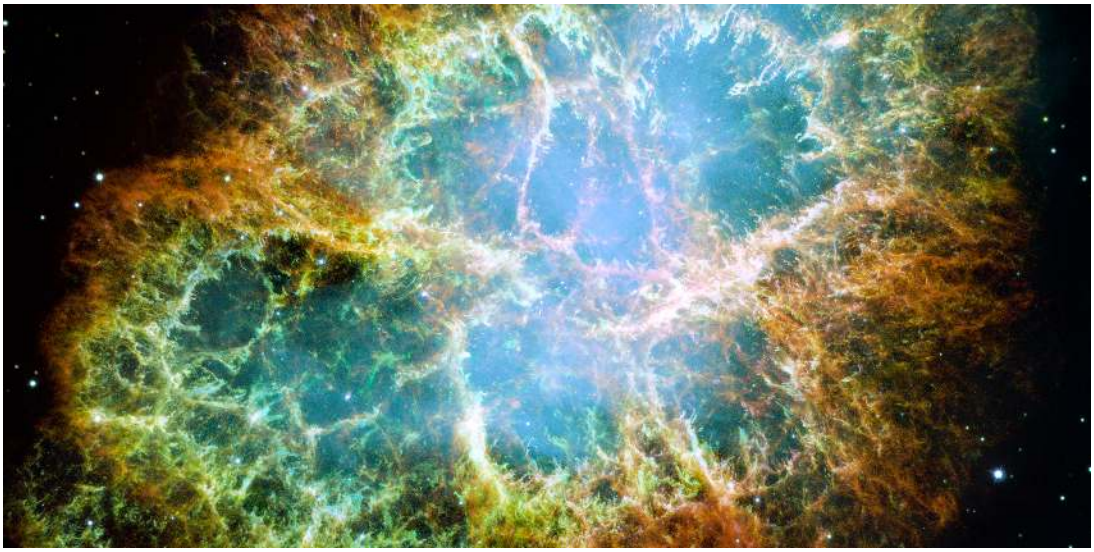


Figure 3.1.1: The Crab Nebula consists of remnants of a supernova (the explosion of a star). NASA’s Hubble Space Telescope produced this composite image. Measurements of the emitted light wavelengths enabled astronomers to identify the elements in the nebula, determining that it contains specific ions including S^{+} (green filaments) and O^{2+} (red filaments). (credit: modification of work by NASA and ESA)

This page titled 3.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **6.0: Introduction** is licensed CC BY 4.0.

3.2: Electromagnetic Radiation

Learning Objectives

- to learn about the characteristics of electromagnetic waves. Light, X-Rays, infrared and microwaves among the types of electromagnetic waves.

Scientists discovered much of what we know about the structure of the atom by observing the interaction of atoms with various forms of radiant, or transmitted, energy, such as the energy associated with the visible light we detect with our eyes, the infrared radiation we feel as heat, the ultraviolet light that causes sunburn, and the x-rays that produce images of our teeth or bones. All these forms of radiant energy should be familiar to you. We begin our discussion of the development of our current atomic model by describing the properties of waves and the various forms of electromagnetic radiation.



Figure 3.2.1: A Wave in Water When a drop of water falls onto a smooth water surface, it generates a set of waves that travel outward in a circular direction.

Properties of Waves

A wave is a periodic oscillation that transmits energy through space. Anyone who has visited a beach or dropped a stone into a puddle has observed waves traveling through water (Figure 3.2.1). These waves are produced when wind, a stone, or some other disturbance, such as a passing boat, transfers energy to the water, causing the surface to oscillate up and down as the energy travels outward from its point of origin. As a wave passes a particular point on the surface of the water, anything floating there moves up and down.

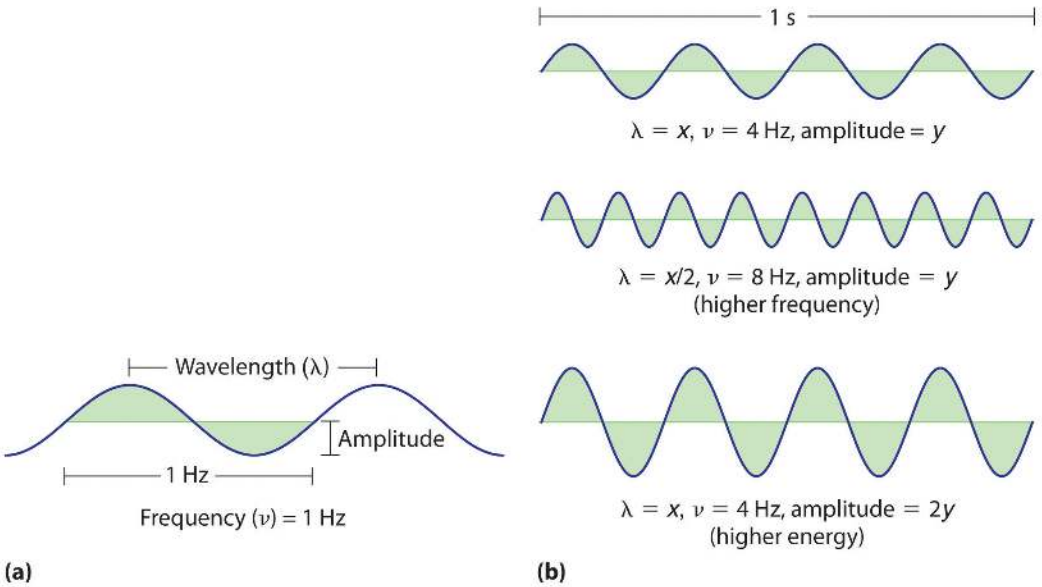


Figure 3.2.2: Important Properties of Waves (a) Wavelength (λ in meters), frequency (ν, in Hz), and amplitude are indicated on this drawing of a wave. (b) The wave with the shortest wavelength has the greatest number of wavelengths per unit time (i.e., the highest frequency). If two waves have the same frequency and speed, the one with the greater amplitude has the higher energy.

Waves have characteristic properties (Figure 3.2.2). As you may have noticed in Figure 3.2.1, waves are periodic, that is, they repeat regularly in both space and time. The distance between two corresponding points in a wave—between the midpoints of two peaks, for example, or two troughs—is the wavelength (λ), distance between two corresponding points in a wave—between the midpoints of two peaks or two troughs. λ is the lowercase Greek lambda, and u is the lowercase Greek nu. Wavelengths are described by a unit of distance, typically meters. The frequency (ν), the number of oscillations (i.e., of a wave) that pass a particular point in a given period of time. of a wave is the number of oscillations that pass a particular point in a given period of time. The usual units are oscillations persecond (1/s = s^{-1}), which in the SI system is called the hertz (Hz). It is named after German physicist Heinrich Hertz (1857–1894),a pioneer in the field of electromagnetic radiation. The amplitude (the vertical height of a wave, which is defined as half the peak-to-trough height), or vertical height, of a wave is defined as half the peak-to-trough height; as the amplitude of a wave with a given frequency increases, so does its energy. As you can see in Figure 3.2.2, two waves can have the same amplitude but different wavelengths and vice versa. The distance traveled by a wave per unit time is its speed (ν), the distance traveled by a wave per unit time, which is typically measured in meters per second (m/s). The speed of a wave is equal to the product of its wavelength and frequency:

$$(\text{wavelength})(\text{frequency}) = speed \tag{3.2.1}$$

$$\lambda u = v \tag{3.2.2}$$

$$\left(\frac{meters}{\cancel{wave}}\right)\left(\frac{\cancel{wave}}{second}\right) = \frac{meters}{second} \tag{3.2.3}$$

Be careful not to confuse the symbols for the speed, v, with the frequency, u. Water waves are slow compared to sound waves, which can travel through solids, liquids, and gases. Whereas water waves may travel a few meters per second, the speed of sound in dry air at 20°C is 343.5 m/s. Ultrasonic waves, which travel at an even higher speed (>1500 m/s) and have a greater frequency, are used in such diverse applications as locating underwater objects and the medical imaging of internal organs.

Electromagnetic Radiation

Water waves transmit energy through space by the periodic oscillation of matter (the water). In contrast, energy that is transmitted, or radiated, through space in the form of periodic oscillations of electric and magnetic fields is known as electromagnetic radiation, which is energy that is transmitted, or radiated, through space in the form of periodic oscillations of electric and magnetic fields. (Figure 3.2.3). Some forms of electromagnetic radiation are shown in Figure 3.2.4. In a vacuum, all forms of electromagnetic radiation—whether microwaves, visible light, or gamma rays—travel at the speed of light (c), which is the speed with which all forms of electromagnetic radiation travel in a vacuum, a fundamental physical constant with a value of 2.99792458 × 10^8 m/s (which is about 3.00 ×10^8 m/s or 1.86 × 10^5 mi/s). This is about a million times faster than the speed of sound.

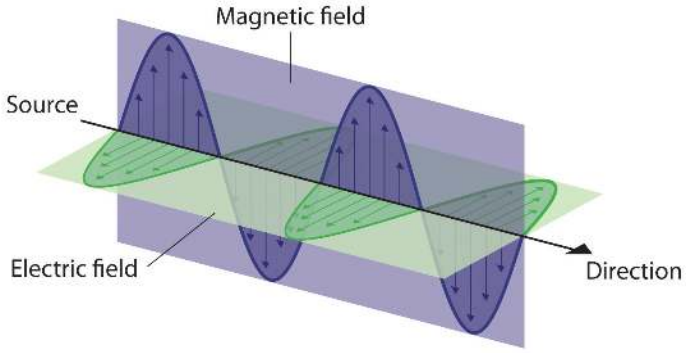


Figure 3.2.3: The Nature of Electromagnetic Radiation. All forms of electromagnetic radiation consist of perpendicular oscillating electric and magnetic fields.

Because the various kinds of electromagnetic radiation all have the same speed (*c*), they differ in only wavelength and frequency. As shown in Figure 3.2.4 and Table 3.2.1, the wavelengths of familiar electromagnetic radiation range from 10^1 m for radio waves to 10^{-12} m for gamma rays, which are emitted by nuclear reactions. By replacing *v* with *c* in Equation 3.2.1, we can show that the frequency of electromagnetic radiation is inversely proportional to its wavelength:

$$\begin{aligned} c &= \lambda u \\ u &= \frac{c}{\lambda} \end{aligned} \tag{3.2.4}$$

For example, the frequency of radio waves is about 10^8 Hz, whereas the frequency of gamma rays is about 10^{20} Hz. Visible light, which is electromagnetic radiation that can be detected by the human eye, has wavelengths between about 7×10^{-7} m (700 nm, or 4.3×10^{14} Hz) and 4×10^{-7} m (400 nm, or 7.5×10^{14} Hz). Note that when frequency increases, wavelength decreases; c being a constant stays the same. Similarly when frequency decreases, the wavelength increases.

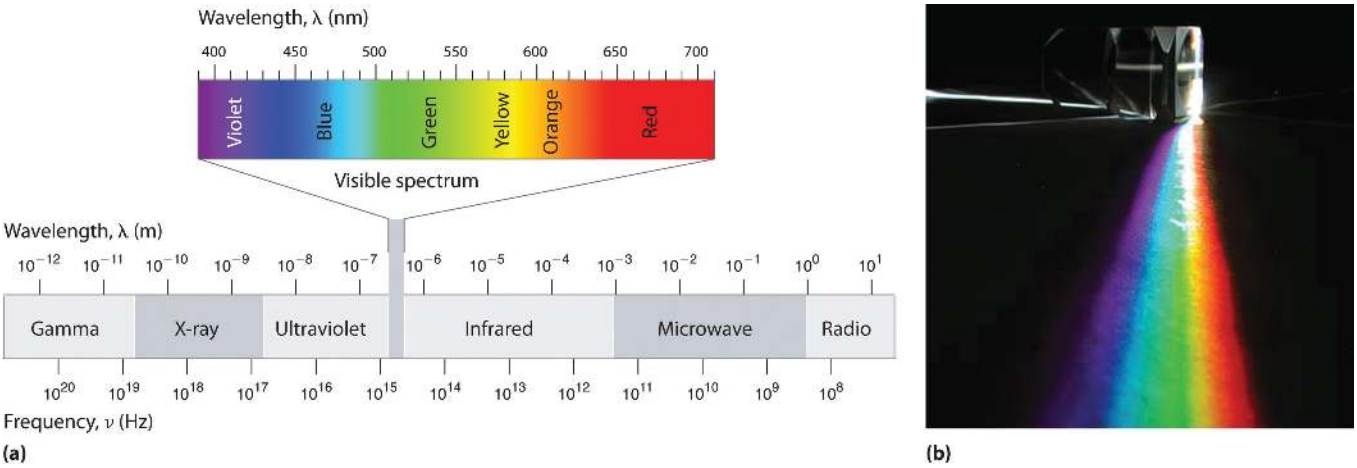


Figure 3.2.4: The Electromagnetic Spectrum. (a) This diagram shows the wavelength and frequency ranges of electromagnetic radiation. The visible portion of the electromagnetic spectrum is the narrow region with wavelengths between about 400 and 700 nm. (b) When white light is passed through a prism, it is split into light of different wavelengths, whose colors correspond to the visible spectrum. (CC BY-NC-SA; anonymous)

Within this visible range our eyes perceive radiation of different wavelengths (or frequencies) as light of different colors, ranging from red to violet in order of decreasing wavelength. The components of white light—a mixture of all the frequencies of visible light—can be separated by a prism, as shown in part (b) in Figure 3.2.4. A similar phenomenon creates a rainbow, where water droplets suspended in the air act as tiny prisms.

Table 3.2.1: Common Wavelength Units for Electromagnetic Radiation

Unit	Symbol	Wavelength (m)	Type of Radiation
picometer	pm	10^{-12}	gamma ray
angstrom	Å	10^{-10}	x-ray
nanometer	nm	10^{-9}	x-ray
micrometer	μm	10^{-6}	infrared
millimeter	mm	10^{-3}	infrared
centimeter	cm	10^{-2}	microwave
meter	m	10^{0}	radio

As you will soon see, the energy of electromagnetic radiation is directly proportional to its frequency and inversely proportional to its wavelength:

$$E \propto u \tag{3.2.5}$$

$$E \propto \frac{1}{\lambda} \tag{3.2.6}$$

Whereas visible light is essentially harmless to our skin, ultraviolet light, with wavelengths of ≤ 400 nm, has enough energy to cause severe damage to our skin in the form of sunburn. Because the ozone layer absorbs sunlight with wavelengths less than 350 nm, it protects us from the damaging effects of highly energetic ultraviolet radiation.

The energy of electromagnetic radiation **increases** with **increasing** frequency and **decreasing** wavelength.

Example 3.2.1

Your favorite FM radio station, WXYZ, broadcasts at a frequency of 101.1 MHz. What is the wavelength of this radiation?

Given: frequency

Asked for: wavelength

Strategy:

Substitute the value for the speed of light in meters per second into Equation 3.2.2 to calculate the wavelength in meters.

Solution:

From Equation 3.2.4, we know that the product of the wavelength and the frequency is the speed of the wave, which for electromagnetic radiation is 2.998 × 10^8 m/s:

$$\lambda\nu = c = 2.998 \times 10^8 m/s \tag{3.2.7}$$

Thus the wavelength λ is given by

$$\lambda = \frac{c}{u} = \left(\frac{2.988 \times 10^8 \ m/\cancel{s}}{101.1 \ \cancel{MHz}}\right)\left(\frac{1 \ \cancel{MHz}}{10^6 \ \cancel{s^{-1}}}\right) = 2.965 \ m \tag{3.2.8}$$

Exercise 3.2.1

As the police officer was writing up your speeding ticket, she mentioned that she was using a state-of-the-art radar gun operating at 35.5 GHz. What is the wavelength of the radiation emitted by the radar gun?

Answer: 8.45 mm



Electromagnetic Radiation: https://youtu.be/TZy7a69pP-w

Summary

Understanding the electronic structure of atoms requires an understanding of the properties of waves and electromagnetic radiation. A basic knowledge of the electronic structure of atoms requires an understanding of the properties of waves and electromagnetic radiation. A **wave** is a periodic oscillation by which energy is transmitted through space. All waves are **periodic**, repeating regularly in both space and time. Waves are characterized by several interrelated properties: **wavelength (λ)**, the distance between successive waves; **frequency (ν)**, the number of waves that pass a fixed point per unit time; **speed (*v*)**, the rate at which the wave propagates through space; and **amplitude**, the magnitude of the oscillation about the mean position. The speed of a wave is equal to the product of its wavelength and frequency. **Electromagnetic radiation** consists of two perpendicular waves, one electric and one magnetic, propagating at the **speed of light (*c*)**. Electromagnetic radiation is radiant energy that includes radio waves, microwaves, visible light, x-rays, and gamma rays, which differ in their frequencies and wavelengths.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

3.2: Electromagnetic Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.1: Electromagnetic Radiation** is licensed CC BY-NC-SA 4.0.

3.3: Atomic Spectra

Learning Objectives

- To know the relationship between atomic spectra and the electronic structure of atoms.

The photoelectric effect provided indisputable evidence for the existence of the photon and thus the particle-like behavior of electromagnetic radiation. The concept of the photon, however, emerged from experimentation with *thermal radiation*, electromagnetic radiation emitted as the result of a source's temperature, which produces a continuous spectrum of energies. More direct evidence was needed to verify the quantized nature of electromagnetic radiation. In this section, we describe how experimentation with visible light provided this evidence.

Line Spectra

Although objects at high temperature emit a continuous spectrum of electromagnetic radiation, a different kind of spectrum is observed when pure samples of individual elements are heated. For example, when a high-voltage electrical discharge is passed through a sample of hydrogen gas at low pressure, the resulting individual isolated hydrogen atoms caused by the dissociation of H_2 emit a red light. Unlike blackbody radiation, the color of the light emitted by the hydrogen atoms does not depend greatly on the temperature of the gas in the tube. When the emitted light is passed through a prism, only a few narrow lines, called a line spectrum, which is a spectrum in which light of only a certain wavelength is emitted or absorbed, rather than a continuous range of wavelengths (Figure 3.3.1), rather than a continuous range of colors. The light emitted by hydrogen atoms is red because, of its four characteristic lines, the most intense line in its spectrum is in the red portion of the visible spectrum, at 656 nm. With sodium, however, we observe a yellow color because the most intense lines in its spectrum are in the yellow portion of the spectrum, at about 589 nm.





Figure 3.3.1: The Emission of Light by Hydrogen Atoms. (a) A sample of excited hydrogen atoms emits a characteristic red light. (b) When the light emitted by a sample of excited hydrogen atoms is split into its component wavelengths by a prism, four characteristic violet, blue, green, and red emission lines can be observed, the most intense of which is at 656 nm.

Such *emission spectra* were observed for many other elements in the late 19th century, which presented a major challenge because classical physics was unable to explain them. Part of the explanation is provided by the observation of only a few values of λ (or ν) in the line spectrum meant that only a few values of E were possible. Thus *the energy levels of a hydrogen atom had to be quantized*; in other words, only states that had certain values of energy were possible, or *allowed*. If a hydrogen atom could have *any* value of energy, then a continuous spectrum would have been observed, similar to blackbody radiation.

In 1885, a Swiss mathematics teacher, Johann Balmer (1825–1898), showed that the frequencies of the lines observed in the visible region of the spectrum of hydrogen fit a simple equation that can be expressed as follows:

$$u = constant\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \tag{3.3.1}$$

where n = 3, 4, 5, 6. As a result, these lines are known as the *Balmer series*. The Swedish physicist Johannes Rydberg (1854–1919) subsequently restated and expanded Balmer's result in the *Rydberg equation*:

$$\frac{1}{\lambda} = \Re\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \tag{3.3.2}$$

where n_1 and n_2 are positive integers, $n_2 > n_1$, and $\Re$ the *Rydberg constant*, has a value of 1.09737 × 10^7 m^{-1}.

Johann Balmer (1825–1898)

A mathematics teacher at a secondary school for girls in Switzerland, Balmer was 60 years old when he wrote the paper on the spectral lines of hydrogen that made him famous.



Balmer published only one other paper on the topic, which appeared when he was 72 years old.

Like Balmer's equation, Rydberg's simple equation described the wavelengths of the visible lines in the emission spectrum of hydrogen (with n_1 = 2, n_2 = 3, 4, 5,…). More important, Rydberg's equation also described the wavelengths of other series of lines that would be observed in the emission spectrum of hydrogen: one in the ultraviolet (n_1 = 1, n_2 = 2, 3, 4,…) and one in the infrared (n_1 = 3, n_2 = 4, 5, 6). Unfortunately, scientists had not yet developed any theoretical justification for an equation of this form.



The Bohr Atom: https://youtu.be/GuFQEOzFOgA

Summary

There is an intimate connection between the atomic structure of an atom and its spectral characteristics. Most light is polychromatic and contains light of many wavelengths. Light that has only a single wavelength is monochromatic and is produced by devices called lasers, which use transitions between two atomic energy levels to produce light in a very narrow range of wavelengths. Atoms can also absorb light of certain energies, resulting in a transition from the ground state or a lower-energy excited state to a higher-energy excited state. This produces an **absorption spectrum**, which has dark lines in the same position as the bright lines in the **emission spectrum** of an element.

Atoms of individual elements emit light at only specific wavelengths, producing a **line spectrum** rather than the continuous spectrum of all wavelengths produced by a hot object. Niels Bohr explained the line spectrum of the hydrogen atom by assuming

that the electron moved in circular orbits and that orbits with only certain radii were allowed. Lines in the spectrum were due to transitions in which an electron moved from a higher-energy orbit with a larger radius to a lower-energy orbit with smaller radius. The orbit closest to the nucleus represented the **ground state** of the atom and was most stable; orbits farther away were higher-energy **excited states**. Transitions from an excited state to a lower-energy state resulted in the emission of light with only a limited number of wavelengths. Bohr's model could not, however, explain the spectra of atoms heavier than hydrogen.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

3.3: Atomic Spectra is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.2: Atomic Spectra** is licensed CC BY-NC-SA 4.0.

3.4: Quantum Theory

Learning Objectives

- To understand how energy is quantized.

By the late 19th century, many physicists thought their discipline was well on the way to explaining most natural phenomena. They could calculate the motions of material objects using Newton's laws of classical mechanics, and they could describe the properties of radiant energy using mathematical relationships known as Maxwell's equations, developed in 1873 by James Clerk Maxwell, a Scottish physicist. The universe appeared to be a simple and orderly place, containing matter, which consisted of particles that had mass and whose location and motion could be accurately described, and electromagnetic radiation, which was viewed as having no mass and whose exact position in space could not be fixed. Thus matter and energy were considered distinct and unrelated phenomena. Soon, however, scientists began to look more closely at a few inconvenient phenomena that could *not* be explained by the theories available at the time.

Blackbody Radiation

One phenomenon that seemed to contradict the theories of classical physics was blackbody radiation, which is electromagnetic radiation whose wavelength and color that depends on the temperature of the object. The wavelength of energy emitted by an object depends on only its temperature, not its surface or composition. Hence an electric stove burner or the filament of a space heater glows dull red or orange when heated, whereas the much hotter tungsten wire in an incandescent light bulb gives off a yellowish light.



Figure 3.4.1: Blackbody Radiation. When heated, all objects emit electromagnetic radiation whose wavelength (and color) depends on the temperature of the object. A relatively low-temperature object, such as a horseshoe forged by a blacksmith, appears red, whereas a higher-temperature object, such as the surface of the sun, appears yellow or white. Images used with permission from Wikipedia.

The *intensity* of radiation is a measure of the energy emitted per unit area. A plot of the intensity of blackbody radiation as a function of wavelength for an object at various temperatures is shown in Figure 3.4.2. One of the major assumptions of classical physics was that energy increased or decreased in a smooth, continuous manner. For example, classical physics predicted that as wavelength decreased, the intensity of the radiation an object emits should increase in a smooth curve without limit at *all* temperatures, as shown by the broken line for 6000 K in Figure 3.4.2. Thus classical physics could not explain the sharp *decrease* in the intensity of radiation emitted at shorter wavelengths (primarily in the ultraviolet region of the spectrum), which was referred to as the "ultraviolet catastrophe." In 1900, however, the German physicist Max Planck (1858–1947) explained the ultraviolet catastrophe by proposing that the energy of electromagnetic waves is *quantized* rather than continuous. This means that for each temperature, there is a maximum intensity of radiation that is emitted in a blackbody object, corresponding to the peaks in Figure

3.4.2, so the intensity does not follow a smooth curve as the temperature increases, as predicted by classical physics. Thus energy could be gained or lost only in integral multiples of some smallest unit of energy, a quantum (the smallest possible unit of energy). Energy can be gained or lost only in integral multiples of a quantum..

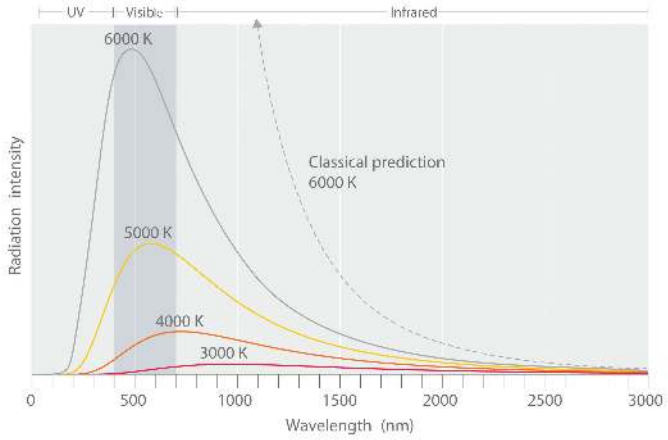
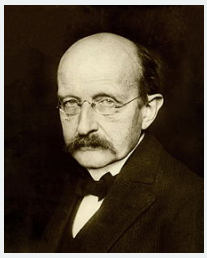


Figure 3.4.2: Relationship between the Temperature of an Object and the Spectrum of Blackbody Radiation It Emits. At relatively low temperatures, most radiation is emitted at wavelengths longer than 700 nm, which is in the infrared portion of the spectrum. The dull red glow of the electric stove element in Figure 3.4.1 is due to the small amount of radiation emitted at wavelengths less than 700 nm, which the eye can detect. As the temperature of the object increases, the maximum intensity shifts to shorter wavelengths, successively resulting in orange, yellow, and finally white light. At high temperatures, all wavelengths of visible light are emitted with approximately equal intensities. The white light spectrum shown for an object at 6000 K closely approximates the spectrum of light emitted by the sun (Figure 3.4.1). Note the sharp decrease in the intensity of radiation emitted at wavelengths below 400 nm, which constituted the ultraviolet catastrophe. The classical prediction fails to fit the experimental curves entirely and does not have a maximum intensity.

Max Planck (1858–1947)

In addition to being a physicist, Planck was a gifted pianist, who at one time considered music as a career. During the 1930s, Planck felt it was his duty to remain in Germany, despite his open opposition to the policies of the Nazi government.



One of his sons was executed in 1944 for his part in an unsuccessful attempt to assassinate Hitler, and bombing during the last weeks of World War II destroyed Planck's home. After WWII, the major German scientific research organization was renamed the Max Planck Society.

Although quantization may seem to be an unfamiliar concept, we encounter it frequently. For example, US money is integral multiples of pennies. Similarly, musical instruments like a piano or a trumpet can produce only certain musical notes, such as C or F sharp. Because these instruments cannot produce a continuous range of frequencies, their frequencies are quantized. Even electrical charge is quantized: an ion may have a charge of −1 or −2 but *not* −1.33 electron charges.

Planck postulated that the energy of a particular quantum of radiant energy could be described Max Planck (1858–1947) by the equation

$$E = hv \tag{3.4.1}$$

where the proportionality constant h is called Planck's constant, one of the most accurately known fundamental constants in science. For our purposes, its value to four significant figures is generally sufficient:

$$h = 6.626 \times 10^{-34}\ \text{J•s (joule-seconds)}$$

As the frequency of electromagnetic radiation increases, the magnitude of the associated quantum of radiant energy increases. By assuming that energy can be emitted by an object only in integral multiples of $h\nu$, Planck devised an equation that fit the experimental data shown in Figure 3.4.2. We can understand Planck's explanation of the ultraviolet catastrophe qualitatively as follows: At low temperatures, radiation with only relatively low frequencies is emitted, corresponding to low-energy quanta. As the temperature of an object increases, there is an increased probability of emitting radiation with higher frequencies, corresponding to higher-energy quanta. At any temperature, however, it is simply more probable for an object to lose energy by emitting a large number of lower-energy quanta than a single very high-energy quantum that corresponds to ultraviolet radiation. The result is a maximum in the plot of intensity of emitted radiation versus wavelength, as shown in Figure 3.4.2, and a shift in the position of the maximum to lower wavelength (higher frequency) with increasing temperature.

At the time he proposed his radical hypothesis, Planck could not explain *why* energies should be quantized. Initially, his hypothesis explained only one set of experimental data—blackbody radiation. If quantization were observed for a large number of different phenomena, then quantization would become a law. In time, a theory might be developed to explain that law. As things turned out, Planck's hypothesis was the seed from which modern physics grew.

The Photoelectric Effect

Only five years after he proposed it, Planck's quantization hypothesis was used to explain a second phenomenon that conflicted with the accepted laws of classical physics. When certain metals are exposed to light, electrons are ejected from their surface (Figure 3.4.3). Classical physics predicted that the number of electrons emitted and their kinetic energy should depend on only the intensity of the light, not its frequency. In fact, however, each metal was found to have a characteristic threshold frequency of light; below that frequency, no electrons are emitted regardless of the light's intensity. Above the threshold frequency, the number of electrons emitted was found to be proportional to the intensity of the light, and their kinetic energy was proportional to the frequency. This phenomenon was called the photoelectric effect (A phenomenon in which electrons are ejected from the surface of a metal that has been exposed to light).

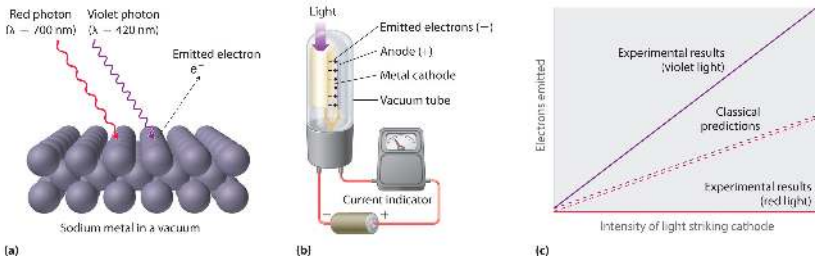


Figure 3.4.3: The Photoelectric Effect *(a) Irradiating a metal surface with photons of sufficiently high energy causes electrons to be ejected from the metal. (b) A photocell that uses the photoelectric effect, similar to those found in automatic door openers. When light strikes the metal cathode, electrons are emitted and attracted to the anode, resulting in a flow of electrical current. If the incoming light is interrupted by, for example, a passing person, the current drops to zero. (c) In contrast to predictions using classical physics, no electrons are emitted when photons of light with energy less than E_o, such as red light, strike the cathode. The energy of violet light is above the threshold frequency, so the number of emitted photons is proportional to the light's intensity.*

Albert Einstein (1879–1955; Nobel Prize in Physics, 1921) quickly realized that Planck's hypothesis about the quantization of radiant energy could also explain the photoelectric effect. The key feature of Einstein's hypothesis was the assumption that radiant energy arrives at the metal surface in particles that we now call photons (a quantum of radiant energy, each of which possesses a particular energy energy E given by Equation 8.3.1 Einstein postulated that each metal has a particular electrostatic attraction for its electrons that must be overcome before an electron can be emitted from its surface ($E_o = h\nu_o$). If photons of light with energy less than E_o strike a metal surface, no single photon has enough energy to eject an electron, so no electrons are emitted regardless of the intensity of the light. If a photon with energy greater than E_o strikes the metal, then part of its energy is used to overcome the forces that hold the electron to the metal surface, and the excess energy appears as the kinetic energy of the ejected electron:

$$kinetic\ energy\ of\ ejected\ electron = E - E_o = h\nu - h\nu_o = h\left(\nu - \nu_o\right) \tag{3.4.2}$$

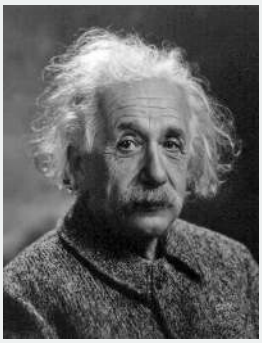
When a metal is struck by light with energy above the threshold energy E_o, the *number* of emitted electrons is proportional to the *intensity* of the light beam, which corresponds to the number of photons per square centimeter, but the *kinetic energy* of the emitted electrons is proportional to the *frequency* of the light. Thus Einstein showed that the energy of the emitted electrons depended on the frequency of the light, contrary to the prediction of classical physics.



Figure 3.4.4: A Beam of Red Light Emitted by a Helium Neon laser reads a bar code. Originally Helium neon lasers, which emit red light at a wavelength of 632.8 nm, were used to read bar codes. Today, smaller, inexpensive diode lasers are used.

Albert Einstein (1879–1955)

In 1900, Einstein was working in the Swiss patent office in Bern. He was born in Germany and throughout his childhood his parents and teachers had worried that he might be developmentally disabled. The patent office job was a low-level civil service position that was not very demanding, but it did allow Einstein to spend a great deal of time reading and thinking about physics.



In 1905, his "miracle year" he published four papers that revolutionized physics. One was on the special theory of relativity, a second on the equivalence of mass and energy, a third on Brownian motion, and the fourth on the photoelectric effect, for which he received the Nobel Prize in 1921, the theory of relativity and energy-matter equivalence being still controversial at the time

Planck's and Einstein's postulate that energy is quantized is in many ways similar to Dalton's description of atoms. Both theories are based on the existence of simple building blocks, atoms in one case and quanta of energy in the other. The work of Planck and Einstein thus suggested a connection between the quantized nature of energy and the properties of individual atoms.

Example 3.4.1

A ruby laser, a device that produces light in a narrow range of wavelengths emits red light at a wavelength of 694.3 nm (Figure 3.4.4). What is the energy in joules of a single photon?

Given: wavelength

Asked for: energy of single photon.

Strategy:

A. Use Equation 6.1.2 and Equation 8.3.1 to calculate the energy in joules.

Solution:

The energy of a single photon is given by $E = h\nu = hc/\lambda$.

Exercise 3.4.1

An x-ray generator, such as those used in hospitals, emits radiation with a wavelength of 1.544 Å.

a. What is the energy in joules of a single photon?
b. How many times more energetic is a single x-ray photon of this wavelength than a photon emitted by a ruby laser?

Answer

a. $1.287 \times 10^{-15}\ J/photon$

b. 4497 times



The Photoelectric Effect: https://youtu.be/mxBMxJLauQk

Summary

- The fundamental building blocks of energy are quanta and of matter are atoms.

The properties of **blackbody radiation**, the radiation emitted by hot objects, could not be explained with classical physics. Max Planck postulated that energy was quantized and could be emitted or absorbed only in integral multiples of a small unit of energy, known as a **quantum**. The energy of a quantum is proportional to the frequency of the radiation; the proportionality constant h is a fundamental constant (Planck's constant). Albert Einstein used Planck's concept of the quantization of energy to explain the **photoelectric effect**, the ejection of electrons from certain metals when exposed to light. Einstein postulated the existence of what today we call **photons**, particles of light with a particular energy, $E = h\nu$. Both energy and matter have fundamental building blocks: quanta and atoms, respectively.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

3.4: Quantum Theory is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.3: Quantum Theory** is licensed CC BY-NC-SA 4.0.

3.5: The Bohr Atom

Learning Objectives

- To know the relationship between atomic spectra and the electronic structure of atoms.

In 1913, a Danish physicist, Niels Bohr (1885–1962; Nobel Prize in Physics, 1922), proposed a theoretical model for the hydrogen atom that explained its emission spectrum. Bohr's model required only one assumption: *The electron moves around the nucleus in circular orbits that can have only certain allowed radii*. Rutherford's earlier model of the atom had also assumed that electrons moved in circular orbits around the nucleus and that the atom was held together by the electrostatic attraction between the positively charged nucleus and the negatively charged electron. Although we now know that the assumption of circular orbits was incorrect, Bohr's insight was to propose that *the electron could occupy only certain regions of space*.

Using classical physics, Niels Bohr showed that the energy of an electron in a particular orbit is given by

$$E_n = \frac{-\Re hc}{n^2} \tag{3.5.1}$$

where $\Re$ is the Rydberg constant, h is Planck's constant, c is the speed of light, and n is a positive integer corresponding to the number assigned to the orbit, with $n = 1$ corresponding to the orbit closest to the nucleus. In this model $n = \infty$ corresponds to the level where the energy holding the electron and the nucleus together is zero. In that level, the electron is unbound from the nucleus and the atom has been separated into a negatively charged (the electron) and a positively charged (the nucleus) ion. In this state the radius of the orbit is also infinite. The atom has been ionized.

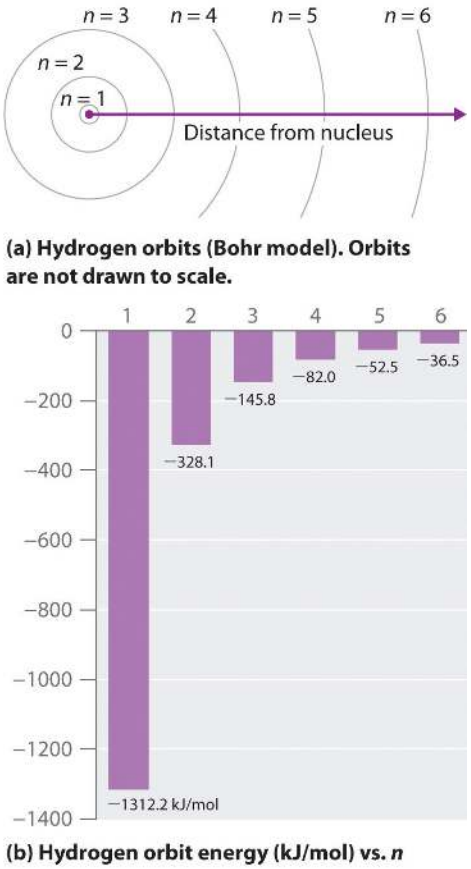
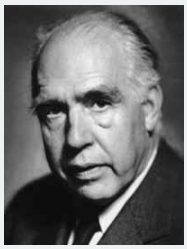


Figure 3.5.2 The Bohr Model of the Hydrogen Atom (a) The distance of the orbit from the nucleus increases with increasing n. (b) The energy of the orbit becomes increasingly less negative with increasing n.

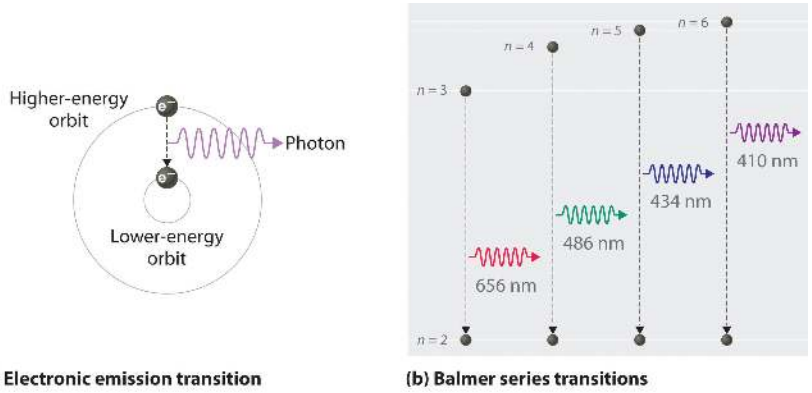
Niels Bohr (1885–1962)

During the Nazi occupation of Denmark in World War II, Bohr escaped to the United States, where he became associated with the Atomic Energy Project.



In his final years, he devoted himself to the peaceful application of atomic physics and to resolving political problems arising from the development of atomic weapons.

As n decreases, the energy holding the electron and the nucleus together becomes increasingly negative, the radius of the orbit shrinks and more energy is needed to ionize the atom. The orbit with n = 1 is the lowest lying and most tightly bound. The negative sign in Equation 3.5.3 indicates that the electron-nucleus pair is more tightly bound when they are near each other than when they are far apart. Because a hydrogen atom with its one electron in this orbit has the lowest possible energy, this is the ground state (the most stable arrangement of electrons for an element or a compound), the most stable arrangement for a hydrogen atom. As n increases, the radius of the orbit increases; the electron is farther from the proton, which results in a less stable arrangement with higher potential energy (Figure 2.10). A hydrogen atom with an electron in an orbit with n > 1 is therefore in an excited state. Any arrangement of electrons that is higher in energy than the ground state.: its energy is higher than the energy of the ground state. When an atom in an excited state undergoes a transition to the ground state in a process called decay, it loses energy by emitting a photon whose energy corresponds to the difference in energy between the two states (Figure 3.5.1).



(a) Electronic emission transition (b) Balmer series transitions

Figure 3.5.3: The Emission of Light by a Hydrogen Atom in an Excited State. (a) Light is emitted when the electron undergoes a transition from an orbit with a higher value of *n* (at a higher energy) to an orbit with a lower value of *n* (at lower energy). (b) The Balmer series of emission lines is due to transitions from orbits with $n \geq 3$ to the orbit with $n = 2$. The differences in energy between these levels corresponds to light in the visible portion of the electromagnetic spectrum.

So the difference in energy (ΔE) between any two orbits or energy levels is given by $\Delta E = E_{n_1} - E_{n_2}$ where n_1 is the final orbit and n_2 the initial orbit. Substituting from Bohr's equation (Equation 3.5.3) for each energy value gives

$$\Delta E = E_{final} - E_{initial} = -\frac{\Re hc}{n_2^2} - \left(-\frac{\Re hc}{n_1^2}\right) = -\Re hc\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) \tag{3.5.2}$$

If $n_2 > n_1$, the transition is from a higher energy state (larger-radius orbit) to a lower energy state (smaller-radius orbit), as shown by the dashed arrow in part (a) in Figure 3.5.3. Substituting hc/λ for ΔE gives

$$\Delta E = \frac{hc}{\lambda} = -\Re hc\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) \tag{3.5.3}$$

Canceling *hc* on both sides gives

$$\frac{1}{\lambda} = -\Re\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) \tag{3.5.4}$$

Except for the negative sign, this is the same equation that Rydberg obtained experimentally. The negative sign in Equation 3.5.5 and Equation 3.5.6 indicates that energy is released as the electron moves from orbit n_2 to orbit n_1 because orbit n_2 is at a higher energy than orbit n_1. Bohr calculated the value of $\Re$ from fundamental constants such as the charge and mass of the electron and Planck's constant and obtained a value of 1.0974×10^7 m^{-1}, the same number Rydberg had obtained by analyzing the emission spectra.

We can now understand the physical basis for the Balmer series of lines in the emission spectrum of hydrogen (part (b) in Figure 2.9). As shown in part (b) in Figure 3.5.3, the lines in this series correspond to transitions from higher-energy orbits (n > 2) to the second orbit (n = 2). Thus the hydrogen atoms in the sample have absorbed energy from the electrical discharge and decayed from a higher-energy excited state (n > 2) to a lower-energy state (n = 2) by emitting a photon of electromagnetic radiation whose energy corresponds exactly to the difference in energy between the two states (part (a) in Figure 3.5.3). The n = 3 to n = 2 transition gives rise to the line at 656 nm (red), the n = 4 to n = 2 transition to the line at 486 nm (green), the n = 5 to n = 2 transition to the line at 434 nm (blue), and the n = 6 to n = 2 transition to the line at 410 nm (violet). Because a sample of hydrogen contains a large number of atoms, the intensity of the various lines in a line spectrum depends on the number of atoms in each excited state. At the temperature in the gas discharge tube, more atoms are in the n = 3 than the n ≥ 4 levels. Consequently, the n = 3 to n = 2 transition is the most intense line, producing the characteristic red color of a hydrogen discharge (part (a) in Figure 3.5.1). Other families of lines are produced by transitions from excited states with n > 1 to the orbit with n = 1 or to orbits with n ≥ 3. These transitions are shown schematically in Figure 3.5.4

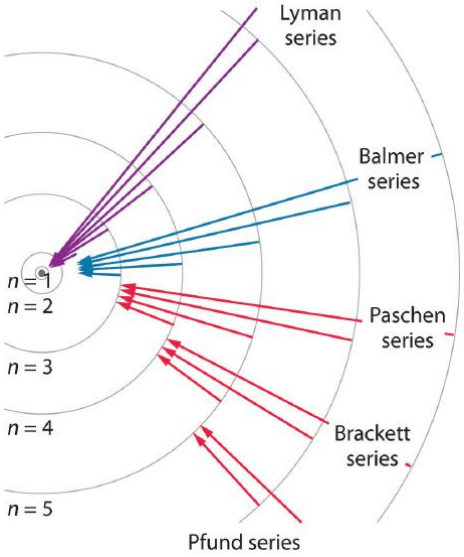


Figure 3.5.4: Electron Transitions Responsible for the Various Series of Lines Observed in the Emission Spectrum of Hydrogen. The Lyman series of lines is due to transitions from higher-energy orbits to the lowest-energy orbit (n = 1); these transitions release a great deal of energy, corresponding to radiation in the ultraviolet portion of the electromagnetic spectrum. The Paschen, Brackett, and Pfund series of lines are due to transitions from higher-energy orbits to orbits with n = 3, 4, and 5, respectively; these transitions release substantially less energy, corresponding to infrared radiation. (Orbits are not drawn to scale.)

In contemporary applications, electron transitions are used in timekeeping that needs to be exact. Telecommunications systems, such as cell phones, depend on timing signals that are accurate to within a millionth of a second per day, as are the devices that control the US power grid. Global positioning system (GPS) signals must be accurate to within a billionth of a second per day, which is equivalent to gaining or losing no more than one second in 1,400,000 years. Quantifying time requires finding an event with an interval that repeats on a regular basis. To achieve the accuracy required for modern purposes, physicists have turned to the atom. The current standard used to calibrate clocks is the cesium atom. Supercooled cesium atoms are placed in a vacuum chamber and bombarded with microwaves whose frequencies are carefully controlled. When the frequency is exactly right, the atoms absorb enough energy to undergo an electronic transition to a higher-energy state. Decay to a lower-energy state emits radiation. The microwave frequency is continually adjusted, serving as the clock's pendulum. In 1967, the second was defined as the duration of 9,192,631,770 oscillations of the resonant frequency of a cesium atom, called the *cesium clock*. Research is currently under way to develop the next generation of atomic clocks that promise to be even more accurate. Such devices would allow scientists to monitor

vanishingly faint electromagnetic signals produced by nerve pathways in the brain and geologists to measure variations in gravitational fields, which cause fluctuations in time, that would aid in the discovery of oil or minerals.

Example 3.5.1: The Lyman Series

The so-called Lyman series of lines in the emission spectrum of hydrogen corresponds to transitions from various excited states to the n = 1 orbit. Calculate the wavelength of the lowest-energy line in the Lyman series to three significant figures. In what region of the electromagnetic spectrum does it occur?

Given: lowest-energy orbit in the Lyman series

Asked for: wavelength of the lowest-energy Lyman line and corresponding region of the spectrum

Strategy:

A. Substitute the appropriate values into Equation 3.5.2 (the Rydberg equation) and solve for λ.
B. Use Figure 3.5.1 to locate the region of the electromagnetic spectrum corresponding to the calculated wavelength.

Solution:

We can use the Rydberg equation to calculate the wavelength:

$$\frac{1}{\lambda} = -\Re\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) \tag{3.5.5}$$

A For the Lyman series, n_1 = 1. The lowest-energy line is due to a transition from the n = 2 to n = 1 orbit because they are the closest in energy.

$$\frac{1}{\lambda} = -\Re\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) = 1.097 \times m^{-1}\left(\frac{1}{1} - \frac{1}{4}\right) = 8.228 \times 10^6\ m^{-1} \tag{3.5.6}$$

It turns out that spectroscopists (the people who study spectroscopy) use cm^{-1} rather than m^{-1} as a common unit. Wavelength is inversely proportional to energy but frequency is directly proportional as shown by Planck's formula, E=hu.

Spectroscopists often talk about energy and frequency as equivalent. The cm^{-1} unit is particularly convenient. The infrared range is roughly 200 - 5,000 cm^{-1}, the visible from 11,000 to 25.000 cm^{-1} and the UV between 25,000 and 100,000 cm^{-1}. The units of cm^{-1} are called wavenumbers, although people often verbalize it as inverse centimeters. We can convert the answer in part A to cm^{-1}.

$$\varpi = \frac{1}{\lambda} = 8.228 \times 10^6\ \cancel{m^{-1}}\left(\frac{\cancel{m}}{100\ cm}\right) = 82,280\ cm^{-1} \tag{3.5.7}$$

and

$$\lambda = 1.215 \times 10^{-7}\ m = 122\ nm \tag{3.5.8}$$

This emission line is called Lyman alpha. It is the strongest atomic emission line from the sun and drives the chemistry of the upper atmosphere of all the planets producing ions by stripping electrons from atoms and molecules. It is completely absorbed by oxygen in the upper stratosphere, dissociating O_2 molecules to O atoms which react with other O_2 molecules to form stratospheric ozone

B This wavelength is in the ultraviolet region of the spectrum.

Exercise 3.5.1: The Pfund Series

The Pfund series of lines in the emission spectrum of hydrogen corresponds to transitions from higher excited states to the n = 5 orbit. Calculate the wavelength of the *second* line in the Pfund series to three significant figures. In which region of the spectrum does it lie?

Answer: 4.65 × 10^3 nm; infrared

Bohr's model of the hydrogen atom gave an exact explanation for its observed emission spectrum. The following are his key contributions to our understanding of atomic structure:

- Electrons can occupy only certain regions of space, called *orbits*.
- Orbits closer to the nucleus are lower in energy.
- Electrons can move from one orbit to another by absorbing or emitting energy, giving rise to characteristic spectra.

Unfortunately, Bohr could not explain *why* the electron should be restricted to particular orbits. Also, despite a great deal of tinkering, such as assuming that orbits could be ellipses rather than circles, his model could not quantitatively explain the emission spectra of any element other than hydrogen (Figure 3.5.5). In fact, Bohr's model worked only for species that contained just one electron: H, He^{+}, Li^{2+}, and so forth. Scientists needed a fundamental change in their way of thinking about the electronic structure of atoms to advance beyond the Bohr model.

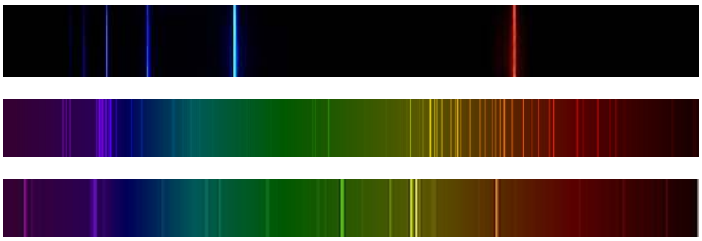


Figure 3.5.5: The Emission Spectra of Elements Compared with Hydrogen. These images show (a) hydrogen gas, which is atomized to hydrogen atoms in the discharge tube; (b) neon; and (c) mercury. The strongest lines in the hydrogen spectrum are in the far UV Lyman series starting at 124 nm and below. The strongest lines in the mercury spectrum are at 181 and 254 nm, also in the UV. These are not shown.

Thus far we have explicitly considered only the emission of light by atoms in excited states, which produces an emission spectrum (a spectrum produced by the emission of light by atoms in excited states). The converse, absorption of light by ground-state atoms to produce an excited state, can also occur, producing an absorption spectrum (a spectrum produced by the absorption of light by ground-state atoms). Because each element has characteristic emission and absorption spectra, scientists can use such spectra to analyze the composition of matter.

> When an atom emits light, it decays to a lower energy state; when an atom absorbs light, it is excited to a higher energy state.

The Energy States of the Hydrogen Atom

If white light is passed through a sample of hydrogen, hydrogen atoms absorb energy as an electron is excited to higher energy levels (orbits with $n \geq 2$). If the light that emerges is passed through a prism, it forms a continuous spectrum with *black* lines (corresponding to no light passing through the sample) at 656, 468, 434, and 410 nm. These wavelengths correspond to the $n = 2$ to $n = 3$, $n = 2$ to $n = 4$, $n = 2$ to $n = 5$, and $n = 2$ to $n = 6$ transitions. Any given element therefore has both a characteristic emission spectrum and a characteristic absorption spectrum, which are essentially complementary images.

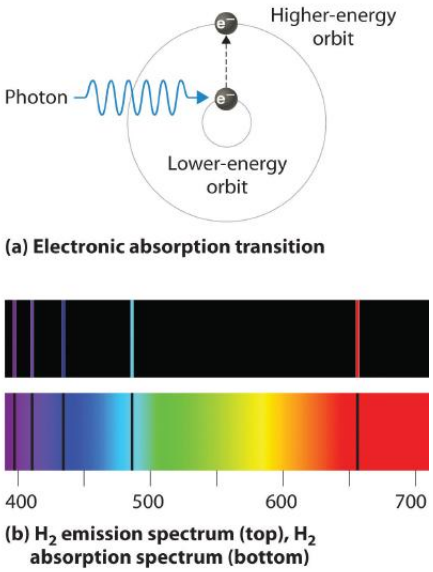


Figure 3.5.6: Absorption and Emission Spectra. Absorption of light by a hydrogen atom. (a) When a hydrogen atom absorbs a photon of light, an electron is excited to an orbit that has a higher energy and larger value of n. (b) Images of the emission and absorption spectra of hydrogen are shown here.

Emission and absorption spectra form the basis of *spectroscopy*, which uses spectra to provide information about the structure and the composition of a substance or an object. In particular, astronomers use emission and absorption spectra to determine the composition of stars and interstellar matter. As an example, consider the spectrum of sunlight shown in Figure 3.5.7 Because the sun is very hot, the light it emits is in the form of a continuous emission spectrum. Superimposed on it, however, is a series of dark lines due primarily to the absorption of specific frequencies of light by cooler atoms in the outer atmosphere of the sun. By comparing these lines with the spectra of elements measured on Earth, we now know that the sun contains large amounts of hydrogen, iron, and carbon, along with smaller amounts of other elements. During the solar eclipse of 1868, the French astronomer Pierre Janssen (1824–1907) observed a set of lines that did not match those of any known element. He suggested that they were due to the presence of a new element, which he named *helium*, from the Greek *helios*, meaning "sun." Helium was finally discovered in uranium ores on Earth in 1895. Alpha particles are helium nuclei. Alpha particles emitted by the radioactive uranium, pick up electrons from the rocks to form helium atoms.

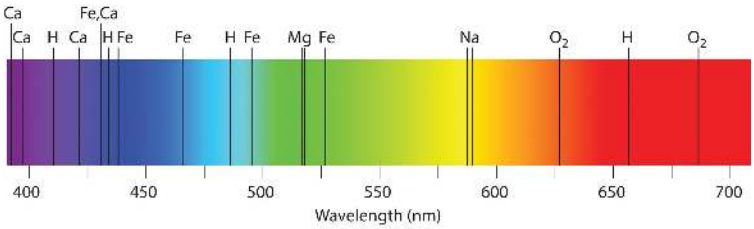


Figure 3.5.7: The Visible Spectrum of Sunlight. The characteristic dark lines are mostly due to the absorption of light by elements that are present in the cooler outer part of the sun's atmosphere; specific elements are indicated by the labels. The lines at 628 and 687 nm, however, are due to the absorption of light by oxygen molecules in Earth's atmosphere.

The familiar red color of "neon" signs used in advertising is due to the emission spectrum of neon shown in part (b) in Figure 3.5.5. Similarly, the blue and yellow colors of certain street lights are caused, respectively, by mercury and sodium discharges. In all these cases, an electrical discharge excites neutral atoms to a higher energy state, and light is emitted when the atoms decay to the ground state. In the case of mercury, most of the emission lines are below 450 nm, which produces a blue light (part (c) in Figure 3.5.5). In the case of sodium, the most intense emission lines are at 589 nm, which produces an intense yellow light.

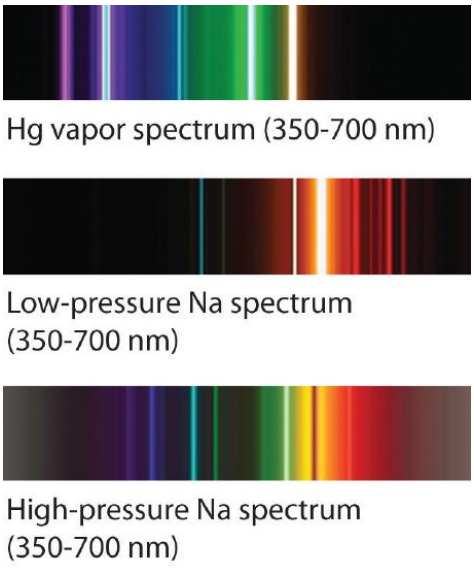


Figure 3.5.8: The emission spectra of sodium and mercury. Sodium and mercury spectra. Many street lights use bulbs that contain sodium or mercury vapor. Due to the very different emission spectra of these elements, they emit light of different colors. The lines in the sodium lamp are broadened by collisions. The dark line in the center of the high pressure sodium lamp where the low pressure lamp is strongest is cause by absorption of light in the cooler outer part of the lamp.



The Bohr Atom: https://youtu.be/GuFQEOzFOgA

The Chemistry of Fireworks

The colors of fireworks are also due to atomic emission spectra. As shown in part (a) in Figure 3.5.9, a typical shell used in a fireworks display contains gunpowder to propel the shell into the air and a fuse to initiate a variety of reactions that produce heat and small explosions. Thermal energy excites the atoms to higher energy states; as they decay to lower energy states, the atoms emit light that gives the familiar colors.



Figure 3.5.9 :The Chemistry of Fireworks *(a) In the "multibreak" shell used for fireworks, the chambers contain mixtures of fuels and oxidizers plus compounds for special effects ("stars") connected by time-delay fuses so that the chambers explode in stages. (b) The finale of a fireworks display usually consists of many shells fired simultaneously to give a dazzling multicolor display. The labels indicate the substances that are responsible for the colors of some of the fireworks shown.*

When oxidant/reductant mixtures listed in Table 3.5.1 are ignited, a flash of white or yellow light is produced along with a loud bang. Achieving the colors shown in part (b) in Figure 3.5.9 requires adding a small amount of a substance that has an emission spectrum in the desired portion of the visible spectrum. For example, sodium is used for yellow because of its 589 nm emission lines. The intense yellow color of sodium would mask most other colors, so potassium and ammonium salts, rather than sodium salts, are usually used as oxidants to produce other colors, which explains the preponderance of such salts in Table 3.5.1. Strontium salts, which are also used in highway flares, emit red light, whereas barium gives a green color. Blue is one of the most difficult colors to achieve. Copper(II) salts emit a pale blue light, but copper is dangerous to use because it forms highly unstable explosive compounds with anions such as chlorate. As you might guess, preparing fireworks with the desired properties is a complex, challenging, and potentially hazardous process. If you have the time here is a NOVA program about how fireworks are made.

Table 3.5.1: Common Chemicals Used in the Manufacture of Fireworks*

Oxidizers	Fuels (reductants)	Special effects
ammonium perchlorate	aluminum	*blue flame*: copper carbonate, copper sulfate, or copper oxide
barium chlorate	antimony sulfide	*red flame*: strontium nitrate or strontium carbonate
barium nitrate	charcoal	*white flame*: magnesium or aluminum
potassium chlorate	magnesium	*yellow flame*: sodium oxalate or cryolite (Na_3AlF_6)
potassium nitrate	sulfur	*green flame*: barium nitrate or barium chlorate
potassium perchlorate	titanium	*white smoke*: potassium nitrate plus sulfur
strontium nitrate		*colored smoke*: potassium chlorate and sulfur, plus organic dye
		whistling noise: potassium benzoate or sodium salicylate
		white sparks: aluminum, magnesium, or titanium
		gold sparks: iron fillings or charcoal

*Almost any combination of an oxidizer and a fuel may be used along with the compounds needed to produce a desired special effect.



Summary

There is an intimate connection between the atomic structure of an atom and its spectral characteristics. Atoms of individual elements emit light at only specific wavelengths, producing a **line spectrum** rather than the continuous spectrum of all wavelengths produced by a hot object. Niels Bohr explained the line spectrum of the hydrogen atom by assuming that the electron moved in circular orbits and that orbits with only certain radii were allowed. Lines in the spectrum were due to transitions in which an electron moved from a higher-energy orbit with a larger radius to a lower-energy orbit with smaller radius. The orbit closest to the nucleus represented the **ground state** of the atom and was most stable; orbits farther away were higher-energy **excited states**. Transitions from an excited state to a lower-energy state resulted in the emission of light with only a limited number of wavelengths. Bohr's model could not, however, explain the spectra of atoms heavier than hydrogen.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

3.5: The Bohr Atom is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.4: The Bohr Atom** is licensed CC BY-NC-SA 4.0.

3.6: Two Ideas Leading to a New Quantum Mechanics

Learning Objectives

- To understand the wave–particle duality of matter.

Einstein's photons of light were individual packets of energy having many of the characteristics of particles. Recall that the collision of an electron (a particle) with a sufficiently energetic photon can eject a *photoelectron* from the surface of a metal. Any excess energy is transferred to the electron and is converted to the kinetic energy of the ejected electron. Einstein's hypothesis that energy is concentrated in localized bundles, however, was in sharp contrast to the classical notion that energy is spread out uniformly in a wave. We now describe Einstein's theory of the relationship between energy and mass, a theory that others built on to develop our current model of the atom.

The Wave Character of Matter

Einstein initially assumed that photons had zero mass, which made them a peculiar sort of particle indeed. In 1905, however, he published his special theory of relativity, which related energy and mass according to the following equation:

$$E = h\nu = h\frac{c}{\lambda} = mc^2 \tag{3.6.1}$$

According to this theory, a photon of wavelength λ and frequency ν has a nonzero mass, which is given as follows:

$$m = \frac{E}{c^2} = \frac{h\nu}{c^2} = \frac{h}{\lambda c} \tag{3.6.2}$$

That is, light, which had always been regarded as a wave, also has properties typical of particles, a condition known as wave–particle duality (a principle that matter and energy have properties typical of both waves and particles). Depending on conditions, light could be viewed as either a wave or a particle.

In 1922, the American physicist Arthur Compton (1892–1962) reported the results of experiments involving the collision of x-rays and electrons that supported the particle nature of light. At about the same time, a young French physics student, Louis de Broglie (1892–1972), began to wonder whether the converse was true: Could particles exhibit the properties of waves? In his PhD dissertation submitted to the Sorbonne in 1924, de Broglie proposed that a particle such as an electron could be described by a wave whose wavelength is given by

$$\lambda = \frac{h}{mv} \tag{3.6.3}$$

where

- is Planck's constant,
- *m* is the mass of the particle, and
- *v* is the velocity of the particle.

This revolutionary idea was quickly confirmed by American physicists Clinton Davisson (1881–1958) and Lester Germer (1896–1971), who showed that beams of electrons, regarded as particles, were diffracted by a sodium chloride crystal in the same manner as x-rays, which were regarded as waves. It was proven experimentally that electrons do exhibit the properties of waves. For his work, de Broglie received the Nobel Prize in Physics in 1929.

If particles exhibit the properties of waves, why had no one observed them before? The answer lies in the numerator of de Broglie's equation, which is an extremely small number. As you will calculate in Example 3.6.1, Planck's constant (6.63×10^{-34} J•s) is so small that the wavelength of a particle with a large mass is too short (less than the diameter of an atomic nucleus) to be noticeable.

Example 3.6.1

Calculate the wavelength of a baseball, which has a mass of 149 g and a speed of 100 mi/h.

Given: mass and speed of object

Asked for: wavelength

Strategy:

A. Convert the speed of the baseball to the appropriate SI units: meters per second.
B. Substitute values into Equation 8.5.3 and solve for the wavelength.

Solution:

The wavelength of a particle is given by λ = *h*/*mv*. We know that *m* = 0.149 kg, so all we need to find is the speed of the baseball:

$$v = \left(\frac{100\ \cancel{mi}}{\cancel{h}}\right)\left(\frac{1\ \cancel{h}}{60\ \cancel{min}}\right)\left(\frac{1.609\ \cancel{km}}{\cancel{mi}}\right)\left(\frac{1000\ m}{\cancel{km}}\right)$$

B Recall that the joule is a derived unit, whose units are (kg•m^2)/s^2. Thus the wavelength of the baseball is

$$\lambda = \frac{6.626 \times 10^{-34}\ J \cdot s}{(0.149\ kg)\,(44.69\ m \cdot s)} = \frac{6.626 \times 10^{-34}\ \cancel{kg} \cdot m^{\cancel{2}} \cdot \cancel{s}^{\cancel{-2}} \cdot \cancel{s}}{\left(0.149\ \cancel{kg}\right)\left(44.69\ \cancel{m} \cdot\ s^{-1}\right)} = 9.95 \times 10^{-35}\ m \tag{3.6.4}$$

(You should verify that the units cancel to give the wavelength in meters.) Given that the diameter of the nucleus of an atom is approximately 10^{-14} m, the wavelength of the baseball is almost unimaginably small.

? Exercise 3.6.1

Calculate the wavelength of a neutron that is moving at 3.00 × 10^3 m/s.

Answer: 1.32 Å, or 132 pm

As you calculated in Example 3.6.1, objects such as a baseball or a neutron have such short wavelengths that they are best regarded primarily as particles. In contrast, objects with very small masses (such as photons) have large wavelengths and can be viewed primarily as waves. Objects with intermediate masses, such as electrons, exhibit the properties of both particles and waves. Although we still usually think of electrons as particles, the wave nature of electrons is employed in an *electron microscope*, which has revealed most of what we know about the microscopic structure of living organisms and materials. Because the wavelength of an electron beam is much shorter than the wavelength of a beam of visible light, this instrument can resolve smaller details than a light microscope can (Figure 3.6.1).

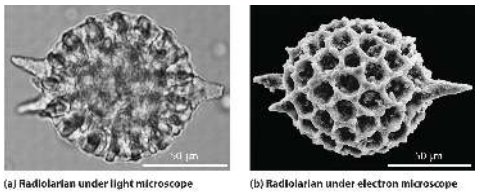


Figure 3.6.1: A Comparison of Images Obtained Using a Light Microscope and an Electron Microscope**.** Because of their shorter wavelength, high-energy electrons have a higher resolving power than visible light. Consequently, an electron microscope (b) is able to resolve finer details than a light microscope (a). (Radiolaria, which are shown here, are unicellular planktonic organisms.)

An Important Wave Property: Phase

A wave is a disturbance that travels in space. The magnitude of the wave at any point in space and time varies sinusoidally. While the absolute value of the magnitude of one wave at any point is not very important, the **relative** displacement of two waves called the phase difference, is vitally important because it determines whether the waves reinforce or interfere with each other. Figure 3.6.2a, on the right shows an arbitrary phase difference between two wave. Figure 3.6.2b shows what happens when the two waves are 180 degrees out of phase. The green line is their sum. Figure 3.6.2c shows what happens when the two lines are in phase, exactly superimposed on each other. Again, the green line is the sum of the intensities.

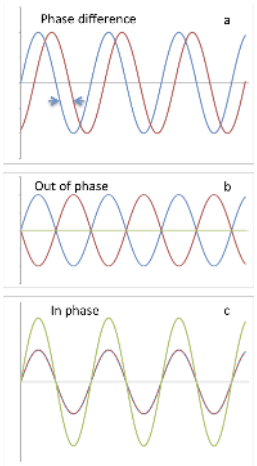


Figure 3.6.2: Phase. Two waves traveling together are displaced by a phase difference. If the phase difference is 0 then they lay on top of each other and reinforce. If the phase difference is 180 they completely cancel each other out.

Standing Waves

De Broglie also investigated why only certain orbits were allowed in Bohr's model of the hydrogen atom. He hypothesized that the electron behaves like a standing wave (a wave that does not travel in space). An example of a standing wave is the motion of a string of a violin or guitar. When the string is plucked, it vibrates at certain fixed frequencies because it is fastened at both ends (Figure 3.6.3). If the length of the string is *L*, then the lowest-energy vibration (the fundamental (the lowest-energy standing wave) has wavelength

$$\frac{\lambda}{2} = L$$
$$\lambda = 2L \tag{3.6.5}$$

Higher-energy vibrations are called overtones (the vibration of a standing wave that is higher in energy than the fundamental vibration) and are produced when the string is plucked more strongly; they have wavelengths given by

$$\lambda = \frac{2L}{n} \tag{3.6.6}$$

where *n* has any integral value. Thus the resonant vibrational energies of the string are quantized. When plucked, all other frequencies die out immediately. Only the resonant frequencies survive and are heard. By analogy we can think of the resonant frequencies as being quantized. Notice in Figure 3.6.3 that all overtones have one or more nodes (the points where the amplitude of a wave is zero), points where the string does not move. The amplitude of the wave at a node is zero.

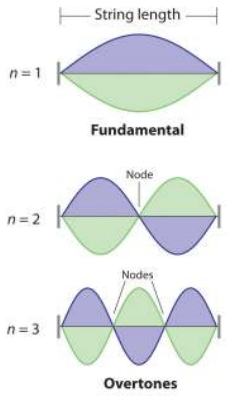


Figure 3.6.3 Standing Waves on a Vibrating String. The vibration with *n* = 1 is the fundamental and contains no nodes. Vibrations with higher values of *n* are called overtones; they contain *n* − 1 nodes.

Quantized vibrations and overtones containing nodes are not restricted to one-dimensional systems, such as strings. A two-dimensional surface, such as a drumhead, also has quantized vibrations. Similarly, when the ends of a string are joined to form a circle, the only allowed vibrations are those with wavelength

$$2\pi r = n\lambda \tag{3.6.7}$$

where *r* is the radius of the circle. De Broglie argued that Bohr's allowed orbits could be understood if the electron behaved like a *standing circular wave* (Figure 3.6.4). The standing wave could exist only if the circumference of the circle was an integral multiple of the wavelength such that the propagated waves were all in phase, thereby increasing the net amplitudes and causing *constructive interference*. Otherwise, the propagated waves would be out of phase, resulting in a net decrease in amplitude and causing *destructive interference.* The non resonant waves interfere with themselves! De Broglie's idea explained Bohr's allowed orbits and energy levels nicely: in the lowest energy level, corresponding to *n* = 1 in Equation 8.5.2, one complete wavelength would close the circle. Higher energy levels would have successively higher values of *n* with a corresponding number of nodes.

Standing waves are often observed on rivers, reservoirs, ponds, and lakes when seismic waves from an earthquake travel through the area. The waves are called *seismic seiches*, a term first used in 1955 when lake levels in England and Norway oscillated from side to side as a result of the Assam earthquake of 1950 in Tibet. They were first described in the Proceedings of the Royal Society in 1755 when they were seen in English harbors and ponds after a large earthquake in Lisbon, Portugal. Seismic seiches were also observed in many places in North America after the Alaska earthquake of March 28, 1964. Those occurring in western reservoirs lasted for two hours or longer, and amplitudes reached as high as nearly 6 ft along the Gulf Coast. The height of seiches is approximately proportional to the thickness of surface sediments; a deeper channel will produce a higher seiche.

Still, as all analogies, although the standing wave model helps us understand much about why Bohr's theory worked, it also, if pushed too far can mislead.

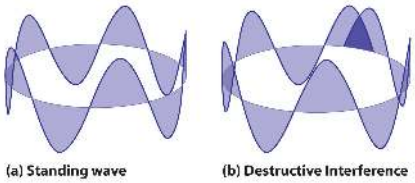


Figure 3.6.4 Standing Circular Wave and Destructive Interference.(a) In a standing circular wave with $n = 5$, the circumference of the circle corresponds to exactly five wavelengths, which results in constructive interference of the wave with itself when overlapping occurs. (b) If the circumference of the circle is not equal to an integral multiple of wavelengths, then the wave does not overlap exactly with itself, and the resulting destructive interference will result in cancellation of the wave. Consequently, a standing wave cannot exist under these conditions.

As you will see, some of de Broglie's ideas are retained in the modern theory of the electronic structure of the atom: the wave behavior of the electron and the presence of nodes that increase in number as the energy level increases. Unfortunately, his (and Bohr's) explanation also contains one major feature that we know to be incorrect: in the currently accepted model, the electron in a given orbit is *not* always at the same distance from the nucleus.

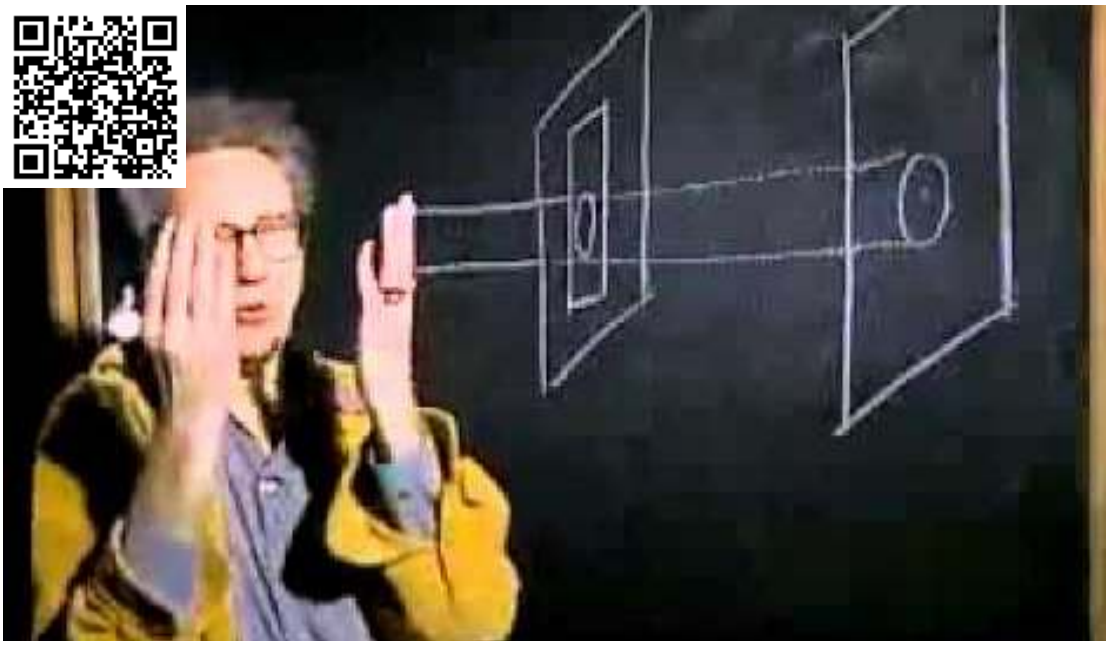


The de Broglie Equation: https://youtu.be/pz0zMHWtK7Q

The Heisenberg Uncertainty Principle

Because a wave is a disturbance that travels in space, it has no fixed position. One might therefore expect that it would also be hard to specify the exact position of a *particle* that exhibits wavelike behavior. A characteristic of light is that is can be bent or spread out by passing through a narrow slit as shown in the video below. You can literally see this by half closing your eyes and looking through your eye lashes. This reduces the brightness of what you are seeing and somewhat fuzzes out the image, but the light bends around your lashes to provide a complete image rather than a bunch of bars across the image. This is called diffraction.

This behavior of waves is captured in Maxwell's equations (1870 or so) for electromagnetic waves and was and is well understood. Heisenberg's uncertainty principle for light is, if you will, merely a conclusion about the nature of electromagnetic waves and nothing new. DeBroglie's idea of wave particle duality means that particles such as electrons which all exhibit wave like characteristics, will also undergo diffraction from slits whose size is of the order of the electron wavelength.



This situation was described mathematically by the German physicist Werner Heisenberg (1901–1976; Nobel Prize in Physics, 1932), who related the position of a particle to its momentum. Referring to the electron, Heisenberg stated that "at every moment the electron has only an inaccurate position and an inaccurate velocity, and between these two inaccuracies there is this uncertainty relation." Mathematically, the Heisenberg uncertainty principle is greater than or equal to Planck's constant *h* divided by 4π: states that the uncertainty in the position of a particle (Δ*x*) multiplied by the uncertainty in its momentum [Δ(*mv*)] is greater than or equal to Planck's constant divided by 4π:

$$(\Delta x)\,(\Delta\,[mv]) \geqslant \frac{h}{4\pi} \tag{3.6.8}$$

Because Planck's constant is a very small number, the Heisenberg uncertainty principle is important only for particles such as electrons that have very low masses. These are the same particles predicted by de Broglie's equation to have measurable wavelengths.

If the precise position *x* of a particle is known absolutely (Δ*x* = 0), then the uncertainty in its momentum must be infinite:

$$(\Delta\,[mv]) = \frac{h}{4\pi\,(\Delta x)} = \frac{h}{4\pi\,(0)} = \infty \tag{3.6.9}$$

Because the mass of the electron at rest (*m*) is both constant and accurately known, the uncertainty in Δ(*mv*) must be due to the Δ*v* term, which would have to be infinitely large for Δ(*mv*) to equal infinity. That is, according to Equation 8.5.8, the more accurately we know the exact position of the electron (as Δ*x* → 0), the less accurately we know the speed and the kinetic energy of the electron (1/2 mv^2) because Δ(*mv*) → ∞. Conversely, the more accurately we know the precise momentum (and the energy) of the electron [as Δ(*mv*) → 0], then Δ*x* → ∞ and we have no idea where the electron is.

Bohr's model of the hydrogen atom violated the Heisenberg uncertainty principle by trying to specify **simultaneously** both the position (an orbit of a particular radius) and the energy (a quantity related to the momentum) of the electron. Moreover, given its mass and wavelike nature, the electron in the hydrogen atom could not possibly orbit the nucleus in a well-defined circular path as predicted by Bohr. You will see, however, that the *most probable radius* of the electron in the hydrogen atom is exactly the one predicted by Bohr's model.

✔ Example 3.6.2

Calculate the minimum uncertainty in the position of the pitched baseball from Example 3.6.1 that has a mass of exactly 149 g and a speed of 100 ± 1 mi/h.

Given: mass and speed of object

Asked for: minimum uncertainty in its position

Strategy:

A. Rearrange the inequality that describes the Heisenberg uncertainty principle (Equation 8.5.7) to solve for the minimum uncertainty in the position of an object (Δ*x*).
B. Find Δ*v* by converting the velocity of the baseball to the appropriate SI units: meters per second.
C. Substitute the appropriate values into the expression for the inequality and solve for Δ*x*.

Solution:

A The Heisenberg uncertainty principle tells us that $(\Delta x)[\Delta(mv)] = h/4\pi$. Rearranging the inequality gives

$$\Delta x \geq \left(\frac{h}{4\pi}\right)\left(\frac{1}{\Delta(mv)}\right)$$

B We know that $h = 6.626 \times 10^{-34}$ J•s and $m = 0.149$ kg. Because there is no uncertainty in the mass of the baseball, $\Delta(mv) = m\Delta v$ and $\Delta v = \pm 1$ mi/h. We have

$$\Delta v = \left(\frac{1\ \cancel{mi}}{\cancel{h}}\right)\left(\frac{1\ \cancel{h}}{60\ \cancel{min}}\right)\left(\frac{1\ \cancel{min}}{60\ s}\right)\left(\frac{1.609\ \cancel{km}}{\cancel{mi}}\right)\left(\frac{1000\ m}{\cancel{km}}\right) = 0.4469\ m/s$$

C Therefore,

$$\Delta x \geqslant \left(\frac{6.626 \times 10^{-34}\ J \cdot s}{4\,(3.1416)}\right)\left(\frac{1}{(0.149\ kg)\,(0.4469\ m \cdot s^{-1})}\right)$$

Inserting the definition of a joule (1 J = 1 kg•m^2/s^2) gives

$$\Delta x \geqslant \left(\frac{6.626 \times 10^{-34}\ \cancel{kg} \cdot m^{\cancel{2}} \cdot \cancel{s}}{4\,(3.1416)\left(\cancel{s^2}\right)}\right)\left(\frac{1\ \cancel{s}}{\left(0.149\ \cancel{kg}\right)(0.4469\ \cancel{m})}\right)$$

$$\Delta x \geqslant 7.92 \pm \times 10^{-34}\ m$$

This is equal to 3.12×10^{-32} inches. We can safely say that if a batter misjudges the speed of a fastball by 1 mi/h (about 1%), he will not be able to blame Heisenberg's uncertainty principle for striking out.

? Exercise 3.6.2

Calculate the minimum uncertainty in the position of an electron traveling at one-third the speed of light, if the uncertainty in its speed is ±0.1%. Assume its mass to be equal to its mass at rest.

Answer: 6×10^{-10} m, or 0.6 nm (about the diameter of a benzene molecule)

Videos and Examples

- de Broglie Waves - Sixty Symbols - a bit of history and explanation (as well as wave particle duality can be explained)
- Calculating the wavelength of a proton - Josh Samson
- Calculating the wavelength of a small car - Prof. Heath
- What is the Uncertainty Principle? - Minute Physics
- AP Chem Heisenberg Uncertainty Principle - Mind Bite
- deBroglie Example - about education
- Heisenberg uncertainty principle quiz worked out - C Craig
- Quantum Chemistry - Ohio State
- Quantum Chemistry Quizzes - mhe education
- AP Chemistry Chapter 7 Review - Science Geek

Answers for these quizzes are included.

Summary

- An electron possesses both particle and wave properties.

The modern model for the electronic structure of the atom is based on recognizing that an electron possesses particle and wave properties, the so-called **wave–particle duality**. Louis de Broglie showed that the wavelength of a particle is equal to Planck's constant divided by the mass times the velocity of the particle.

$$\lambda = \frac{h}{mv} \tag{8.5.3}$$

The electron in Bohr's circular orbits could thus be described as a **standing wave**, one that does not move through space. Standing waves are familiar from music: the lowest-energy standing wave is the **fundamental** vibration, and higher-energy vibrations are **overtones** and have successively more **nodes**, points where the amplitude of the wave is always zero. Werner Heisenberg's **uncertainty principle** states that it is impossible to precisely describe both the location and the speed of particles that exhibit wavelike behavior.

$$(\Delta x)\,(\Delta\,[mv]) \geqslant \frac{h}{4\pi} \tag{8.5.7}$$

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

3.6: Two Ideas Leading to a New Quantum Mechanics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.5: Two Ideas Leading to a New Quantum Mechanics** is licensed CC BY-NC-SA 4.0.

3.7: Wave Mechanics

Learning Objectives

- To apply the results of quantum mechanics to chemistry.

The paradox described by Heisenberg's uncertainty principle and the wavelike nature of subatomic particles such as the electron made it impossible to use the equations of classical physics to describe the motion of electrons in atoms. Scientists needed a new approach that took the wave behavior of the electron into account. In 1926, an Austrian physicist, Erwin Schrödinger (1887–1961; Nobel Prize in Physics, 1933), developed *wave mechanics*, a mathematical technique that describes the relationship between the motion of a particle that exhibits wavelike properties (such as an electron) and its allowed energies. In doing so, Schrödinger developed the theory of quantum mechanics, which is used today to describe the energies and spatial distributions of electrons in atoms and molecules.

Schrödinger (1887–1961)

Schrödinger's unconventional approach to atomic theory was typical of his unconventional approach to life. He was notorious for his intense dislike of memorizing data and learning from books. When Hitler came to power in Germany, Schrödinger escaped to Italy. He then worked at Princeton University in the United States but eventually moved to the Institute for Advanced Studies in Dublin, Ireland, where he remained until his retirement in 1955.

Although quantum mechanics uses sophisticated mathematics, you do not need to understand the mathematical details to follow our discussion of its general conclusions. We focus on the properties of the *wave functions* that are the solutions of Schrödinger's equations.

Wave Functions

A wave function (Ψ) is a mathematical function that relates the location of an electron at a given point in space (identified by x, y, and z coordinates) to the amplitude of its wave, which corresponds to its energy. Thus each wave function is associated with a particular energy E. The properties of wave functions derived from quantum mechanics are summarized here:

- **A wave function uses three variables to describe the position of an electron.** A fourth variable is usually required to fully describe the location of objects in motion. Three specify the position in space (as with the Cartesian coordinates x, y, and z), and one specifies the time at which the object is at the specified location. For example, if you wanted to intercept an enemy submarine, you would need to know its latitude, longitude, and depth, as well as the time at which it was going to be at this position (Figure 3.7.1). For electrons, we can ignore the time dependence because we will be using standing waves, which by definition do not change with time, to describe the position of an electron.

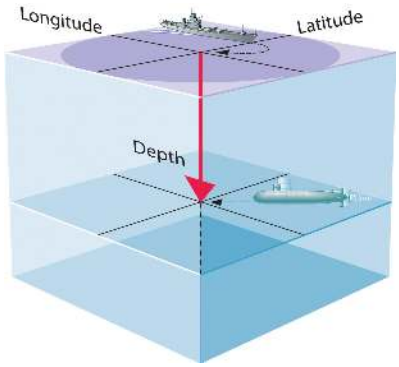


Figure 3.7.1: The Four Variables (Latitude, Longitude, Depth, and Time) Required to Precisely Locate an Object

If you are the captain of a ship trying to intercept an enemy submarine, you need to deliver your depth charge to the right location at the right time.

- **The magnitude of the wave function at a particular point in space is proportional to the amplitude of the wave at that point.** Many wave functions are complex functions, which is a mathematical term indicating that they contain $\sqrt{-1}$,

represented as i. Hence the amplitude of the wave has no real physical significance. In contrast, the sign of the wave function (either positive or negative) corresponds to the phase of the wave, which will be important in our discussion of chemical bonding. The sign of the wave function should *not* be confused with a positive or negative electrical charge.

- **The square of the wave function at a given point is proportional to the probability of finding an electron at that point, which leads to a distribution of probabilities in space.** The square of the wave function (Ψ^2) is always a real quantity [recall that that $\sqrt{-1}^2 = -1$] is proportional to the probability of finding an electron at a given point. More accurately, the probability is given by the product of the wave function Ψ and its complex conjugate Ψ*, in which all terms that contain *i* are replaced by $-i$. We use probabilities because, according to Heisenberg's uncertainty principle, we cannot precisely specify the position of an electron. The probability of finding an electron at any point in space depends on several factors, including the distance from the nucleus and, in many cases, the atomic equivalent of latitude and longitude. As one way of graphically representing the probability distribution, the probability of finding an electron is indicated by the density of colored dots, as shown for the ground state of the hydrogen atom in Figure 3.7.2.

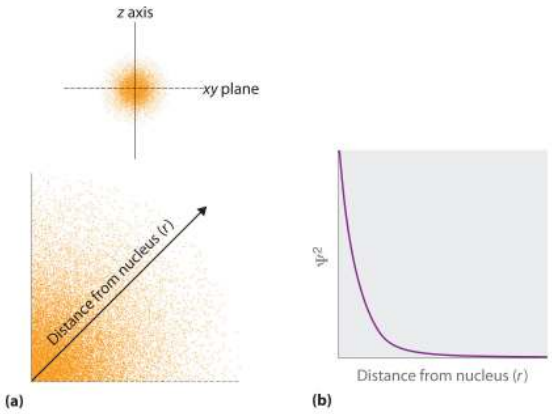


Figure 3.7.2: Probability of Finding the Electron in the Ground State of the Hydrogen Atom at Different Points in Space. (a) The density of the dots shows electron probability. (b) In this plot of Ψ^2 versus *r* for the ground state of the hydrogen atom, the electron probability density is greatest at *r* = 0 (the nucleus) and falls off with increasing *r*. Because the line never actually reaches the horizontal axis, the probability of finding the electron at very large values of *r* is very small but *not* zero.

- **Describing the electron distribution as a standing wave leads to sets of *quantum numbers* that are characteristic of each wave function.** From the patterns of one- and two-dimensional standing waves shown in Figure 6.18 and Figure 6.19, you might expect (correctly) that the patterns of three-dimensional standing waves would be complex. Fortunately, however, in the 18th century, a French mathematician, Adrien Legendre (1752–1783), developed a set of equations to describe the motion of tidal waves on the surface of a flooded planet. Schrödinger incorporated Legendre's equations into his wave functions. The requirement that the waves must be in phase with one another to avoid cancellation and produce a standing wave results in a limited number of solutions (wave functions), each of which is specified by a set of numbers called quantum numbers.
- **Each wave function is associated with a particular energy.** As in Bohr's model, the energy of an electron in an atom is quantized; it can have only certain allowed values. The major difference between Bohr's model and Schrödinger's approach is that Bohr had to impose the idea of quantization arbitrarily, whereas in Schrödinger's approach, quantization is a natural consequence of describing an electron as a standing wave.

Quantum Numbers

Schrödinger's approach uses three quantum numbers (*n*, *l*, and m_l) to specify any wave function. The quantum numbers provide information about the spatial distribution of an electron. Although *n* can be any positive integer, only certain values of *l* and m_l are allowed for a given value of *n*.



Introduction to Quantum Numbers: https://youtu.be/07JpBeaPxL8

The Principal Quantum Number

The **principal quantum number** (n) tells the average relative distance of an electron from the nucleus:

$$n = 1, 2, 3, 4, \ldots \tag{3.7.1}$$

As *n* increases for a given atom, so does the average distance of an electron from the nucleus. A negatively charged electron that is, on average, closer to the positively charged nucleus is attracted to the nucleus more strongly than an electron that is farther out in space. This means that electrons with higher values of *n* are easier to remove from an atom. All wave functions that have the same value of *n* are said to constitute a principal shell because those electrons have similar average distances from the nucleus. As you will see, the principal quantum number *n* corresponds to the *n* used by Bohr to describe electron orbits and by Rydberg to describe atomic energy levels.

The Azimuthal Quantum Number

The second quantum number is often called the azimuthal quantum number (*l*). The value of *l* describes the *shape* of the region of space occupied by the electron. The allowed values of *l* depend on the value of *n* and can range from 0 to *n* − 1:

$$l = 0, 1, 2, \ldots, n - 1 \tag{3.7.2}$$

For example, if *n* = 1, *l* can be only 0; if *n* = 2, *l* can be 0 or 1; and so forth. For a given atom, all wave functions that have the same values of both *n* and *l* form a subshell. The regions of space occupied by electrons in the same subshell usually have the same shape, but they are oriented differently in space.



Principal quantum number (n) & Orbital angular momentum (l): The Orbital Subshell: https://youtu.be/ms7WR149fAY

The Magnetic Quantum Number

The third quantum number is the magnetic quantum number (m_l). The value of m_l describes the *orientation* of the region in space occupied by an electron with respect to an applied magnetic field. The allowed values of m_l depend on the value of l: m_l can range from $-l$ to l in integral steps:

$$m_l = -l, -l+1, \ldots, 0, \ldots, l-1, l \tag{3.7.3}$$

For example, if $l = 0$, m_l can be only 0; if l = 1, m_l can be −1, 0, or +1; and if l = 2, m_l can be −2, −1, 0, +1, or +2.

Each wave function with an allowed combination of n, l, and m_l values describes an atomic orbital, a particular spatial distribution for an electron. For a given set of quantum numbers, each principal shell has a fixed number of subshells, and each subshell has a fixed number of orbitals.

Example 3.7.1: n=4 Shell Structure

How many subshells and orbitals are contained within the principal shell with n = 4?

Given: value of n

Asked for: number of subshells and orbitals in the principal shell

Strategy:

A. Given n = 4, calculate the allowed values of l. From these allowed values, count the number of subshells.
B. For each allowed value of l, calculate the allowed values of m_l. The sum of the number of orbitals in each subshell is the number of orbitals in the principal shell.

Solution:

A We know that l can have all integral values from 0 to n − 1. If n = 4, then l can equal 0, 1, 2, or 3. Because the shell has four values of l, it has four subshells, each of which will contain a different number of orbitals, depending on the allowed values of m_l.

B For l = 0, m_l can be only 0, and thus the l = 0 subshell has only one orbital. For l = 1, m_l can be 0 or ±1; thus the l = 1 subshell has three orbitals. For l = 2, m_l can be 0, ±1, or ±2, so there are five orbitals in the l = 2 subshell. The last allowed value of l is l = 3, for which m_l can be 0, ±1, ±2, or ±3, resulting in seven orbitals in the l = 3 subshell. The total number of orbitals in the n = 4 principal shell is the sum of the number of orbitals in each subshell and is equal to n^2:

Exercise 3.7.1: n=3 Shell Structure

How many subshells and orbitals are in the principal shell with n = 3?

Answer: three subshells; nine orbitals

Rather than specifying all the values of n and l every time we refer to a subshell or an orbital, chemists use an abbreviated system with lowercase letters to denote the value of l for a particular subshell or orbital:

l =	0	1	2	3
Designation	*s*	*p*	*d*	*f*

The principal quantum number is named first, followed by the letter *s*, *p*, *d*, or *f* as appropriate. These orbital designations are derived from corresponding spectroscopic characteristics: *s*harp, *p*rinciple, *d*iffuse, and *f*undamental. A 1*s* orbital has n = 1 and l = 0; a 2*p* subshell has n = 2 and l = 1 (and has three 2*p* orbitals, corresponding to m_l = −1, 0, and +1); a 3*d* subshell has n = 3 and l = 2 (and has five 3*d* orbitals, corresponding to m_l = −2, −1, 0, +1, and +2); and so forth.

We can summarize the relationships between the quantum numbers and the number of subshells and orbitals as follows (Table 3.7.1):

- Each principal shell has n subshells. For n = 1, only a single subshell is possible (1*s*); for n = 2, there are two subshells (2*s* and 2*p*); for n = 3, there are three subshells (3*s*, 3*p*, and 3*d*); and so forth. Every shell has an *ns* subshell, any shell with $n \geq 2$ also

has an *np* subshell, and any shell with $n \geq 3$ also has an *nd* subshell. Because a 2*d* subshell would require both $n = 2$ and $l = 2$, which is not an allowed value of *l* for $n = 2$, a 2*d* subshell does not exist.

- Each subshell has $2l + 1$ orbitals. This means that all *ns* subshells contain a single *s* orbital, all *np* subshells contain three *p* orbitals, all *nd* subshells contain five *d* orbitals, and all *nf* subshells contain seven *f* orbitals.

> **Note**
>
> Each principal shell has *n* subshells, and each subshell has $2l + 1$ orbitals.

Table 3.7.1 Values of n, l, and ml through n = 4

n	*l*	Subshell Designation	m_l	Number of Orbitals in Subshell	Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	4
	1	2*p*	−1, 0, 1	3	
3	0	3*s*	0	1	9
	1	3*p*	−1, 0, 1	3	
	2	3*d*	−2, −1, 0, 1, 2	5	
4	0	4*s*	0	1	16
	1	4*p*	−1, 0, 1	3	
	2	4*d*	−2, −1, 0, 1, 2	5	
	3	4*f*	−3, −2, −1, 0, 1, 2, 3	7	



Magnetic Quantum Number (m_l) & Spin Quantum Number (m_s): https://youtu.be/gbmGVUXBOBk

Summary

- There is a relationship between the motions of electrons in atoms and molecules and their energies that is described by quantum mechanics.

Because of wave–particle duality, scientists must deal with the probability of an electron being at a particular point in space. To do so required the development of **quantum mechanics**, which uses **wave functions (Ψ)** to describe the mathematical relationship between the motion of electrons in atoms and molecules and their energies. Wave functions have five important properties:

1. the wave function uses three variables (Cartesian axes *x*, *y*, and *z*) to describe the position of an electron;

2. the magnitude of the wave function is proportional to the intensity of the wave;
3. the probability of finding an electron at a given point is proportional to the square of the wave function at that point, leading to a distribution of probabilities in space that is often portrayed as an **electron density** plot;
4. describing electron distributions as standing waves leads naturally to the existence of sets of **quantum numbers** characteristic of each wave function; and
5. each spatial distribution of the electron described by a wave function with a given set of quantum numbers has a particular energy.

Quantum numbers provide important information about the energy and spatial distribution of an electron. The **principal quantum number** n can be any positive integer; as n increases for an atom, the average distance of the electron from the nucleus also increases. All wave functions with the same value of n constitute a **principal shell** in which the electrons have similar average distances from the nucleus. The **azimuthal quantum number** l can have integral values between 0 and $n - 1$; it describes the shape of the electron distribution. Wave functions that have the same values of both n and l constitute a **subshell**, corresponding to electron distributions that usually differ in orientation rather than in shape or average distance from the nucleus. The **magnetic quantum number** m_l can have $2l + 1$ integral values, ranging from $-l$ to $+l$, and describes the orientation of the electron distribution. Each wave function with a given set of values of n, l, and m_l describes a particular spatial distribution of an electron in an atom, an **atomic orbital**.

3.7: Wave Mechanics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.6: Wave Mechanics** is licensed CC BY-NC-SA 4.0.

3.8: Electron Spin- A Fourth Quantum Number

Learning Objectives

- To introduce the concept of electron spin and demonstrated it is quantized

The process of describing each atom's electronic structure consists, essentially, of beginning with hydrogen and adding one proton and one electron at a time to create the next heavier element in the table. All stable nuclei other than hydrogen also contain one or more neutrons. Because neutrons have no electrical charge, however, they can be ignored in the following discussion. Before demonstrating how to do this, however, we must introduce the concept of electron spin.

The Stern-Gerlach Experiment

The quantum numbers n, l, m are not sufficient to fully characterize the physical state of the electrons in an atom. In 1926, Otto Stern and Walther Gerlach carried out an experiment that could not be explained in terms of the three quantum numbers n, l, m and showed that there is, in fact, another quantum-mechanical degree of freedom that needs to be included in the theory. The experiment is illustrated in the figure 8.9.1. A beam of atoms (e.g. hydrogen or silver atoms) is sent through a spatially inhomogeneous magnetic field with a definite field gradient toward one of the poles. It is observed that the beam splits into two beams as it passes through the field region.

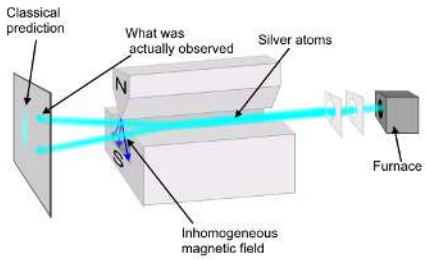


Figure 3.8.1: The Stern-Gerlach apparatus. from Wikipedia

The fact that the beam splits into 2 beams suggests that the electrons in the atoms have a degree of freedom capable of coupling to the magnetic field. That is, an electron has an *intrinsic magnetic moment* M arising from a degree of freedom that has no classical analog. The magnetic moment must take on only 2 values according to the Stern-Gerlach experiment. The intrinsic property that gives rise to the magnetic moment must have some analog to a *spin*, S; unlike position and momentum, which have clear classical analogs, spin does not.

The implication of the Stern-Gerlach experiment is that we need to include a fourth quantum number, m_s in our description of the physical state of the electron. That is, in addition to give its principle, angular, and magnetic quantum numbers, we also need to say if it is a spin-up electron or a spin-down electron.

Electron Spin

When scientists analyzed the emission and absorption spectra of the elements more closely, they saw that for elements having more than one electron, nearly all the lines in the spectra were actually *pairs* of very closely spaced lines. Because each line represents an energy level available to electrons in the atom, there are twice as many energy levels available as would be predicted solely based on the quantum numbers *n*, *l*, and m_l. Scientists also discovered that applying a magnetic field caused the lines in the pairs to split farther apart. In 1925, two graduate students in physics in the Netherlands, George Uhlenbeck (1900–1988) and Samuel Goudsmit (1902–1978), proposed that the splittings were caused by an electron spinning about its axis, much as Earth spins about its axis. When an electrically charged object spins, it produces a magnetic moment parallel to the axis of rotation, making it behave like a magnet. Although the electron cannot be viewed solely as a particle, spinning or otherwise, it is indisputable that it does have a magnetic moment. This magnetic moment is called electron spin.

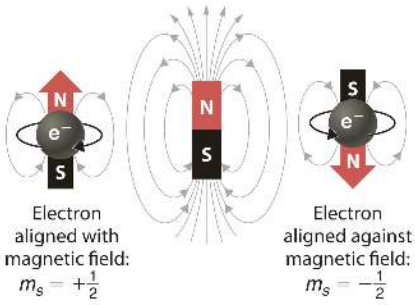


Figure 3.8.2: Electron Spin. In a magnetic field, an electron has two possible orientations with different energies, one with spin up, aligned with the magnetic field, and one with spin down, aligned against it. All other orientations are forbidden.

In an external magnetic field, the electron has two possible orientations (Figure 3.8.2). These are described by a fourth quantum number (m_s), which for any electron can have only two possible values, designated +½ (up) and −½ (down) to indicate that the two orientations are opposites; the subscript *s* is for spin. An electron behaves like a magnet that has one of two possible orientations, aligned either with the magnetic field or against it.

The implications of electron spin for chemistry were recognized almost immediately by an Austrian physicist, Wolfgang Pauli (1900–1958; Nobel Prize in Physics, 1945), who determined that each orbital can contain no more than two electrons. He developed the Pauli exclusion principle: *No two electrons in an atom can have the same values of all four quantum numbers (n, l, m_l, m_s).*

By giving the values of n, l, and m_l, we also specify a particular orbital (e.g., 1*s* with $n = 1$, $l = 0$, $m_l = 0$). Because m_s has only two possible values (+½ or −½), two electrons, *and only two electrons*, can occupy any given orbital, one with spin up and one with spin down. With this information, we can proceed to construct the entire periodic table, which was originally based on the physical and chemical properties of the known elements.

Example 3.8.2

List all the allowed combinations of the four quantum numbers (n, l, m_l, m_s) for electrons in a 2*p* orbital and predict the maximum number of electrons the 2*p* subshell can accommodate.

Given: orbital

Asked for: allowed quantum numbers and maximum number of electrons in orbital

Strategy:

A. List the quantum numbers (n, l, m_l) that correspond to an $n = 2p$ orbital. List all allowed combinations of (n, l, m_l).
B. Build on these combinations to list all the allowed combinations of (n, l, m_l, m_s).
C. Add together the number of combinations to predict the maximum number of electrons the 2*p* subshell can accommodate.

Solution:

A For a 2*p* orbital, we know that $n = 2$, $l = n - 1 = 1$, and $m_l = -l, (-l + 1), \ldots, (l - 1), l$. There are only three possible combinations of (n, l, m_l): (2, 1, 1), (2, 1, 0), and (2, 1, −1).

B Because m_s is independent of the other quantum numbers and can have values of only +½ and −½, there are six possible combinations of (n, l, m_l, m_s): (2, 1, 1, +½), (2, 1, 1, −½), (2, 1, 0, +½), (2, 1, 0, −½), (2, 1, −1, +½), and (2, 1, −1, −½).

C Hence the 2*p* subshell, which consists of three 2*p* orbitals ($2p_x$, $2p_y$, and $2p_z$), can contain a total of six electrons, two in each orbital.

Exercise 3.8.2

List all the allowed combinations of the four quantum numbers (n, l, m_l, m_s) for a 6*s* orbital, and predict the total number of electrons it can contain.

Answer: (6, 0, 0, +½), (6, 0, 0, −½); two electrons

Summary

In addition to the three quantum numbers (n, l, m_l) dictated by quantum mechanics, a fourth quantum number is required to explain certain properties of atoms. This is the **electron spin** quantum number (m_s), which can have values of +½ or −½ for any electron, corresponding to the two possible orientations of an electron in a magnetic field. The concept of electron spin has important consequences for chemistry because the **Pauli exclusion principle** implies that no orbital can contain more than two electrons (with opposite spin).

3.8: Electron Spin- A Fourth Quantum Number is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.9: Electron Spin: A Fourth Quantum Number** is licensed CC BY-NC-SA 4.0.

3.9: The Shape of Atomic Orbitals

Learning Objectives

- To understand the 3D representation of electronic orbitals

An orbital is the quantum mechanical refinement of Bohr's orbit. In contrast to his concept of a simple circular orbit with a fixed radius, orbitals are mathematically derived regions of space with different **probabilities** of containing an electron.

One way of representing electron probability distributions was illustrated previously for the 1**s** orbital of hydrogen. Because Ψ^2 gives the probability of finding an electron in a given volume of space (such as a cubic picometer), a plot of Ψ^2 versus distance from the nucleus (**r**) is a plot of the **probability density**. The 1**s** orbital is spherically symmetrical, so the probability of finding a 1**s** electron at any given point depends **only** on its distance from the nucleus. The probability density is greatest at $r = 0$ (at the nucleus) and decreases steadily with increasing distance. At very large values of **r**, the electron probability density is very small but **not** zero.

In contrast, we can calculate the **radial probability** (the probability of finding a 1s electron at a distance r from the nucleus) by adding together the probabilities of an electron being at all points on a series of **x** spherical shells of radius $\mathbf{r}_1$, $\mathbf{r}_2$, $\mathbf{r}_3$,…, $\mathbf{r}_{x-1}$, $\mathbf{r}_x$. In effect, we are dividing the atom into very thin concentric shells, much like the layers of an onion (Figure $3.9.1a$), and calculating the probability of finding an electron on each spherical shell. Recall that the electron probability density is greatest at r = 0 (Figure $3.9.1b$), so the density of dots is greatest for the smallest spherical shells in part (a) in Figure $3.9.1$. In contrast, the surface area of each spherical shell is equal to $4\pi r^2$, which increases very rapidly with increasing r (Figure $3.9.1c$). Because the surface area of the spherical shells increases more rapidly with increasing r than the electron probability density decreases, the plot of radial probability has a maximum at a particular distance (Figure $3.9.1d$). Most important, when r is very small, the surface area of a spherical shell is so small that the **total** probability of finding an electron close to the nucleus is very low; at the nucleus, the electron probability vanishes (Figure $3.9.1d$).

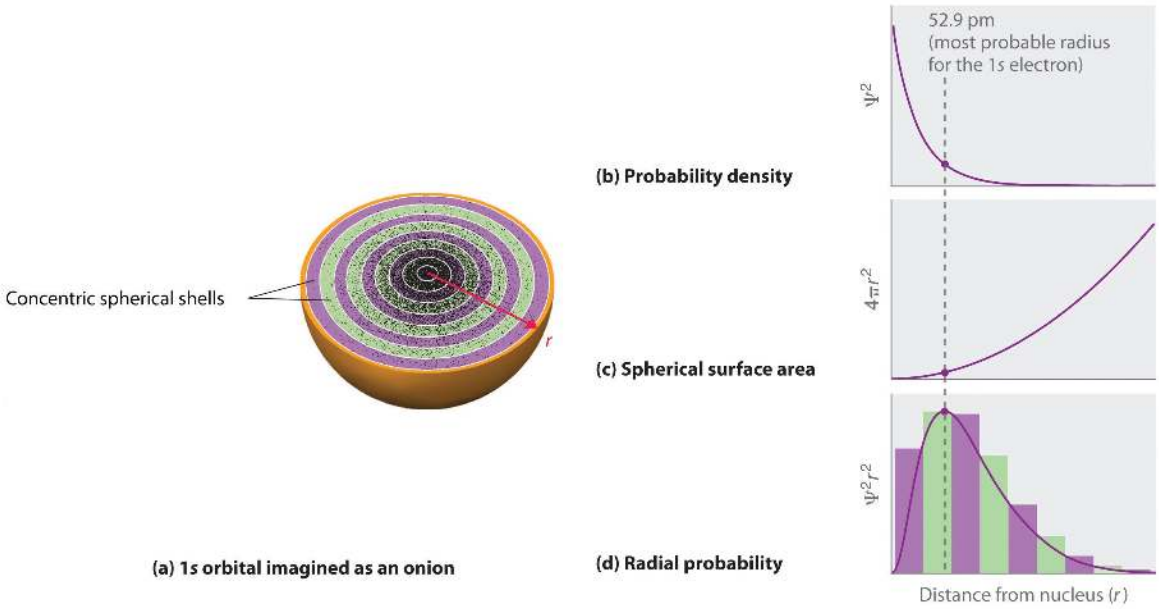


Figure 3.9.1: Most Probable Radius for the Electron in the Ground State of the Hydrogen Atom. (a) Imagine dividing the atom's total volume into very thin concentric shells as shown in the onion drawing. (b) A plot of electron probability density Ψ^2 versus r shows that the electron probability density is greatest at r = 0 and falls off smoothly with increasing r. The density of the dots is therefore greatest in the innermost shells of the onion. (c) The surface area of each shell, given by $4\pi\mathbf{r}^2$, increases rapidly with increasing r. (d) If we count the number of dots in each spherical shell, we obtain the total probability of finding the electron at a given value of r. Because the surface area of each shell increases more rapidly with increasing r than the electron probability density decreases, a plot of electron probability versus r (the **radial probability**) shows a peak. This peak corresponds to the most probable radius for the electron, 52.9 pm, which is exactly the radius predicted by Bohr's model of the hydrogen atom.

For the hydrogen atom, the peak in the radial probability plot occurs at r = 0.529 Å (52.9 pm), which is exactly the radius calculated by Bohr for the **n** = 1 orbit. Thus the **most probable radius** obtained from quantum mechanics is identical to the radius

calculated by classical mechanics. In Bohr's model, however, the electron was assumed to be at this distance 100% of the time, whereas in the Schrödinger model, it is at this distance only some of the time. The difference between the two models is attributable to the wavelike behavior of the electron and the Heisenberg uncertainty principle.

Figure 3.9.2 compares the electron probability densities for the hydrogen 1**s**, 2**s**, and 3**s** orbitals. Note that all three are spherically symmetrical. For the 2**s** and 3**s** orbitals, however (and for all other **s** orbitals as well), the electron probability density does not fall off smoothly with increasing r. Instead, a series of minima and maxima are observed in the radial probability plots (Figure 3.9.2*c*). The minima correspond to spherical **nodes** (regions of zero electron probability), which alternate with spherical regions of nonzero electron probability. The existence of these nodes is a consequence of changes of wave **phase** in the wavefunction Ψ.

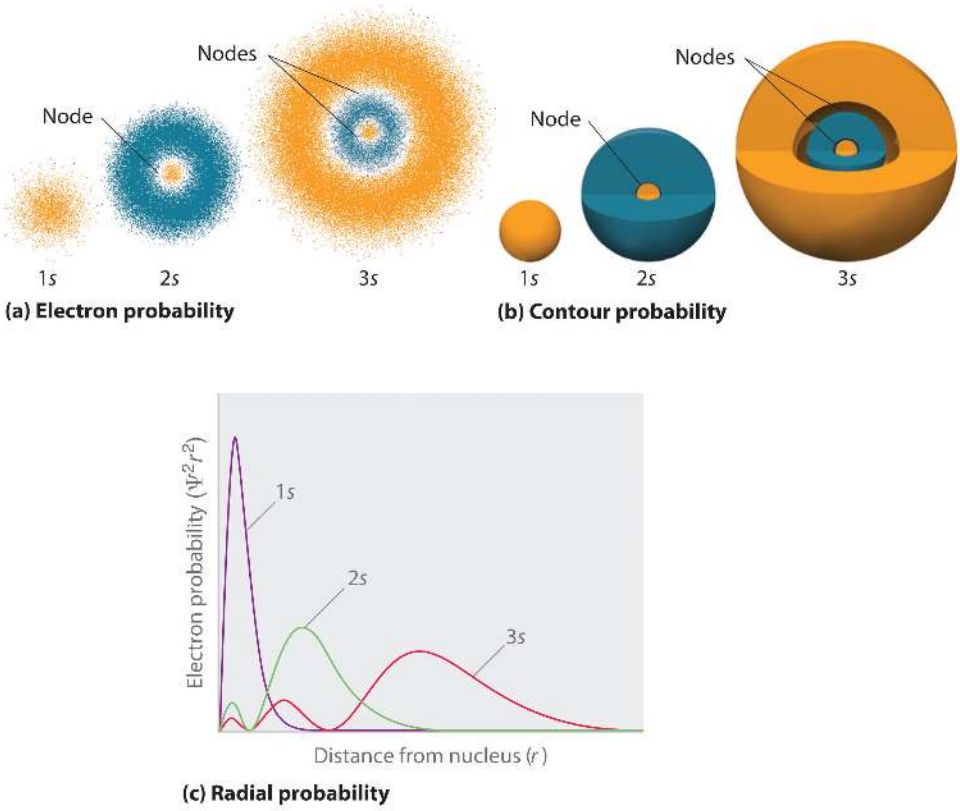


Figure 3.9.2: Probability Densities for the 1**s**, 2**s**, and 3**s** Orbitals of the Hydrogen Atom. (a) The electron probability density in any plane that contains the nucleus is shown. Note the presence of circular regions, or nodes, where the probability density is zero. (b) Contour surfaces enclose 90% of the electron probability, which illustrates the different sizes of the 1**s**, 2**s**, and 3**s** orbitals. The cutaway drawings give partial views of the internal spherical nodes. The orange color corresponds to regions of space where the phase of the wave function is positive, and the blue color corresponds to regions of space where the phase of the wave function is negative. (c) In these plots of electron probability as a function of distance from the nucleus (**r**) in all directions (radial probability), the most probable radius increases as **n** increases, but the 2**s** and 3**s** orbitals have regions of significant electron probability at small values of r.

s Orbitals (l=0)

Three things happen to **s** orbitals as **n** increases (Figure 3.9.2):

1. They become larger, extending farther from the nucleus.
2. They contain more nodes. This is similar to a standing wave that has regions of significant amplitude separated by nodes, points with zero amplitude.
3. For a given atom, the **s** orbitals also become higher in energy as **n** increases because of their increased distance from the nucleus.

Orbitals are generally drawn as three-dimensional surfaces that enclose 90% of the electron density, as was shown for the hydrogen 1**s**, 2**s**, and 3**s** orbitals in part (b) in Figure 3.9.2. Although such drawings show the relative sizes of the orbitals, they do not normally show the spherical nodes in the 2**s** and 3**s** orbitals because the spherical nodes lie inside the 90% surface. Fortunately, the positions of the spherical nodes are not important for chemical bonding.

p Orbitals (l=1)

Only **s** orbitals are spherically symmetrical. As the value of **l** increases, the number of orbitals in a given subshell increases, and the shapes of the orbitals become more complex. Because the 2**p** subshell has **l** = 1, with three values of $\mathbf{m_l}$ (−1, 0, and +1), there are three 2**p** orbitals.

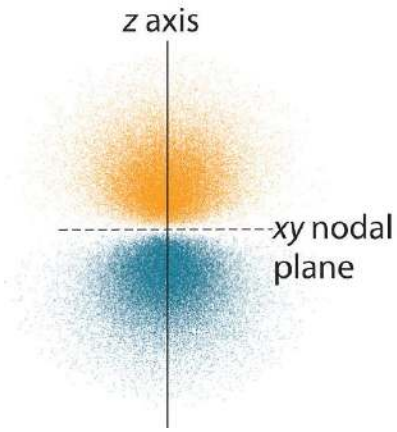


Figure 3.9.3: Electron Probability Distribution for a Hydrogen 2**p** Orbital. The nodal plane of zero electron density separates the two lobes of the 2**p** orbital. As in Figure 3.9.2, the colors correspond to regions of space where the phase of the wave function is positive (orange) and negative (blue).

The electron probability distribution for one of the hydrogen 2**p** orbitals is shown in Figure 3.9.3. Because this orbital has two lobes of electron density arranged along the **z** axis, with an electron density of zero in the **xy** plane (i.e., the **xy** plane is a nodal plane), it is a $2p_z$ orbital. As shown in Figure 3.9.4, the other two 2**p** orbitals have identical shapes, but they lie along the **x** axis ($2p_x$) and y axis ($2p_y$), respectively. Note that each **p** orbital has just one **nodal plane**. In each case, the phase of the wave function for each of the 2**p** orbitals is positive for the lobe that points along the positive axis and negative for the lobe that points along the negative axis. It is important to emphasize that these signs correspond to the **phase** of the wave that describes the electron motion, **not** to positive or negative charges.

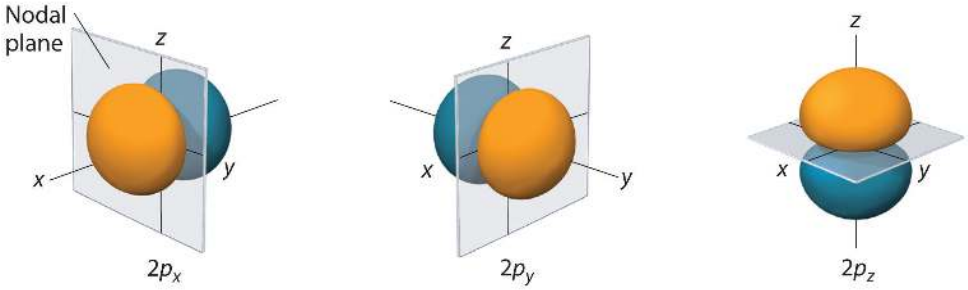


Figure 3.9.4 The Three Equivalent 2**p** Orbitals of the Hydrogen Atom

The surfaces shown enclose 90% of the total electron probability for the $2\mathbf{p_x}$, $2\mathbf{p_y}$, and $2\mathbf{p_z}$ orbitals. Each orbital is oriented along the axis indicated by the subscript and a nodal plane that is perpendicular to that axis bisects each 2**p** orbital. The phase of the wave function is positive (orange) in the region of space where **x**, **y**, or **z** is positive and negative (blue) where **x**, **y**, or **z** is negative. Just as with the **s** orbitals, the size and complexity of the **p** orbitals for any atom increase as the principal quantum number **n** increases. The shapes of the 90% probability surfaces of the 3**p**, 4**p**, and higher-energy **p** orbitals are, however, essentially the same as those shown in Figure 3.9.4.

d Orbitals (l=2)

Subshells with **l** = 2 have five **d** orbitals; the first principal shell to have a **d** subshell corresponds to **n** = 3. The five **d** orbitals have $\mathbf{m_l}$ values of −2, −1, 0, +1, and +2.

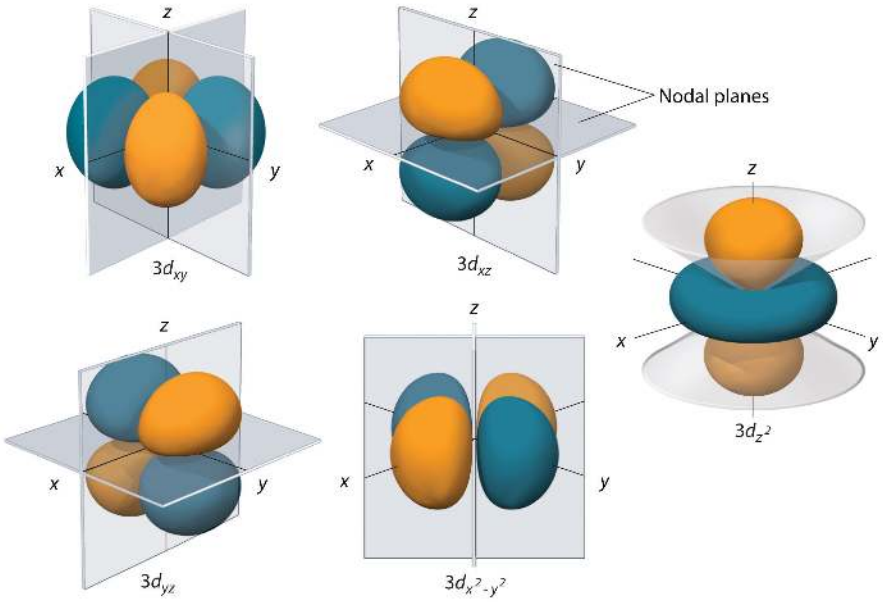


Figure 3.9.5: The Five Equivalent 3d Orbitals of the Hydrogen Atom. The surfaces shown enclose 90% of the total electron probability for the five hydrogen 3d orbitals. Four of the five 3d orbitals consist of four lobes arranged in a plane that is intersected by two perpendicular nodal planes. These four orbitals have the same shape but different orientations. The fifth 3d orbital, $3d_{z^2}$, has a distinct shape even though it is mathematically equivalent to the others. The phase of the wave function for the different lobes is indicated by color: orange for positive and blue for negative.

The 3 d orbitals can be broken down into 3 d x y, 3 d x z, 3 d y squared, 3 d x squared minus y squared, 3 d z squared.

The hydrogen 3d orbitals, shown in Figure 3.9.5, have more complex shapes than the 2p orbitals. All five 3d orbitals contain two nodal surfaces, as compared to one for each p orbital and zero for each s orbital. In three of the d orbitals, the lobes of electron density are oriented between the x and y, x and z, and y and z planes; these orbitals are referred to as the $3d_{xy}$, \)3d_{xz}\), and $3d_{yz}$ orbitals, respectively. A fourth d orbital has lobes lying along the x and y axes; this is the $3d_{x^2-y^2}$ orbital. The fifth 3d orbital, called the $3d_{z^2}$ orbital, has a unique shape: it looks like a $2p_z$ orbital combined with an additional doughnut of electron probability lying in the xy plane. Despite its peculiar shape, the $3d_{z^2}$ orbital is mathematically equivalent to the other four and has the same energy. In contrast to p orbitals, the phase of the wave function for d orbitals is the same for opposite pairs of lobes. As shown in Figure 3.9.5, the phase of the wave function is positive for the two lobes of the dz^2 orbital that lie along the z axis, whereas the phase of the wave function is negative for the doughnut of electron density in the xy plane. Like the s and p orbitals, as n increases, the size of the d orbitals increases, but the overall shapes remain similar to those depicted in Figure 3.9.5.

f Orbitals (l=3)

Principal shells with **n** = 4 can have subshells with **l** = 3 and $\mathbf{m_l}$ values of −3, −2, −1, 0, +1, +2, and +3. These subshells consist of seven **f** orbitals. Each **f** orbital has three nodal surfaces, so their shapes are complex. Because **f** orbitals are not particularly important for our purposes, we do not discuss them further, and orbitals with higher values of **l** are not discussed at all.

Orbital Energies

Although we have discussed the shapes of orbitals, we have said little about their comparative energies. We begin our discussion of orbital energies by considering atoms or ions with only a single electron (such as H or He^+).

The relative energies of the atomic orbitals with **n** ≤ 4 for a hydrogen atom are plotted in Figure 3.9.6; note that the orbital energies depend on **only** the principal quantum number **n**. Consequently, the energies of the 2**s** and 2**p** orbitals of hydrogen are the same; the energies of the 3**s**, 3**p**, and 3**d** orbitals are the same; and so forth. Quantum mechanics predicts that in the hydrogen atom, all orbitals with the same value of **n** (e.g., the three 2**p** orbitals) are **degenerate**, meaning that they have the same energy. The orbital energies obtained for hydrogen using quantum mechanics are exactly the same as the allowed energies calculated by Bohr. In contrast to Bohr's model, however, which allowed only one orbit for each energy level, quantum mechanics predicts that there are **4** orbitals with different electron density distributions in the **n** = 2 principal shell (one 2**s** and three 2**p** orbitals), **9** in the **n** = 3 principal shell, and **16** in the **n** = 4 principal shell.The different values of **l** and $\mathbf{m_l}$ for the individual orbitals within a given principal shell are not important for understanding the emission or absorption spectra of the hydrogen atom under most conditions,

but they do explain the splittings of the main lines that are observed when hydrogen atoms are placed in a magnetic field. Figure 3.9.6 shows that the energy levels become closer and closer together as the value of **n** increases, as expected because of the $1/\mathbf{n}^2$ dependence of orbital energies.

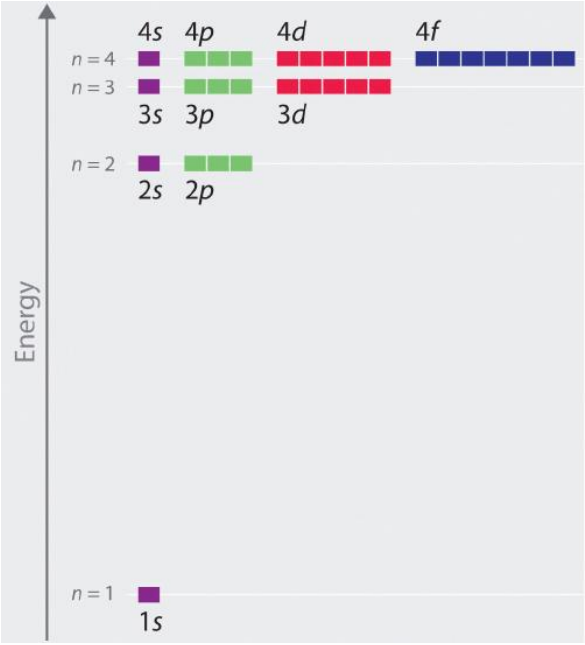


Figure 3.9.6: Orbital Energy Level Diagram for the Hydrogen Atom with a single electron. Each box corresponds to one orbital. Note that the difference in energy between orbitals decreases rapidly with increasing values of **n**.

The energies of the orbitals in any species with only one electron can be calculated by a minor variation of Bohr's equation, which can be extended to other single-electron species by incorporating the nuclear charge Z (the number of protons in the nucleus):

$$E = -\frac{Z^2}{n^2}Rhc \tag{3.9.1}$$

In general, both energy and radius decrease as the nuclear charge increases. Thus the most stable orbitals (those with the lowest energy) are those closest to the nucleus. For example, in the ground state of the hydrogen atom, the single electron is in the 1s orbital, whereas in the first excited state, the atom has absorbed energy and the electron has been promoted to one of the n = 2 orbitals. In ions with only a single electron, the energy of a given orbital depends on only n, and all subshells within a principal shell, such as the p_x, p_y, and p_z orbitals, are degenerate.

Summary

The four chemically important types of atomic orbital correspond to values of $\ell = 0$, 1, 2, and 3. Orbitals with $\ell = 0$ are s orbitals and are spherically symmetrical, with the greatest probability of finding the electron occurring at the nucleus. All orbitals with values of $n > 1$ and $ell = 0$ contain one or more nodes. Orbitals with $\ell = 1$ are p orbitals and contain a nodal plane that includes the nucleus, giving rise to a dumbbell shape. Orbitals with $\ell = 2$ are d orbitals and have more complex shapes with at least two nodal surfaces. Orbitals with $\ell = 3$ are f orbitals, which are still more complex.

Because its average distance from the nucleus determines the energy of an electron, each atomic orbital with a given set of quantum numbers has a particular energy associated with it, the **orbital energy**.

$$E = -\frac{Z^2}{n^2}Rhc$$

In atoms or ions with only a single electron, all orbitals with the same value of n have the same energy (they are degenerate), and the energies of the principal shells increase smoothly as n increases. An atom or ion with the electron(s) in the lowest-energy orbital(s) is said to be in its ground state, whereas an atom or ion in which one or more electrons occupy higher-energy orbitals is said to be in an excited state.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)

3.9: The Shape of Atomic Orbitals is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **6.6: 3D Representation of Orbitals** is licensed CC BY-NC-SA 3.0.

3.10: Multielectron Atoms

Learning Objectives

- To understand the basics of adding electrons to atomic orbitals
- To understand the basics of the Aufbau principle

The electron configuration of an element is the arrangement of its electrons in its atomic orbitals. By knowing the electron configuration of an element, we can predict and explain a great deal of its chemistry.

The Aufbau Principle

We construct the periodic table by following the aufbau principle (from German, meaning "building up"). First we determine the number of electrons in the atom; then we add electrons one at a time to the lowest-energy orbital available *without violating the Pauli principle*. We use the orbital energy diagram of Figure 6.29, recognizing that each orbital can hold two electrons, one with spin up ↑, corresponding to $m_s = +½$, which is arbitrarily written first, and one with spin down ↓, corresponding to $m_s = −½$. A filled orbital is indicated by ↑↓, in which the electron spins are said to be *paired*. Here is a schematic orbital diagram for a hydrogen atom in its ground state:

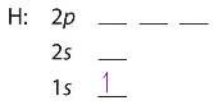
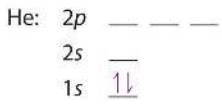


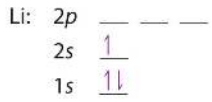
Figure 6.29"), and the electron configuration is written as $1s^1$ and read as "one-s-one."

A neutral helium atom, with an atomic number of 2 ($Z = 2$), has two electrons. We place one electron in the orbital that is lowest in energy, the 1*s* orbital. From the Pauli exclusion principle, we know that an orbital can contain two electrons with opposite spin, so we place the second electron in the same orbital as the first but pointing down, so that the electrons are paired. The orbital diagram for the helium atom is therefore



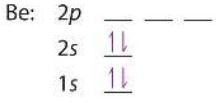
written as $1s^2$, where the superscript 2 implies the pairing of spins. Otherwise, our configuration would violate the Pauli principle.

The next element is lithium, with $Z = 3$ and three electrons in the neutral atom. We know that the 1*s* orbital can hold two of the electrons with their spins paired. Figure 6.29 tells us that the next lowest energy orbital is 2*s*, so the orbital diagram for lithium is

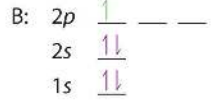


This electron configuration is written as $1s^22s^1$.

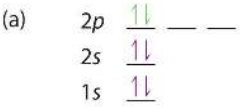
The next element is beryllium, with $Z = 4$ and four electrons. We fill both the 1*s* and 2*s* orbitals to achieve a $1s^22s^2$ electron configuration:

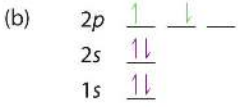


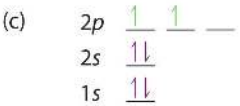
When we reach boron, with $Z = 5$ and five electrons, we must place the fifth electron in one of the 2*p* orbitals. Because all three 2*p* orbitals are degenerate, it doesn't matter which one we select. The electron configuration of boron is $1s^22s^22p^1$:



At carbon, with $Z = 6$ and six electrons, we are faced with a choice. Should the sixth electron be placed in the same $2p$ orbital that already has an electron, or should it go in one of the empty $2p$ orbitals? If it goes in an empty $2p$ orbital, will the sixth electron have its spin aligned with or be opposite to the spin of the fifth? In short, which of the following three orbital diagrams is correct for carbon, remembering that the $2p$ orbitals are degenerate?

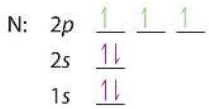
(a)


(b)


(c)


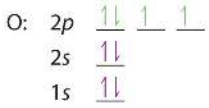
Because of electron-electron repulsions, it is more favorable energetically for an electron to be in an unoccupied orbital than in one that is already occupied; hence we can eliminate choice a. Similarly, experiments have shown that choice b is slightly higher in energy (less stable) than choice c because electrons in degenerate orbitals prefer to line up with their spins parallel; thus, we can eliminate choice b. Choice c illustrates Hund's rule (named after the German physicist Friedrich H. Hund, 1896–1997), which today says that the lowest-energy electron configuration for an atom is the one that has the maximum number of electrons with parallel spins in degenerate orbitals. By Hund's rule, the electron configuration of carbon, which is $1s^2 2s^2 2p^2$, is understood to correspond to the orbital diagram shown in c. Experimentally, it is found that the ground state of a neutral carbon atom does indeed contain two unpaired electrons.

When we get to nitrogen ($Z = 7$, with seven electrons), Hund's rule tells us that the lowest-energy arrangement is

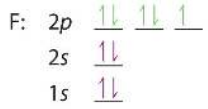


with three unpaired electrons. The electron configuration of nitrogen is thus $1s^2 2s^2 2p^3$.

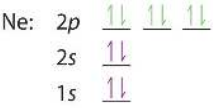
At oxygen, with $Z = 8$ and eight electrons, we have no choice. One electron must be paired with another in one of the $2p$ orbitals, which gives us two unpaired electrons and a $1s^2 2s^2 2p^4$ electron configuration. Because all the $2p$ orbitals are degenerate, it doesn't matter which one has the pair of electrons.



Similarly, fluorine has the electron configuration $1s^2 2s^2 2p^5$:



When we reach neon, with $Z = 10$, we have filled the $2p$ subshell, giving a $1s^2 2s^2 2p^6$ electron configuration:



Notice that for neon, as for helium, all the orbitals through the $2p$ level are completely filled. This fact is very important in dictating both the chemical reactivity and the bonding of helium and neon, as you will see.



Electron Configuration of Atoms: https://youtu.be/LlY-O3-bfnk

Valence Electrons

As we continue through the periodic table in this way, writing the electron configurations of larger and larger atoms, it becomes tedious to keep copying the configurations of the filled inner subshells. In practice, chemists simplify the notation by using a bracketed noble gas symbol to represent the configuration of the noble gas from the preceding row because all the orbitals in a noble gas are filled. For example, [Ne] represents the $1s^2 2s^2 2p^6$ electron configuration of neon (Z = 10), so the electron configuration of sodium, with Z = 11, which is $1s^2 2s^2 2p^6 3s^1$, is written as [Ne]$3s^1$:

Neon	Z = 10	$1s^2 2s^2 2p^6$
Sodium	Z = 11	$1s^2 2s^2 2p^6 \mathbf{3}s^1$ = [Ne]$\mathbf{3}s^1$

Because electrons in filled inner orbitals are closer to the nucleus and more tightly bound to it, they are rarely involved in chemical reactions. This means that the chemistry of an atom depends mostly on the electrons in its outermost shell, which are called the valence electrons. The simplified notation allows us to see the valence-electron configuration more easily. Using this notation to compare the electron configurations of sodium and lithium, we have:

Sodium	$1s^2 2s^2 2p^6 3s^1$ = [Ne]$3s^1$
Lithium	$1s^2 2s^1$ = [He]$2s^1$

It is readily apparent that both sodium and lithium have one *s* electron in their valence shell. We would therefore predict that sodium and lithium have very similar chemistry, which is indeed the case.

As we continue to build the eight elements of period 3, the 3*s* and 3*p* orbitals are filled, one electron at a time. This row concludes with the noble gas argon, which has the electron configuration [Ne]$3s^2 3p^6$, corresponding to a filled valence shell.

✔ Example 3.10.2

Draw an orbital diagram and use it to derive the electron configuration of phosphorus, Z = 15. What is its valence electron configuration?

Given: atomic number

Asked for: orbital diagram and valence electron configuration for phosphorus

Strategy:

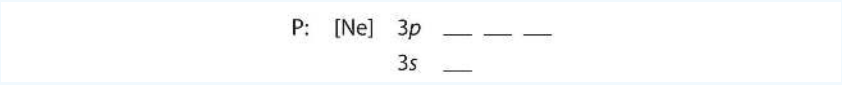
A. Locate the nearest noble gas preceding phosphorus in the periodic table. Then subtract its number of electrons from those in phosphorus to obtain the number of valence electrons in phosphorus.

B. Referring to Figure 6.29, draw an orbital diagram to represent those valence orbitals. Following Hund's rule, place the valence electrons in the available orbitals, beginning with the orbital that is lowest in energy. Write the electron configuration from your orbital diagram.
C. Ignore the inner orbitals (those that correspond to the electron configuration of the nearest noble gas) and write the valence electron configuration for phosphorus.

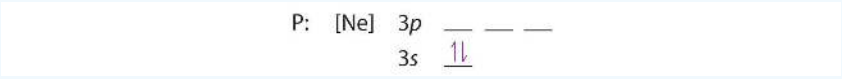
Solution:

A Because phosphorus is in the third row of the periodic table, we know that it has a [Ne] closed shell with 10 electrons. We begin by subtracting 10 electrons from the 15 in phosphorus.

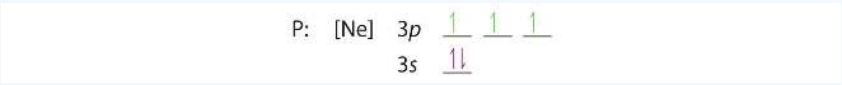
B The additional five electrons are placed in the next available orbitals, which Figure 6.29 tells us are the 3*s* and 3*p* orbitals:



Because the 3*s* orbital is lower in energy than the 3*p* orbitals, we fill it first:



Hund's rule tells us that the remaining three electrons will occupy the degenerate 3*p* orbitals separately but with their spins aligned:



The electron configuration is $[Ne]3s^2 3p^3$.

C We obtain the valence electron configuration by ignoring the inner orbitals, which for phosphorus means that we ignore the [Ne] closed shell. This gives a valence-electron configuration of $3s^2 3p^3$.

? Exercise 3.10.2

Draw an orbital diagram and use it to derive the electron configuration of chlorine, $Z = 17$. What is its valence electron configuration?

Answer: $[Ne]3s^2 3p^5$; $3s^2 3p^5$



Definition of Valence Electrons: https://youtu.be/_ldxOYwM2VM

The general order in which orbitals are filled is depicted in Figure 3.10.1. Subshells corresponding to each value of *n* are written from left to right on successive horizontal lines, where each row represents a row in the periodic table. The order in which the orbitals are filled is indicated by the diagonal lines running from the upper right to the lower left. Accordingly, the 4*s* orbital is filled prior to the 3*d* orbital because of shielding and penetration effects. Consequently, the electron configuration of potassium,

which begins the fourth period, is [Ar]$4s^1$, and the configuration of calcium is [Ar]$4s^2$. Five 3*d* orbitals are filled by the next 10 elements, the transition metals, followed by three 4*p* orbitals. Notice that the last member of this row is the noble gas krypton (Z = 36), [Ar]$4s^2 3d^{10} 4p^6$ = [Kr], which has filled 4*s*, 3*d*, and 4*p* orbitals. The fifth row of the periodic table is essentially the same as the fourth, except that the 5*s*, 4*d*, and 5*p* orbitals are filled sequentially.

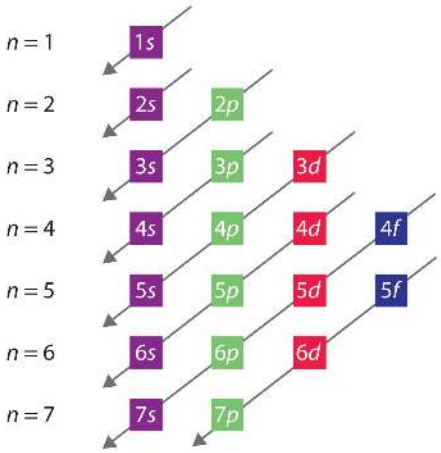


Figure 3.10.1 Predicting the Order in Which Orbitals Are Filled in Multielectron Atoms. If you write the subshells for each value of the principal quantum number on successive lines, the observed order in which they are filled is indicated by a series of diagonal lines running from the upper right to the lower left.

The sixth row of the periodic table will be different from the preceding two because the 4*f* orbitals, which can hold 14 electrons, are filled between the 6*s* and the 5*d* orbitals. The elements that contain 4*f* orbitals in their valence shell are the lanthanides. When the 6*p* orbitals are finally filled, we have reached the next (and last known) noble gas, radon (Z = 86), [Xe]$6s^2 4f^{14} 5d^{10} 6p^6$ = [Rn]. In the last row, the 5*f* orbitals are filled between the 7*s* and the 6*d* orbitals, which gives the 14 actinide elements. Because the large number of protons makes their nuclei unstable, all the actinides are radioactive.

Example 3.10.3

Write the electron configuration of mercury (Z = 80), showing all the inner orbitals.

Given: atomic number

Asked for: complete electron configuration

Strategy:

Using the orbital diagram in Figure 3.10.1 and the periodic table as a guide, fill the orbitals until all 80 electrons have been placed.

Solution:

By placing the electrons in orbitals following the order shown in Figure 3.10.1 and using the periodic table as a guide, we obtain

$1s^2$	row 1	2 electrons
$2s^2 2p^6$	row 2	8 electrons
$3s^2 3p^6$	row 3	8 electrons
$4s^2 3d^{10} 4p^6$	row 4	18 electrons
$5s^2 4d^{10} 5p^6$	row 5	18 electrons
	row 1–5	54 electrons

After filling the first five rows, we still have 80 − 54 = 26 more electrons to accommodate. According to Figure 3.10.2, we need to fill the 6*s* (2 electrons), 4*f* (14 electrons), and 5*d* (10 electrons) orbitals. The result is mercury's electron configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} = \text{Hg} = [\text{Xe}]6s^2 4f^{14} 5d^{10}$$

with a filled 5*d* subshell, a $6s^2 4f^{14} 5d^{10}$ valence shell configuration, and a total of 80 electrons. (You should always check to be sure that the total number of electrons equals the atomic number.)

? Exercise 3.10.3

Although element 114 is not stable enough to occur in nature, two isotopes of element 114 were created for the first time in a nuclear reactor in 1999 by a team of Russian and American scientists. Write the complete electron configuration for element 114.

Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10} 7p^2$

The electron configurations of the elements are presented in Figure 3.10.3, which lists the orbitals in the order in which they are filled. In several cases, the ground state electron configurations are different from those predicted by Figure 3.10.1. Some of these anomalies occur as the 3*d* orbitals are filled. For example, the observed ground state electron configuration of chromium is $[\text{Ar}]4s^1 3d^5$ rather than the predicted $[\text{Ar}]4s^2 3d^4$. Similarly, the observed electron configuration of copper is $[\text{Ar}]4s^1 3d^{10}$ instead of $[\text{Ar}]s^2 3d^9$. The actual electron configuration may be rationalized in terms of an added stability associated with a half-filled (ns^1, np^3, nd^5, nf^7) or filled (ns^2, np^6, nd^{10}, nf^{14}) subshell. Given the small differences between higher energy levels, this added stability is enough to shift an electron from one orbital to another. In heavier elements, other more complex effects can also be important, leading to some of the additional anomalies indicated in Figure 3.10.3. For example, cerium has an electron configuration of $[\text{Xe}]6s^2 4f^1 5d^1$, which is impossible to rationalize in simple terms. In most cases, however, these apparent anomalies do not have important chemical consequences.

Note

Additional stability is associated with half-filled or filled subshells.

Summary

Based on the Pauli principle and a knowledge of orbital energies obtained using hydrogen-like orbitals, it is possible to construct the periodic table by filling up the available orbitals beginning with the lowest-energy orbitals (the **aufbau principle**), which gives rise to a particular arrangement of electrons for each element (its **electron configuration**). **Hund's rule** says that the lowest-energy arrangement of electrons is the one that places them in degenerate orbitals with their spins parallel. For chemical purposes, the most important electrons are those in the outermost principal shell, the **valence electrons**.

3.10: Multielectron Atoms is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.10: Multielectron Atoms** is licensed CC BY-NC-SA 4.0.

3.11: Key Terms

actinide
inner transition metal in the bottom of the bottom two rows of the periodic table

alkali metal
element in group 1

alkaline earth metal
element in group 2

amplitude
extent of the displacement caused by a wave

atomic orbital
mathematical function that describes the behavior of an electron in an atom (also called the wavefunction)

Aufbau principle
procedure in which the electron configuration of the elements is determined by “building” them in order of atomic numbers, adding one proton to the nucleus and one electron to the proper subshell at a time

blackbody
idealized perfect absorber of all incident electromagnetic radiation; such bodies emit electromagnetic radiation in characteristic continuous spectra called blackbody radiation

Bohr’s model of the hydrogen atom
structural model in which an electron moves around the nucleus only in circular orbits, each with a specific allowed radius

chalcogen
element in group 16

continuous spectrum
electromagnetic radiation given off in an unbroken series of wavelengths (e.g., white light from the sun)

core electron
electron in an atom that occupies the orbitals of the inner shells

covalent bond
attractive force between the nuclei of a molecule’s atoms and pairs of electrons between the atoms

covalent compound
(also, molecular compound) composed of molecules formed by atoms of two or more different elements

covalent radius
one-half the distance between the nuclei of two identical atoms when they are joined by a covalent bond

***d* orbital**
region of space with high electron density that is either four lobed or contains a dumbbell and torus shape; describes orbitals with $l = 2$.

degenerate orbitals
orbitals that have the same energy

effective nuclear charge
charge that leads to the Coulomb force exerted by the nucleus on an electron, calculated as the nuclear charge minus shielding

electromagnetic radiation
energy transmitted by waves that have an electric-field component and a magnetic-field component

electromagnetic spectrum
range of energies that electromagnetic radiation can comprise, including radio, microwaves, infrared, visible, ultraviolet, X-rays, and gamma rays

electron affinity
energy change associated with addition of an electron to a gaseous atom or ion

electron configuration
listing that identifies the electron occupancy of an atom's shells and subshells

electron density
a measure of the probability of locating an electron in a particular region of space, it is equal to the squared absolute value of the wave function ψ

endothermic
processes that increase the energy of an atom and involve the absorption of light

excited state
state having an energy greater than the ground-state energy

exothermic
processes that decrease the energy of an atom and involve the emission of light

***f* orbital**
multilobed region of space with high electron density, describes orbitals with $l = 3$

frequency (ν)
number of wave cycles (peaks or troughs) that pass a specified point in space per unit time

ground state
state in which the electrons in an atom, ion, or molecule have the lowest energy possible

group
vertical column of the periodic table

halogen
element in group 17

Heisenberg uncertainty principle
rule stating that it is impossible to exactly determine both certain conjugate dynamical properties such as the momentum and the position of a particle at the same time. The uncertainty principle is a consequence of quantum particles exhibiting wave–particle duality

hertz (Hz)
the unit of frequency, which is the number of cycles per second, s^{-1}

Hund's rule
every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin

inert gas
(also, noble gas) element in group 18

inner transition metal
(also, lanthanide or actinide) element in the bottom two rows; if in the first row, also called lanthanide, or if in the second row, also called actinide

intensity
property of wave-propagated energy related to the amplitude of the wave, such as brightness of light or loudness of sound

interference pattern
pattern typically consisting of alternating bright and dark fringes; it results from constructive and destructive interference of waves

ionic bond
electrostatic forces of attraction between the oppositely charged ions of an ionic compound

ionic compound
compound composed of cations and anions combined in ratios, yielding an electrically neutral substance

ionization energy
energy required to remove an electron from a gaseous atom or ion

isoelectronic
group of ions or atoms that have identical electron configurations

lanthanide
inner transition metal in the top of the bottom two rows of the periodic table

line spectrum
electromagnetic radiation emitted at discrete wavelengths by a specific atom (or atoms) in an excited state

magnetic quantum number (m_l)
quantum number signifying the orientation of an atomic orbital around the nucleus

main-group element
(also, representative element) element in groups 1, 2, and 13–18

metal
element that is shiny, malleable, good conductor of heat and electricity

metalloid
element that conducts heat and electricity moderately well, and possesses some properties of metals and some properties of nonmetals

molecular compound
(also, covalent compound) composed of molecules formed by atoms of two or more different elements

monatomic ion
ion composed of a single atom

noble gas
(also, inert gas) element in group 18

node
any point of a standing wave with zero amplitude

nonmetal

element that appears dull, poor conductor of heat and electricity

orbital diagram
pictorial representation of the electron configuration showing each orbital as a box and each electron as an arrow

oxyanion
polyatomic anion composed of a central atom bonded to oxygen atoms

***p* orbital**
dumbbell-shaped region of space with high electron density, describes orbitals with $l = 1$

Pauli exclusion principle
specifies that no two electrons in an atom can have the same value for all four quantum numbers

period
(also, series) horizontal row of the periodic table

periodic law
properties of the elements are periodic function of their atomic numbers

periodic table
table of the elements that places elements with similar chemical properties close together

photon
smallest possible packet of electromagnetic radiation, a particle of light

pnictogen
element in group 15

polyatomic ion
ion composed of more than one atom

principal quantum number (*n*)
quantum number specifying the shell an electron occupies in an atom

quantization
limitation of some property to specific discrete values, not continuous

quantum mechanics
field of study that includes quantization of energy, wave-particle duality, and the Heisenberg uncertainty principle to describe matter

quantum number
number having only specific allowed values and used to characterize the arrangement of electrons in an atom

representative element
(also, main-group element) element in columns 1, 2, and 12–18

***s* orbital**
spherical region of space with high electron density, describes orbitals with $l = 0$

secondary (angular momentum) quantum number (*l*)
quantum number distinguishing the different shapes of orbitals; it is also a measure of the orbital angular momentum

series
(also, period) horizontal row of the period table

shell

atomic orbitals with the same principal quantum number, n

spin quantum number (m_s)

number specifying the electron spin direction, either $+\frac{1}{2}$ or $-\frac{1}{2}$

standing wave

(also, stationary wave) localized wave phenomenon characterized by discrete wavelengths determined by the boundary conditions used to generate the waves; standing waves are inherently quantized

subshell

atomic orbitals with the same values of n and l

transition metal

element in groups 3–12 (more strictly defined, 3–11; see chapter on transition metals and coordination chemistry)

valence electrons

electrons in the high energy outer shell(s) of an atom

valence shell

high energy outer shell(s) of an atom

wave

oscillation of a property over time or space; can transport energy from one point to another

wave-particle duality

observation that elementary particles can exhibit both wave-like and particle-like properties

wavefunction (ψ)

mathematical description of an atomic orbital that describes the shape of the orbital; it can be used to calculate the probability of finding the electron at any given location in the orbital, as well as dynamical variables such as the energy and the angular momentum

wavelength (λ)

distance between two consecutive peaks or troughs in a wave

This page titled 3.11: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

3.12: Key Equations

$c = \lambda\nu$
$E = h\nu = \frac{hc}{\lambda}$, where $h = 6.626 \times 10^{-34}$ J s
$\frac{1}{\lambda} = R_\infty \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$

$E_n = -\frac{kZ^2}{n^2}$, $n = 1,\ 2,\ 3,\ \ldots$
$\Delta E = kZ^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$
$r = (n^2/z)a_0 r$

This page titled 3.12: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

3.13: Summary

Light and other forms of electromagnetic radiation move through a vacuum with a constant speed, c, of 2.998×10^8 m s^{-1}. This radiation shows wavelike behavior, which can be characterized by a frequency, ν, and a wavelength, λ, such that $c = \lambda\nu$. Light is an example of a travelling wave. Other important wave phenomena include standing waves, periodic oscillations, and vibrations. Standing waves exhibit quantization, since their wavelengths are limited to discrete integer multiples of some characteristic lengths. Electromagnetic radiation that passes through two closely spaced narrow slits having dimensions roughly similar to the wavelength will show an interference pattern that is a result of constructive and destructive interference of the waves. Electromagnetic radiation also demonstrates properties of particles called photons. The energy of a photon is related to the frequency (or alternatively, the wavelength) of the radiation as $E = h\nu$ (or $E =$), where h is Planck's constant. That light demonstrates both wavelike and particle-like behavior is known as wave-particle duality. All forms of electromagnetic radiation share these properties, although various forms including X-rays, visible light, microwaves, and radio waves interact differently with matter and have very different practical applications. Electromagnetic radiation can be generated by exciting matter to higher energies, such as by heating it. The emitted light can be either continuous (incandescent sources like the sun) or discrete (from specific types of excited atoms). Continuous spectra often have distributions that can be approximated as blackbody radiation at some appropriate temperature. The line spectrum of hydrogen can be obtained by passing the light from an electrified tube of hydrogen gas through a prism. This line spectrum was simple enough that an empirical formula called the Rydberg formula could be derived from the spectrum. Three historically important paradoxes from the late 19th and early 20th centuries that could not be explained within the existing framework of classical mechanics and classical electromagnetism were the blackbody problem, the photoelectric effect, and the discrete spectra of atoms. The resolution of these paradoxes ultimately led to quantum theories that superseded the classical theories.

Bohr incorporated Planck's and Einstein's quantization ideas into a model of the hydrogen atom that resolved the paradox of atom stability and discrete spectra. The Bohr model of the hydrogen atom explains the connection between the quantization of photons and the quantized emission from atoms. Bohr described the hydrogen atom in terms of an electron moving in a circular orbit about a nucleus. He postulated that the electron was restricted to certain orbits characterized by discrete energies. Transitions between these allowed orbits result in the absorption or emission of photons. When an electron moves from a higher-energy orbit to a more stable one, energy is emitted in the form of a photon. To move an electron from a stable orbit to a more excited one, a photon of energy must be absorbed. Using the Bohr model, we can calculate the energy of an electron and the radius of its orbit in any one-electron system.

Macroscopic objects act as particles. Microscopic objects (such as electrons) have properties of both a particle and a wave. Their exact trajectories cannot be determined. The quantum mechanical model of atoms describes the three-dimensional position of the electron in a *probabilistic* manner according to a mathematical function called a wavefunction, often denoted as ψ. Atomic wavefunctions are also called orbitals. The squared magnitude of the wavefunction describes the distribution of the probability of finding the electron in a particular region in space. Therefore, atomic orbitals describe the areas in an atom where electrons are most likely to be found.

An atomic orbital is characterized by three quantum numbers. The principal quantum number, n, can be any positive integer. The general region for value of energy of the orbital and the average distance of an electron from the nucleus are related to n. Orbitals having the same value of n are said to be in the same shell. The secondary (angular momentum) quantum number, l, can have any integer value from 0 to $n - 1$. This quantum number describes the shape or type of the orbital. Orbitals with the same principal quantum number and the same l value belong to the same subshell. The magnetic quantum number, m_l, with $2l + 1$ values ranging from $-l$ to $+l$, describes the orientation of the orbital in space. In addition, each electron has a spin quantum number, m_s, that can be equal to $\pm \frac{1}{2}$. No two electrons in the same atom can have the same set of values for all the four quantum numbers.

The relative energy of the subshells determine the order in which atomic orbitals are filled (1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, and so on). Electron configurations and orbital diagrams can be determined by applying the Pauli exclusion principle (no two electrons can have the same set of four quantum numbers) and Hund's rule (whenever possible, electrons retain unpaired spins in degenerate orbitals).

Electrons in the outermost orbitals, called valence electrons, are responsible for most of the chemical behavior of elements. In the periodic table, elements with analogous valence electron configurations usually occur within the same group. There are some exceptions to the predicted filling order, particularly when half-filled or completely filled orbitals can be formed. The periodic table

can be divided into three categories based on the orbital in which the last electron to be added is placed: main group elements (*s* and *p* orbitals), transition elements (*d* orbitals), and inner transition elements (*f* orbitals).

Electron configurations allow us to understand many periodic trends. Covalent radius increases as we move down a group because the *n* level (orbital size) increases. Covalent radius mostly decreases as we move left to right across a period because the effective nuclear charge experienced by the electrons increases, and the electrons are pulled in tighter to the nucleus. Anionic radii are larger than the parent atom, while cationic radii are smaller, because the number of valence electrons has changed while the nuclear charge has remained constant. Ionization energy (the energy associated with forming a cation) decreases down a group and mostly increases across a period because it is easier to remove an electron from a larger, higher energy orbital. Electron affinity (the energy associated with forming an anion) is more favorable (exothermic) when electrons are placed into lower energy orbitals, closer to the nucleus. Therefore, electron affinity becomes increasingly negative as we move left to right across the periodic table and decreases as we move down a group. For both IE and electron affinity data, there are exceptions to the trends when dealing with completely filled or half-filled subshells.

The discovery of the periodic recurrence of similar properties among the elements led to the formulation of the periodic table, in which the elements are arranged in order of increasing atomic number in rows known as periods and columns known as groups. Elements in the same group of the periodic table have similar chemical properties. Elements can be classified as metals, metalloids, and nonmetals, or as a main-group elements, transition metals, and inner transition metals. Groups are numbered 1–18 from left to right. The elements in group 1 are known as the alkali metals; those in group 2 are the alkaline earth metals; those in 15 are the pnictogens; those in 16 are the chalcogens; those in 17 are the halogens; and those in 18 are the noble gases.

Metals (particularly those in groups 1 and 2) tend to lose the number of electrons that would leave them with the same number of electrons as in the preceding noble gas in the periodic table. By this means, a positively charged ion is formed. Similarly, nonmetals (especially those in groups 16 and 17, and, to a lesser extent, those in Group 15) can gain the number of electrons needed to provide atoms with the same number of electrons as in the next noble gas in the periodic table. Thus, nonmetals tend to form negative ions. Positively charged ions are called cations, and negatively charged ions are called anions. Ions can be either monatomic (containing only one atom) or polyatomic (containing more than one atom).

Compounds that contain ions are called ionic compounds. Ionic compounds generally form from metals and nonmetals. Compounds that do not contain ions, but instead consist of atoms bonded tightly together in molecules (uncharged groups of atoms that behave as a single unit), are called covalent compounds. Covalent compounds usually form from two nonmetals.

This page titled 3.13: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

3.14: Exercises

1. The light produced by a red neon sign is due to the emission of light by excited neon atoms. Qualitatively describe the spectrum produced by passing light from a neon lamp through a prism.
2. An FM radio station found at 103.1 on the FM dial broadcasts at a frequency of $1.031 \times 10^8\ s^{-1}$ (103.1 MHz). What is the wavelength of these radio waves in meters?
3. FM-95, an FM radio station, broadcasts at a frequency of $9.51 \times 10^7\ s^{-1}$ (95.1 MHz). What is the wavelength of these radio waves in meters?

4. A bright violet line occurs at 435.8 nm in the emission spectrum of mercury vapor. What amount of energy, in joules, must be released by an electron in a mercury atom to produce a photon of this light?

5. Light with a wavelength of 614.5 nm looks orange. What is the energy, in joules, per photon of this orange light? What is the energy in eV (1 eV = 1.602×10^{-19} J)?

6. Heated lithium atoms emit photons of light with an energy of 2.961×10^{-19} J. Calculate the frequency and wavelength of one of these photons. What is the total energy in 1 mole of these photons? What is the color of the emitted light?

7. A photon of light produced by a surgical laser has an energy of 3.027×10^{-19} J. Calculate the frequency and wavelength of the photon. What is the total energy in 1 mole of photons? What is the color of the emitted light?

8. When rubidium ions are heated to a high temperature, two lines are observed in its line spectrum at wavelengths (a) 7.9×10^{-7} m and (b) 4.2×10^{-7} m. What are the frequencies of the two lines? What color do we see when we heat a rubidium compound?

9. The emission spectrum of cesium contains two lines whose frequencies are (a) 3.45×10^{14} Hz and (b) 6.53×10^{14} Hz. What are the wavelengths and energies per photon of the two lines? What color are the lines?

10. Photons of infrared radiation are responsible for much of the warmth we feel when holding our hands before a fire. These photons will also warm other objects. How many infrared photons with a wavelength of 1.5×10^{-6} m must be absorbed by the water to warm a cup of water (175 g) from 25.0 °C to 40 °C?

11. One of the radiographic devices used in a dentist's office emits an X-ray of wavelength 2.090×10^{-11} m. What is the energy, in joules, and frequency of this X-ray?

12. The eyes of certain reptiles pass a single visual signal to the brain when the visual receptors are struck by photons of a wavelength of 850 nm. If a total energy of 3.15×10^{-14} J is required to trip the signal, what is the minimum number of photons that must strike the receptor?

13. RGB color television and computer displays use cathode ray tubes that produce colors by mixing red, green, and blue light. If we look at the screen with a magnifying glass, we can see individual dots turn on and off as the colors change. Using a spectrum of visible light, determine the approximate wavelength of each of these colors. What is the frequency and energy of a photon of each of these colors?

14. Answer the following questions about a Blu-ray laser:

(a) The laser on a Blu-ray player has a wavelength of 405 nm. In what region of the electromagnetic spectrum is this radiation? What is its frequency?

(b) A Blu-ray laser has a power of 5 milliwatts (1 watt = 1 J s^{-1}). How many photons of light are produced by the laser in 1 hour?

(c) The ideal resolution of a player using a laser (such as a Blu-ray player), which determines how close together data can be stored on a compact disk, is determined using the following formula: Resolution = $0.60(\lambda/NA)$, where λ is the wavelength of the laser and NA is the numerical aperture. Numerical aperture is a measure of the size of the spot of light on the disk; the larger the NA, the smaller the spot. In a typical Blu-ray system, NA = 0.95. If the 405-nm laser is used in a Blu-ray player, what is the closest that information can be stored on a Blu-ray disk?

(d) The data density of a Blu-ray disk using a 405-nm laser is 1.5×10^7 bits mm^{-2}. Disks have an outside diameter of 120 mm and a hole of 15-mm diameter. How many data bits can be contained on the disk? If a Blu-ray disk can hold 9,400,000 pages of text,

how many data bits are needed for a typed page? (Hint: Determine the area of the disk that is available to hold data. The area inside a circle is given by $A = \pi r^2$, where the radius r is one-half of the diameter.)

15. What is the threshold frequency for sodium metal if a photon with frequency $6.66 \times 10^{14}\ s^{-1}$ ejects an electron with 7.74×10^{-20} J kinetic energy? Will the photoelectric effect be observed if sodium is exposed to orange light?

16. Why is the electron in a Bohr hydrogen atom bound less tightly when it has a quantum number of 3 than when it has a quantum number of 1?

17. What does it mean to say that the energy of the electrons in an atom is quantized?

18. Using the Bohr model, determine the energy, in joules, necessary to ionize a ground-state hydrogen atom. Show your calculations.

19. The electron volt (eV) is a convenient unit of energy for expressing atomic-scale energies. It is the amount of energy that an electron gains when subjected to a potential of 1 volt; $1\ eV = 1.602 \times 10^{-19}$ J. Using the Bohr model, determine the energy, in electron volts, of the photon produced when an electron in a hydrogen atom moves from the orbit with $n = 5$ to the orbit with $n = 2$. Show your calculations.

20. Using the Bohr model, determine the lowest possible energy, in joules, for the electron in the Li^{2+} ion.

21. Using the Bohr model, determine the lowest possible energy for the electron in the He^{+} ion.

22. Using the Bohr model, determine the energy of an electron with $n = 6$ in a hydrogen atom.

23. Using the Bohr model, determine the energy of an electron with $n = 8$ in a hydrogen atom.

24. How far from the nucleus in angstroms (1 angstrom = 1×10^{-10} m) is the electron in a hydrogen atom if it has an energy of -8.72×10^{-20} J?

25. What is the radius, in angstroms, of the orbital of an electron with $n = 8$ in a hydrogen atom?

26. Using the Bohr model, determine the energy in joules of the photon produced when an electron in a He^{+} ion moves from the orbit with $n = 5$ to the orbit with $n = 2$.

27. Using the Bohr model, determine the energy in joules of the photon produced when an electron in a Li^{2+} ion moves from the orbit with $n = 2$ to the orbit with $n = 1$.

28. Consider a large number of hydrogen atoms with electrons randomly distributed in the $n = 1$, 2, 3, and 4 orbits.

(a) How many different wavelengths of light are emitted by these atoms as the electrons fall into lower-energy orbits?

(b) Calculate the lowest and highest energies of light produced by the transitions described in part (a).

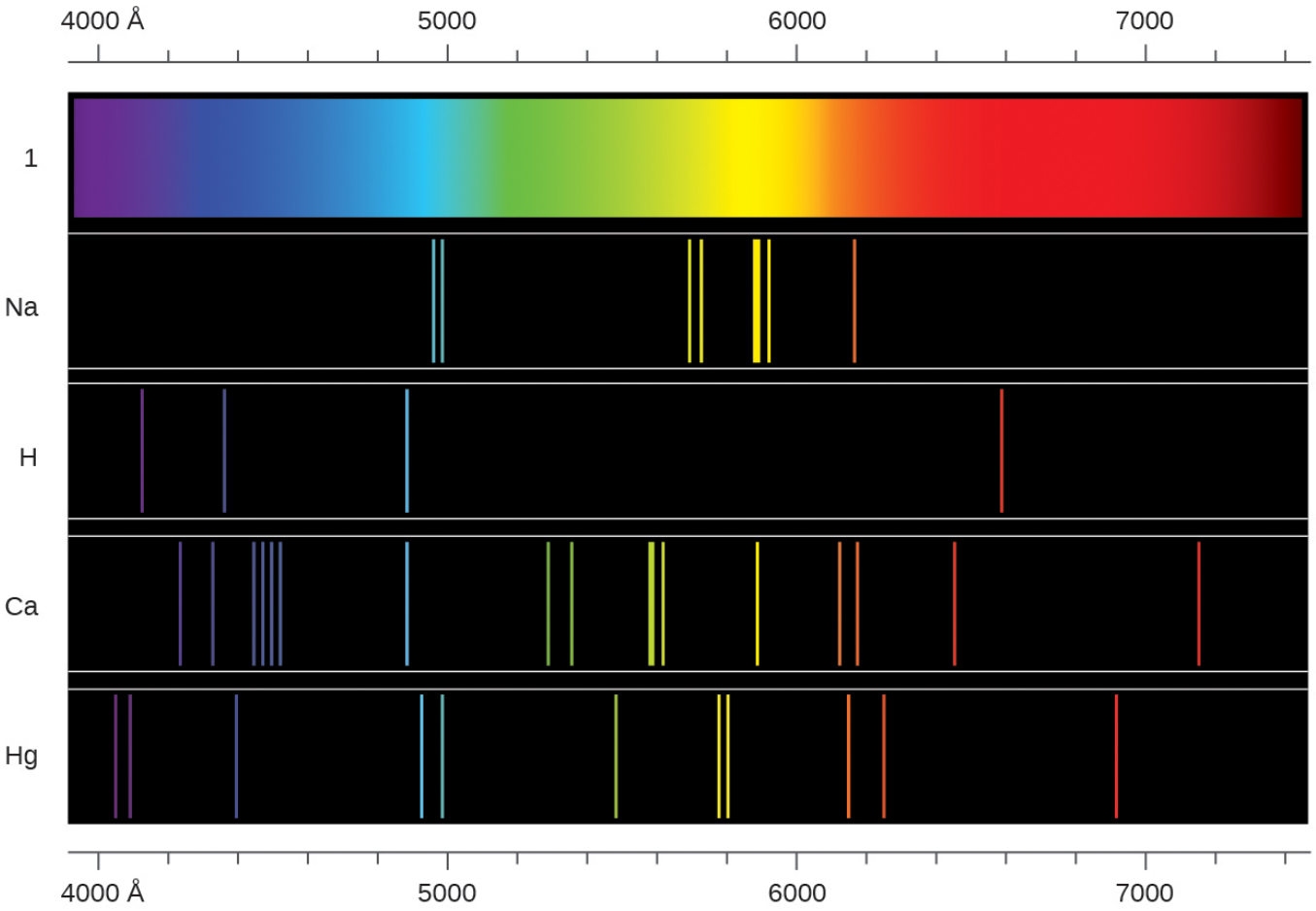
(c) Calculate the frequencies and wavelengths of the light produced by the transitions described in part (b).

29.

How are the Bohr model and the Rutherford model of the atom similar? How are they different?

30.

The spectra of hydrogen and of calcium are shown here.



What causes the lines in these spectra? Why are the colors of the lines different? Suggest a reason for the observation that the spectrum of calcium is more complicated than the spectrum of hydrogen.

31.

How are the Bohr model and the quantum mechanical model of the hydrogen atom similar? How are they different?

32.

What are the allowed values for each of the four quantum numbers: n, l, m_l, and m_s?

33.

Describe the properties of an electron associated with each of the following four quantum numbers: n, l, m_l, and m_s.

34.

Answer the following questions:

(a) Without using quantum numbers, describe the differences between the shells, subshells, and orbitals of an atom.

(b) How do the quantum numbers of the shells, subshells, and orbitals of an atom differ?

35.

Identify the subshell in which electrons with the following quantum numbers are found:

(a) $n = 2$, $l = 1$

(b) $n = 4$, $l = 2$

(c) $n = 6$, $l = 0$

36.

Which of the subshells described in the previous question contain degenerate orbitals? How many degenerate orbitals are in each?

37.

Identify the subshell in which electrons with the following quantum numbers are found:

(a) $n = 3$, $l = 2$

(b) $n = 1$, $l = 0$

(c) $n = 4$, $l = 3$

38.

Which of the subshells described in the previous question contain degenerate orbitals? How many degenerate orbitals are in each?

39.

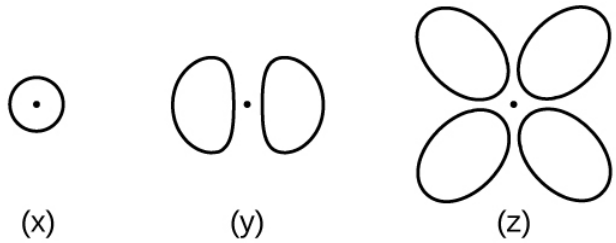
Sketch the boundary surface of a d_{x^2} and a p_y orbital. Be sure to show and label the axes.

40.

Sketch the p_x and d_{xz} orbitals. Be sure to show and label the coordinates.

41.

Consider the orbitals shown here in outline.



(a) What is the maximum number of electrons contained in an orbital of type (x)? Of type (y)? Of type (z)?

(b) How many orbitals of type (x) are found in a shell with $n = 2$? How many of type (y)? How many of type (z)?

(c) Write a set of quantum numbers for an electron in an orbital of type (x) in a shell with $n = 4$. Of an orbital of type (y) in a shell with $n = 2$. Of an orbital of type (z) in a shell with $n = 3$.

(d) What is the smallest possible n value for an orbital of type (x)? Of type (y)? Of type (z)?

(e) What are the possible l and m_l values for an orbital of type (x)? Of type (y)? Of type (z)?

42.

State the Heisenberg uncertainty principle. Describe briefly what the principle implies.

43.

How many electrons could be held in the second shell of an atom if the spin quantum number m_s could have three values instead of just two? (Hint: Consider the Pauli exclusion principle.)

44.

Which of the following equations describe particle-like behavior? Which describe wavelike behavior? Do any involve both types of behavior? Describe the reasons for your choices.

(a) $c = \lambda\nu$

(b) $E =$

(c) $r =$

(d) $E = h\nu$

(e) $\lambda =$

45.

Write a set of quantum numbers for each of the electrons with an n of 4 in a Se atom.

46.

Read the labels of several commercial products and identify monatomic ions of at least four transition elements contained in the products. Write the complete electron configurations of these cations.

47.

Read the labels of several commercial products and identify monatomic ions of at least six main group elements contained in the products. Write the complete electron configurations of these cations and anions.

48.

Using complete subshell notation (not abbreviations, $1s^2 2s^2 2p^6$, and so forth), predict the electron configuration of each of the following atoms:

(a) C

(b) P

(c) V

(d) Sb

(e) Sm

49.

Using complete subshell notation ($1s^2 2s^2 2p^6$, and so forth), predict the electron configuration of each of the following atoms:

(a) N

(b) Si

(c) Fe

(d) Te

(e) Tb

50.

Is $1s^2 2s^2 2p^6$ the symbol for a macroscopic property or a microscopic property of an element? Explain your answer.

51.

What additional information do we need to answer the question “Which ion has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$”?

52.

Draw the orbital diagram for the valence shell of each of the following atoms:

(a) C

(b) P

(c) V

(d) Sb

(e) Ru

53.

Use an orbital diagram to describe the electron configuration of the valence shell of each of the following atoms:

(a) N

(b) Si

(c) Fe

(d) Te

(e) Mo

54.

Using complete subshell notation ($1s^2 2s^2 2p^6$, and so forth), predict the electron configurations of the following ions.

(a) N^{3-}

(b) Ca^{2+}

(c) S^{-}

(d) Cs^{2+}

(e) Cr^{2+}

(f) Gd^{3+}

55.

Which atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$?

56.

Which atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$?

57.

Which ion with a +1 charge has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$? Which ion with a –2 charge has this configuration?

58.

Which of the following atoms contains only three valence electrons: Li, B, N, F, Ne?

59.

Which of the following has two unpaired electrons?

(a) Mg

(b) Si

(c) S

(d) Both Mg and S

(e) Both Si and S.

60.

Which atom would be expected to have a half-filled 6*p* subshell?

61.

Which atom would be expected to have a half-filled 4*s* subshell?

62.

In one area of Australia, the cattle did not thrive despite the presence of suitable forage. An investigation showed the cause to be the absence of sufficient cobalt in the soil. Cobalt forms cations in two oxidation states, Co^{2+} and Co^{3+}. Write the electron structure of the two cations.

63.

Thallium was used as a poison in the Agatha Christie mystery story “The Pale Horse.” Thallium has two possible cationic forms, +1 and +3. The +1 compounds are the more stable. Write the electron structure of the +1 cation of thallium.

64.

Write the electron configurations for the following atoms or ions:

(a) B^{3+}

(b) O^{-}

(c) Cl^{3+}

(d) Ca^{2+}

(e) Ti

65.

Cobalt–60 and iodine–131 are radioactive isotopes commonly used in nuclear medicine. How many protons, neutrons, and electrons are in atoms of these isotopes? Write the complete electron configuration for each isotope.

66.

Write a set of quantum numbers for each of the electrons with an *n* of 3 in a Sc atom.

67.

Based on their positions in the periodic table, predict which has the smallest atomic radius: Mg, Sr, Si, Cl, I.

68.

Based on their positions in the periodic table, predict which has the largest atomic radius: Li, Rb, N, F, I.

69.

Based on their positions in the periodic table, predict which has the largest first ionization energy: Mg, Ba, B, O, Te.

70.

Based on their positions in the periodic table, predict which has the smallest first ionization energy: Li, Cs, N, F, I.

71.

Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: F, Li, N, Rb

72.

Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: Mg, O, S, Si

73.

Atoms of which group in the periodic table have a valence shell electron configuration of ns^2np^3?

74.

Atoms of which group in the periodic table have a valence shell electron configuration of ns^2?

75.

Based on their positions in the periodic table, list the following atoms in order of increasing radius: Mg, Ca, Rb, Cs.

76.

Based on their positions in the periodic table, list the following atoms in order of increasing radius: Sr, Ca, Si, Cl.

77.

Based on their positions in the periodic table, list the following ions in order of increasing radius: K^+, Ca^{2+}, Al^{3+}, Si^{4+}.

78.

List the following ions in order of increasing radius: Li^+, Mg^{2+}, Br^-, Te^{2-}.

79.

Which atom and/or ion is (are) isoelectronic with Br^+: Se^{2+}, Se, As^-, Kr, Ga^{3+}, Cl^-?

80.

Which of the following atoms and ions is (are) isoelectronic with S^{2+}: Si^{4+}, Cl^{3+}, Ar, As^{3+}, Si, Al^{3+}?

81.

Compare both the numbers of protons and electrons present in each to rank the following ions in order of increasing radius: As^{3-}, Br^-, K^+, Mg^{2+}.

82.

Of the five elements Al, Cl, I, Na, Rb, which has the most exothermic reaction? (E represents an atom.) What name is given to the energy for the reaction? Hint: Note the process depicted does *not* correspond to electron affinity.)
$E^+(g) + e^- \longrightarrow E(g)$

83.

Of the five elements Sn, Si, Sb, O, Te, which has the most endothermic reaction? (E represents an atom.) What name is given to the energy for the reaction?

$\mathrm{E}(g) \longrightarrow \mathrm{E}+(g) + \mathrm{e}^-$

84.

The ionic radii of the ions S^{2-}, Cl^-, and K^+ are 184, 181, 138 pm respectively. Explain why these ions have different sizes even though they contain the same number of electrons.

85.

Which main group atom would be expected to have the lowest second ionization energy?

86.

Explain why Al is a member of group 13 rather than group 3?

87.

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

(a) uranium

(b) bromine

(c) strontium

(d) neon

(e) gold

(f) americium

(g) rhodium

(h) sulfur

(i) carbon

(j) potassium

88.

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

(a) cobalt

(b) europium

(c) iodine

(d) indium

(e) lithium

(f) oxygen

(g) cadmium

(h) terbium

(i) rhenium

89.

Using the periodic table, identify the lightest member of each of the following groups:

(a) noble gases

(b) alkaline earth metals

(c) alkali metals

(d) chalcogens

90.

Using the periodic table, identify the heaviest member of each of the following groups:

(a) alkali metals

(b) chalcogens

(c) noble gases

(d) alkaline earth metals

91.

Use the periodic table to give the name and symbol for each of the following elements:

(a) the noble gas in the same period as germanium

(b) the alkaline earth metal in the same period as selenium

(c) the halogen in the same period as lithium

(d) the chalcogen in the same period as cadmium

92.

Use the periodic table to give the name and symbol for each of the following elements:

(a) the halogen in the same period as the alkali metal with 11 protons

(b) the alkaline earth metal in the same period with the neutral noble gas with 18 electrons

(c) the noble gas in the same row as an isotope with 30 neutrons and 25 protons

(d) the noble gas in the same period as gold

93.

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

(a) the alkali metal with 11 protons and a mass number of 23

(b) the noble gas element with 75 neutrons in its nucleus and 54 electrons in the neutral atom

(c) the isotope with 33 protons and 40 neutrons in its nucleus

(d) the alkaline earth metal with 88 electrons and 138 neutrons

94.

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

(a) the chalcogen with a mass number of 125

(b) the halogen whose longest-lived isotope is radioactive

(c) the noble gas, used in lighting, with 10 electrons and 10 neutrons

(d) the lightest alkali metal with three neutrons

95.

Using the periodic table, predict whether the following chlorides are ionic or covalent: KCl, NCl_3, ICl, $MgCl_2$, PCl_5, and CCl_4.

96.

Using the periodic table, predict whether the following chlorides are ionic or covalent: $SiCl_4$, PCl_3, $CaCl_2$, CsCl, $CuCl_2$, and $CrCl_3$.

97.

For each of the following compounds, state whether it is ionic or covalent. If it is ionic, write the symbols for the ions involved:

(a) NF_3

(b) BaO,

(c) $(NH_4)_2CO_3$

(d) $Sr(H_2PO_4)_2$

(e) IBr

(f) Na_2O

98.

For each of the following compounds, state whether it is ionic or covalent, and if it is ionic, write the symbols for the ions involved:

(a) $KClO_4$

(b) $Mg(C_2H_3O_2)_2$

(c) H_2S

(d) Ag_2S

(e) N_2Cl_4

(f) $Co(NO_3)_2$

99.

For each of the following pairs of ions, write the symbol for the formula of the compound they will form.

(a) Ca^{2+}, S^{2-}

(b) $NH_4{}^+$, $SO_4{}^{2-}$

(c) Al^{3+}, Br^-

(d) Na^+, $HPO_4{}^{2-}$

(e) Mg^{2+}, $PO_4{}^{3-}$

100.

For each of the following pairs of ions, write the symbol for the formula of the compound they will form.

(a) K^+, O^{2-}

(b) $NH_4{}^+$, $PO_4{}^{3-}$

(c) Al^{3+}, O^{2-}

(d) Na^+, $CO_3{}^{2-}$

(e) Ba^{2+}, $PO_4{}^{3-}$

This page titled 3.14: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

4: The Periodic Table and Some Atomic Properties

4.1: Classifying the Elements- The Periodic Law and the Periodic Table
4.2: Metals and Nonmetals and their Ions
4.3: Sizes of Atoms and Ions
4.4: Ionization Energy
4.5: Electron Affinity
4.6: Magnetic Properties
4.7: Periodic Properties of the Elements
4.8: Key Terms
4.9: Key Equations
4.10: Summary
4.11: Exercises

4: The Periodic Table and Some Atomic Properties is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.1: Classifying the Elements- The Periodic Law and the Periodic Table

Learning Objectives

- To become familiar with the history of the periodic table.

The modern periodic table has evolved through a long history of attempts by chemists to arrange the elements according to their properties as an aid in predicting chemical behavior. One of the first to suggest such an arrangement was the German chemist Johannes Dobereiner (1780–1849), who noticed that many of the known elements could be grouped in triads (a set of three elements that have similar properties)—for example, chlorine, bromine, and iodine; or copper, silver, and gold. Dobereiner proposed that all elements could be grouped in such triads, but subsequent attempts to expand his concept were unsuccessful. We now know that portions of the periodic table—the *d* block in particular—contain triads of elements with substantial similarities. The middle three members of most of the other columns, such as sulfur, selenium, and tellurium in group 16 or aluminum, gallium, and indium in group 13, also have remarkably similar chemistry.

By the mid-19th century, the atomic masses of many of the elements had been determined. The English chemist John Newlands (1838–1898), hypothesizing that the chemistry of the elements might be related to their masses, arranged the known elements in order of increasing atomic mass and discovered that every seventh element had similar properties (Figure 4.1.1). (The noble gases were still unknown.) Newlands therefore suggested that the elements could be classified into octaves A group of seven elements, corresponding to the horizontal rows in the main group elements (not counting the noble gases, which were unknown at the time)., corresponding to the *horizontal* rows in the main group elements. Unfortunately, Newlands's "law of octaves" did not seem to work for elements heavier than calcium, and his idea was publicly ridiculed. At one scientific meeting, Newlands was asked why he didn't arrange the elements in alphabetical order instead of by atomic mass, since that would make just as much sense! Actually, Newlands was on the right track—with only a few exceptions, atomic mass does increase with atomic number, and similar properties occur every time a set of ns^2np^6 subshells is filled. Despite the fact that Newlands's table had no logical place for the *d*-block elements, he was honored for his idea by the Royal Society of London in 1887.

Note: John Newlands (1838–1898)

John Alexander Reina Newlands was an English chemist who worked on the development of the periodic table. He noticed that elemental properties repeated every seventh (or multiple of seven) element, as musical notes repeat every eighth note.



No.	No.	No.	No.	No.	No.	No.	No.
H 1	F 8	Cl 15	Co & Ni 22	Br 29	Pd 36	I 42	Pt & Ir 50
Li 2	Na 9	K 16	Cu 23	Rb 30	Ag 37	Cs 44	Tl 51
G 3	Mg 10	Ca 17	Zn 25	Sr 31	Bd [sic Cd] 38	Ba & V 45	Pb 54
Bo 4	Al 11	Cr 19	Y 24	Ce & La 33	U 40	Ta 46	Th 56
C 5	Si 12	Ti 18	In 26	Zr 32	Sn 39	W 47	Hg 52
N 6	P 13	Mn 20	As 27	Di & Mo 34	Sb 41	Nb 48	Bi 55
O 7	S 14	Fe 21	Se 28	Ro & Ru 35	Te 43	Au 49	Os 51

NOTE. Where two elements happen to have the same equivalent, both are designated by the same number.

Figure 4.1.1: The Arrangement of the Elements into Octaves as Proposed by Newlands. The table shown here accompanied a letter from a 27-year-old Newlands to the editor of the journal *Chemical News* in which he wrote: "If the elements are arranged in the order of their equivalents, with a few slight transpositions, as in the accompanying table, it will be observed that elements belonging to the same group usually appear on the same horizontal line. It will also be seen that the numbers of analogous elements generally differ either by 7 or by some multiple of seven; in other words, members of the same group stand to each other in the same relation as the extremities of one or more octaves in music. Thus, in the nitrogen group, between nitrogen and phosphorus there are 7 elements; between phosphorus and arsenic, 14; between arsenic and antimony, 14; and lastly, between antimony and bismuth, 14 also. This peculiar relationship I propose to provisionally term the *Law of Octaves*. I am, &c. John A. R. Newlands, F.C.S. Laboratory, 19, Great St. Helen's, E.C., August 8, 1865."

The periodic table achieved its modern form through the work of the German chemist Julius Lothar Meyer (1830–1895) and the Russian chemist Dimitri Mendeleev (1834–1907), both of whom focused on the relationships between atomic mass and various physical and chemical properties. In 1869, they independently proposed essentially identical arrangements of the elements. Meyer aligned the elements in his table according to periodic variations in simple atomic properties, such as "atomic volume" (Figure 4.1.2), which he obtained by dividing the atomic mass (molar mass) in grams per mole by the density (ρ) of the element in grams per cubic centimeter. This property is equivalent to what is today defined as molar volume, the molar mass of an element divided by its density, (measured in cubic centimeters per mole):

$$\frac{molar\ mass\left(\not{g}/mol\right)}{density\left(\not{g}/cm^3\right)} = molar\ volume\left(cm^3/mol\right) \tag{4.1.1}$$

As shown in Figure 4.1.2, the alkali metals have the highest molar volumes of the solid elements. In Meyer's plot of atomic volume versus atomic mass, the nonmetals occur on the rising portion of the graph, and metals occur at the peaks, in the valleys, and on the downslopes.

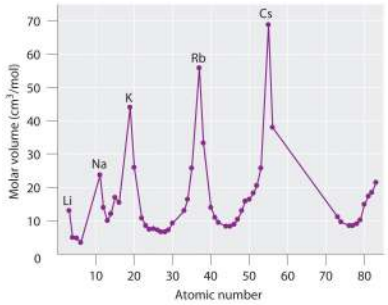
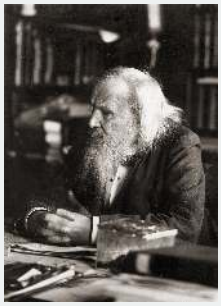


Figure 4.1.2: Variation of Atomic Volume with Atomic Number, Adapted from Meyer's Plot of 1870. *Note the periodic increase and decrease in atomic volume. Because the noble gases had not yet been discovered at the time this graph was formulated, the peaks correspond to the alkali metals (group 1).*

Note: Dimitri Mendeleev (1834–1907)

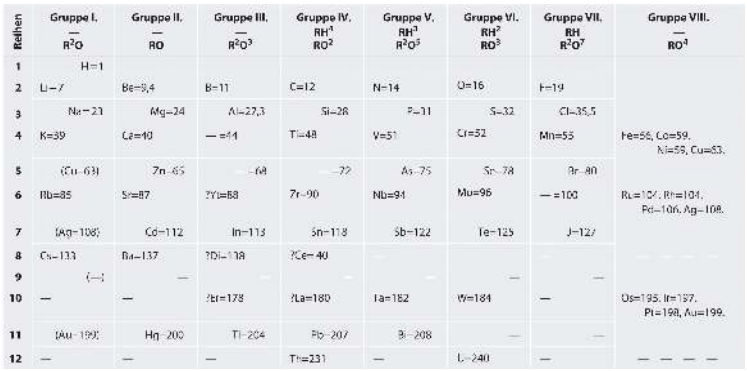
When his family's glass factory was destroyed by fire, Mendeleev moved to St. Petersburg, Russia, to study science. He became ill and was not expected to recover, but he finished his PhD with the help of his professors and fellow students.



In addition to the periodic table, another of Mendeleev's contributions to science was an outstanding textbook, *The Principles of Chemistry*, which was used for many years.

Mendeleev's Periodic Table

Mendeleev, who first published his periodic table in 1869 (Figure 4.1.3), is usually credited with the origin of the modern periodic table. The key difference between his arrangement of the elements and that of Meyer and others is that Mendeleev did not assume that all the elements had been discovered (actually, only about two-thirds of the naturally occurring elements were known at the time). Instead, he deliberately left blanks in his table at atomic masses 44, 68, 72, and 100, in the expectation that elements with those atomic masses would be discovered. Those blanks correspond to the elements we now know as scandium, gallium, germanium, and technetium.

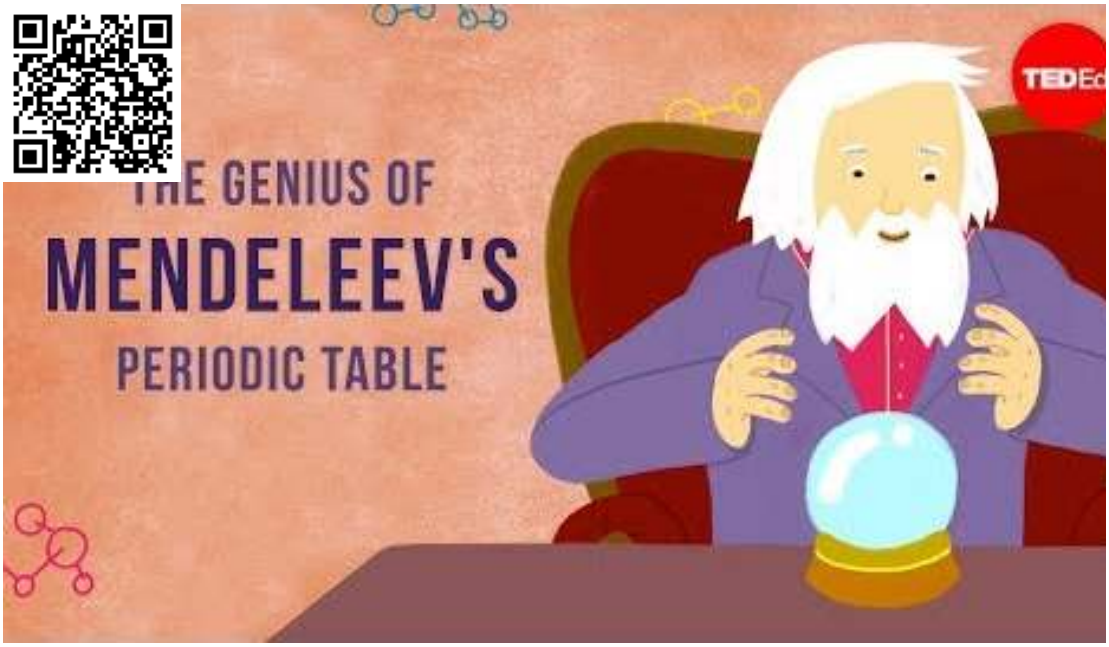


Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	— =44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	— =100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140				
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Figure 4.1.3: Mendeleev's Periodic Table, as Published in the German Journal *Annalen der Chemie und Pharmacie* in 1872. The column headings "Reihen" and "Gruppe" are German for "row" and "group." Formulas indicate the type of compounds formed by each group, with "R" standing for "any element" and superscripts used where we now use subscripts. Atomic masses are shown after equal signs and increase across each row from left to right.

The groups in Mendeleev's table are determined by how many oxygen or hydrogen atoms are needed to form compounds with each element. For example, in Group I, two atoms of hydrogen, lithium, Li, sodium, Na, and potassium form compounds with one atom of oxygen. In Group VII, one atom of fluorine, F, chlorine, Cl, and bromine, Br, react with one atom of hydrogen. Notice how this approach has trouble with the transition metals. Until roughly 1960, a rectangular table developed from Mendeleev's table and based on reactivity was standard at the front of chemistry lecture halls.

The most convincing evidence in support of Mendeleev's arrangement of the elements was the discovery of two previously unknown elements whose properties closely corresponded with his predictions (Table 4.1.1). Two of the blanks Mendeleev had left in his original table were below aluminum and silicon, awaiting the discovery of two as-yet-unknown elements, *eka*-aluminum and *eka*-silicon (from the Sanskrit *eka*, meaning "one," as in "one beyond aluminum"). The observed properties of gallium and germanium matched those of *eka*-aluminum and *eka*-silicon so well that once they were discovered, Mendeleev's periodic table rapidly gained acceptance.



***Video* 4.1.1**: *The genius of Mendeleev's periodic table.*

When the chemical properties of an element suggested that it might have been assigned the wrong place in earlier tables, Mendeleev carefully reexamined its atomic mass. He discovered, for example, that the atomic masses previously reported for beryllium, indium, and uranium were incorrect. The atomic mass of indium had originally been reported as 75.6, based on an assumed stoichiometry of InO for its oxide. If this atomic mass were correct, then indium would have to be placed in the middle of the nonmetals, between arsenic (atomic mass 75) and selenium (atomic mass 78). Because elemental indium is a silvery-white *metal*, however, Mendeleev postulated that the stoichiometry of its oxide was really In_2O_3 rather than InO. This would mean that indium's atomic mass was actually 113, placing the element between two other metals, cadmium and tin.

Table 4.1.1: Comparison of the Properties Predicted by Mendeleev in 1869 for eka-Aluminum and eka-Silicon with the Properties of Gallium (Discovered in 1875) and Germanium (Discovered in 1886)

Property	*eka*-Aluminum (predicted)	Gallium (observed)	*eka*-Silicon (predicted)	Germanium (observed)
atomic mass	68	69.723	72	72.64
element	metal	metal	dirty-gray metal	gray-white metal
	low mp*	mp = 29.8°C	high mp	mp = 938°C
	ρ = 5.9 g/cm^3	ρ = 5.91 g/cm^3	ρ = 5.5 g/cm^3	ρ = 5.323 g/cm^3
oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
	ρ = 5.5 g/cm^3	ρ = 6.0 g/cm^3	ρ = 4.7 g/cm^3	ρ = 4.25 g/cm^3
chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$
	volatile	mp = 78°C bp* = 201°C	bp < 100°C	bp = 87°C

*mp = melting point; bp = boiling point.

One group of elements that absent from Mendeleev's table is the noble gases, all of which were discovered more than 20 years later, between 1894 and 1898, by Sir William Ramsay (1852–1916; Nobel Prize in Chemistry 1904). Initially, Ramsay did not know where to place these elements in the periodic table. Argon, the first to be discovered, had an atomic mass of 40. This was greater than chlorine's and comparable to that of potassium, so Ramsay, using the same kind of reasoning as Mendeleev, decided to place the noble gases between the halogens and the alkali metals.

The Role of the Atomic Number in the Periodic Table

Despite its usefulness, Mendeleev's periodic table was based entirely on empirical observation supported by very little understanding. It was not until 1913, when a young British physicist, H. G. J. Moseley (1887–1915), while analyzing the frequencies of x-rays emitted by the elements, discovered that the underlying foundation of the order of the elements was by the *atomic number*, not the atomic mass. Moseley hypothesized that the placement of each element in his series corresponded to its atomic number Z, which is the number of positive charges (protons) in its nucleus. Argon, for example, although having an atomic mass greater than that of potassium (39.9 amu versus 39.1 amu, respectively), was placed *before* potassium in the periodic table. While analyzing the frequencies of the emitted x-rays, Moseley noticed that the atomic number of argon is 18, whereas that of potassium is 19, which indicated that they were indeed placed correctly. Moseley also noticed three gaps in his table of x-ray frequencies, so he predicted the existence of three unknown elements: technetium (Z = 43), discovered in 1937; promethium (Z = 61), discovered in 1945; and rhenium (Z = 75), discovered in 1925.

Note: H. G. J. Moseley (1887–1915)

Moseley left his research work at the University of Oxford to join the British army as a telecommunications officer during World War I. He was killed during the Battle of Gallipoli in Turkey.



Example 4.1.1

Before its discovery in 1999, some theoreticians believed that an element with a Z of 114 existed in nature. Use Mendeleev's reasoning to name element 114 as *eka*-_____; then identify the known element whose chemistry you predict would be most similar to that of element 114.

Given: atomic number

Asked for: name using prefix *eka*-

Strategy:

A. Using the periodic table locate the n = 7 row. Identify the location of the unknown element with Z = 114; then identify the known element that is directly above this location.
B. Name the unknown element by using the prefix *eka*- before the name of the known element.

Solution:

A The n = 7 row can be filled in by assuming the existence of elements with atomic numbers greater than 112, which is underneath mercury (Hg). Counting three boxes to the right gives element 114, which lies directly below lead (Pb). **B** If Mendeleev were alive today, he would call element 114 *eka*-lead.

Exercise 4.1.1

Use Mendeleev's reasoning to name element 112 as *eka*-_____; then identify the known element whose chemistry you predict would be most similar to that of element 112.

Answer: *eka*-mercury

Summary

- The elements in the periodic table are arranged according to their properties, and the periodic table serves as an aid in predicting chemical behavior.

The periodic table arranges the elements according to their electron configurations, such that elements in the same column have the same valence electron configurations. Periodic variations in size and chemical properties are important factors in dictating the types of chemical reactions the elements undergo and the kinds of chemical compounds they form. The modern periodic table was based on empirical correlations of properties such

as atomic mass; early models using limited data noted the existence of **triads** and **octaves** of elements with similar properties. The periodic table achieved its current form through the work of Dimitri Mendeleev and Julius Lothar Meyer, who both focused on the relationship between atomic mass and chemical properties. Meyer arranged the elements by their atomic volume, which today is equivalent to the **molar volume**, defined as molar mass divided by molar density. The correlation with the electronic structure of atoms was made when H. G. J. Moseley showed that the periodic arrangement of the elements was determined by atomic number, not atomic mass.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

4.1: Classifying the Elements- The Periodic Law and the Periodic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **9.1: Classifying the Elements: The Periodic Law and the Periodic Table** is licensed CC BY-NC-SA 4.0.

4.2: Metals and Nonmetals and their Ions

Learning Objectives

- To understand the basic properties separating Metals from Nonmetals and Metalloids

An element is the simplest form of matter that cannot be split into simpler substances or built from simpler substances by any ordinary chemical or physical method. There are 118 elements known to us, out of which 92 are naturally occurring, while the rest have been prepared artificially. Elements are further classified into metals, non-metals, and metalloids based on their properties, which are correlated with their placement in the periodic table.

Table 4.2.1: Characteristic properties of metallic and non-metallic elements:

Metallic Elements	Nonmetallic elements
Distinguishing luster (shine)	Non-lustrous, various colors
Malleable and ductile (flexible) as solids	Brittle, hard or soft
Conduct heat and electricity	Poor conductors
Metallic oxides are basic, ionic	Nonmetallic oxides are acidic, covalent
Form cations in aqueous solution	Form anions, oxyanions in aqueous solution

Metals

With the exception of hydrogen, all elements that form positive ions by losing electrons during chemical reactions are called metals. Thus metals are electropositive elements with relatively low ionization energies. They are characterized by bright luster, hardness, ability to resonate sound and are excellent conductors of heat and electricity. Metals are solids under normal conditions except for Mercury.

Physical Properties of Metals

Metals are lustrous, malleable, ductile, good conductors of heat and electricity. Other properties include:

- **State**: Metals are solids at room temperature with the exception of mercury, which is liquid at room temperature (Gallium is liquid on hot days).
- **Luster**: Metals have the quality of reflecting light from their surface and can be polished e.g., gold, silver and copper.
- **Malleability:** Metals have the ability to withstand hammering and can be made into thin sheets known as foils. For example, a sugar cube sized chunk of gold can be pounded into a thin sheet that will cover a football field.
- **Ductility:** Metals can be drawn into wires. For example, 100 g of silver can be drawn into a thin wire about 200 meters long.
- **Hardness:** All metals are hard except sodium and potassium, which are soft and can be cut with a knife.
- **Valency:** Metals typically have 1 to 3 electrons in the outermost shell of their atoms.
- **Conduction**: Metals are good conductors because they have free electrons. Silver and copper are the two best conductors of heat and electricity. Lead is the poorest conductor of heat. Bismuth, mercury and iron are also poor conductors
- **Density**: Metals have high density and are very heavy. Iridium and osmium have the highest densities whereas lithium has the lowest density.
- **Melting and Boiling Points**: Metals have high melting and boiling points. Tungsten has the highest melting and boiling points whereas mercury has the lowest. Sodium and potassium also have low melting points.

Chemical Properties of Metals

Metals are electropositive elements that generally form *basic* or *amphoteric* oxides with oxygen. Other chemical properties include:

- **Electropositive Character**: Metals tend to have low ionization energies, and *typically lose electrons (i.e. are **oxidized**) when they undergo chemical reactions* They normally do not accept electrons. For example:
  - Alkali metals are always 1^+ (lose the electron in *s* subshell)
  - Alkaline earth metals are always 2^+ (lose both electrons in *s* subshell)

- Transition metal ions do not follow an obvious pattern, 2^+ is common (lose both electrons in *s* subshell), and 1^+ and 3^+ are also observed

$$\mathrm{Na^0 \rightarrow Na^+ + e^-} \tag{4.2.1}$$

$$\mathrm{Mg^0 \rightarrow Mg^{2+} + 2\,e^-} \tag{4.2.2}$$

$$\mathrm{Al^0 \rightarrow Al^{3+} + 3\,e^-} \tag{4.2.3}$$

Compounds of metals with non-metals tend to be ***ionic*** in nature. Most metal oxides are basic oxides and dissolve in water to form ***metal hydroxides***:

$$\mathrm{Na_2O(s) + H_2O(l) \rightarrow 2\,NaOH(aq)} \tag{4.2.4}$$

$$\mathrm{CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)} \tag{4.2.5}$$

Metal oxides exhibit their ***basic*** chemical nature by reacting with ***acids*** to form metal ***salts*** and water:

$$\mathrm{MgO(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)} \tag{4.2.6}$$

$$\mathrm{NiO(s) + H_2SO_4(aq) \rightarrow NiSO_4(aq) + H_2O(l)} \tag{4.2.7}$$

Example 4.2.1

What is the chemical formula for aluminum oxide?

Solution

Al has a 3+ charge, the oxide ion is O^{2-}, thus Al_2O_3.

Example 4.2.2

Would you expect it to be solid, liquid or gas at room temperature?

Solutions

Oxides of metals are characteristically solid at room temperature

Example 4.2.3

Write the balanced chemical equation for the reaction of aluminum oxide with nitric acid:

Solution

Metal oxide + acid -> salt + water

$$\mathrm{Al_2O_3(s) + 6\,HNO_3(aq) \rightarrow 2\,Al(NO_3)_3(aq) + 3\,H_2O(l)}$$

Nonmetals

Elements that tend to gain electrons to form anions during chemical reactions are called non-metals. These are electronegative elements with high ionization energies. They are non-lustrous, brittle and poor conductors of heat and electricity (except graphite). Non-metals can be gases, liquids or solids.

Physical Properties of Nonmetals

- **Physical State**: Most of the non-metals exist in two of the three states of matter at room temperature: gases (oxygen) and solids (carbon). Only bromine exists as a liquid at room temperature.
- **Non-Malleable and Ductile**: Non-metals are very brittle, and cannot be rolled into wires or pounded into sheets.
- **Conduction**: They are poor conductors of heat and electricity.
- **Luster:** These have no metallic luster and do not reflect light.
- **Melting and Boiling Points**: The melting points of non-metals are *generally* lower than metals, but are highly variable.
- Seven non-metals exist under standard conditions as *diatomic molecules*: $\mathrm{H_2(g)}$, $\mathrm{N_2(g)}$, $\mathrm{O_2(g)}$, $\mathrm{F_2(g)}$, $\mathrm{Cl_2(g)}$, $\mathrm{Br_2(l)}$, $\mathrm{I_2(s)}$.

Chemical Properties of Nonmetals

Non-metals have a tendency to gain or share electrons with other atoms. They are electronegative in character. Nonmetals, when reacting with metals, tend to gain electrons (typically *attaining noble gas electron configuration)* and become ***anions:***

$$3\,Br_2(l) + 2\,Al(s) \rightarrow 2\,AlBr_3(s)$$

Compounds composed entirely of nonmetals are covalent substances. They generally form acidic or neutral oxides with oxygen that that dissolve in water to form acids:

$$CO_2(g) + H_2O(l) \rightarrow \underset{\text{carbonic acid}}{H_2CO_3(aq)}$$

As you may know, carbonated water is slightly acidic (carbonic acid).

Nonmetal oxides can combine with bases to form salts.

$$CO_2(g) + 2\,NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

Metalloids

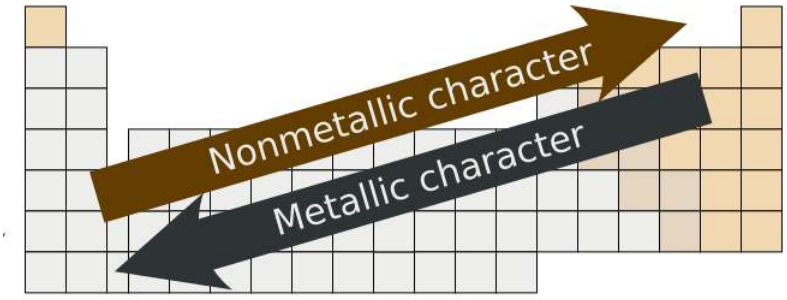
Metalloids have properties intermediate between the metals and nonmetals. Metalloids are useful in the semiconductor industry. Metalloids are all solid at room temperature. They can form alloys with other metals. Some metalloids, such as silicon and germanium, can act as electrical conductors under the right conditions, thus they are called semiconductors. Silicon for example appears lustrous, but is *not* malleable nor ductile (it is *brittle* - a characteristic of some nonmetals). It is a much poorer conductor of heat and electricity than the metals. The physical properties of metalloids tend to be metallic, but their chemical properties tend to be non-metallic. The oxidation number of an element in this group can range from +5 to -2, depending on the group in which it is located.

Table 4.2.2: Elements categorized into metals, non-metals and metalloids.

Metals	Non-metals	Metalloids
Gold	Oxygen	Silicon
Silver	Carbon	Boron
Copper	Hydrogen	Arsenic
Iron	Nitrogen	Antimony
Mercury	Sulfur	Germanium
Zinc	Phosphorus	

Trends in Metallic and Nonmetallic Character

Metallic character is strongest for the elements in the leftmost part of the periodic table, and tends to decrease as we move to the right in any period (nonmetallic character increases with increasing electronegativity and ionization energy values). Within any group of elements (columns), the metallic character increases from top to bottom (the electronegativity and ionization energy values generally decrease as we move down a group). This general trend is not necessarily observed with the transition metals.



Non metallic character increases bottom to top and left to right with the most nonmetal element in the top right corner. metallic character shows the opposite trend with the most metallic element being in the bottom left.

Contributors and Attributions

- Mike Blaber (Florida State University)
- Binod Shrestha (University of Lorraine)

4.2: Metals and Nonmetals and their Ions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **7.6: Metals, Nonmetals, and Metalloids** is licensed CC BY-NC-SA 3.0.

4.3: Sizes of Atoms and Ions

Learning Objectives

- To understand periodic trends in atomic radii.
- To predict relative ionic sizes within an isoelectronic series.

Although some people fall into the trap of visualizing atoms and ions as small, hard spheres similar to miniature table-tennis balls or marbles, the quantum mechanical model tells us that their shapes and boundaries are much less definite than those images suggest. As a result, atoms and ions cannot be said to have exact sizes. In this section, we discuss how atomic and ion "sizes" are defined and obtained.

Atomic Radii

Recall that the probability of finding an electron in the various available orbitals falls off slowly as the distance from the nucleus increases. This point is illustrated in Figure 4.3.1 which shows a plot of total electron density for all occupied orbitals for three noble gases as a function of their distance from the nucleus. Electron density diminishes gradually with increasing distance, which makes it impossible to draw a sharp line marking the boundary of an atom.

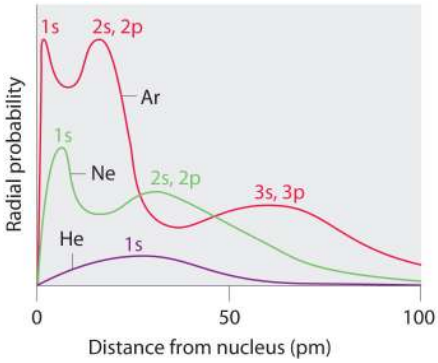


Figure 4.3.1: Plots of Radial Probability as a Function of Distance from the Nucleus for He, Ne, and Ar. In He, the 1*s* electrons have a maximum radial probability at ≈30 pm from the nucleus. In Ne, the 1*s* electrons have a maximum at ≈8 pm, and the 2*s* and 2*p* electrons combine to form another maximum at ≈35 pm (the $n = 2$ shell). In Ar, the 1*s* electrons have a maximum at ≈2 pm, the 2*s* and 2*p* electrons combine to form a maximum at ≈18 pm, and the 3*s* and 3*p* electrons combine to form a maximum at ≈70 pm.

Figure 4.3.1 also shows that there are distinct peaks in the total electron density at particular distances and that these peaks occur at different distances from the nucleus for each element. Each peak in a given plot corresponds to the electron density in a given principal shell. Because helium has only one filled shell ($n = 1$), it shows only a single peak. In contrast, neon, with filled $n = 1$ and 2 principal shells, has two peaks. Argon, with filled $n = 1$, 2, and 3 principal shells, has three peaks. The peak for the filled $n = 1$ shell occurs at successively shorter distances for neon ($Z = 10$) and argon ($Z = 18$) because, with a greater number of protons, their nuclei are more positively charged than that of helium. Because the $1s^2$ shell is closest to the nucleus, its electrons are very poorly shielded by electrons in filled shells with larger values of n. Consequently, the two electrons in the $n = 1$ shell experience nearly the full nuclear charge, resulting in a strong electrostatic interaction between the electrons and the nucleus. The energy of the $n = 1$ shell also decreases tremendously (the filled 1*s* orbital becomes more stable) as the nuclear charge increases. For similar reasons, the filled $n = 2$ shell in argon is located closer to the nucleus and has a lower energy than the $n = 2$ shell in neon.

Figure 4.3.1 illustrates the difficulty of measuring the dimensions of an individual atom. Because distances between the nuclei in pairs of covalently bonded atoms can be measured quite precisely, however, chemists use these distances as a basis for describing the approximate sizes of atoms. For example, the internuclear distance in the diatomic Cl_2 molecule is known to be 198 pm. We assign half of this distance to each chlorine atom, giving chlorine a covalent atomic radius (r_{cov}), which is half the distance between the nuclei of two like atoms joined by a covalent bond in the same molecule, of 99 pm or 0.99 Å (part (a) in Figure 4.3.2). Atomic radii are often measured in angstroms (Å), a non-SI unit: $1\ Å = 1 \times 10^{-10}$ m = 100 pm.

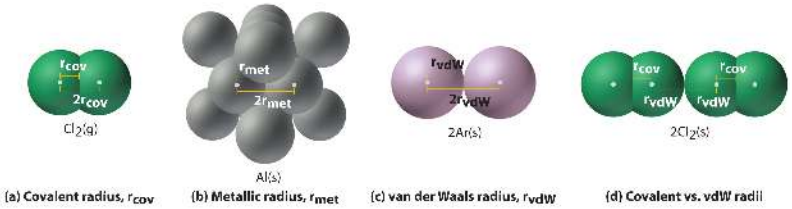


Figure 4.3.2: Definitions of the Atomic Radius. (a) The covalent atomic radius, r_{cov}, is half the distance between the nuclei of two like atoms joined by a covalent bond in the same molecule, such as Cl_2. (b) The metallic atomic radius, r_{met}, is half the distance between the nuclei of two adjacent atoms in a pure solid metal, such as aluminum. (c) The van der Waals atomic radius, r_{vdW}, is half the distance between the nuclei of two like atoms, such as argon, that are closely packed but not bonded. (d) This is a depiction of covalent versus van der Waals radii of chlorine. The covalent radius of Cl_2 is the distance between the two chlorine atoms in a single molecule of Cl_2. The van der Waals radius is the distance between chlorine nuclei in two different but touching Cl_2 molecules. Which do you think is larger? Why?

In a similar approach, we can use the lengths of carbon–carbon single bonds in organic compounds, which are remarkably uniform at 154 pm, to assign a value of 77 pm as the covalent atomic radius for carbon. If these values do indeed reflect the actual sizes of the atoms, then we should be able to predict the lengths of covalent bonds formed between different elements by adding them. For example, we would predict a carbon–chlorine distance of 77 pm + 99 pm = 176 pm for a C–Cl bond, which is very close to the average value observed in many organochlorine compounds.A similar approach for measuring the size of ions is discussed later in this section.

Covalent atomic radii can be determined for most of the nonmetals, but how do chemists obtain atomic radii for elements that do not form covalent bonds? For these elements, a variety of other methods have been developed. With a metal, for example, the metallic atomic radius (r_{met}) is defined as half the distance between the nuclei of two adjacent metal atoms (part (b) in Figure 4.3.2). For elements such as the noble gases, most of which form no stable compounds, we can use what is called the van der Waals atomic radius (r_{vdW}), which is half the internuclear distance between two nonbonded atoms in the solid (part (c) in Figure 4.3.2). This is somewhat difficult for helium which does not form a solid at any temperature. An atom such as chlorine has both a covalent radius (the distance between the two atoms in a Cl_2 molecule) and a van der Waals radius (the distance between two Cl atoms in different molecules in, for example, $Cl_{2(s)}$ at low temperatures). These radii are generally not the same (part (d) in Figure 4.3.2).

Periodic Trends in Atomic Radii

Because it is impossible to measure the sizes of both metallic and nonmetallic elements using any one method, chemists have developed a self-consistent way of calculating atomic radii using the quantum mechanical functions. Although the radii values obtained by such calculations are not identical to any of the experimentally measured sets of values, they do provide a way to compare the intrinsic sizes of all the elements and clearly show that atomic size varies in a periodic fashion (Figure 4.3.3).

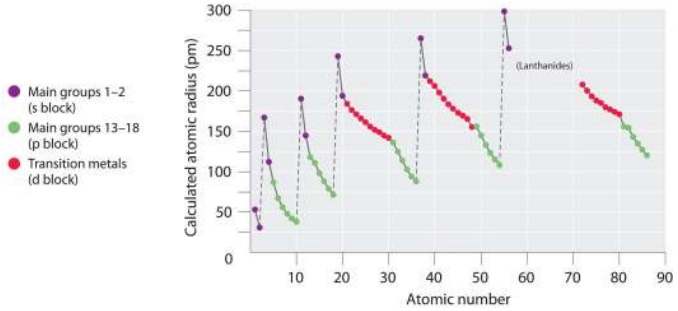


Figure 4.3.3: A Plot of Periodic Variation of Atomic Radius with Atomic Number for the First Six Rows of the Periodic Table.

In the periodic table, atomic radii decrease from left to right across a row and increase from top to bottom down a column. Because of these two trends, the largest atoms are found in the lower left corner of the periodic table, and the smallest are found in the upper right corner (Figure 4.3.4).

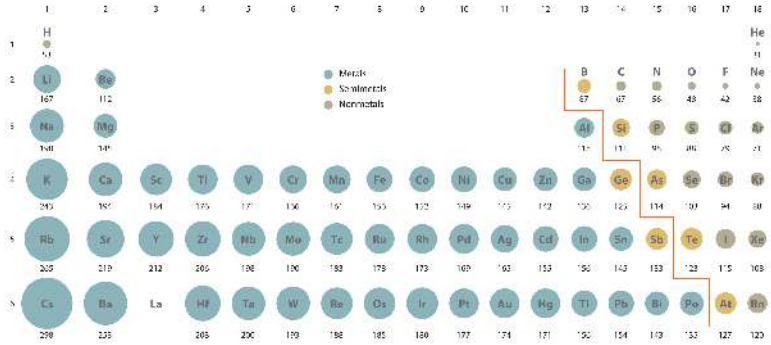


Figure 4.3.4 *Calculated Atomic Radii (in Picometers) of the s-, p-, and d-Block Elements. The sizes of the circles illustrate the relative sizes of the atoms. The calculated values are based on quantum mechanical wave functions. Source: http://www.webelements.com. Web Elements is an excellent on line source for looking up atomic properties.*

Note

Atomic radii decrease from left to right across a row and increase from top to bottom down a column.

Trends in atomic size result from differences in the **effective nuclear charges (Z_{eff})** experienced by electrons in the outermost orbitals of the elements. For all elements except H, the effective nuclear charge is always *less* than the actual nuclear charge because of shielding effects. The greater the effective nuclear charge, the more strongly the outermost electrons are attracted to the nucleus and the smaller the atomic radius.

The atoms in the second row of the periodic table (Li through Ne) illustrate the effect of electron shielding. All have a filled $1s^2$ inner shell, but as we go from left to right across the row, the nuclear charge increases from +3 to +10. Although electrons are being added to the 2*s* and 2*p* orbitals, *electrons in the same principal shell are not very effective at shielding one another from the nuclear charge*. Thus the single 2*s* electron in lithium experiences an effective nuclear charge of approximately +1 because the electrons in the filled $1s^2$ shell effectively neutralize two of the three positive charges in the nucleus. (More detailed calculations give a value of Z_{eff} = +1.26 for Li.) In contrast, the two 2*s* electrons in beryllium do not shield each other very well, although the filled $1s^2$ shell effectively neutralizes two of the four positive charges in the nucleus. This means that the effective nuclear charge experienced by the 2*s* electrons in beryllium is between +1 and +2 (the calculated value is +1.66). Consequently, beryllium is significantly smaller than lithium. Similarly, as we proceed across the row, the increasing nuclear charge is not effectively neutralized by the electrons being added to the 2*s* and 2*p* orbitals. The result is a steady increase in the effective nuclear charge and a steady decrease in atomic size.

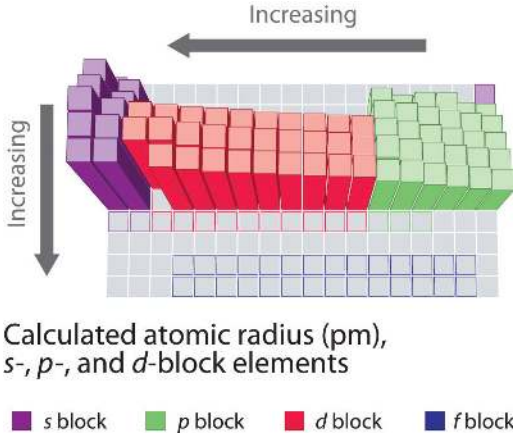


Figure 4.3.5: *The Atomic Radius of the Elements. The atomic radius of the elements increases as we go from right to left across a period and as we go down the periods in a group.*

The increase in atomic size going down a column is also due to electron shielding, but the situation is more complex because the principal quantum number *n* is not constant. As we saw in Chapter 2, the size of the orbitals increases as *n* increases, *provided the nuclear charge remains the same*. In group 1, for example, the size of the atoms increases substantially going down the column. It may at first seem reasonable to attribute this effect to the successive addition of electrons to *ns* orbitals with increasing values of *n*.

However, it is important to remember that the radius of an orbital depends dramatically on the nuclear charge. As we go down the column of the group 1 elements, the principal quantum number *n* increases from 2 to 6, but the nuclear charge increases from +3 to +55!

As a consequence the radii of the **lower electron orbitals** in Cesium are much smaller than those in lithium and the electrons in those orbitals experience a much larger force of attraction to the nucleus. That force depends on the effective nuclear charge experienced by the the inner electrons. If the outermost electrons in cesium experienced the full nuclear charge of +55, a cesium atom would be very small indeed. In fact, the effective nuclear charge felt by the outermost electrons in cesium is much less than expected (6 rather than 55). This means that cesium, with a $6s^1$ valence electron configuration, is much larger than lithium, with a $2s^1$ valence electron configuration. The effective nuclear charge changes relatively little for electrons in the outermost, or valence shell, from lithium to cesium because *electrons in filled inner shells are highly effective at shielding electrons in outer shells from the nuclear charge*. Even though cesium has a nuclear charge of +55, it has 54 electrons in its filled $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$ shells, abbreviated as $[Xe]5s^2 4d^{10} 5p^6$, which effectively neutralize most of the 55 positive charges in the nucleus. The same dynamic is responsible for the steady increase in size observed as we go down the other columns of the periodic table. Irregularities can usually be explained by variations in effective nuclear charge.

Note

Electrons in the same principal shell are not very effective at shielding one another from the nuclear charge, whereas electrons in filled inner shells are highly effective at shielding electrons in outer shells from the nuclear charge.

Example 4.3.1

On the basis of their positions in the periodic table, arrange these elements in order of increasing atomic radius: aluminum, carbon, and silicon.

Given: three elements

Asked for: arrange in order of increasing atomic radius

Strategy:

A. Identify the location of the elements in the periodic table. Determine the relative sizes of elements located in the same column from their principal quantum number *n*. Then determine the order of elements in the same row from their effective nuclear charges. If the elements are not in the same column or row, use pairwise comparisons.
B. List the elements in order of increasing atomic radius.

Solution:

A These elements are not all in the same column or row, so we must use pairwise comparisons. Carbon and silicon are both in group 14 with carbon lying above, so carbon is smaller than silicon (C < Si). Aluminum and silicon are both in the third row with aluminum lying to the left, so silicon is smaller than aluminum (Si < Al) because its effective nuclear charge is greater. **B** Combining the two inequalities gives the overall order: C < Si < Al.

Exercise 4.3.1

On the basis of their positions in the periodic table, arrange these elements in order of increasing size: oxygen, phosphorus, potassium, and sulfur.

Answer: O < S < P < K



Atomic Radius: https://youtu.be/ZYKB8SNrGVY

Ionic Radii and Isoelectronic Series

An ion is formed when either one or more electrons are removed from a neutral atom (cations) to form a positive ion or when additional electrons attach themselves to neutral atoms (anions) to form a negative one. The designations cation or anion come from the early experiments with electricity which found that positively charged particles were attracted to the negative pole of a battery, the cathode, while negatively charged ones were attracted to the positive pole, the anode.

Ionic compounds consist of regular repeating arrays of alternating positively charged cations and negatively charges anions. Although it is not possible to measure an ionic radius directly for the same reason it is not possible to directly measure an atom's radius, it *is* possible to measure the distance between the nuclei of a cation and an adjacent anion in an ionic compound to determine the ionic radius (the radius of a cation or anion) of one or both. As illustrated in Figure 4.3.6, the internuclear distance corresponds to the *sum* of the radii of the cation and anion. A variety of methods have been developed to divide the experimentally measured distance proportionally between the smaller cation and larger anion. These methods produce sets of ionic radii that are internally consistent from one ionic compound to another, although each method gives slightly different values. For example, the radius of the Na^+ ion is essentially the same in NaCl and Na_2S, as long as the same method is used to measure it. Thus despite minor differences due to methodology, certain trends can be observed.

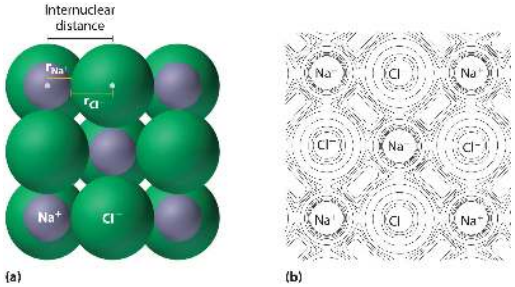


Figure 4.3.6: Definition of Ionic Radius. (a) The internuclear distance is apportioned between adjacent cations (positively charged ions) and anions (negatively charged ions) in the ionic structure, as shown here for Na^+ and Cl^- in sodium chloride. (b) This depiction of electron density contours for a single plane of atoms in the NaCl structure shows how the lines connect points of equal electron density. Note the relative sizes of the electron density contour lines around Cl^- and Na^+.

A comparison of ionic radii with atomic radii (Figure 4.3.7) *cation, having lost an electron, is always smaller than its parent neutral atom, and an anion, having gained an electron, is always larger than the parent neutral atom*. When one or more electrons is removed from a neutral atom, two things happen: (1) repulsions between electrons in the same principal shell decrease because fewer electrons are present, and (2) the effective nuclear charge felt by the remaining electrons increases because there are fewer electrons to shield one another from the nucleus. Consequently, the size of the region of space occupied by electrons decreases (compare Li at 167 pm with Li^+ at 76 pm). If different numbers of electrons can be removed to produce ions with different charges, the ion with the greatest positive charge is the smallest (compare Fe^{2+} at 78 pm with Fe^{3+} at 64.5 pm). Conversely, adding one or more electrons to a neutral atom causes electron–electron repulsions to increase and the effective nuclear charge to decrease, so the size of the probability region increases (compare F at 42 pm with F^- at 133 pm).

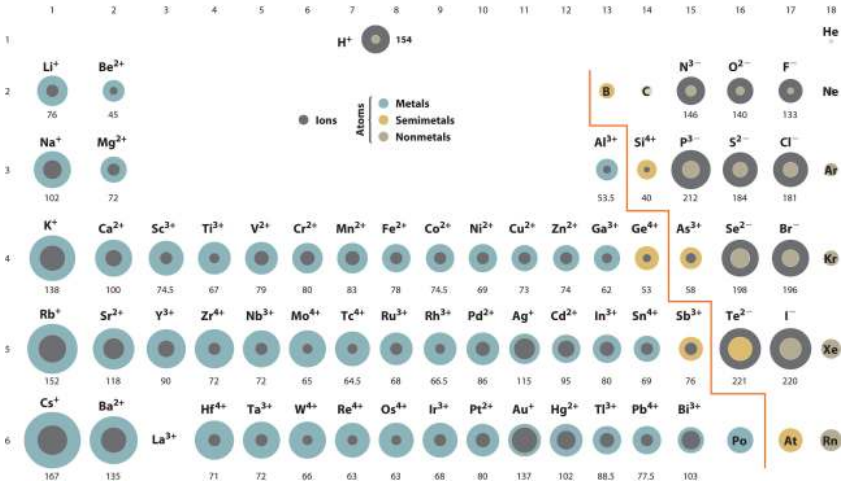


Figure 3.7. Source: Ionic radius data from R. D. Shannon, “Revised effective ionic radii and systematic studies of interatomic distances in halides and chalcogenides,” *Acta Crystallographica* 32, no. 5 (1976): 751–767.

Note

Cations are **always** smaller than the neutral atom, and anions are **always** larger.

Because most elements form either a cation or an anion but not both, there are few opportunities to compare the sizes of a cation and an anion derived from the same neutral atom. A few compounds of sodium, however, contain the Na^- ion, allowing comparison of its size with that of the far more familiar Na^+ ion, which is found in *many* compounds. The radius of sodium in each of its three known oxidation states is given in Table 4.3.1. All three species have a nuclear charge of +11, but they contain 10 (Na^+), 11 (Na^0), and 12 (Na^-) electrons. The Na^+ ion is significantly smaller than the neutral Na atom because the $3s^1$ electron has been removed to give a closed shell with $n = 2$. The Na^- ion is larger than the parent Na atom because the additional electron produces a $3s^2$ valence electron configuration, while the nuclear charge remains the same.

Table 4.3.1: Experimentally Measured Values for the Radius of Sodium in Its Three Known Oxidation States

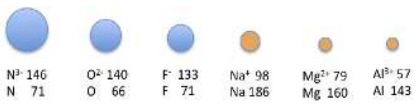
	Na^+	Na^0	Na^-
Electron Configuration	$1s^22s^22p^6$	$1s^22s^22p^63s^1$	$1s^22s^22p^63s^2$
Radius (pm)	102	154*	202†

*The metallic radius measured for Na(s).

†Source: M. J. Wagner and J. L. Dye, “Alkalides, Electrides, and Expanded Metals,” Annual Review of Materials Science 23 (1993): 225–253.

Ionic radii follow the same vertical trend as atomic radii; that is, for ions with the same charge, the ionic radius increases going down a column. The reason is the same as for atomic radii: shielding by filled inner shells produces little change in the effective nuclear charge felt by the outermost electrons. Again, principal shells with larger values of n lie at successively greater distances from the nucleus.

Because elements in different columns tend to form ions with different charges, it is not possible to compare ions of the same charge across a row of the periodic table. Instead, elements that are next to each other tend to form ions with the same number of electrons but with different overall charges because of their different atomic numbers. Such a set of species is known as an isoelectronic series. For example, the isoelectronic series of species with the neon closed-shell configuration ($1s^22s^22p^6$) is shown in Table 7.3



https://chem.libretexts.org/@go/page/518185

The sizes of the ions in this series decrease smoothly from N^{3-} to Al^{3+}. All six of the ions contain 10 electrons in the 1*s*, 2*s*, and 2*p* orbitals, but the nuclear charge varies from +7 (N) to +13 (Al). As the positive charge of the nucleus increases while the number of electrons remains the same, there is a greater electrostatic attraction between the electrons and the nucleus, which causes a decrease in radius. Consequently, the ion with the greatest nuclear charge (Al^{3+}) is the smallest, and the ion with the smallest nuclear charge (N^{3-}) is the largest. The neon atom in this isoelectronic series is not listed in Table 4.3.3, because neon forms no covalent or ionic compounds and hence its radius is difficult to measure.

Table 9.3.3 Radius of Ions with the Neon Closed-Shell Electron Configuration. Source: R. D. Shannon, "Revised effective ionic radii and systematic studies of interatomic distances in halides and chalcogenides," Acta Crystallographica 32, no. 5 (1976): 751–767.

Ion	Radius (pm)	Atomic Number
N^{3-}	146	7
O^{2-}	140	8
F^{-}	133	9
Na^{+}	98	11
Mg^{2+}	79	12
Al^{3+}	57	13

Example 4.3.2

Based on their positions in the periodic table, arrange these ions in order of increasing radius: Cl^{-}, K^{+}, S^{2-}, and Se^{2-}.

Given: four ions

Asked for: order by increasing radius

Strategy:

A. Determine which ions form an isoelectronic series. Of those ions, predict their relative sizes based on their nuclear charges. For ions that do not form an isoelectronic series, locate their positions in the periodic table.
B. Determine the relative sizes of the ions based on their principal quantum numbers *n* and their locations within a row.

Solution:

A We see that S and Cl are at the right of the third row, while K and Se are at the far left and right ends of the fourth row, respectively. K^{+}, Cl^{-}, and S^{2-} form an isoelectronic series with the [Ar] closed-shell electron configuration; that is, all three ions contain 18 electrons but have different nuclear charges. Because K^{+} has the greatest nuclear charge (Z = 19), its radius is smallest, and S^{2-} with Z = 16 has the largest radius. Because selenium is directly below sulfur, we expect the Se^{2-} ion to be even larger than S^{2-}. **B** The order must therefore be $K^{+} < Cl^{-} < S^{2-} < Se^{2-}$.

Exercise 4.3.2

Based on their positions in the periodic table, arrange these ions in order of increasing size: Br^{-}, Ca^{2+}, Rb^{+}, and Sr^{2+}.

Answer: $Ca^{2+} < Sr^{2+} < Rb^{+} < Br^{-}$

Summary

- Ionic radii share the same vertical trend as atomic radii, but the horizontal trends differ due to differences in ionic charges.

A variety of methods have been established to measure the size of a single atom or ion. The **covalent atomic radius (r_{cov})** is half the internuclear distance in a molecule with two identical atoms bonded to each other, whereas the **metallic atomic radius (r_{met})** is defined as half the distance between the nuclei of two adjacent atoms in a metallic element. The **van der Waals radius (r_{vdW})** of an element is half the internuclear distance between two nonbonded atoms in a solid. Atomic radii decrease from left to right across a row because of the increase in effective nuclear charge due to poor electron screening by other electrons in the same principal shell. Moreover, atomic radii increase from top to bottom down a column because the effective nuclear charge remains relatively constant as the principal quantum number increases. The **ionic radii** of cations and anions are always smaller or larger,

respectively, than the parent atom due to changes in electron–electron repulsions, and the trends in ionic radius parallel those in atomic size. A comparison of the dimensions of atoms or ions that have the same number of electrons but different nuclear charges, called an **isoelectronic series**, shows a clear correlation between increasing nuclear charge and decreasing size.

Contributors and Attributions

Modified by Joshua Halpern (Howard University)

4.3: Sizes of Atoms and Ions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **9.3: Sizes of Atoms and Ions** is licensed CC BY-NC-SA 4.0.

4.4: Ionization Energy

Learning Objectives

- To correlate ionization energies with the chemistry of the elements

We have seen that when elements react, they often gain or lose enough electrons to achieve the valence electron configuration of the nearest noble gas. In this section, we develop a more quantitative approach to predicting such reactions by examining periodic trends in the energy changes that accompany ion formation.

Ionization Energies

Because atoms do not spontaneously lose electrons, energy is required to remove an electron from an atom to form a cation. Chemists define the ionization energy (I) of an element as the amount of energy needed to remove an electron from the gaseous atom E in its ground state. I is therefore the energy required for the reaction

$$E_{(g)} \rightarrow E^+_{(g)} + e^- \qquad \text{energy required=I} \tag{4.4.1}$$

Because an input of energy is required, the ionization energy is always positive ($I > 0$) for the reaction as written in Equation 9.4.1. Larger values of *I* mean that the electron is more tightly bound to the atom and harder to remove. Typical units for ionization energies are kilojoules/mole (kJ/mol) or electron volts (eV):

$$1\ eV/atom = 96.49\ kJ/mol \tag{4.4.2}$$

If an atom possesses more than one electron, the amount of energy needed to remove successive electrons increases steadily. We can define a first ionization energy (I_1), a second ionization energy (I_2), and in general an *n*th ionization energy (I_n) according to the following reactions:

$$E_{(g)} \rightarrow E^+_{(g)} + e^- \qquad I_1 = \text{1st ionization energy} \tag{4.4.3}$$

$$E_{(g)} \rightarrow E^+_{(g)} + e^- \qquad I_2 = \text{2nd ionization energy} \tag{4.4.4}$$

$$E^+_{(g)} \rightarrow E^{2+}_{(g)} + e^- \qquad I_3 = \text{3rd ionization energy} \tag{4.4.5}$$

Values for the ionization energies of Li and Be listed in Table 4.4.1 show that successive ionization energies for an element increase steadily; that is, it takes more energy to remove the second electron from an atom than the first, and so forth. There are two reasons for this trend. First, the second electron is being removed from a positively charged species rather than a neutral one, so in accordance with Coulomb's law, more energy is required. Second, removing the first electron reduces the repulsive forces among the remaining electrons, so the attraction of the remaining electrons to the nucleus is stronger.

Note

Successive ionization energies for an element **increase** steadily.

Table 4.4.1: Ionization Energies (in kJ/mol) for Removing Successive Electrons from Li and Be. Source: Data from CRC Handbook of Chemistry and Physics (2004).

Reaction	Electronic Transition	I (4.4.6)	Reaction	Electronic Transition	I (4.4.7)
$Li_{(g)} \rightarrow Li^+_{(g)} + e^-$	$1s^2 2s^1 \rightarrow 1s^2$	I_1 = 520.2	$Be_{(g)} \rightarrow Be^+_{(g)} + e^-$	$1s^2 2s^2 \rightarrow 1s^2 2s^1$	I_1 = 899.5
$Li^+_{(g)} \rightarrow Li^{2+}_{(g)} + e^-$	$1s^2 \rightarrow 1s^1$	I_2 = 7298.2	$Be^+_{(g)} \rightarrow Be^{2+}_{(g)} + e^-$	$1s^2 2s^1 \rightarrow 1s^2$	I_2 = 1757.1
$Li^{2+}_{(g)} \rightarrow Li^{3+}_{(g)} + e^-$	$1s^1 \rightarrow 1s^0$	I_3 = 11,815.0	$Be^{2+}_{(g)} \rightarrow Be^{3+}_{(g)} + e^-$	$1s^2 \rightarrow 1s^1$	I_3 = 14,848.8
			$Be^{3+}_{(g)} \rightarrow Be^{4+}_{(g)} + e^-$	$1s^1 \rightarrow 1s^0$	I_4 = 21,006.6

The most important consequence of the values listed in Table 4.4.1 is that the chemistry of Li is dominated by the Li^+ ion, while the chemistry of Be is dominated by the +2 oxidation state. The energy required to remove the *second* electron from Li

$$Li^+_{(g)} \rightarrow Li^{2+}_{(g)} + e^- \tag{4.4.8}$$

is more than 10 times greater than the energy needed to remove the first electron. Similarly, the energy required to remove the *third* electron from Be

$$Be^{2+}_{(g)} \rightarrow Be^{3+}_{(g)} + e^- \tag{4.4.9}$$

is about 15 times greater than the energy needed to remove the first electron and around 8 times greater than the energy required to remove the second electron. Both Li^+ and Be^{2+} have $1s^2$ closed-shell configurations, and much more energy is required to remove an electron from the $1s^2$ core than from the 2*s* valence orbital of the same element. The chemical consequences are enormous: lithium (and all the alkali metals) forms compounds with the 1+ ion but not the 2+ or 3+ ions. Similarly, beryllium (and all the alkaline earth metals) forms compounds with the 2+ ion but not the 3+ or 4+ ions. *The energy required to remove electrons from a filled core is prohibitively large and simply cannot be achieved in normal chemical reactions.*

Note

The energy required to remove electrons from a filled core is prohibitively large under normal reaction conditions.

Ionization Energies of *s*- and *p*-Block Elements

Ionization energies of the elements in the third row of the periodic table exhibit the same pattern as those of Li and Be (Table 4.4.2): successive ionization energies increase steadily as electrons are removed from the valence orbitals (3*s* or 3*p*, in this case), followed by an especially large increase in ionization energy when electrons are removed from filled core levels as indicated by the bold diagonal line in Table 4.4.2. Thus in the third row of the periodic table, the largest increase in ionization energy corresponds to removing the fourth electron from Al, the fifth electron from Si, and so forth—that is, removing an electron from an ion that has the valence electron configuration of the preceding noble gas. This pattern explains why the chemistry of the elements normally involves only valence electrons. Too much energy is required to either remove or share the inner electrons.

Table 4.4.2: Successive Ionization Energies (in kJ/mol) for the Elements in the Third Row of the Periodic Table.Source: Data from CRC Handbook of Chemistry and Physics (2004).

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495.8	**4562.4***	—	—	—	—	—
Mg	737.7	1450.7	**7732.7**	—	—	—	—
Al	579.4.4	1816.7	2744.8	**11,579.4.4**	—	—	—
Si	786.5	1577.1	3231.6	4355.5	**16,090.6**	—	—
P	1011.8	1909.4.4	2914.1	4963.6	6274.0	**21,269.4.3**	—
S	999.6	2251.8	3357	4556.2	7004.3	8495.8	**27,109.4.3**
Cl	1251.2	2297.7	3822	5158.6	6540	9362	11,018.2
Ar	1520.6	2665.9	3931	5771	7238	8781.0	11,995.3

*Inner-shell electron

Example 4.4.1: Highest Fourth Ionization Energy

From their locations in the periodic table, predict which of these elements has the highest fourth ionization energy: B, C, or N.

Given: three elements

Asked for: element with highest fourth ionization energy

Strategy:

a. List the electron configuration of each element.
b. Determine whether electrons are being removed from a filled or partially filled valence shell. Predict which element has the highest fourth ionization energy, recognizing that the highest energy corresponds to the removal of electrons from a filled electron core.

Solution:

A These elements all lie in the second row of the periodic table and have the following electron configurations:

- B: [He]$2s^2 2p^1$
- C: [He]$2s^2 2p^2$
- N: [He]$2s^2 2p^3$

B The fourth ionization energy of an element (I_4) is defined as the energy required to remove the fourth electron:

$$E^{3+}_{(g)} \rightarrow E^{4+}_{(g)} + e^- \tag{4.4.10}$$

Because carbon and nitrogen have four and five valence electrons, respectively, their fourth ionization energies correspond to removing an electron from a partially filled valence shell. The fourth ionization energy for boron, however, corresponds to removing an electron from the filled $1s^2$ subshell. This should require much more energy. The actual values are as follows: B, 25,026 kJ/mol; C, 6223 kJ/mol; and N, 7475 kJ/mol.

? Exercise 4.4.1: Lowest Second Ionization Energy

From their locations in the periodic table, predict which of these elements has the lowest second ionization energy: Sr, Rb, or Ar.

Answer: Sr

The first column of data in Table 4.4.2 shows that first ionization energies tend to increase across the third row of the periodic table. This is because the valence electrons do not screen each other very well, allowing the effective nuclear charge to increase steadily across the row. The valence electrons are therefore attracted more strongly to the nucleus, so atomic sizes decrease and ionization energies increase. These effects represent two sides of the same coin: stronger electrostatic interactions between the electrons and the nucleus further increase the energy required to remove the electrons.

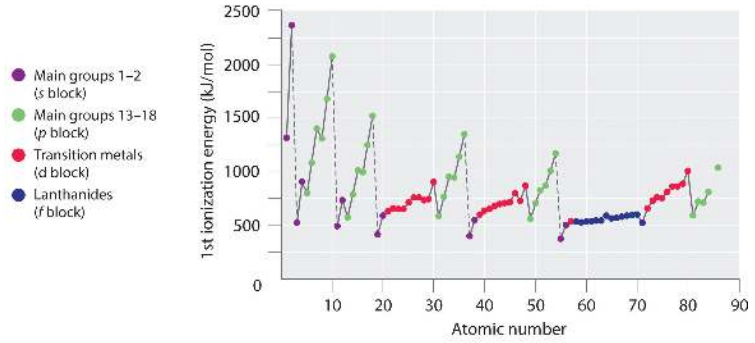


Figure 4.4.1: A Plot of Periodic Variation of First Ionization Energy with Atomic Number for the First Six Rows of the Periodic Table. There is a decrease in ionization energy within a group (most easily seen here for groups 1 and 18).

However, the first ionization energy decreases at Al ([Ne]$3s^2 3p^1$) and at S ([Ne]$3s^2 3p^4$). The electrons in aluminum's filled $3s^2$ subshell are better at screening the $3p^1$ electron than they are at screening each other from the nuclear charge, so the *s* electrons penetrate closer to the nucleus than the *p* electron does. The decrease at S occurs because the two electrons in the same *p* orbital repel each other. This makes the S atom slightly less stable than would otherwise be expected, as is true of all the group 16 elements.

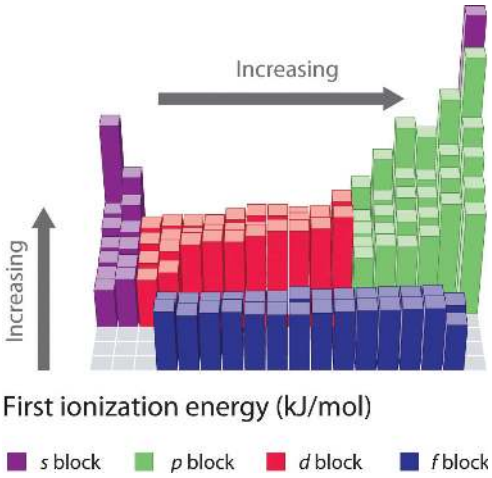


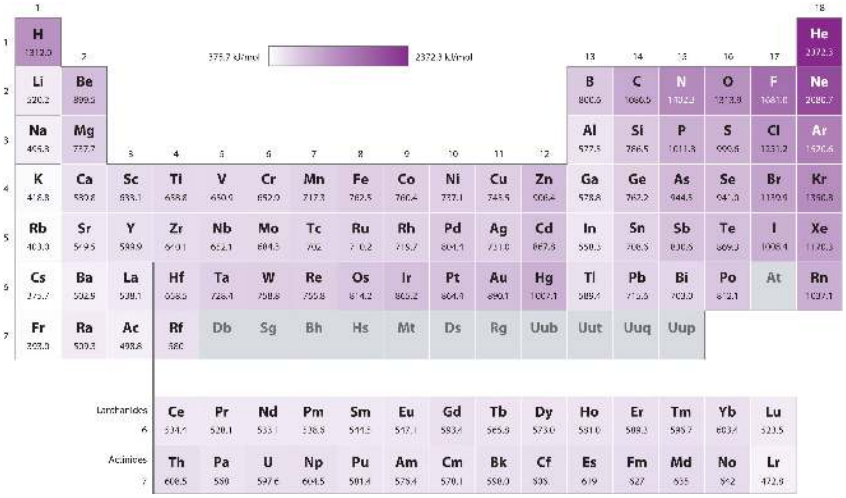
Figure* 4.4.2*: *First Ionization Energies of the s-, p-, d-, and f-Block Elements*

The first ionization energies of the elements in the first six rows of the periodic table are plotted in Figure 4.4.1 and are presented numerically and graphically in Figure 4.4.2. These figures illustrate three important trends:

1. The changes seen in the second (Li to Ne), fourth (K to Kr), fifth (Rb to Xe), and sixth (Cs to Rn) rows of the *s* and *p* blocks follow a pattern similar to the pattern described for the third row of the periodic table. The transition metals are included in the fourth, fifth, and sixth rows, however, and the lanthanides are included in the sixth row. The first ionization energies of the transition metals are somewhat similar to one another, as are those of the lanthanides. Ionization energies increase from left to right across each row, with discrepancies occurring at ns^2np^1 (group 13), ns^2np^4 (group 16), and $ns^2(n-1)d^{10}$ (group 12) electron configurations.
2. First ionization energies generally decrease down a column. Although the principal quantum number *n* increases down a column, filled inner shells are effective at screening the valence electrons, so there is a relatively small increase in the effective nuclear charge. Consequently, the atoms become larger as they acquire electrons. Valence electrons that are farther from the nucleus are less tightly bound, making them easier to remove, which causes ionization energies to decrease. *A larger radius corresponds to a lower ionization energy*.
3. Because of the first two trends, the elements that form positive ions most easily (have the lowest ionization energies) lie in the lower left corner of the periodic table, whereas those that are hardest to ionize lie in the upper right corner of the periodic table. Consequently, ionization energies generally increase diagonally from lower left (Cs) to upper right (He).

Note

Generally, I_1 increases diagonally from the lower left of the periodic table to the upper right.



https://chem.libretexts.org/@go/page/518186

The darkness of the shading inside the cells of the table indicates the relative magnitudes of the ionization energies. Elements in gray have undetermined first ionization energies. Source: Data from CRC Handbook of Chemistry and Physics (2004).

Gallium (Ga), which is the first element following the first row of transition metals, has the following electron configuration: [Ar]$4s^2 3d^{10} 4p^1$. Its first ionization energy is significantly lower than that of the immediately preceding element, zinc, because the filled $3d^{10}$ subshell of gallium lies inside the 4*p* subshell, screening the single 4*p* electron from the nucleus. Experiments have revealed something of even greater interest: the second and third electrons that are removed when gallium is ionized come from the $4s^2$ orbital, *not* the $3d^{10}$ subshell. The chemistry of gallium is dominated by the resulting Ga^{3+} ion, with its [Ar]$3d^{10}$ electron configuration. This and similar electron configurations are particularly stable and are often encountered in the heavier *p*-block elements. They are sometimes referred to as pseudo noble gas configurations. In fact, for elements that exhibit these configurations, *no chemical compounds are known in which electrons are removed from the* $(n - 1)d^{10}$ *filled subshell*.

Ionization Energies of Transition Metals & Lanthanides

As we noted, the first ionization energies of the transition metals and the lanthanides change very little across each row. Differences in their second and third ionization energies are also rather small, in sharp contrast to the pattern seen with the *s*- and *p*-block elements. The reason for these similarities is that the transition metals and the lanthanides form cations by losing the *ns* electrons before the $(n - 1)d$ or $(n - 2)f$ electrons, respectively. This means that transition metal cations have $(n - 1)d^n$ valence electron configurations, and lanthanide cations have $(n - 2)f^n$ valence electron configurations. Because the $(n - 1)d$ and $(n - 2)f$ shells are closer to the nucleus than the *ns* shell, the $(n - 1)d$ and $(n - 2)f$ electrons screen the *ns* electrons quite effectively, reducing the effective nuclear charge felt by the *ns* electrons. As *Z* increases, the increasing positive charge is largely canceled by the electrons added to the $(n - 1)d$ or $(n - 2)f$ orbitals.

That the *ns* electrons are removed before the $(n - 1)d$ or $(n - 2)f$ electrons may surprise you because the orbitals were filled in the reverse order. In fact, the *ns*, the $(n - 1)d$, and the $(n - 2)f$ orbitals are so close to one another in energy, and interpenetrate one another so extensively, that very small changes in the effective nuclear charge can change the order of their energy levels. As the *d* orbitals are filled, the effective nuclear charge causes the 3*d* orbitals to be slightly lower in energy than the 4*s* orbitals. The [Ar]$3d^2$ electron configuration of Ti^{2+} tells us that the 4*s* electrons of titanium are lost before the 3*d* electrons; this is confirmed by experiment. A similar pattern is seen with the lanthanides, producing cations with an $(n - 2)f^n$ valence electron configuration.

Because their first, second, and third ionization energies change so little across a row, these elements have important *horizontal* similarities in chemical properties in addition to the expected vertical similarities. For example, all the first-row transition metals except scandium form stable compounds as M^{2+} ions, whereas the lanthanides primarily form compounds in which they exist as M^{3+} ions.

✔ Example 4.4.2: Lowest First Ionization Energy

Use their locations in the periodic table to predict which element has the lowest first ionization energy: Ca, K, Mg, Na, Rb, or Sr.

Given: six elements

Asked for: element with lowest first ionization energy

Strategy:

Locate the elements in the periodic table. Based on trends in ionization energies across a row and down a column, identify the element with the lowest first ionization energy.

Solution:

These six elements form a rectangle in the two far-left columns of the periodic table. Because we know that ionization energies increase from left to right in a row and from bottom to top of a column, we can predict that the element at the bottom left of the rectangle will have the lowest first ionization energy: Rb.

? Exercise 4.4.2: Highest First Ionization Energy

Use their locations in the periodic table to predict which element has the highest first ionization energy: As, Bi, Ge, Pb, Sb, or Sn.

Answer: As



Ionization Energy: https://youtu.be/k7j-u02ifzo

Summary

- Generally, the first ionization energy and electronegativity values increase diagonally from the lower left of the periodic table to the upper right, and electron affinities become more negative across a row.

The tendency of an element to lose is one of the most important factors in determining the kind of compounds it forms. Periodic behavior is most evident for **ionization energy (*I*)**, the energy required to remove an electron from a gaseous atom. The energy required to remove successive electrons from an atom increases steadily, with a substantial increase occurring with the removal of an electron from a filled inner shell. Consequently, only valence electrons can be removed in chemical reactions, leaving the filled inner shell intact. Ionization energies explain the common oxidation states observed for the elements. Ionization energies increase diagonally from the lower left of the periodic table to the upper right. Minor deviations from this trend can be explained in terms of particularly stable electronic configurations, called **pseudo noble gas configurations**, in either the parent atom or the resulting ion.

4.4: Ionization Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **9.4: Ionization Energy** is licensed CC BY-NC-SA 4.0.

4.5: Electron Affinity

Learning Objectives

- To master the concept of electron affinity as a measure of the energy required to adding an electron to an atom or ion.
- To recognize the inverse relationship of ionization energies and electron affinities

The electron affinity (EA) of an element E is defined as the energy change that occurs when an electron is added to a gaseous atom:

$$E_{(g)} + e^- \rightarrow E^-_{(g)} \qquad \text{energy change} = EA \tag{4.5.1}$$

Unlike ionization energies, which are always positive for a neutral atom because energy is required to remove an electron, electron affinities can be negative (energy is released when an electron is added), positive (energy must be added to the system to produce an anion), or zero (the process is energetically neutral). This sign convention is consistent with a negative value corresponded to the energy change for an exothermic process, which is one in which heat is released.

The chlorine atom has the most negative electron affinity of any element, which means that more energy is released when an electron is added to a gaseous chlorine atom than to an atom of any other element:

$$Cl_{(g)} + e^- \rightarrow Cl^-_{(g)} \qquad EA = -346\ kJ/mol \tag{4.5.2}$$

In contrast, beryllium does not form a stable anion, so its effective electron affinity is

$$Be_{(g)} + e^- \rightarrow Be^-_{(g)} \qquad EA \geq 0 \tag{4.5.3}$$

Nitrogen is unique in that it has an electron affinity of approximately zero. Adding an electron neither releases nor requires a significant amount of energy:

$$N_{(g)} + e^- \rightarrow N^-_{(g)} \qquad EA \approx 0 \tag{4.5.4}$$

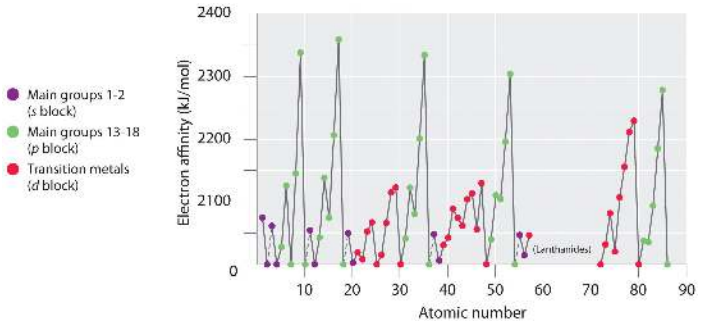


Figure 4.5.1: A Plot of Periodic Variation of Electron Affinity with Atomic Number for the First Six Rows of the Periodic Table

Note

Generally, electron affinities become more negative across a row of the periodic table.

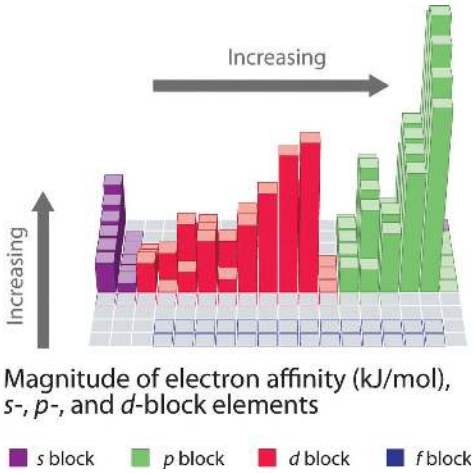
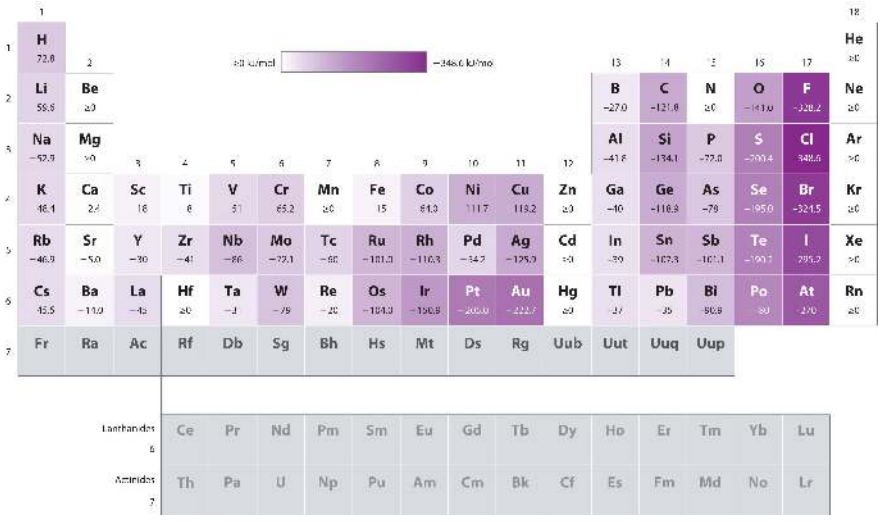


Figure 4.5.2 Electron Affinities (in kJ/mol) of the *s*-, *p*-, and *d*-Block Elements.

In general, electron affinities of the main-group elements become less negative as we proceed down a column. This is because as *n* increases, the extra electrons enter orbitals that are increasingly far from the nucleus. Atoms with the largest radii, which have the lowest ionization energies (affinity for their own valence electrons), also have the lowest affinity for an added electron. There are, however, two major exceptions to this trend:

1. The electron affinities of elements B through F in the second row of the periodic table are *less* negative than those of the elements immediately below them in the third row. Apparently, the increased electron–electron repulsions experienced by electrons confined to the relatively small 2*p* orbitals overcome the increased electron–nucleus attraction at short nuclear distances. Fluorine, therefore, has a lower affinity for an added electron than does chlorine. Consequently, the elements of the *third* row ($n = 3$) have the most negative electron affinities. Farther down a column, the attraction for an added electron decreases because the electron is entering an orbital more distant from the nucleus. Electron–electron repulsions also decrease because the valence electrons occupy a greater volume of space. These effects tend to cancel one another, so the changes in electron affinity within a family are much smaller than the changes in ionization energy.
2. The electron affinities of the alkaline earth metals become more negative from Be to Ba. The energy separation between the filled ns^2 and the empty *np* subshells decreases with increasing *n*, so that formation of an anion from the heavier elements becomes energetically more favorable.



There are many more exceptions to the trends across rows and down columns than with first ionization energies. Elements that do not form stable ions, such as the noble gases, are assigned an effective electron affinity that is greater than or equal to zero. Elements for which no data are available are shown in gray. Source: Data from Journal of Physical and Chemical Reference Data 28, no. 6 (1999).

Note

In general, electron affinities become more negative across a row and less negative down a column.

The equations for second and higher electron affinities are analogous to those for second and higher ionization energies:

$$E_{(g)} + e^- \rightarrow E^-_{(g)} \qquad \text{energy change} = EA_1 \tag{4.5.5}$$

$$E^-_{(g)} + e^- \rightarrow E^{2-}_{(g)} \qquad \text{energy change} = EA_2 \tag{4.5.6}$$

As we have seen, the first electron affinity can be greater than or equal to zero or negative, depending on the electron configuration of the atom. In contrast, the second electron affinity is *always* positive because the increased electron–electron repulsions in a dianion are far greater than the attraction of the nucleus for the extra electrons. For example, the first electron affinity of oxygen is −141 kJ/mol, but the second electron affinity is +744 kJ/mol:

$$O_{(g)} + e^- \rightarrow O^-_{(g)} \qquad EA_1 = -141\ kJ/mol \tag{4.5.7}$$

$$O^-_{(g)} + e^- \rightarrow O^{2-}_{(g)} \qquad EA_2 = +744\ kJ/mol \tag{4.5.8}$$

Thus the formation of a gaseous oxide (O^{2-}) ion is energetically quite *unfavorable* (estimated by adding both steps):

$$O_{(g)} + 2e^- \rightarrow O^{2-}_{(g)} \qquad EA = +603\ kJ/mol \tag{4.5.9}$$

Similarly, the formation of all common dianions (such as S^{2-}) or trianions (such as P^{3-}) is energetically unfavorable in the gas phase.

Note

While first electron affinities can be negative, positive, or zero, second electron affinities are always positive.

If energy is required to form both positively charged ions and monatomic polyanions, why do ionic compounds such as MgO, Na_2S, and Na_3P form at all? The key factor in the formation of stable ionic compounds is the favorable electrostatic interactions between the cations and the anions *in the crystalline salt*.

Example 4.5.1: Contrasting Electron Affinities of Sb, Se, and Te

Based on their positions in the periodic table, which of Sb, Se, or Te would you predict to have the most negative electron affinity?

Given: three elements

Asked for: element with most negative electron affinity

Strategy:

A. Locate the elements in the periodic table. Use the trends in electron affinities going down a column for elements in the same group. Similarly, use the trends in electron affinities from left to right for elements in the same row.
B. Place the elements in order, listing the element with the most negative electron affinity first.

Solution:

A We know that electron affinities become less negative going down a column (except for the anomalously low electron affinities of the elements of the second row), so we can predict that the electron affinity of Se is more negative than that of Te. We also know that electron affinities become more negative from left to right across a row, and that the group 15 elements tend to have values that are less negative than expected. Because Sb is located to the left of Te and belongs to group 15, we predict that the electron affinity of Te is more negative than that of Sb. The overall order is Se < Te < Sb, so Se has the most negative electron affinity among the three elements.

Exercise 4.5.1: Contrasting Electron Affinities of Rb, Sr, and Xe

Based on their positions in the periodic table, which of Rb, Sr, or Xe would you predict to most likely form a gaseous anion?

Answer: Rb



Electron Affinity: https://youtu.be/M5MsRM2Xtds

Summary

The **electron affinity** (*EA*) of an element is the energy change that occurs when an electron is added to a gaseous atom to give an anion. In general, elements with the most negative electron affinities (the highest affinity for an added electron) are those with the smallest size and highest ionization energies and are located in the upper right corner of the periodic table.

4.5: Electron Affinity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **9.5: Electron Affinity** is licensed CC BY-NC-SA 4.0.

4.6: Magnetic Properties

Learning Objectives

- To predict the magnetic properties of atoms and molecules based on their electronic configurations.

The magnetic moment of a system measures the strength and the direction of its magnetism. The term itself usually refers to the magnetic dipole moment. Anything that is magnetic, like a bar magnet or a loop of electric current, has a magnetic moment. A magnetic moment is a vector quantity, with a magnitude and a direction. An electron has an electron magnetic dipole moment, generated by the electron's intrinsic spin property, making it an electric charge in motion. There are many different magnetic forms: including paramagnetism, and diamagnetism, ferromagnetism, and anti-ferromagnetism. Only paramagnetism, and diamagnetism are discussed here.

Paramagnetism

Paramagnetism refers to the magnetic state of an atom with one or more unpaired electrons. The unpaired electrons are attracted by a magnetic field due to the electrons' magnetic dipole moments. Hund's Rule states that electrons must occupy every orbital singly before any orbital is doubly occupied. This may leave the atom with many unpaired electrons. Because unpaired electrons can orient in either direction, they exhibit magnetic moments that can align with a magnet. This capability allows paramagnetic atoms to be attracted to magnetic fields. Diatomic oxygen, O_2 is a good example of paramagnetism (that is best understood with molecular orbital theory). The following video shows liquid oxygen attracted into a magnetic field created by a strong magnet:



Video* 4.6.1*: *A chemical demonstration of the paramagnetism of molecular oxygen, as shown by the attraction of liquid oxygen to magnets.*

As shown in Video 4.6.1, since molecular oxygen (O_2 is paramagnetic, it is attracted to the magnet. In contrast, molecular nitrogen, N_2, has no unpaired electrons and it is diamagnetic (discussed below); it is therefore unaffected by the magnet.

Note

Paramagnetism is a form of magnetism whereby materials are **attracted** by an externally applied magnetic field.

There are some exceptions to the paramagnetism rule; these concern some transition metals, in which the unpaired electron is not in a d-orbital. Examples of these metals include Sc^{3+}, Ti^{4+}, Zn^{2+}, and Cu^{+}. These metals are the not defined as paramagnetic: they are considered diamagnetic because all d-electrons are paired. Paramagnetic compounds sometimes display bulk magnetic properties due to the clustering of the metal atoms. This phenomenon is known as ferromagnetism, but this property is not discussed here.

Diamagnetism

Diamagnetic substances are characterized by paired electrons—except in the previously-discussed case of transition metals, there are no unpaired electrons. According to the Pauli Exclusion Principle which states that no two identical electrons may take up the same quantum state at the same time, the electron spins are oriented in opposite directions. This causes the magnetic fields of the electrons to cancel out; thus there is no net

magnetic moment, and the atom cannot be attracted into a magnetic field. In fact, diamagnetic substances are weakly ***repelled*** by a magnetic field as demonstrated with the pyrolytic carbon sheet in Figure 4.6.1.

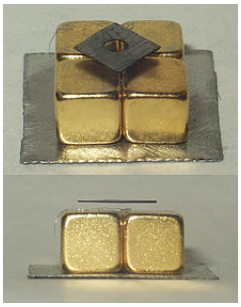


Figure 4.6.1: Levitating pyrolytic carbon: A small (~6 mm) piece of pyrolytic graphite levitating over a permanent neodymium magnet array (5 mm cubes on a piece of steel). Note that the poles of the magnets are aligned vertically and alternate (two with north facing up, and two with south facing up, diagonally). from Wikipedia.

Note

Diamagnetic materials are **repelled** by the applied magnetic field.

Diamagnetism, to a greater or lesser degree, is a property of all materials and always makes a weak contribution to the material's response to a magnetic field. For materials that show some other form of magnetism (such paramagntism), the diamagnetic contribution becomes negligible.

How to tell if a substance is paramagnetic or diamagnetic

The magnetic form of a substance can be determined by examining its electron configuration: if it shows unpaired electrons, then the substance is paramagnetic; if all electrons are paired, the substance is diamagnetic. This process can be broken into four steps:

1. Find the electron configuration
2. Draw the valence orbitals
3. Look for unpaired electrons
4. Determine whether the substance is paramagnetic (one or more electrons unpaired) or diamagnetic (all electrons paired)



Determining Magnetic Properties from Orbital Diagrams: https://youtu.be/lun_w5VKD8k

Example 4.6.1: Chlorine atoms

Are chlorine atoms paramagnetic or diamagnetic?

Solution

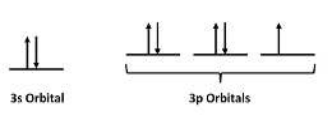
Follow the four steps outlines above.

Step 1: Find the electron configuration

For Cl atoms, the electron configuration is $3s^2 3p^5$

Step 2: Draw the valence orbitals

Ignore the core electrons and focus on the valence electrons only.



Step 3: Look for unpaired electrons

There is one unpaired electron.

Step 4: Determine whether the substance is paramagnetic or diamagnetic

Since there is an unpaired electron, Cl atoms are paramagnetic (but weakly since only one electron is unpaired).

Exercise 4.6.1

Indicate whether boron atoms are paramagnetic or diamagnetic.

Answer:

The B atom has $2s^2 2p^1$ as the electron configuration. Because it has one unpaired electron, it is paramagnetic.

Example 4.6.2: Zinc Atoms

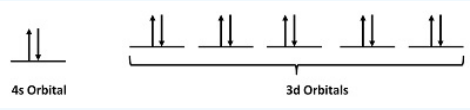
Are zinc atoms paramagnetic or diamagnetic?

Solution

Step 1: Find the electron configuration

For Zn atoms, the electron configuration is $4s^2 3d^{10}$

Step 2: Draw the valence orbitals



Step 3: Look for unpaired electrons

There are no unpaired electrons.

Step 4: Determine whether the substance is paramagnetic or diamagnetic

Because there are no unpaired electrons, Zn atoms are diamagnetic.

Exercise 4.6.2

Indicate whether F^- ions are paramagnetic or diamagnetic.

Answer

The F^- ion has $2s^2 2p^6$ has the electron configuration. Because it has no unpaired electrons, it is diamagnetic.

Problems

1. How many unpaired electrons are found in oxygen atoms ?
2. How many unpaired electrons are found in bromine atoms?
3. Indicate whether Fe^{2+} ions are paramagnetic or diamagnetic.

Answers

1. The O atom has $2s^2 2p^4$ as the electron configuration. Therefore, O has 2 unpaired electrons.
2. The Br atom has $4s^2 3d^{10} 4p^5$ as the electron configuration. Therefore, Br has 1 unpaired electron.
3. The Fe^{2+} ion has $3d^6$ has the electron configuration. Because it has 4 unpaired electrons, it is paramagnetic.

References

1. Pettrucci, Ralph H. General Chemistry: Principles and Modern Applications. 9th. Upper Saddle River: Pearson Prentice Hall, 2007
2. Sherman, Alan, Sharon J. Sherman, and Leonard Russikoff. Basic Concepts of Chemistry Fifth Edition. Boston, MA: Houghton Mifflin Company, 1992. Print.

4.6: Magnetic Properties is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **9.6: Magnetic Properties** is licensed CC BY-NC-SA 4.0.

4.7: Periodic Properties of the Elements

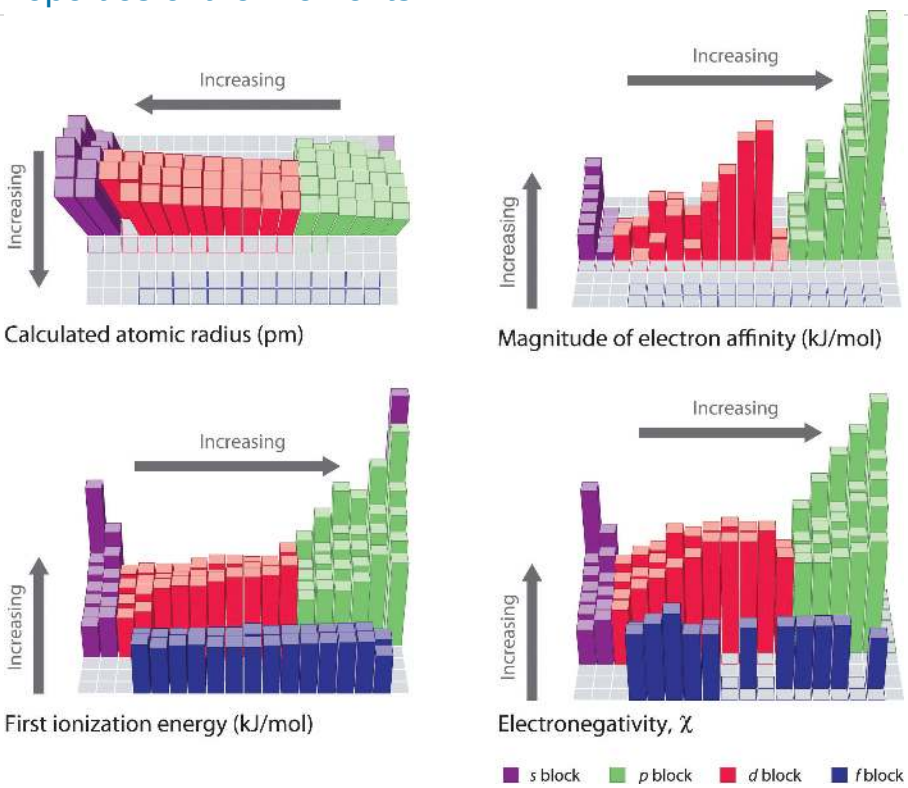


Figure 4.7.1 Summary of Major Periodic Trends. The general trends for the first ionization energy, electron affinity, and electronegativity are opposite to the general trend for covalent atomic radius.

4.7: Periodic Properties of the Elements is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.8: Key Terms

amplitude | extent of the displacement caused by a wave

atomic orbital | mathematical function that describes the behavior of an electron in an atom (also called the wavefunction)

Aufbau principle | procedure in which the electron configuration of the elements is determined by "building" them in order of atomic numbers, adding one proton to the nucleus and one electron to the proper subshell at a time

blackbody | idealized perfect absorber of all incident electromagnetic radiation; such bodies emit electromagnetic radiation in characteristic continuous spectra called blackbody radiation

Bohr's model of the hydrogen atom | structural model in which an electron moves around the nucleus only in circular orbits, each with a specific allowed radius

continuous spectrum | electromagnetic radiation given off in an unbroken series of wavelengths (e.g., white light from the sun)

core electron | electron in an atom that occupies the orbitals of the inner shells

covalent radius | one-half the distance between the nuclei of two identical atoms when they are joined by a covalent bond

***d* orbital** | region of space with high electron density that is either four lobed or contains a dumbbell and torus shape; describes orbitals with $l = 2$.

degenerate orbitals | orbitals that have the same energy

effective nuclear charge | charge that leads to the Coulomb force exerted by the nucleus on an electron, calculated as the nuclear charge minus shielding

electromagnetic radiation | energy transmitted by waves that have an electric-field component and a magnetic-field component

electromagnetic spectrum | range of energies that electromagnetic radiation can comprise, including radio, microwaves, infrared, visible, ultraviolet, X-rays, and gamma rays

electron affinity | energy change associated with addition of an electron to a gaseous atom or ion

electron configuration | listing that identifies the electron occupancy of an atom's shells and subshells

electron density | a measure of the probability of locating an electron in a particular region of space, it is equal to the squared absolute value of the wave function ψ

excited state | state having an energy greater than the ground-state energy

***f* orbital** | multilobed region of space with high electron density, describes orbitals with $l = 3$

frequency (ν) | number of wave cycles (peaks or troughs) that pass a specified point in space per unit time

ground state | state in which the electrons in an atom, ion, or molecule have the lowest energy possible

Heisenberg uncertainty principle | rule stating that it is impossible to exactly determine both certain conjugate dynamical properties such as the momentum and the position of a particle at the same time. The uncertainty principle is a consequence of quantum particles exhibiting wave–particle duality

hertz (Hz) | the unit of frequency, which is the number of cycles per second, s^{-1}

Hund's rule | every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin

intensity | property of wave-propagated energy related to the amplitude of the wave, such as brightness of light or loudness of sound

interference pattern | pattern typically consisting of alternating bright and dark fringes; it results from constructive and destructive interference of waves

ionization energy | energy required to remove an electron from a gaseous atom or ion

isoelectronic | group of ions or atoms that have identical electron configurations

line spectrum | electromagnetic radiation emitted at discrete wavelengths by a specific atom (or atoms) in an excited state

magnetic quantum number (m_l) | quantum number signifying the orientation of an atomic orbital around the nucleus

node | any point of a standing wave with zero amplitude

orbital diagram | pictorial representation of the electron configuration showing each orbital as a box and each electron as an arrow

***p* orbital** | dumbbell-shaped region of space with high electron density, describes orbitals with $l = 1$

Pauli exclusion principle | specifies that no two electrons in an atom can have the same value for all four quantum numbers

photon | smallest possible packet of electromagnetic radiation, a particle of light

principal quantum number (n) | quantum number specifying the shell an electron occupies in an atom

quantization | limitation of some property to specific discrete values, not continuous

quantum mechanics | field of study that includes quantization of energy, wave-particle duality, and the Heisenberg uncertainty principle to describe matter

quantum number | number having only specific allowed values and used to characterize the arrangement of electrons in an atom

***s* orbital** | spherical region of space with high electron density, describes orbitals with $l = 0$

secondary (angular momentum) quantum number (l) | quantum number distinguishing the different shapes of orbitals; it is also a measure of the orbital angular momentum

shell | atomic orbitals with the same principal quantum number, n

spin quantum number (m_s) | number specifying the electron spin direction, either or

standing wave | (also, stationary wave) localized wave phenomenon characterized by discrete wavelengths determined by the boundary conditions used to generate the waves; standing waves are inherently quantized

subshell | atomic orbitals with the same values of n and l

valence electrons | electrons in the high energy outer shell(s) of an atom

valence shell | high energy outer shell(s) of an atom

wave | oscillation of a property over time or space; can transport energy from one point to another

wave-particle duality | observation that elementary particles can exhibit both wave-like and particle-like properties

wavefunction (ψ) | mathematical description of an atomic orbital that describes the shape of the orbital; it can be used to calculate the probability of finding the electron at any given location in the orbital, as well as dynamical variables such as the energy and the angular momentum

wavelength (λ) | distance between two consecutive peaks or troughs in a wave

---

4.8: Key Terms is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **6.6: Key Terms** is licensed CC BY 4.0.

4.9: Key Equations

$c = \lambda v$
where $h = 6.626\ 10^{-34}$ J s

4.9: Key Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **6.7: Key Equations** is licensed CC BY 4.0.

4.10: Summary

Light and other forms of electromagnetic radiation move through a vacuum with a constant speed, c, of 2.998 10^8 m s^{-1}. This radiation shows wavelike behavior, which can be characterized by a frequency, ν, and a wavelength, λ, such that $c = \lambda\nu$. Light is an example of a travelling wave. Other important wave phenomena include standing waves, periodic oscillations, and vibrations. Standing waves exhibit quantization, since their wavelengths are limited to discrete integer multiples of some characteristic lengths. Electromagnetic radiation that passes through two closely spaced narrow slits having dimensions roughly similar to the wavelength will show an interference pattern that is a result of constructive and destructive interference of the waves. Electromagnetic radiation also demonstrates properties of particles called photons. The energy of a photon is related to the frequency (or alternatively, the wavelength) of the radiation as $E = h\nu$ (or), where h is Planck's constant. That light demonstrates both wavelike and particle-like behavior is known as wave-particle duality. All forms of electromagnetic radiation share these properties, although various forms including X-rays, visible light, microwaves, and radio waves interact differently with matter and have very different practical applications. Electromagnetic radiation can be generated by exciting matter to higher energies, such as by heating it. The emitted light can be either continuous (incandescent sources like the sun) or discrete (from specific types of excited atoms). Continuous spectra often have distributions that can be approximated as blackbody radiation at some appropriate temperature. The line spectrum of hydrogen can be obtained by passing the light from an electrified tube of hydrogen gas through a prism. This line spectrum was simple enough that an empirical formula called the Rydberg formula could be derived from the spectrum. Three historically important paradoxes from the late 19th and early 20th centuries that could not be explained within the existing framework of classical mechanics and classical electromagnetism were the blackbody problem, the photoelectric effect, and the discrete spectra of atoms. The resolution of these paradoxes ultimately led to quantum theories that superseded the classical theories.

Bohr incorporated Planck's and Einstein's quantization ideas into a model of the hydrogen atom that resolved the paradox of atom stability and discrete spectra. The Bohr model of the hydrogen atom explains the connection between the quantization of photons and the quantized emission from atoms. Bohr described the hydrogen atom in terms of an electron moving in a circular orbit about a nucleus. He postulated that the electron was restricted to certain orbits characterized by discrete energies. Transitions between these allowed orbits result in the absorption or emission of photons. When an electron moves from a higher-energy orbit to a more stable one, energy is emitted in the form of a photon. To move an electron from a stable orbit to a more excited one, a photon of energy must be absorbed. Using the Bohr model, we can calculate the energy of an electron and the radius of its orbit in any one-electron system.

Macroscopic objects act as particles. Microscopic objects (such as electrons) have properties of both a particle and a wave. Their exact trajectories cannot be determined. The quantum mechanical model of atoms describes the three-dimensional position of the electron in a *probabilistic* manner according to a mathematical function called a wavefunction, often denoted as ψ. Atomic wavefunctions are also called orbitals. The squared magnitude of the wavefunction describes the distribution of the probability of finding the electron in a particular region in space. Therefore, atomic orbitals describe the areas in an atom where electrons are most likely to be found.

An atomic orbital is characterized by three quantum numbers. The principal quantum number, n, can be any positive integer. The general region for value of energy of the orbital and the average distance of an electron from the nucleus are related to n. Orbitals having the same value of n are said to be in the same shell. The secondary (angular momentum) quantum number, l, can have any integer value from 0 to $n - 1$. This quantum number describes the shape or type of the orbital. Orbitals with the same principal quantum number and the same l value belong to the same subshell. The magnetic quantum number, m_l, with $2l + 1$ values ranging from $-l$ to $+l$, describes the orientation of the orbital in space. In addition, each electron has a spin quantum number, m_s, that can be equal to No two electrons in the same atom can have the same set of values for all the four quantum numbers.

The relative energy of the subshells determine the order in which atomic orbitals are filled (1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, and so on). Electron configurations and orbital diagrams can be determined by applying the Pauli exclusion principle (no two electrons can have the same set of four quantum numbers) and Hund's rule (whenever possible, electrons retain unpaired spins in degenerate orbitals).

Electrons in the outermost orbitals, called valence electrons, are responsible for most of the chemical behavior of elements. In the periodic table, elements with analogous valence electron configurations usually occur within the same group. There are some exceptions to the predicted filling order, particularly when half-filled or completely filled orbitals can be formed. The periodic table

can be divided into three categories based on the orbital in which the last electron to be added is placed: main group elements (*s* and *p* orbitals), transition elements (*d* orbitals), and inner transition elements (*f* orbitals).

Electron configurations allow us to understand many periodic trends. Covalent radius increases as we move down a group because the *n* level (orbital size) increases. Covalent radius mostly decreases as we move left to right across a period because the effective nuclear charge experienced by the electrons increases, and the electrons are pulled in tighter to the nucleus. Anionic radii are larger than the parent atom, while cationic radii are smaller, because the number of valence electrons has changed while the nuclear charge has remained constant. Ionization energy (the energy associated with forming a cation) decreases down a group and mostly increases across a period because it is easier to remove an electron from a larger, higher energy orbital. Electron affinity (the energy associated with forming an anion) is more favorable (exothermic) when electrons are placed into lower energy orbitals, closer to the nucleus. Therefore, electron affinity becomes increasingly negative as we move left to right across the periodic table and decreases as we move down a group. For both IE and electron affinity data, there are exceptions to the trends when dealing with completely filled or half-filled subshells.

4.10: Summary is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **6.8: Summary** is licensed CC BY 4.0.

4.11: Exercises

1. Arrange K, F, Ba, Pb, B, and I in order of decreasing atomic radius.
2. Arrange Ag, Pt, Mg, C, Cu, and Si in order of increasing atomic radius.
3. Using the periodic table, arrange Li, Ga, Ba, Cl, and Ni in order of increasing atomic radius.
4. Element M is a metal that forms compounds of the type MX_2, MX_3, and MX_4, where X is a halogen. What is the expected trend in the ionic radius of M in these compounds? Arrange these compounds in order of decreasing ionic radius of M.
5. The atomic radii of Na and Cl are 190 and 79 pm, respectively, but the distance between sodium and chlorine in NaCl is 282 pm. Explain this discrepancy.
6. Plot the ionic charge versus ionic radius using the following data for Mo: Mo3+, 69 pm; Mo4+, 65 pm; and Mo5+, 61 pm. Then use this plot to predict the ionic radius of Mo6+. Is the observed trend consistent with the general trends discussed in the chapter? Why or why not?
7. Are shielding effects on the atomic radius more pronounced across a row or down a group? Why?
8. What two factors influence the size of an ion relative to the size of its parent atom? Would you expect the ionic radius of S^{2-} to be the same in both MgS and Na_2S? Why or why not?
9. Arrange Br^-, Al^{3+}, Sr^{2+}, F^-, O^{2-}, and I^- in order of increasing ionic radius.
10. Arrange P^{3-}, N^{3-}, Cl^-, In^{3+}, and S^{2-} in order of decreasing ionic radius.
11. How is an isoelectronic series different from a series of ions with the same charge? Do the cations in magnesium, strontium, and potassium sulfate form an isoelectronic series? Why or why not?
12. What isoelectronic series arises from fluorine, nitrogen, magnesium, and carbon? Arrange the ions in this series by

    a) increasing nuclear charge.

    b) increasing size.

13. What would be the charge and electron configuration of an ion formed from calcium that is isoelectronic with

    a) a chloride ion?

    b) Ar^+?

14. Arrange the gaseous species Mg^{2+}, P^{3-}, Br^{-}, S^{2-}, F^{-}, and N^{3-} in order of increasing radius and justify your decisions.

15. The following table gives values of the first and third ionization energies for selected elements:

Number of Electrons	Element	I_1 ($E \rightarrow E^+ + e^-$, kJ/mol)	Element	I_3 ($E^{2+} \rightarrow E^{3+} + e^-$, kJ/mol)
11	Na	495.9	Al	2744.8
12	Mg	737.8	Si	3231.6
13	Al	577.6	P	2914.1
14	Si	786.6	S	3357
15	P	1011.9	Cl	3822
16	S	999.6	Ar	3931
17	Cl	1251.2	K	4419.6
18	Ar	1520.6	Ca	4912.4

16. Plot the ionization energies versus the number of electrons. Explain why the slopes of the I1 and I3 plots are different, even though the species in each row of the table have the same electron configurations.

17. Would you expect the third ionization energy of iron, corresponding to the removal of an electron from a gaseous Fe2+ ion, to be larger or smaller than the fourth ionization energy, corresponding to the removal of an electron from a gaseous Fe3+ ion? Why? How would these ionization energies compare to the first ionization energy of Ca?

18. Which would you expect to have the highest first ionization energy: Mg, Al, or Si? Which would you expect to have the highest third ionization energy. Why?

19. Use the values of the first ionization energies given in Figure 7.11 to construct plots of first ionization energy versus atomic number for (a) boron through oxygen in the second period; and (b) oxygen through tellurium in group 16. Which plot shows more variation? Explain the reason for the variation in first ionization energies for this group of elements.

20. Arrange Ga, In, and Zn in order of increasing first ionization energies. Would the order be the same for second and third ionization energies? Explain your reasoning.

21. Arrange each set of elements in order of increasing magnitude of electron affinity.

a) Pb, Bi, and Te

b) Na, K, and Rb

c) P, C, and Ge

22. Arrange each set of elements in order of decreasing magnitude of electron affinity.

a) As, Bi, and N

b) O, F, and Ar

c) Cs, Ba, and Rb

23. Of the species F, O^-, Al^{3+}, and Li^+, which has the highest electron affinity? Explain your reasoning.

24. Of the species O^-, N^{2-}, Hg^{2+}, and H^+, which has the highest electron affinity? Which has the lowest electron affinity? Justify your answers.

25. Based on their valence electron configurations, classify the following elements as either electrical insulators, electrical conductors, or substances with intermediate conductivity: S, Ba, Fe, Al, Te, Be, O, C, P, Sc, W, Na, B, and Rb.

26. Using the data in Problem 10, what conclusions can you draw with regard to the relationship between electronegativity and electrical properties? Estimate the approximate electronegativity of a pure element that is very dense, lustrous, and malleable.

27. Of the elements Al, Mg, O_2, Ti, I_2, and H_2, which, if any, would you expect to be a good reductant? Explain your reasoning.

28. Of the elements Zn, B, Li, Se, Co, and Br_2, which if any, would you expect to be a good oxidant? Explain your reasoning.

29. Determine whether each species is a good oxidant, a good reductant, or neither.

a) Ba

b) Mo

c) Al

d) Ni

e) O_2

f) Xe

30. Determine whether each species is a good oxidant, a good reductant, or neither.

a) Ir

b) Cs

c) Be

d) B

e) N

f) Po

g) Ne

31. Of the species I_2, O^-, Zn, Sn^{2+}, and K^+, choose which you would expect to be a good oxidant. Then justify your answer.

32. Based on the valence electron configuration of the noble gases, would you expect them to have positive or negative electron affinities? What does this imply about their most likely oxidation states? their reactivity?

Q6.5.1

Based on their positions in the periodic table, predict which has the smallest atomic radius: Mg, Sr, Si, Cl, I.

S6.5.1

Cl

Q6.5.2

Based on their positions in the periodic table, predict which has the largest atomic radius: Li, Rb, N, F, I.

Q6.5.3

Based on their positions in the periodic table, predict which has the largest first ionization energy: Mg, Ba, B, O, Te.

S6.5.3

O

Q6.5.4

Based on their positions in the periodic table, predict which has the smallest first ionization energy: Li, Cs, N, F, I.

Q6.5.5

Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: F, Li, N, Rb

S6.5.5

Rb < Li < N < F

Q6.5.6

Based on their positions in the periodic table, rank the following atoms or compounds in order of increasing first ionization energy: Mg, O, S, Si

Q6.5.7

Atoms of which group in the periodic table have a valence shell electron configuration of ns^2np^3?

S6.5.7

15 (5A)

Q6.5.8

Atoms of which group in the periodic table have a valence shell electron configuration of ns^2?

Q6.5.9

Based on their positions in the periodic table, list the following atoms in order of increasing radius: Mg, Ca, Rb, Cs.

S6.5.9

Mg < Ca < Rb < Cs

Q6.5.10

Based on their positions in the periodic table, list the following atoms in order of increasing radius: Sr, Ca, Si, Cl.

Q6.5.11

Based on their positions in the periodic table, list the following ions in order of increasing radius: K^+, Ca^{2+}, Al^{3+}, Si^{4+}.

S6.5.11

$Si^{4+} < Al^{3+} < Ca^{2+} < K^+$

Q6.5.12

List the following ions in order of increasing radius: Li^+, Mg^{2+}, Br^-, Te^{2-}.

Q6.5.13

Which atom and/or ion is (are) isoelectronic with Br^+: Se^{2+}, Se, As^-, Kr, Ga^{3+}, Cl^-?

S6.5.13

Se, As^{-}

Q6.5.14

Which of the following atoms and ions is (are) isoelectronic with S^{2+}: Si^{4+}, Cl^{3+}, Ar, As^{3+}, Si, Al^{3+}?

Q6.5.15

Compare both the numbers of protons and electrons present in each to rank the following ions in order of increasing radius: As^{3-}, Br^{-}, K^{+}, Mg^{2+}.

S6.5.15

$Mg^{2+} < K^{+} < Br^{-} < As^{3-}$

Q6.5.16

Of the five elements Al, Cl, I, Na, Rb, which has the most exothermic reaction? (E represents an atom.) What name is given to the energy for the reaction? Hint: note the process depicted does *not* correspond to electron affinityE+(g)+e−⟶E(g)" id="MathJax-Element-160-Frame" role="presentation" style="position:relative;" tabindex="0">E+(*g*)+e−⟶E(*g*)

Q6.5.17

Of the five elements Sn, Si, Sb, O, Te, which has the most endothermic reaction? (E represents an atom.) What name is given to the energy for the reaction?

(6.14)E(g)⟶E+(g)+e−" id="MathJax-Element-161-Frame" role="presentation" style="position:relative;" tabindex="0">(*g*)⟶E+(*g*)+e−(6.14

S6.5.17

O, IE_1

Q6.5.18

The ionic radii of the ions S^{2-}, Cl^{-}, and K^{+} are 184, 181, 138 pm respectively. Explain why these ions have different sizes even though they contain the same number of electrons.

Q6.5.19

Which main group atom would be expected to have the lowest second ionization energy?

S6.5.19

Ra

This page titled 4.11: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by OpenStax via source content that was edited to the style and standards of the LibreTexts platform.

CHAPTER OVERVIEW

5: Chemical Bonding and Molecular Geometry

A chemical bond is an attraction between atoms that allows the formation of chemical substances that contain two or more atoms. The bond is caused by the electrostatic force of attraction between opposite charges, either between electrons and nuclei, or as the result of a dipole attraction. All bonds can be explained by quantum theory, but, in practice, simplification rules allow chemists to predict the strength, directionality, and polarity of bonds. The octet rule and VSEPR theory are two examples. More sophisticated theories are valence bond theory which includes orbital hybridization and resonance, and the linear combination of atomic orbitals molecular orbital method. Electrostatics are used to describe bond polarities and the effects they have on chemical substances.

Faculty Feature

Dr. Xiao-an Zhang's research focuses on how molecular geometry and chemical bonding affect nanomaterial characteristics. His research on the structural and electrical characteristics of these materials provides important insights into how atomic arrangements influence molecular behaviour. Zhang's research on nanotechnology emphasizes how chemical bonds, such as covalent, ionic, and metallic, influence the stability and reactivity of nanomaterials. This understanding of bonding and geometry is critical for developing more efficient catalysts, sensors, and drug delivery devices. His research combines fundamental chemistry and nanotechnology, driving advances in materials science.

Read more about Dr. Xiao-an Zhang and his research here: https://www.utsc.utoronto.ca/people/xazhang/



5.1: Introduction
5.2: Ionic Bonding
5.3: Covalent Bonding
5.4: Chemical Nomenclature
5.5: Lewis Symbols and Structures
5.6: Formal Charges and Resonance
5.7: Molecular Structure and Polarity
5.8: Key Terms
5.9: Key Equations
5.10: Summary
5.11: Exercises

This page titled 5: Chemical Bonding and Molecular Geometry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

5.1: Introduction

It has long been known that pure carbon occurs in different forms (allotropes) including graphite and diamonds. But it was not until 1985 that a new form of carbon was recognized: buckminsterfullerene. This molecule was named after the architect and inventor R. Buckminster Fuller (1895–1983), whose signature architectural design was the geodesic dome, characterized by a lattice shell structure supporting a spherical surface. Experimental evidence revealed the formula, C_{60}, and then scientists determined how 60 carbon atoms could form one symmetric, stable molecule. They were guided by bonding theory—the topic of this chapter—which explains how individual atoms connect to form more complex structures.



Figure 5.1.1: Nicknamed "buckyballs," buckminsterfullerene molecules (C_{60}) contain only carbon atoms (left) arranged to form a geometric framework of hexagons and pentagons, similar to the pattern on a soccer ball (center). This molecular structure is named after architect R. Buckminster Fuller, whose innovative designs combined simple geometric shapes to create large, strong structures such as this weather radar dome near Tucson, Arizona (right). (credit middle: modification of work by "Petey21"/Wikimedia Commons; credit right: modification of work by Bill Morrow)

Faculty Feature

Dr. Xiao-an Zhang's research focuses on how molecular geometry and chemical bonding affect nanomaterial characteristics. His research on the structural and electrical characteristics of these materials provides important insights into how atomic arrangements influence molecular behaviour. Zhang's research on nanotechnology emphasizes how chemical bonds, such as covalent, ionic, and metallic, influence the stability and reactivity of nanomaterials. This understanding of bonding and geometry is critical for developing more efficient catalysts, sensors, and drug delivery devices. His research combines fundamental chemistry and nanotechnology, driving advances in materials science.

Read more about Dr. Xiao-an Zhang and his research here: https://www.utsc.utoronto.ca/people/xazhang/



This page titled 5.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **7.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.2: Ionic Bonding

Learning Objectives

- To describe the characteristics of ionic bonding.
- To quantitatively describe the energetic factors involved in the formation of an ionic bond.

Ions are atoms or molecules which are electrically charged. **Cations** are positively charged and **anions** carry a negative charge. Ions form when atoms gain or lose electrons. Since electrons are negatively charged, an atom that loses one or more electrons will become positively charged; an atom that gains one or more electrons becomes negatively charged. **Ionic bonding** is the attraction between positively- and negatively-charged ions. These oppositely charged ions attract each other to form ionic networks (or **lattices**). Electrostatics explains why this happens: opposite charges attract and like charges repel. When many ions attract each other, they form large, ordered, crystal lattices in which each ion is surrounded by ions of the opposite charge. Generally, when metals react with non-metals, electrons are transferred from the metals to the non-metals. The metals form positively-charged ions and the non-metals form negatively-charged ions.

Generating Ionic Bonds

Ionic bonds form when metals and non-metals chemically react. By definition, a metal is relatively stable if it loses electrons to form a complete valence shell and becomes positively charged. Likewise, a non-metal becomes stable by gaining electrons to complete its valence shell and become negatively charged. When metals and non-metals react, the metals lose electrons by transferring them to the non-metals, which gain them. Consequently, ions are formed, which instantly attract each other—ionic bonding.

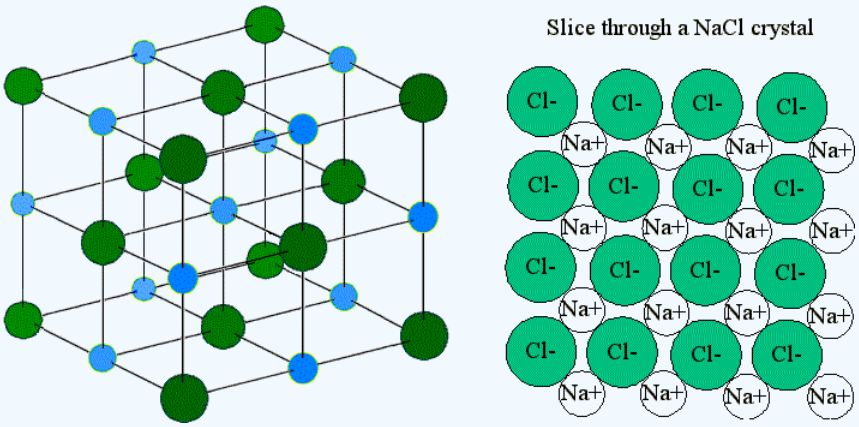
In the overall ionic compound, positive and negative charges must be balanced, because electrons cannot be created or destroyed, only transferred. Thus, the total number of electrons lost by the cationic species must equal the total number of electrons gained by the anionic species.

Example 5.2.1: Sodium Chloride

For example, in the reaction of Na (sodium) and Cl (chlorine), each Cl atom takes one electron from a Na atom. Therefore each Na becomes a Na^+ cation and each Cl atom becomes a Cl^- anion. Due to their opposite charges, they attract each other to form an ionic lattice. The formula (ratio of positive to negative ions) in the lattice is $NaCl$.

$$2\,Na(s) + Cl2\,(g) \rightarrow 2\,NaCl(s)$$

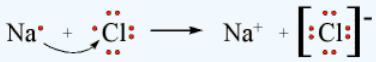
These ions are arranged in solid NaCl in a regular three-dimensional arrangement (or lattice):



NaCl lattice. (left) 3-D structure and (right) simple 2D slice through lattes. Images used with permission from Wikipedia and Mike Blaber.

The chlorine has a high affinity for electrons, and the sodium has a low ionization energy. Thus the chlorine gains an electron from the sodium atom. This can be represented using ewis dot symbols (here we will consider one chlorine atom, rather than

Cl_2):



Chlorine becomes more negative when it takes the one valence electron from sodium when it makes an octet.

The arrow indicates the transfer of the electron from sodium to chlorine to form the Na^+ metal ion and the Cl^- chloride ion. Each ion now has an **octet** of electrons in its valence shell:

- Na^+: $2s^2 2p^6$
- Cl^-: $3s^2 3p^6$

Energetics of Ionic Bond Formation

Ionic bonds are formed when positively and negatively charged ions are held together by electrostatic forces. Consider a single pair of ions, one cation and one anion. How strong will the force of their attraction be? According to **Coulomb's Law**, the energy of the electrostatic attraction (E) between two charged particles is proportional to the magnitude of the charges and inversely proportional to the internuclear distance between the particles (r):

$$E \propto \frac{Q_1 Q_2}{r} \tag{5.2.1}$$

$$E = k\frac{Q_1 Q_2}{r} \tag{5.2.2}$$

where each ion's charge is represented by the symbol Q. The proportionality constant k is equal to 2.31×10^{-28} J·m. This value of k includes the charge of a single electron (1.6022×10^{-19} C) for each ion. The equation can also be written using the charge of each ion, expressed in coulombs (C), incorporated in the constant. In this case, the proportionality constant, k, equals 8.999 × 109 J·m/C^2. In the example given, $Q_1 = +1(1.6022 \times 10^{-19}$ C) and $Q_2 = -1(1.6022 \times 10^{-19}$ C). If Q_1 and Q_2 have opposite signs (as in NaCl, for example, where Q_1 is +1 for Na^+ and Q_2 is −1 for Cl^-), then E is negative, which means that energy is *released* when oppositely charged ions are brought together from an infinite distance to form an isolated ion pair.

> *Energy is always released when a bond is formed and correspondingly, it always requires energy to break a bond.*

As shown by the green curve in the lower half of Figure 5.2.1, the maximum energy would be released when the ions are infinitely close to each other, at $r = 0$. Because ions occupy space and have a structure with the positive nucleus being surrounded by electrons, however, they cannot be infinitely close together. At very short distances, repulsive electron–electron interactions between electrons on adjacent ions become stronger than the attractive interactions between ions with opposite charges, as shown by the red curve in the upper half of Figure 5.2.1. The total energy of the system is a balance between the attractive and repulsive interactions. The purple curve in Figure 5.2.1 shows that the total energy of the system reaches a minimum at r_0, the point where the electrostatic repulsions and attractions are exactly balanced. This distance is the same as the experimentally measured **bond distance**.

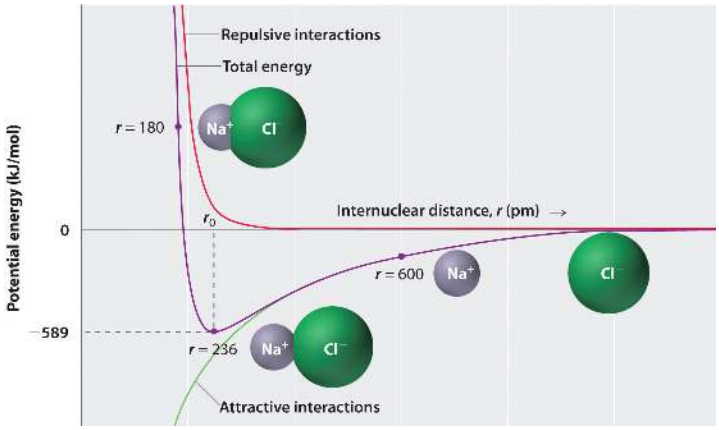


Figure 5.2.1: A Plot of Potential Energy versus Internuclear Distance for the Interaction between a Gaseous Na^+ Ion and a Gaseous Cl^- Ion. The energy of the system reaches a minimum at a particular distance (r_0) when the attractive and repulsive interactions are balanced.

Consider the energy released when a gaseous Na^+ ion and a gaseous Cl^- ion are brought together from $r = \infty$ to $r = r_0$. Given that the observed gas-phase internuclear distance is 236 pm, the energy change associated with the formation of an ion pair from an $Na^+_{(g)}$ ion and a $Cl^-_{(g)}$ ion is as follows:

$$E = k\frac{Q_1 Q_2}{r_0}$$

$$= (2.31 \times 10^{-28}\text{J} \cdot \cancel{\text{m}}) \left(\frac{(+1)(-1)}{236 \ \cancel{\text{pm}} \times 10^{-12} \ \text{m}/\cancel{\text{pm}}} \right)$$

$$= -9.79 \times 10^{-19} \ J/ion \ pair$$

The negative value indicates that energy is released. Our convention is that if a chemical process provides energy to the outside world, the energy change is negative. If it requires energy, the energy change is positive. To calculate the energy change in the formation of a mole of NaCl pairs, we need to multiply the energy per ion pair by Avogadro's number:

$$E = \left(-9.79 \times 10^{-19} \ J/ \ \cancel{ionpair}\right) \left(6.022 \times 10^{23} \ \ \cancel{ion \ pair}/mol\right) = -589 \ kJ/mol \tag{5.2.3}$$

This is the energy released when 1 mol of gaseous ion pairs is formed, *not* when 1 mol of positive and negative ions condenses to form a crystalline lattice. Because of long-range interactions in the lattice structure, this energy does not correspond directly to the **lattice energy** of the crystalline solid. However, the large negative value indicates that bringing positive and negative ions together is energetically very favorable, whether an ion pair or a crystalline lattice is formed.

We summarize the important points about ionic bonding:

- At r_0, the ions are more stable (have a lower potential energy) than they are at an infinite internuclear distance. When oppositely charged ions are brought together from $r = \infty$ to $r = r_0$, the energy of the system is lowered (energy is released).
- Because of the low potential energy at r_0, energy must be added to the system to separate the ions. The amount of energy needed is the bond energy.
- The energy of the system reaches a minimum at a particular internuclear distance (the bond distance).

✔ Example 5.2.2: LiF

Calculate the amount of energy released when 1 mol of gaseous Li^+F^- ion pairs is formed from the separated ions. The observed internuclear distance in the gas phase is 156 pm.

Given: cation and anion, amount, and internuclear distance

Asked for: energy released from formation of gaseous ion pairs

Strategy:

Substitute the appropriate values into Equation 5.2.2 to obtain the energy released in the formation of a single ion pair and then multiply this value by Avogadro's number to obtain the energy released per mole.

Solution:

Inserting the values for Li^+F^- into Equation 5.2.2 (where Q_1 = +1, Q_2 = −1, and r = 156 pm), we find that the energy associated with the formation of a single pair of Li^+F^- ions is

$$E = k\frac{Q_1 Q_2}{r_0}$$

$$= \left(2.31 \times 10^{-28} J \cdot \cancel{m}\right)\left(\frac{(+1)(-1)}{156\ pm \times 10^{-12}\ m/\cancel{pm}}\right)$$

$$= -1.48 \times 10^{-18}$$

Then the energy released per mole of Li^+F^- ion pairs is

$$E = \left(-1.48 \times 10^{-18} J/\ \text{ion}\ \cancel{\text{pair}}\right)\left(6.022 \times 10^{23}\ \text{ion}\ \cancel{\text{pair}}/mol\right)$$

$$-891\ kJ/mol$$

Because Li^+ and F^- are smaller than Na^+ and Cl^- (see Section 7.3), the internuclear distance in LiF is shorter than in NaCl. Consequently, in accordance with Equation 5.2.2, much more energy is released when 1 mol of gaseous Li^+F^- ion pairs is formed (−891 kJ/mol) than when 1 mol of gaseous Na^+Cl^- ion pairs is formed (−589 kJ/mol).

? Exercise 5.2.2: Magnesium oxide

Calculate the amount of energy released when 1 mol of gaseous MgO ion pairs is formed from the separated ions. The internuclear distance in the gas phase is 175 pm.

Answer

−3180 kJ/mol = -3.18×10^3 kJ/mol

Electron Configuration of Ions

How does the energy released in lattice formation compare to the energy required to strip away a second electron from the Na^+ ion? Since the Na^+ ion has a noble gas electron configuration, stripping away the next electron from this stable arrangement would require more energy than what is released during lattice formation (Sodium I_2 = 4,560 kJ/mol). Thus, sodium is present in ionic compounds as Na^+ and not Na^{2+}. Likewise, adding an electron to fill a valence shell (and achieve noble gas electron configuration) is exothermic or only slightly endothermic. To add an additional electron into a ***new*** subshell requires tremendous energy - more than the lattice energy. Thus, we find Cl^- in ionic compounds, but not Cl^{2-}.

Table 5.2.1: Lattice energies range from around 700 kJ/mol to 4000 kJ/mol:

Compound	Lattice Energy (kJ/mol)
LiF	1024
LiI	744
NaF	911
NaCl	788
NaI	693
KF	815
KBr	682
KI	641

Compound	Lattice Energy (kJ/mol)
MgF_2	2910
$SrCl_2$	2130
MgO	3938

This amount of energy can compensate for values as large as I_3 for valence electrons (i.e. can strip away up to 3 valence electrons). Because most transition metals would require the removal of more than 3 electrons to attain a noble gas core, they are not found in ionic compounds with a noble gas core. A transition metal always loses electrons first from the higher 's' subshell, before losing from the underlying 'd' subshell. (The remaining electrons in the unfilled d subshell are the reason for the bright colors observed in many transition metal compounds!) For example, iron ions will **not** form a noble gas core:

- Fe: [Ar]$4s^2 3d^6$
- Fe^{2+}: [Ar] $3d^6$
- Fe^{3+}: [Ar] $3d^5$

Some metal ions can form a pseudo noble gas core (and be colorless), for example:

- Ag: [Kr]$5s^1 4d^{10}$ Ag^+ [Kr]$4d^{10}$ Compound: AgCl
- Cd: [Kr]$5s^2 4d^{10}$ Cd^{2+} [Kr]$4d^{10}$ Compound: CdS

The valence electrons do not adhere to the "octet rule" in this case (a limitation of the usefulness of this rule). Note: The silver and cadmium atoms lost the 5s electrons in achieving the ionic state.

> *When a positive ion is formed from an atom, electrons are **always** lost first from the subshell with the largest principle quantum number*

Polyatomic Ions

Not all ionic compounds are formed from only two elements. Many **polyatomic ions** exist, in which two or more atoms are bound together by covalent bonds. They form a stable grouping which carries a charge (positive or negative). The group of atoms as a whole acts as a charged species in forming an ionic compound with an oppositely charged ion. Polyatomic ions may be either positive or negative, for example:

- NH_4^+ (ammonium) = cation
- SO_4^{2-} (sulfate) = anion

The principles of ionic bonding with polyatomic ions are the same as those with monatomic ions. Oppositely charged ions come together to form a crystalline lattice, releasing a lattice energy. Based on the shapes and charges of the polyatomic ions, these compounds may form crystalline lattices with interesting and complex structures.

Summary

The amount of energy needed to separate a gaseous ion pair is its bond energy. The formation of ionic compounds are usually ***extremely exothermic***. The strength of the electrostatic attraction between ions with opposite charges is directly proportional to the magnitude of the charges on the ions and inversely proportional to the internuclear distance. The total energy of the system is a balance between the repulsive interactions between electrons on adjacent ions and the attractive interactions between ions with opposite charges.

Contributors and Attributions

- Wikibooks [en.wikibooks.org]
- Mike Blaber (Florida State University)
- Modified by Joshua Halpern (Howard University)

5.2: Ionic Bonding is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **8.2: Ionic Bonding** is licensed CC BY-NC-SA 3.0.

5.3: Covalent Bonding

Learning Objectives

By the end of this section, you will be able to:

- Describe the formation of covalent bonds
- Define electronegativity and assess the polarity of covalent bonds

Ionic bonding results from the electrostatic attraction of oppositely charged ions that are typically produced by the transfer of electrons between metallic and nonmetallic atoms. A different type of bonding results from the mutual attraction of atoms for a "shared" pair of electrons. Such bonds are called covalent bonds. Covalent bonds are formed between two atoms when both have similar tendencies to attract electrons to themselves (i.e., when both atoms have identical or fairly similar ionization energies and electron affinities). For example, two hydrogen atoms bond covalently to form an H_2 molecule; each hydrogen atom in the H_2 molecule has two electrons stabilizing it, giving each atom the same number of valence electrons as the noble gas He.

Compounds that contain covalent bonds exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds. In fact, many covalent compounds are liquids or gases at room temperature, and, in their solid states, they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good conductors of electricity when dissolved in water, most covalent compounds are insoluble in water; since they are electrically neutral, they are poor conductors of electricity in any state.

Formation of Covalent Bonds

Nonmetal atoms frequently form covalent bonds with other nonmetal atoms. For example, the hydrogen molecule, H_2, contains a covalent bond between its two hydrogen atoms. Figure $5.3.1$ illustrates why this bond is formed. Starting on the far right, we have two separate hydrogen atoms with a particular potential energy, indicated by the red line. Along the *x*-axis is the distance between the two atoms. As the two atoms approach each other (moving left along the *x*-axis), their valence orbitals (1*s*) begin to overlap. The single electrons on each hydrogen atom then interact with both atomic nuclei, occupying the space around both atoms. The strong attraction of each shared electron to both nuclei stabilizes the system, and the potential energy decreases as the bond distance decreases. If the atoms continue to approach each other, the positive charges in the two nuclei begin to repel each other, and the potential energy increases. The bond length is determined by the distance at which the lowest potential energy is achieved.

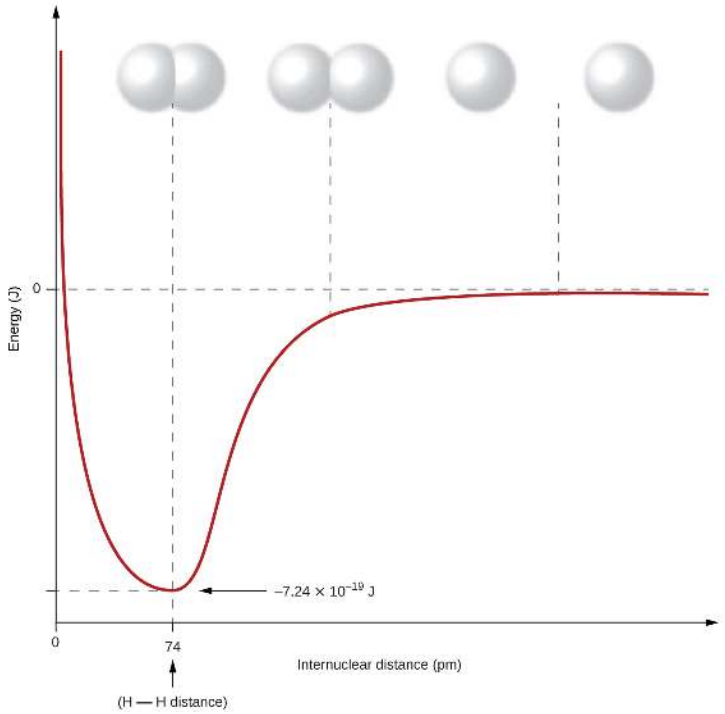


Figure 5.3.1: The potential energy of two separate hydrogen atoms (right) decreases as they approach each other, and the single electrons on each atom are shared to form a covalent bond. The bond length is the internuclear distance at which the lowest potential energy is achieved.

It is essential to remember that energy must be added to break chemical bonds (an endothermic process), whereas forming chemical bonds releases energy (an exothermic process). In the case of H_2, the covalent bond is very strong; a large amount of energy, 436 kJ, must be added to break the bonds in one mole of hydrogen molecules and cause the atoms to separate:

$$\mathrm{H_2(g)} \longrightarrow 2\,\mathrm{H(g)} \quad \Delta H = 436\,\mathrm{kJ}$$

Conversely, the same amount of energy is released when one mole of H_2 molecules forms from two moles of H atoms:

$$2\,\mathrm{H(g)} \longrightarrow \mathrm{H_2(g)} \quad \Delta H = -436\,\mathrm{kJ}$$

Pure vs. Polar Covalent Bonds

If the atoms that form a covalent bond are identical, as in H_2, Cl_2, and other diatomic molecules, then the electrons in the bond must be shared equally. We refer to this as a pure covalent bond. Electrons shared in pure covalent bonds have an equal probability of being near each nucleus.

In the case of Cl_2, each atom starts off with seven valence electrons, and each Cl shares one electron with the other, forming one covalent bond:

$$\mathrm{Cl + Cl} \longrightarrow \mathrm{Cl_2}$$

The total number of electrons around each individual atom consists of six nonbonding electrons and two shared (i.e., bonding) electrons for eight total electrons, matching the number of valence electrons in the noble gas argon. Since the bonding atoms are identical, Cl_2 also features a pure covalent bond.

When the atoms linked by a covalent bond are different, the bonding electrons are shared, but no longer equally. Instead, the bonding electrons are more attracted to one atom than the other, giving rise to a shift of electron density toward that atom. This unequal distribution of electrons is known as a polar covalent bond, characterized by a partial positive charge on one atom and a partial negative charge on the other. The atom that attracts the electrons more strongly acquires the partial negative charge and vice versa. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom. Thus, in an HCl molecule, the chlorine atom carries a partial negative charge and the hydrogen atom has a partial positive charge. Figure 5.3.2 shows the distribution of electrons in the H–Cl bond. Note that the shaded area around Cl is

much larger than it is around H. Compare this to Figure 5.3.1, which shows the even distribution of electrons in the H_2 nonpolar bond.

We sometimes designate the positive and negative atoms in a polar covalent bond using a lowercase Greek letter "delta," δ, with a plus sign or minus sign to indicate whether the atom has a partial positive charge (δ+) or a partial negative charge (δ–). This symbolism is shown for the H–Cl molecule in Figure 5.3.2.

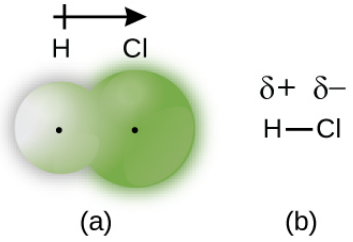


Figure 5.3.2: (a) The distribution of electron density in the HCl molecule is uneven. The electron density is greater around the chlorine nucleus. The small, black dots indicate the location of the hydrogen and chlorine nuclei in the molecule. (b) Symbols δ+ and δ– indicate the polarity of the H–Cl bond.

Electronegativity

Whether a bond is nonpolar or polar covalent is determined by a property of the bonding atoms called electronegativity. Electronegativity is a measure of the tendency of an atom to attract electrons (or electron density) towards itself. It determines how the shared electrons are distributed between the two atoms in a bond. The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a polar covalent bond are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarized the electron distribution and the larger the partial charges of the atoms.

Figure 5.3.3 shows the electronegativity values of the elements as proposed by one of the most famous chemists of the twentieth century: Linus Pauling (Figure 5.3.4). In general, electronegativity increases from left to right across a period in the periodic table and decreases down a group. Thus, the nonmetals, which lie in the upper right, tend to have the highest electronegativities, with fluorine the most electronegative element of all (EN = 4.0). Metals tend to be less electronegative elements, and the group 1 metals have the lowest electronegativities. Note that noble gases are excluded from this figure because these atoms usually do not share electrons with others atoms since they have a full valence shell. (While noble gas compounds such as XeO_2 do exist, they can only be formed under extreme conditions, and thus they do not fit neatly into the general model of electronegativity.)

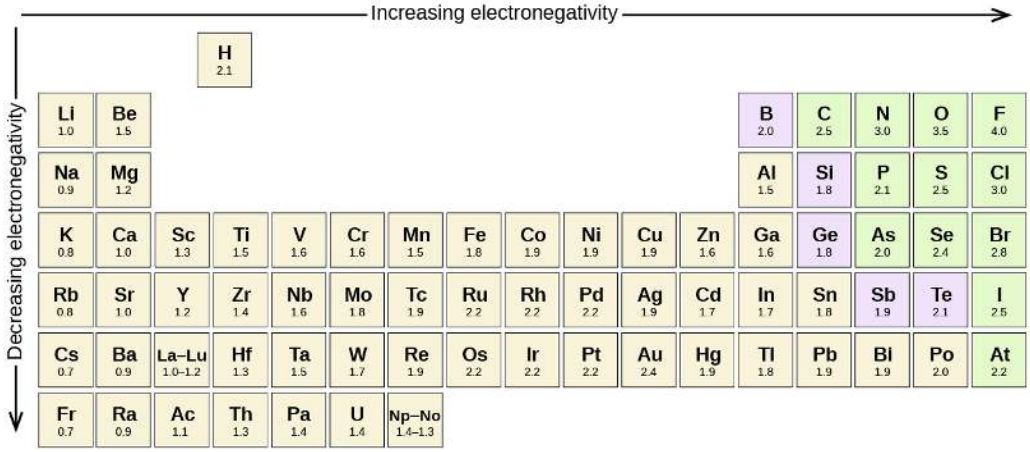


Figure 5.3.3: The electronegativity values derived by Pauling follow predictable periodic trends, with the higher electronegativities toward the upper right of the periodic table.

Electronegativity versus Electron Affinity

We must be careful not to confuse electronegativity and electron affinity. The electron affinity of an element is a measurable physical quantity, namely, the energy released or absorbed when an isolated gas-phase atom acquires an electron, measured in kJ/mol. Electronegativity, on the other hand, describes how tightly an atom attracts electrons in a bond. It is a dimensionless

quantity that is calculated, not measured. Pauling derived the first electronegativity values by comparing the amounts of energy required to break different types of bonds. He chose an arbitrary relative scale ranging from 0 to 4.

Portrait of a Chemist: Linus Pauling

Linus Pauling, shown in Figure 5.3.4, is the only person to have received two unshared (individual) Nobel Prizes: one for chemistry in 1954 for his work on the nature of chemical bonds and one for peace in 1962 for his opposition to weapons of mass destruction. He developed many of the theories and concepts that are foundational to our current understanding of chemistry, including electronegativity and resonance structures.

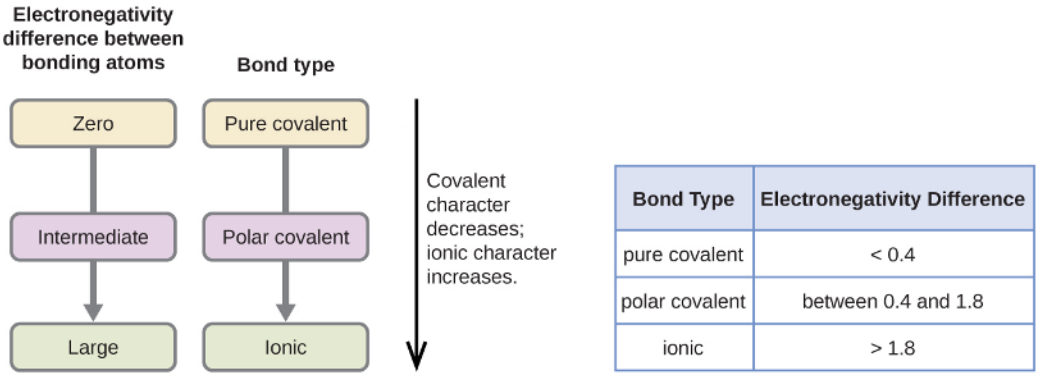


Figure 5.3.4: Linus Pauling (1901–1994) made many important contributions to the field of chemistry. He was also a prominent activist, publicizing issues related to health and nuclear weapons.

Pauling also contributed to many other fields besides chemistry. His research on sickle cell anemia revealed the cause of the disease—the presence of a genetically inherited abnormal protein in the blood—and paved the way for the field of molecular genetics. His work was also pivotal in curbing the testing of nuclear weapons; he proved that radioactive fallout from nuclear testing posed a public health risk.

Electronegativity and Bond Type

The absolute value of the difference in electronegativity (ΔEN) of two bonded atoms provides a rough measure of the polarity to be expected in the bond and, thus, the bond type. When the difference is very small or zero, the bond is covalent and nonpolar. When it is large, the bond is polar covalent or ionic. The absolute values of the electronegativity differences between the atoms in the bonds H–H, H–Cl, and Na–Cl are 0 (nonpolar), 0.9 (polar covalent), and 2.1 (ionic), respectively. The degree to which electrons are shared between atoms varies from completely equal (pure covalent bonding) to not at all (ionic bonding). Figure 5.3.5 shows the relationship between electronegativity difference and bond type.



Bond Type	Electronegativity Difference
pure covalent	< 0.4
polar covalent	between 0.4 and 1.8
ionic	> 1.8

Figure 5.3.5: As the electronegativity difference increases between two atoms, the bond becomes more ionic.

A rough approximation of the electronegativity differences associated with covalent, polar covalent, and ionic bonds is shown in Figure 5.3.5. This table is just a general guide, however, with many exceptions. For example, the H and F atoms in HF have an electronegativity difference of 1.9, and the N and H atoms in NH_3 a difference of 0.9, yet both of these compounds form bonds that are considered polar covalent. Likewise, the Na and Cl atoms in NaCl have an electronegativity difference of 2.1, and the Mn and I atoms in MnI_2 have a difference of 1.0, yet both of these substances form ionic compounds.

The best guide to the covalent or ionic character of a bond is to consider the types of atoms involved and their relative positions in the periodic table. Bonds between two nonmetals are generally covalent; bonding between a metal and a nonmetal is often ionic.

Some compounds contain both covalent and ionic bonds. The atoms in polyatomic ions, such as OH^-, and are held together by polar covalent bonds. However, these polyatomic ions form ionic compounds by combining with ions of opposite charge. For example, potassium nitrate, KNO_3, contains the K^+ cation and the polyatomic anion. Thus, bonding in potassium nitrate is ionic, resulting from the electrostatic attraction between the ions K^+ and as well as covalent between the nitrogen and oxygen atoms in

Example 5.3.1: Electronegativity and Bond Polarity

Bond polarities play an important role in determining the structure of proteins. Using the electronegativity values in Figure 5.3.3, arrange the following covalent bonds—all commonly found in amino acids—in order of increasing polarity. Then designate the positive and negative atoms using the symbols δ+ and δ–:

C–H, C–N, C–O, N–H, O–H, S–H

Solution

The polarity of these bonds increases as the absolute value of the electronegativity difference increases. The atom with the δ– designation is the more electronegative of the two. Table 5.3.1 shows these bonds in order of increasing polarity.

Table 5.3.1: Bond Polarity and Electronegativity Difference

Bond	ΔEN	Polarity
C–H	0.4	
S–H	0.4	
C–N	0.5	
N–H	0.9	
C–O	1.0	
O–H	1.4	

Exercise 5.3.1

Silicones are polymeric compounds containing, among others, the following types of covalent bonds: Si–O, Si–C, C–H, and C–C. Using the electronegativity values in Figure 5.3.3, arrange the bonds in order of increasing polarity and designate the positive and negative atoms using the symbols δ+ and δ–.

Answer

Bond	Electronegativity Difference	Polarity
C–C	0.0	nonpolar
C–H	0.4	
Si–C	0.7	
Si–O	1.7	

This page titled 5.3: Covalent Bonding is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **7.2: Covalent Bonding** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.4: Chemical Nomenclature

Learning Objectives

By the end of this section, you will be able to:

- Derive names for common types of inorganic compounds using a systematic approach

Nomenclature, a collection of rules for naming things, is important in science and in many other situations. This module describes an approach that is used to name simple ionic and molecular compounds, such as NaCl, $CaCO_3$, and N_2O_4. The simplest of these are **binary compounds**, those containing only two elements, but we will also consider how to name ionic compounds containing polyatomic ions, and one specific, very important class of compounds known as acids (subsequent chapters in this text will focus on these compounds in great detail). We will limit our attention here to inorganic compounds, compounds that are composed principally of elements other than carbon, and will follow the nomenclature guidelines proposed by IUPAC. The rules for organic compounds, in which carbon is the principle element, will be treated in a later chapter on organic chemistry.

Ionic Compounds

To name an inorganic compound, we need to consider the answers to several questions. First, is the compound ionic or molecular? If the compound is ionic, does the metal form ions of only one type (fixed charge) or more than one type (variable charge)? Are the ions monatomic or polyatomic? If the compound is molecular, does it contain hydrogen? If so, does it also contain oxygen? From the answers we derive, we place the compound in an appropriate category and then name it accordingly.

Compounds Containing Only Monatomic Ions

The name of a binary compound containing monatomic ions consists of the name of the cation (the name of the metal) followed by the name of the anion (the name of the nonmetallic element with its ending replaced by the suffix *–ide*). Some examples are given in Table 5.4.1.

Table 5.4.1: Names of Some Ionic Compounds

NaCl, sodium chloride	Na_2O, sodium oxide
KBr, potassium bromide	CdS, cadmium sulfide
CaI_2, calcium iodide	Mg_3N_2, magnesium nitride
CsF, cesium fluoride	Ca_3P_2, calcium phosphide
LiCl, lithium chloride	Al_4C_3, aluminum carbide

Compounds Containing Polyatomic Ions

Compounds containing polyatomic ions are named similarly to those containing only monatomic ions, i.e. by naming first the cation and then the anion. Examples are shown in Table 5.4.2.

Table 5.4.2: Names of Some Polyatomic Ionic Compounds

$KC_2H_3O_2$, potassium acetate	NH_4Cl, ammonium chloride
$NaHCO_3$, sodium bicarbonate	$CaSO_4$, calcium sulfate
$Al_2(CO_3)_3$, aluminum carbonate	$Mg_3(PO_4)_2$, magnesium phosphate

Chemistry in Everyday Life: Ionic Compounds in Your Cabinet

Every day you encounter and use a large number of ionic compounds. Some of these compounds, where they are found, and what they are used for are listed in Table 5.4.3. Look at the label or ingredients list on the various products that you use during the next few days, and see if you run into any of those in this table, or find other ionic compounds that you could now name or write as a formula.

Table 5.4.3: Everyday Ionic Compounds

Ionic Compound	Use
NaCl, sodium chloride	ordinary table salt
KI, potassium iodide	added to “iodized” salt for thyroid health
NaF, sodium fluoride	ingredient in toothpaste
$NaHCO_3$, sodium bicarbonate	baking soda; used in cooking (and as antacid)
Na_2CO_3, sodium carbonate	washing soda; used in cleaning agents
NaOCl, sodium hypochlorite	active ingredient in household bleach
$CaCO_3$ calcium carbonate	ingredient in antacids
$Mg(OH)_2$, magnesium hydroxide	ingredient in antacids
$Al(OH)_3$, aluminum hydroxide	ingredient in antacids
NaOH, sodium hydroxide	lye; used as drain cleaner
K_3PO_4, potassium phosphate	food additive (many purposes)
$MgSO_4$, magnesium sulfate	added to purified water
Na_2HPO_4, sodium hydrogen phosphate	anti-caking agent; used in powdered products
Na_2SO_3, sodium sulfite	preservative

Compounds Containing a Metal Ion with a Variable Charge

Most of the transition metals and some main group metals can form two or more cations with different charges. Compounds of these metals with nonmetals are named with the same method as compounds in the first category, except the charge of the metal ion is specified by a Roman numeral in parentheses after the name of the metal. The charge of the metal ion is determined from the formula of the compound and the charge of the anion. For example, consider binary ionic compounds of iron and chlorine. Iron typically exhibits a charge of either 2+ or 3+ (see Figure 2.29), and the two corresponding compound formulas are $FeCl_2$ and $FeCl_3$. The simplest name, “iron chloride,” will, in this case, be ambiguous, as it does not distinguish between these two compounds. In cases like this, the charge of the metal ion is included as a Roman numeral in parentheses immediately following the metal name. These two compounds are then unambiguously named iron(II) chloride and iron(III) chloride, respectively. Other examples are provided in Table 5.4.4.

Table 5.4.4: Some Ionic Compounds with Variably Charged Metal Ions

Compound	Name
$FeCl_2$	iron(II) chloride
$FeCl_3$	iron(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
SnF_2	tin(II) fluoride
SnF_4	tin(IV) fluoride

Out-of-date nomenclature used the suffixes *–ic* and *–ous* to designate metals with higher and lower charges, respectively: Iron(III) chloride, $FeCl_3$, was previously called ferric chloride, and iron(II) chloride, $FeCl_2$, was known as ferrous chloride. Though this naming convention has been largely abandoned by the scientific community, it remains in use by some segments of industry. For example, you may see the words *stannous fluoride* on a tube of toothpaste. This represents the formula SnF_2, which is more

properly named tin(II) fluoride. The other fluoride of tin is SnF_4, which was previously called stannic fluoride but is now named tin(IV) fluoride.

Ionic Hydrates

Ionic compounds that contain water molecules as integral components of their crystals are called **hydrates**. The name for an ionic hydrate is derived by adding a term to the name for the anhydrous (meaning "not hydrated") compound that indicates the number of water molecules associated with each formula unit of the compound. The added word begins with a Greek prefix denoting the number of water molecules (see Table 5.4.5) and ends with "hydrate." For example, the anhydrous compound copper(II) sulfate also exists as a hydrate containing five water molecules and named copper(II) sulfate pentahydrate. Washing soda is the common name for a hydrate of sodium carbonate containing 10 water molecules; the systematic name is sodium carbonate decahydrate.

Formulas for ionic hydrates are written by appending a vertically centered dot, a coefficient representing the number of water molecules, and the formula for water. The two examples mentioned in the previous paragraph are represented by the formulas

$$\text{copper(II) sulfate pentahydrate} \quad \mathrm{CuSO_4 \cdot 5\,H_2O}$$

$$\text{sodium carbonate decahydrate} \quad \mathrm{Na_2CO_3 \cdot 10\,H_2O}$$

Table 5.4.5: Nomenclature Prefixes

Number	Prefix	Number	Prefix
1 (sometimes omitted)	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Example 5.4.1: Naming Ionic Compounds

Name the following ionic compounds

a. Fe_2S_3
b. CuSe
c. GaN
d. $MgSO_4{\cdot}7H_2O$
e. $Ti_2(SO_4)_3$

Solution

The anions in these compounds have a fixed negative charge (S^{2-}, Se^{2-}, N^{3-}, and and the compounds must be neutral. Because the total number of positive charges in each compound must equal the total number of negative charges, the positive ions must be Fe^{3+}, Cu^{2+}, Ga^{3+}, Mg^{2+}, and Ti^{3+}. These charges are used in the names of the metal ions:

a. iron(III) sulfide
b. copper(II) selenide
c. gallium(III) nitride
d. magnesium sulfate heptahydrate
e. titanium(III) sulfate

Exercise 5.4.1

Write the formulas of the following ionic compounds:

a. chromium(III) phosphide
b. mercury(II) sulfide
c. manganese(II) phosphate

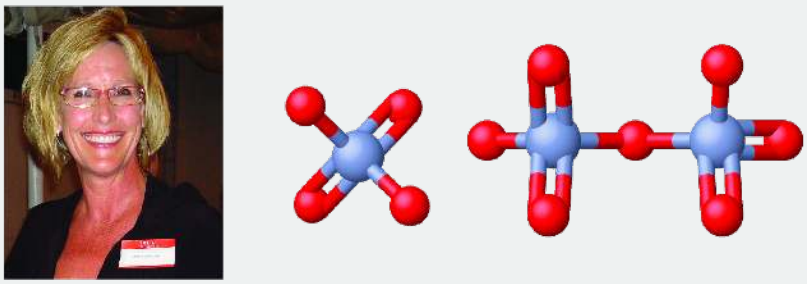
d. copper(I) oxide
e. iron(III) chloride dihydrate

Answer

(a) CrP; (b) HgS; (c) $Mn_3(PO_4)_2$; (d) Cu_2O; (e) $FeCl_3 \cdot 2H_2O$

Chemistry in Everyday Life: Erin Brockovich and Chromium Contamination

In the early 1990s, legal file clerk Erin Brockovich (Figure 5.4.1) discovered a high rate of serious illnesses in the small town of Hinckley, California. Her investigation eventually linked the illnesses to groundwater contaminated by Cr(VI) used by Pacific Gas & Electric (PG&E) to fight corrosion in a nearby natural gas pipeline. As dramatized in the film *Erin Brockovich* (for which Julia Roberts won an Oscar), Erin and lawyer Edward Masry sued PG&E for contaminating the water near Hinckley in 1993. The settlement they won in 1996—$333 million—was the largest amount ever awarded for a direct-action lawsuit in the US at that time.



(a) (b)

Figure 5.4.1: (a) Erin Brockovich found that Cr(VI), used by PG&E, had contaminated the Hinckley, California, water supply. (b) The Cr(VI) ion is often present in water as the polyatomic ions chromate, (left), and dichromate, (right).

Chromium compounds are widely used in industry, such as for chrome plating, in dye-making, as preservatives, and to prevent corrosion in cooling tower water, as occurred near Hinckley. In the environment, chromium exists primarily in either the Cr(III) or Cr(VI) forms. Cr(III), an ingredient of many vitamin and nutritional supplements, forms compounds that are not very soluble in water, and it has low toxicity. But Cr(VI) is much more toxic and forms compounds that are reasonably soluble in water. Exposure to small amounts of Cr(VI) can lead to damage of the respiratory, gastrointestinal, and immune systems, as well as the kidneys, liver, blood, and skin.

Despite cleanup efforts, Cr(VI) groundwater contamination remains a problem in Hinckley and other locations across the globe. A 2010 study by the Environmental Working Group found that of 35 US cities tested, 31 had higher levels of Cr(VI) in their tap water than the public health goal of 0.02 parts per billion set by the California Environmental Protection Agency.

Molecular (Covalent) Compounds

The bonding characteristics of inorganic molecular compounds are different from ionic compounds, and they are named using a different system as well. The charges of cations and anions dictate their ratios in ionic compounds, so specifying the names of the ions provides sufficient information to determine chemical formulas. However, because covalent bonding allows for significant variation in the combination ratios of the atoms in a molecule, the names for molecular compounds must explicitly identify these ratios.

Compounds Composed of Two Elements

When two nonmetallic elements form a molecular compound, several combination ratios are often possible. For example, carbon and oxygen can form the compounds CO and CO_2. Since these are different substances with different properties, they cannot both have the same name (they cannot both be called carbon oxide). To deal with this situation, we use a naming method that is somewhat similar to that used for ionic compounds, but with added prefixes to specify the numbers of atoms of each element. The name of the more metallic element (the one farther to the left and/or bottom of the periodic table) is first, followed by the name of the more nonmetallic element (the one farther to the right and/or top) with its ending changed to the suffix *–ide*. The numbers of atoms of each element are designated by the Greek prefixes shown in Table 5.4.5.

When only one atom of the first element is present, the prefix *mono-* is usually deleted from that part. Thus, CO is named carbon monoxide, and CO_2 is called carbon dioxide. When two vowels are adjacent, the *a* in the Greek prefix is usually dropped. Some other examples are shown in Table 5.4.6.

Table 5.4.6: Names of Some Molecular Compounds Composed of Two Elements

Compound	Name	Compound	Name
SO_2	sulfur dioxide	BCl_3	boron trichloride
SO_3	sulfur trioxide	SF_6	sulfur hexafluoride
NO_2	nitrogen dioxide	PF_5	phosphorus pentafluoride
N_2O_4	dinitrogen tetroxide	P_4O_{10}	tetraphosphorus decaoxide
N_2O_5	dinitrogen pentoxide	IF_7	iodine heptafluoride

There are a few common names that you will encounter as you continue your study of chemistry. For example, although NO is often called nitric oxide, its proper name is nitrogen monoxide. Similarly, N_2O is known as nitrous oxide even though our rules would specify the name dinitrogen monoxide. (And H_2O is usually called water, not dihydrogen monoxide.) You should commit to memory the common names of compounds as you encounter them.

Example 5.4.2: Naming Covalent Compounds

Name the following covalent compounds:

a. SF_6
b. N_2O_3
c. Cl_2O_7
d. P_4O_6

Solution

Because these compounds consist solely of nonmetals, we use prefixes to designate the number of atoms of each element:

a. sulfur hexafluoride
b. dinitrogen trioxide
c. dichlorine heptoxide
d. tetraphosphorus hexoxide

Exercise 5.4.2

Write the formulas for the following compounds:

a. phosphorus pentachloride
b. dinitrogen monoxide
c. iodine heptafluoride
d. carbon tetrachloride

Answer

(a) PCl_5; (b) N_2O; (c) IF_7; (d) CCl_4

Link to Learning

The following website provides practice with naming chemical compounds and writing chemical formulas. You can choose binary, polyatomic, and variable charge ionic compounds, as well as molecular compounds.

Binary Acids

Some compounds containing hydrogen are members of an important class of substances known as **acids**. The chemistry of these compounds is explored in more detail in later chapters of this text, but for now, it will suffice to note that many acids release hydrogen ions, H^+, when dissolved in water. To denote this distinct chemical property, a mixture of water with an acid is given a name derived from the compound's name. If the compound is a binary acid (comprised of hydrogen and one other nonmetallic element):

1. The word "hydrogen" is changed to the prefix *hydro-*
2. The other nonmetallic element name is modified by adding the suffix *-ic*
3. The word "acid" is added as a second word

For example, when the gas HCl (hydrogen chloride) is dissolved in water, the solution is called *hydrochloric acid*. Several other examples of this nomenclature are shown in Table 5.4.7.

Table 5.4.7: Names of Some Simple Acids

Name of Gas	Name of Acid
HF(*g*), hydrogen fluoride	HF(*aq*), hydrofluoric acid
HCl(*g*), hydrogen chloride	HCl(*aq*), hydrochloric acid
HBr(*g*), hydrogen bromide	HBr(*aq*), hydrobromic acid
HI(*g*), hydrogen iodide	HI(*aq*), hydroiodic acid
H_2S(*g*), hydrogen sulfide	H_2S(*aq*), hydrosulfuric acid

Oxyacids

Many compounds containing three or more elements (such as organic compounds or coordination compounds) are subject to specialized nomenclature rules that you will learn later. However, we will briefly discuss the important compounds known as **oxyacids**, compounds that contain hydrogen, oxygen, and at least one other element, and are bonded in such a way as to impart acidic properties to the compound (you will learn the details of this in a later chapter). Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion. To name oxyacids:

1. Omit "hydrogen"
2. Start with the root name of the anion
3. Replace *–ate* with *–ic*, or *–ite* with *–ous*
4. Add "acid"

For example, consider H_2CO_3 (which you might be tempted to call "hydrogen carbonate"). To name this correctly, "hydrogen" is omitted; the *–ate* of carbonate is replace with *–ic*; and acid is added—so its name is carbonic acid. Other examples are given in Table 5.4.8. There are some exceptions to the general naming method (e.g., H_2SO_4 is called sulfuric acid, not sulfic acid, and H_2SO_3 is sulfurous, not sulfous, acid).

Table 5.4.8: Names of Common Oxyacids

Formula	Anion Name	Acid Name
$HC_2H_3O_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
$HClO_4$	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid

Formula	Anion Name	Acid Name
H_3PO_4	phosphate	phosphoric acid

This page titled 5.4: Chemical Nomenclature is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **2.7: Chemical Nomenclature** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.5: Lewis Symbols and Structures

Learning Objectives

By the end of this section, you will be able to:

- Write Lewis symbols for neutral atoms and ions
- Draw Lewis structures depicting the bonding in simple molecules

Thus far in this chapter, we have discussed the various types of bonds that form between atoms and/or ions. In all cases, these bonds involve the sharing or transfer of valence shell electrons between atoms. In this section, we will explore the typical method for depicting valence shell electrons and chemical bonds, namely Lewis symbols and Lewis structures.

Lewis Symbols

We use Lewis symbols to describe valence electron configurations of atoms and monatomic ions. A Lewis symbol consists of an elemental symbol surrounded by one dot for each of its valence electrons:



Figure 5.5.1: shows the Lewis symbols for the elements of the third period of the periodic table.

Atoms	Electronic Configuration	Lewis Symbol
sodium	$[Ne]3s^1$	Na·
magnesium	$[Ne]3s^2$	·Mg·
aluminum	$[Ne]3s^23p^1$	·Al·
silicon	$[Ne]3s^23p^2$	·Si·
phosphorus	$[Ne]3s^23p^3$	·P·
sulfur	$[Ne]3s^23p^4$	:S·
chlorine	$[Ne]3s^23p^5$	:Cl·
argon	$[Ne]3s^23p^6$	:Ar:

Figure 5.5.1: Lewis symbols illustrating the number of valence electrons for each element in the third period of the periodic table.

Lewis symbols can also be used to illustrate the formation of cations from atoms, as shown here for sodium and calcium:



Likewise, they can be used to show the formation of anions from atoms, as shown here for chlorine and sulfur:



Figure 5.5.2 demonstrates the use of Lewis symbols to show the transfer of electrons during the formation of ionic compounds.

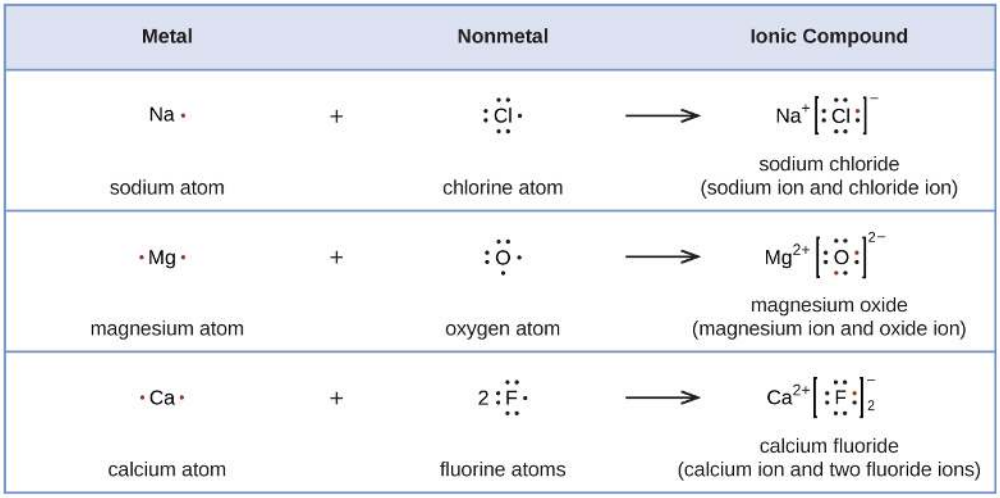
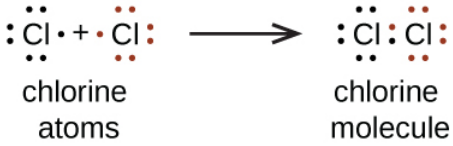


Figure 5.5.2: Cations are formed when atoms lose electrons, represented by fewer Lewis dots, whereas anions are formed by atoms gaining electrons. The total number of electrons does not change.

Lewis Structures

We also use Lewis symbols to indicate the formation of covalent bonds, which are shown in Lewis structures, drawings that describe the bonding in molecules and polyatomic ions. For example, when two chlorine atoms form a chlorine molecule, they share one pair of electrons:



The Lewis structure indicates that each Cl atom has three pairs of electrons that are not used in bonding (called lone pairs) and one shared pair of electrons (written between the atoms). A dash (or line) is sometimes used to indicate a shared pair of electrons:

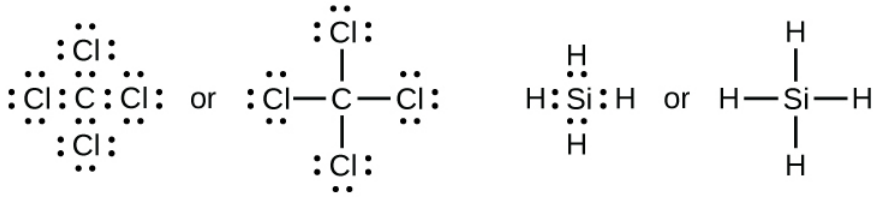


A single shared pair of electrons is called a single bond. Each Cl atom interacts with eight valence electrons: the six in the lone pairs and the two in the single bond.

The Octet Rule

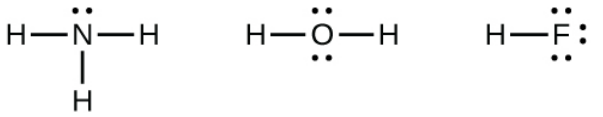
The other halogen molecules (F_2, Br_2, I_2, and At_2) form bonds like those in the chlorine molecule: one single bond between atoms and three lone pairs of electrons per atom. This allows each halogen atom to have a noble gas electron configuration. The tendency of main group atoms to form enough bonds to obtain eight valence electrons is known as the octet rule.

The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet (eight valence electrons); this is especially true of the nonmetals of the second period of the periodic table (C, N, O, and F). For example, each atom of a group 14 element has four electrons in its outermost shell and therefore requires four more electrons to reach an octet. These four electrons can be gained by forming four covalent bonds, as illustrated here for carbon in CCl_4 (carbon tetrachloride) and silicon in SiH_4 (silane). Because hydrogen only needs two electrons to fill its valence shell, it is an exception to the octet rule. The transition elements and inner transition elements also do not follow the octet rule:



carbon tetrachloride silane

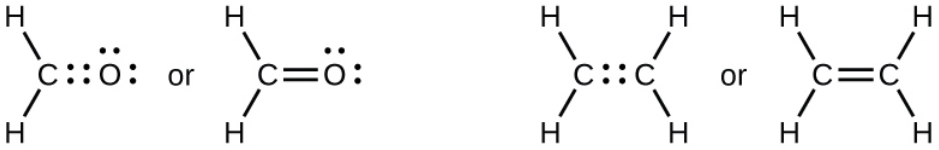
Group 15 elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To obtain an octet, these atoms form three covalent bonds, as in NH_3 (ammonia). Oxygen and other atoms in group 16 obtain an octet by forming two covalent bonds:



ammonia Water hydrogen fluoride

Double and Triple Bonds

As previously mentioned, when a pair of atoms shares one pair of electrons, we call this a single bond. However, a pair of atoms may need to share more than one pair of electrons in order to achieve the requisite octet. A double bond forms when two pairs of electrons are shared between a pair of atoms, as between the carbon and oxygen atoms in CH_2O (formaldehyde) and between the two carbon atoms in C_2H_4 (ethylene):



formaldehyde ethylene

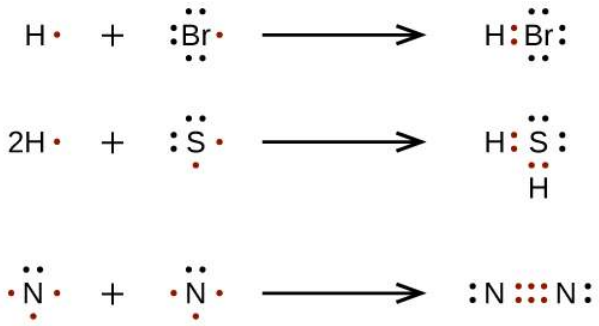
A triple bond forms when three electron pairs are shared by a pair of atoms, as in carbon monoxide (CO) and the cyanide ion (CN^-):



carbon monoxide cyanide ion

Writing Lewis Structures with the Octet Rule

For very simple molecules and molecular ions, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms. See these examples:



For more complicated molecules and molecular ions, it is helpful to follow the step-by-step procedure outlined here.

Writing Lewis Structures

1. Determine the total number of valence (outer shell) electrons. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.
2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom. (Generally, the least electronegative element should be placed in the center.) Connect each atom to the central atom with a single bond (one electron pair).
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen), completing an octet around each atom.
4. Place all remaining electrons on the central atom.
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

Let us determine the Lewis structures of SiH_4, CHO_2^-, NO^+, and OF_2 as examples in following this procedure:

1. Determine the total number of valence (outer shell) electrons in the molecule or ion.
   - For a molecule, we add the number of valence electrons on each atom in the molecule:

$$\begin{array}{rr} & \mathrm{SiH_4} \\ \mathrm{Si:}\ 4\ \text{valence electrons/atom} \times\ 1\ \text{atom} & = 4 \\ +\quad \mathrm{H:}\ 1\ \text{valence electron/atom} \times\ 4\ \text{atoms} & = 4 \\ \hline \text{total valence electrons} & = 8 \end{array}$$

   - For a *negative ion*, such as CHO_2^-, we add the number of valence electrons on the atoms to the number of negative charges on the ion (one electron is gained for each single negative charge):

$$\begin{array}{rr} & \mathrm{CHO_2^-} \\ \mathrm{C:}\ 4\ \text{valence electrons/atom} \times\ 1\ \text{atom} & = 4 \\ \mathrm{H:}\ 1\ \text{valence electron/atom} \times\ 1\ \text{atoms} & = 1 \\ \mathrm{O:}\ 6\ \text{valence electron/atom} \times\ 2\ \text{atoms} & = 12 \\ +\quad 1\ \text{electron} & = 1 \\ \hline \text{total valence electrons} & = 18 \end{array}$$

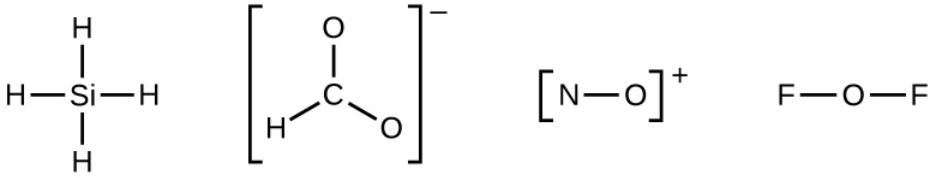
   - For a *positive ion*, such as NO^+, we add the number of valence electrons on the atoms in the ion and then subtract the number of positive charges on the ion (one electron is lost for each single positive charge) from the total number of valence electrons:

$$\begin{array}{rr} & \mathrm{NO^+} \\ \mathrm{N:}\ 5\ \text{valence electrons/atom} \times\ 1\ \text{atom} & = 5 \\ \mathrm{O:}\ 6\ \text{valence electron/atom} \times\ 1\ \text{atoms} & = 6 \\ +\quad \text{e: -1 electron (positive charge)} \times\ 1\ \text{electron} & = -1 \\ \hline \text{total valence electrons} & = 10 \end{array}$$

   - Since OF_2 is a neutral molecule, we simply add the number of valence electrons:

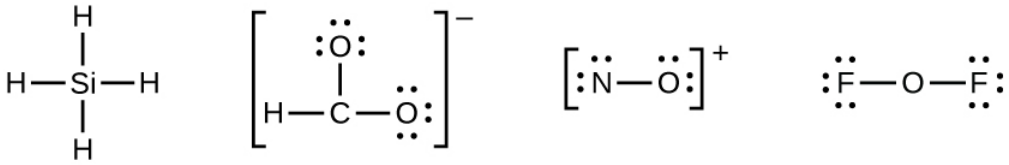
$$\begin{array}{rr} & \mathrm{OF_2} \\ \mathrm{O:\ 6\ valence\ electrons/atom} \times & 1\ \mathrm{atom} = 6 \\ + \quad \mathrm{F:\ 7\ valence\ electron/atom} \times & 2\ \mathrm{atoms} = 14 \\ \hline & \mathrm{total\ valence\ electrons} = 20 \end{array}$$

2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom and connecting each atom to the central atom with a single (one electron pair) bond. (Note that we denote ions with brackets around the structure, indicating the charge outside the brackets:)

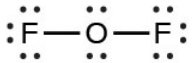


When several arrangements of atoms are possible, as for $\mathrm{CHO_2^-}$, we must use experimental evidence to choose the correct one. In general, the less electronegative elements are more likely to be central atoms. In $\mathrm{CHO_2^-}$, the less electronegative carbon atom occupies the central position with the oxygen and hydrogen atoms surrounding it. Other examples include P in $\mathrm{POCl_3}$, S in $\mathrm{SO_2}$, and Cl in $\mathrm{ClO_4^-}$. An exception is that hydrogen is almost never a central atom. As the most electronegative element, fluorine also cannot be a central atom.

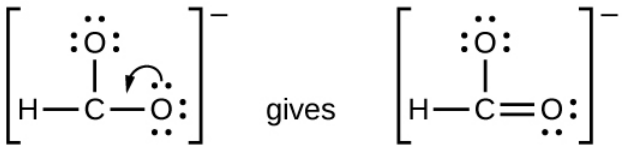
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons.
   - There are no remaining electrons on $\mathrm{SiH_4}$, so it is unchanged:



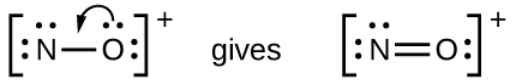
4. Place all remaining electrons on the central atom.
   - For $\mathrm{SiH_4}$, $\mathrm{CHO_2^-}$, and $\mathrm{NO^+}$, there are no remaining electrons; we already placed all of the electrons determined in Step 1.
   - For $\mathrm{OF_2}$, we had 16 electrons remaining in Step 3, and we placed 12, leaving 4 to be placed on the central atom:



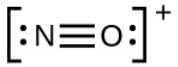
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.
   - $\mathrm{SiH_4}$: Si already has an octet, so nothing needs to be done.
   - $\mathrm{CHO2^-}$: We have distributed the valence electrons as lone pairs on the oxygen atoms, but the carbon atom lacks an octet:



   - $\mathrm{NO^+}$: For this ion, we added eight valence electrons, but neither atom has an octet. We cannot add any more electrons since we have already used the total that we found in Step 1, so we must move electrons to form a multiple bond:



This still does not produce an octet, so we must move another pair, forming a triple bond:



- In OF_2, each atom has an octet as drawn, so nothing changes.

Example 5.5.1: Writing Lewis Structures

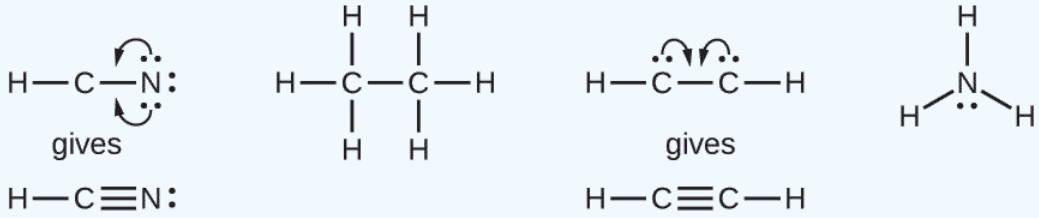
NASA's Cassini-Huygens mission detected a large cloud of toxic hydrogen cyanide (HCN) on Titan, one of Saturn's moons. Titan also contains ethane (H_3CCH_3), acetylene (HCCH), and ammonia (NH_3). What are the Lewis structures of these molecules?

Solution

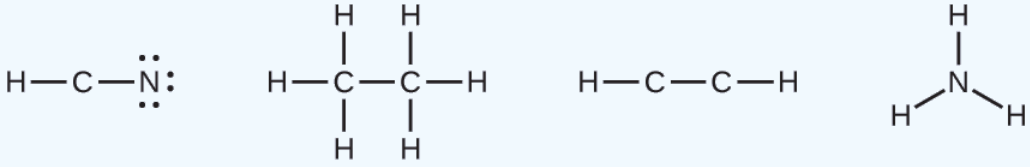
Step 1. *Calculate the number of valence electrons.*

HCN: (1 × 1) + (4 × 1) + (5 × 1) = 10
H_3CCH_3: (1 × 3) + (2 × 4) + (1 × 3) = 14
HCCH: (1 × 1) + (2 × 4) + (1 × 1) = 10
NH_3: (5 × 1) + (3 × 1) = 8

Step 2 *Draw a skeleton and connect the atoms with single bonds.* Remember that H is never a central atom:

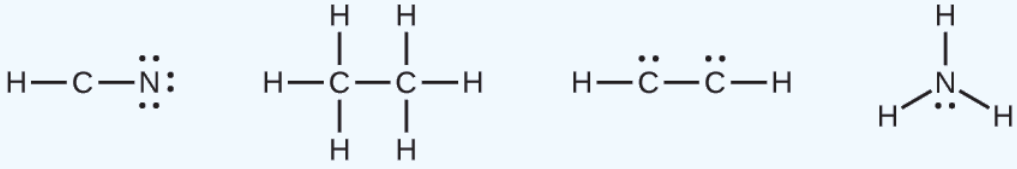


Step 3. *Where needed, distribute electrons to the terminal atoms:*



HCN: six electrons placed on N
H_3CCH_3: no electrons remain
HCCH: no terminal atoms capable of accepting electrons
NH_3: no terminal atoms capable of accepting electrons

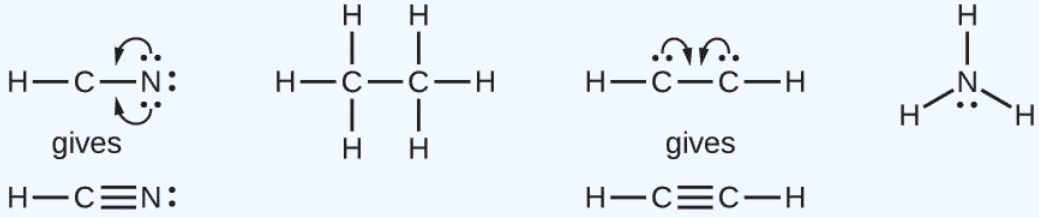
Step 4. *Where needed, place remaining electrons on the central atom:*



HCN: no electrons remain
H_3CCH_3: no electrons remain
HCCH: four electrons placed on carbon
NH_3: two electrons placed on nitrogen

Step 5. *Where needed, rearrange electrons to form multiple bonds in order to obtain an octet on each atom:*

HCN: form two more C–N bonds
H_3CCH_3: all atoms have the correct number of electrons
HCCH: form a triple bond between the two carbon atoms
NH_3: all atoms have the correct number of electrons



Exercise 5.5.1

Both carbon monoxide, CO, and carbon dioxide, CO_2, are products of the combustion of fossil fuels. Both of these gases also cause problems: CO is toxic and CO_2 has been implicated in global climate change. What are the Lewis structures of these two molecules?

Answer



How Sciences Interconnect: Fullerene Chemistry

Carbon, in various forms and compounds, has been known since prehistoric times, . Soot has been used as a pigment (often called carbon black) for thousands of years. Charcoal, high in carbon content, has likewise been critical to human development. Carbon is the key additive to iron in the steelmaking process, and diamonds have a unique place in both culture and industry. With all this usage came significant study, particularly with the emergence of organic chemistry. And even with all the known forms and functions of the element, scientists began to uncover the potential for even more varied and extensive carbon structures.

As early as the 1960s, chemists began to observe complex carbon structures, but they had little evidence to support their concepts, or their work did not make it into the mainstream. Eiji Osawa predicted a spherical form based on observations of a similar structure, but his work was not widely known outside Japan. In a similar manner, the most comprehensive advance was likely computational chemist Elena Galpern's, who in 1973 predicted a highly stable, 60-carbon molecule; her work was also isolated to her native Russia. Still later, Harold Kroto, working with Canadian radio astronomers, sought to uncover the nature of long carbon chains that had been discovered in interstellar space.

Kroto sought to use a machine developed by Richard Smalley's team at Rice University to learn more about these structures. Together with Robert Curl, who had introduced them, and three graduate students—James Heath, Sean O'Brien, and Yuan Liu —they performed an intensive series of experiments that led to a major discovery.

In 1996, the Nobel Prize in Chemistry was awarded to Richard Smalley (Figure 5.5.3), Robert Curl, and Harold Kroto for their work in discovering a new form of carbon, the C_{60} buckminsterfullerene molecule. An entire class of compounds, including spheres and tubes of various shapes, were discovered based on C_{60}. This type of molecule, called a fullerene, shows promise in a variety of applications. Because of their size and shape, fullerenes can encapsulate other molecules, so they have shown

potential in various applications from hydrogen storage to targeted drug delivery systems. They also possess unique electronic and optical properties that have been put to good use in solar powered devices and chemical sensors.

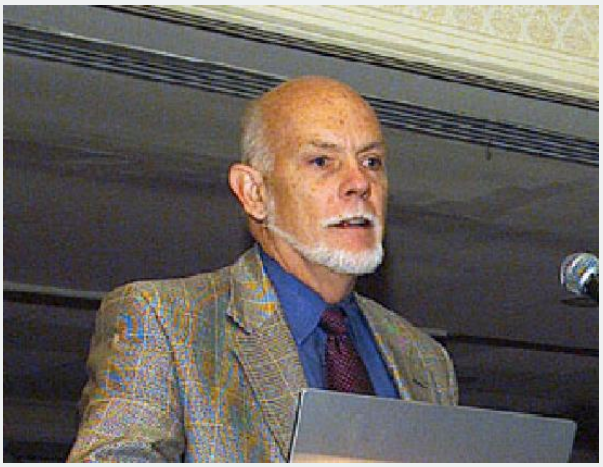


Figure 5.5.3: Richard Smalley (1943–2005), a professor of physics, chemistry, and astronomy at Rice University, was one of the leading advocates for fullerene chemistry. Upon his death in 2005, the US Senate honored him as the "Father of Nanotechnology." (credit: United States Department of Energy)

Exceptions to the Octet Rule

Many covalent molecules have central atoms that do not have eight electrons in their Lewis structures. These molecules fall into three categories:

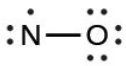
- Odd-electron molecules have an odd number of valence electrons, and therefore have an unpaired electron.
- Electron-deficient molecules have a central atom that has fewer electrons than needed for a noble gas configuration.
- Hypervalent molecules have a central atom that has more electrons than needed for a noble gas configuration.

Odd-electron Molecules

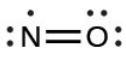
We call molecules that contain an odd number of electrons free radicals. Nitric oxide, NO, is an example of an odd-electron molecule; it is produced in internal combustion engines when oxygen and nitrogen react at high temperatures.

To draw the Lewis structure for an odd-electron molecule like NO, we follow the same five steps we would for other molecules, but with a few minor changes:

1. *Determine the total number of valence (outer shell) electrons*. The sum of the valence electrons is 5 (from N) + 6 (from O) = 11. The odd number immediately tells us that we have a free radical, so we know that not every atom can have eight electrons in its valence shell.
2. *Draw a skeleton structure of the molecule*. We can easily draw a skeleton with an N–O single bond:
   N–O
3. *Distribute the remaining electrons as lone pairs on the terminal atoms*. In this case, there is no central atom, so we distribute the electrons around both atoms. We give eight electrons to the more electronegative atom in these situations; thus oxygen has the filled valence shell:

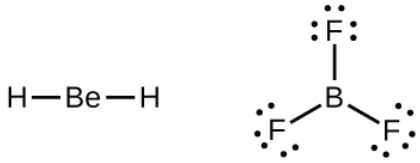


4. *Place all remaining electrons on the central atom.* Since there are no remaining electrons, this step does not apply.
5. *Rearrange the electrons to make multiple bonds with the central atom in order to obtain octets wherever possible.* We know that an odd-electron molecule cannot have an octet for every atom, but we want to get each atom as close to an octet as possible. In this case, nitrogen has only five electrons around it. To move closer to an octet for nitrogen, we take one of the lone pairs from oxygen and use it to form a NO double bond. (We cannot take another lone pair of electrons on oxygen and form a triple bond because nitrogen would then have nine electrons:)

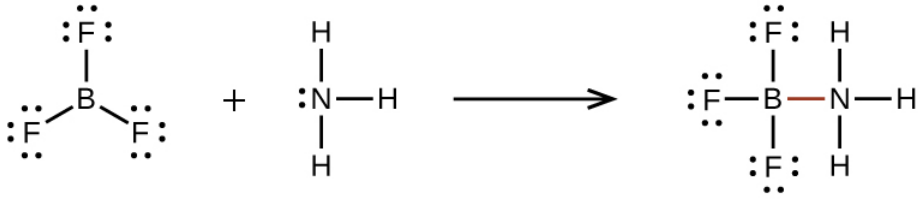


Electron-deficient Molecules

We will also encounter a few molecules that contain central atoms that do not have a filled valence shell. Generally, these are molecules with central atoms from groups 2 and 13, outer atoms that are hydrogen, or other atoms that do not form multiple bonds. For example, in the Lewis structures of beryllium dihydride, BeH_2, and boron trifluoride, BF_3, the beryllium and boron atoms each have only four and six electrons, respectively. It is possible to draw a structure with a double bond between a boron atom and a fluorine atom in BF_3, satisfying the octet rule, but experimental evidence indicates the bond lengths are closer to that expected for B–F single bonds. This suggests the best Lewis structure has three B–F single bonds and an electron deficient boron. The reactivity of the compound is also consistent with an electron deficient boron. However, the B–F bonds are slightly shorter than what is actually expected for B–F single bonds, indicating that some double bond character is found in the actual molecule.



An atom like the boron atom in BF_3, which does not have eight electrons, is very reactive. It readily combines with a molecule containing an atom with a lone pair of electrons. For example, NH_3 reacts with BF_3 because the lone pair on nitrogen can be shared with the boron atom:



Hypervalent Molecules

Elements in the second period of the periodic table ($n = 2$) can accommodate only eight electrons in their valence shell orbitals because they have only four valence orbitals (one 2*s* and three 2*p* orbitals). Elements in the third and higher periods ($n \geq 3$) have more than four valence orbitals and can share more than four pairs of electrons with other atoms because they have empty *d* orbitals in the same shell. Molecules formed from these elements are sometimes called hypervalent molecules. Figure 5.5.4: shows the Lewis structures for two hypervalent molecules, PCl_5 and SF_6.

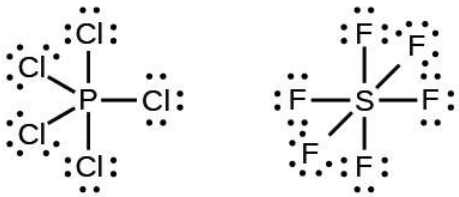
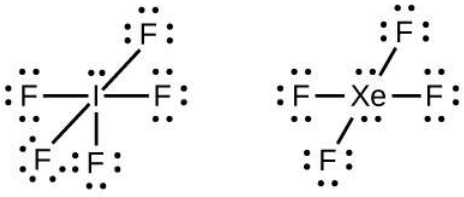


Figure 5.5.4: In PCl_5, the central atom phosphorus shares five pairs of electrons. In SF_6, sulfur shares six pairs of electrons.

In some hypervalent molecules, such as IF_5 and XeF_4, some of the electrons in the outer shell of the central atom are lone pairs:



When we write the Lewis structures for these molecules, we find that we have electrons left over after filling the valence shells of the outer atoms with eight electrons. These additional electrons must be assigned to the central atom.

Example 5.5.2: Writing Lewis Structures - Octet Rule Violations

Xenon is a noble gas, but it forms a number of stable compounds. We examined XeF_4 earlier. What are the Lewis structures of XeF_2 and XeF_6?

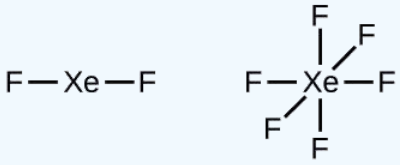
Solution

We can draw the Lewis structure of any covalent molecule by following the six steps discussed earlier. In this case, we can condense the last few steps, since not all of them apply.

Step 1. Calculate the number of valence electrons:

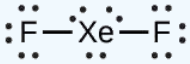
XeF_2: 8 + (2 × 7) = 22
XeF_6: 8 + (6 × 7) = 50

Step 2. Draw a skeleton joining the atoms by single bonds. Xenon will be the central atom because fluorine cannot be a central atom:

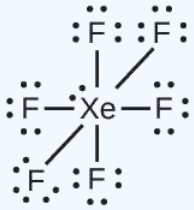


Step 3. Distribute the remaining electrons.

XeF_2: We place three lone pairs of electrons around each F atom, accounting for 12 electrons and giving each F atom 8 electrons. Thus, six electrons (three lone pairs) remain. These lone pairs must be placed on the Xe atom. This is acceptable because Xe atoms have empty valence shell *d* orbitals and can accommodate more than eight electrons. The Lewis structure of XeF_2 shows two bonding pairs and three lone pairs of electrons around the Xe atom:



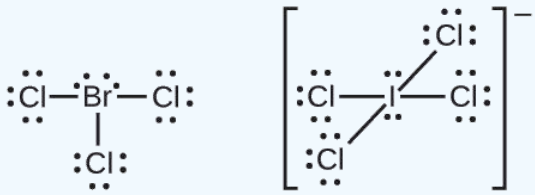
XeF_6: We place three lone pairs of electrons around each F atom, accounting for 36 electrons. Two electrons remain, and this lone pair is placed on the Xe atom:



Exercise 5.5.2: Interhalogens

The halogens form a class of compounds called the interhalogens, in which halogen atoms covalently bond to each other. Write the Lewis structures for the interhalogens $BrCl_3$ and ICl_4^-.

Answer



This page titled 5.5: Lewis Symbols and Structures is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **7.3: Lewis Symbols and Structures** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.6: Formal Charges and Resonance

Learning Objectives

By the end of this section, you will be able to:

- Compute formal charges for atoms in any Lewis structure
- Use formal charges to identify the most reasonable Lewis structure for a given molecule
- Explain the concept of resonance and draw Lewis structures representing resonance forms for a given molecule

In the previous section, we discussed how to write Lewis structures for molecules and polyatomic ions. As we have seen, however, in some cases, there is seemingly more than one valid structure for a molecule. We can use the concept of formal charges to help us predict the most appropriate Lewis structure when more than one is reasonable.

Calculating Formal Charge

The formal charge of an atom in a molecule is the *hypothetical* charge the atom would have if we could redistribute the electrons in the bonds evenly between the atoms. Another way of saying this is that formal charge results when we take the number of valence electrons of a neutral atom, subtract the nonbonding electrons, and then subtract the number of bonds connected to that atom in the Lewis structure.

Thus, we calculate formal charge as follows:

$$\text{formal charge } = \#\text{ valence shell electrons (free atom) } - \#\text{ lone pair electrons } - \frac{1}{2}\#\text{ bonding electrons}$$

We can double-check formal charge calculations by determining the sum of the formal charges for the whole structure. The sum of the formal charges of all atoms in a molecule must be zero; the sum of the formal charges in an ion should equal the charge of the ion.

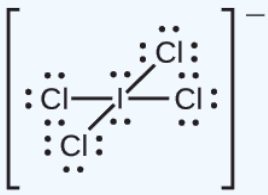
We must remember that the formal charge calculated for an atom is not the *actual* charge of the atom in the molecule. Formal charge is only a useful bookkeeping procedure; it does not indicate the presence of actual charges.

Example 5.6.1: Calculating Formal Charge from Lewis Structures

Assign formal charges to each atom in the interhalogen ion

Solution

Step 1. *We divide the bonding electron pairs equally for all I–Cl bonds:*



Step 2. *We assign lone pairs of electrons to their atoms*. Each Cl atom now has seven electrons assigned to it, and the I atom has eight.

Step 3. *Subtract this number from the number of valence electrons for the neutral atom:*

I: 7 – 8 = –1
Cl: 7 – 7 = 0

The sum of the formal charges of all the atoms equals –1, which is identical to the charge of the ion (–1).

? Exercise 5.6.1

Calculate the formal charge for each atom in the carbon monoxide molecule:



Answer

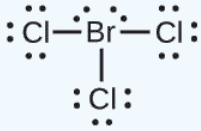
C −1, O +1

✓ Example 5.6.2: Calculating Formal Charge from Lewis Structures

Assign formal charges to each atom in the interhalogen molecule $BrCl_3$.

Solution

Step 1. *Assign one of the electrons in each Br–Cl bond to the Br atom and one to the Cl atom in that bond:*



Step 2. *Assign the lone pairs to their atom.* Now each Cl atom has seven electrons and the Br atom has seven electrons.

Step 3. *Subtract this number from the number of valence electrons for the neutral atom.* This gives the formal charge:

Br: 7 – 7 = 0
Cl: 7 – 7 = 0

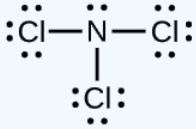
All atoms in $BrCl_3$ have a formal charge of zero, and the sum of the formal charges totals zero, as it must in a neutral molecule.

? Exercise 5.6.2

Determine the formal charge for each atom in NCl_3.

Answer

N: 0; all three Cl atoms: 0

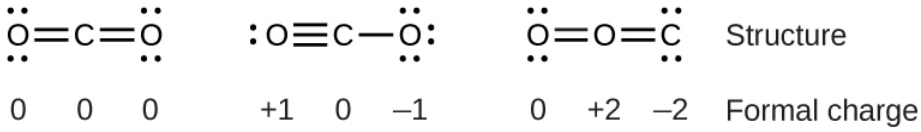


Using Formal Charge to Predict Molecular Structure

The arrangement of atoms in a molecule or ion is called its **molecular structure**. In many cases, following the steps for writing Lewis structures may lead to more than one possible molecular structure—different multiple bond and lone-pair electron placements or different arrangements of atoms, for instance. A few guidelines involving formal charge can be helpful in deciding which of the possible structures is most likely for a particular molecule or ion:

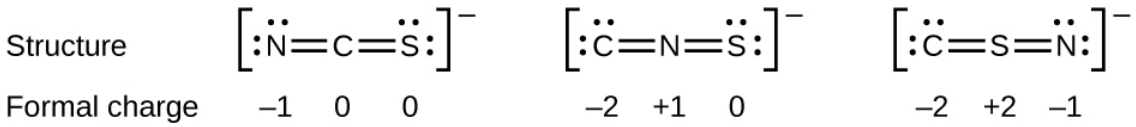
1. A molecular structure in which all formal charges are zero is preferable to one in which some formal charges are not zero.
2. If the Lewis structure must have nonzero formal charges, the arrangement with the smallest nonzero formal charges is preferable.
3. Lewis structures are preferable when adjacent formal charges are zero or of the opposite sign.
4. When we must choose among several Lewis structures with similar distributions of formal charges, the structure with the negative formal charges on the more electronegative atoms is preferable.

To see how these guidelines apply, let us consider some possible structures for carbon dioxide, CO_2. We know from our previous discussion that the less electronegative atom typically occupies the central position, but formal charges allow us to understand *why* this occurs. We can draw three possibilities for the structure: carbon in the center and double bonds, carbon in the center with a single and triple bond, and oxygen in the center with double bonds:



Comparing the three formal charges, we can definitively identify the structure on the left as preferable because it has only formal charges of zero (Guideline 1).

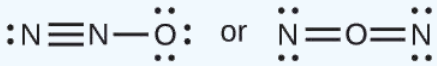
As another example, the thiocyanate ion, an ion formed from a carbon atom, a nitrogen atom, and a sulfur atom, could have three different molecular structures: NCS^-, CNS^-, or CSN^-. The formal charges present in each of these molecular structures can help us pick the most likely arrangement of atoms. Possible Lewis structures and the formal charges for each of the three possible structures for the thiocyanate ion are shown here:



Note that the sum of the formal charges in each case is equal to the charge of the ion (–1). However, the first arrangement of atoms is preferred because it has the lowest number of atoms with nonzero formal charges (Guideline 2). Also, it places the least electronegative atom in the center, and the negative charge on the more electronegative element (Guideline 4).

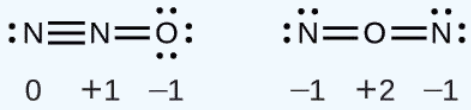
Example 5.6.3: Using Formal Charge to Determine Molecular Structure

Nitrous oxide, N_2O, commonly known as laughing gas, is used as an anesthetic in minor surgeries, such as the routine extraction of wisdom teeth. Which is the more likely structure for nitrous oxide?

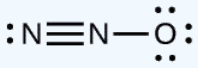


Solution

Determining formal charge yields the following:



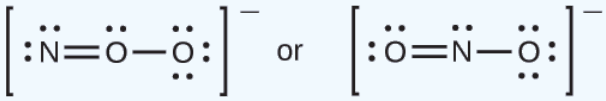
The structure with a terminal oxygen atom best satisfies the criteria for the most stable distribution of formal charge:



The number of atoms with formal charges are minimized (Guideline 2), there is no formal charge with a magnitude greater than one (Guideline 2), the negative formal charge is on the more electronegative element (Guideline 4), and the less electronegative atom is in the center position.

Exercise 5.6.3

Which is the most likely molecular structure for the nitrite ion?



Answer

ONO$^-$

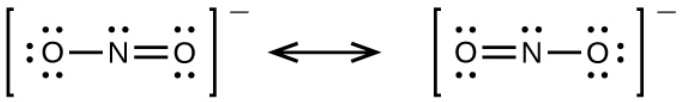
Resonance

Notice that the more likely structure for the nitrite anion in Example 5.6.3 may actually be drawn in two different ways, distinguished by the locations of the N-O and N=O bonds:



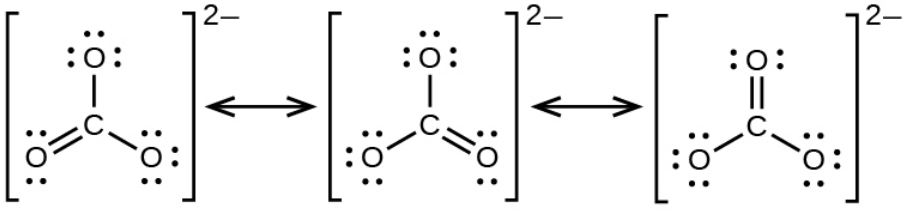
If nitrite ions do indeed contain a single and a double bond, we would expect for the two bond lengths to be different. A double bond between two atoms is shorter (and stronger) than a single bond between the same two atoms. Experiments show, however, that both N–O bonds in have the same strength and length, and are identical in all other properties.

It is not possible to write a single Lewis structure for in which nitrogen has an octet and both bonds are equivalent. Instead, we use the concept of **resonance**: if two or more Lewis structures with the same arrangement of atoms can be written for a molecule or ion, the actual distribution of electrons is an *average* of that shown by the various Lewis structures. The actual distribution of electrons in each of the nitrogen-oxygen bonds in is the average of a double bond and a single bond. We call the individual Lewis structures **resonance forms**. The actual electronic structure of the molecule (the average of the resonance forms) is called a **resonance hybrid** of the individual resonance forms. A double-headed arrow between Lewis structures indicates that they are resonance forms.



We should remember that a molecule described as a resonance hybrid *never* possesses an electronic structure described by either resonance form. It does not fluctuate between resonance forms; rather, the actual electronic structure is *always* the average of that shown by all resonance forms. George Wheland, one of the pioneers of resonance theory, used a historical analogy to describe the relationship between resonance forms and resonance hybrids. A medieval traveler, having never before seen a rhinoceros, described it as a hybrid of a dragon and a unicorn because it had many properties in common with both. Just as a rhinoceros is neither a dragon sometimes nor a unicorn at other times, a resonance hybrid is neither of its resonance forms at any given time. Like a rhinoceros, it is a real entity that experimental evidence has shown to exist. It has some characteristics in common with its resonance forms, but the resonance forms themselves are convenient, imaginary images (like the unicorn and the dragon).

The carbonate anion, provides a second example of resonance:



One oxygen atom must have a double bond to carbon to complete the octet on the central atom. All oxygen atoms, however, are equivalent, and the double bond could form from any one of the three atoms. This gives rise to three resonance forms of the carbonate ion. Because we can write three identical resonance structures, we know that the actual arrangement of electrons in the carbonate ion is the average of the three structures. Again, experiments show that all three C–O bonds are exactly the same.

Link to Learning

Use this online quiz to practice your skills in drawing resonance structures and estimating formal charges.

This page titled 5.6: Formal Charges and Resonance is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **7.4: Formal Charges and Resonance** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.7: Molecular Structure and Polarity

Learning Objectives

By the end of this section, you will be able to:

- Predict the structures of small molecules using valence shell electron pair repulsion (VSEPR) theory
- Explain the concepts of polar covalent bonds and molecular polarity
- Assess the polarity of a molecule based on its bonding and structure

Thus far, we have used two-dimensional Lewis structures to represent molecules. However, molecular structure is actually three-dimensional, and it is important to be able to describe molecular bonds in terms of their distances, angles, and relative arrangements in space (Figure 5.7.1). A bond angle is the angle between any two bonds that include a common atom, usually measured in degrees. A bond distance (or bond length) is the distance between the nuclei of two bonded atoms along the straight line joining the nuclei. Bond distances are measured in Ångstroms (1 Å = 10^{-10} m) or picometers (1 pm = 10^{-12} m, 100 pm = 1 Å).

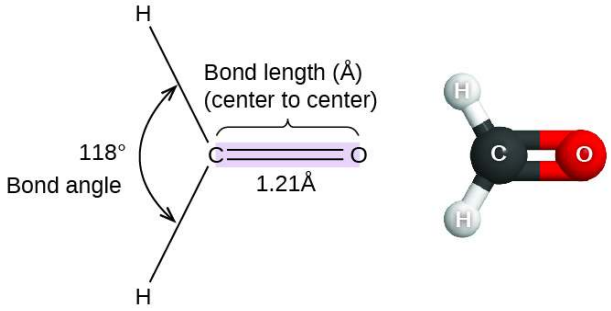


Figure 5.7.1: Bond distances (lengths) and angles are shown for the formaldehyde molecule, H_2CO.

VSEPR Theory

Valence shell electron-pair repulsion theory (VSEPR theory) enables us to predict the molecular structure, including approximate bond angles around a central atom, of a molecule from an examination of the number of bonds and lone electron pairs in its Lewis structure. The VSEPR model assumes that electron pairs in the valence shell of a central atom will adopt an arrangement that minimizes repulsions between these electron pairs by maximizing the distance between them. The electrons in the valence shell of a central atom form either bonding pairs of electrons, located primarily between bonded atoms, or lone pairs. The electrostatic repulsion of these electrons is reduced when the various regions of high electron density assume positions as far from each other as possible.

VSEPR theory predicts the arrangement of electron pairs around each central atom and, usually, the correct arrangement of atoms in a molecule. We should understand, however, that the theory only considers electron-pair repulsions. Other interactions, such as nuclear-nuclear repulsions and nuclear-electron attractions, are also involved in the final arrangement that atoms adopt in a particular molecular structure.

As a simple example of VSEPR theory, let us predict the structure of a gaseous BeF_2 molecule. The Lewis structure of BeF_2 (Figure 5.7.2) shows only two electron pairs around the central beryllium atom. With two bonds and no lone pairs of electrons on the central atom, the bonds are as far apart as possible, and the electrostatic repulsion between these regions of high electron density is reduced to a minimum when they are on opposite sides of the central atom. The bond angle is 180° (Figure 5.7.2).

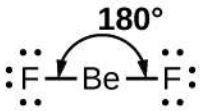


Figure 5.7.2: The BeF_2 molecule adopts a linear structure in which the two bonds are as far apart as possible, on opposite sides of the Be atom.

Figure 5.7.3 illustrates this and other electron-pair geometries that minimize the repulsions among regions of high electron density (bonds and/or lone pairs). Two regions of electron density around a central atom in a molecule form a linear geometry; three regions form a trigonal planar geometry; four regions form a tetrahedral geometry; five regions form a trigonal bipyramidal geometry; and six regions form an octahedral geometry.

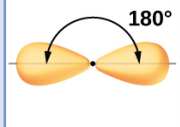
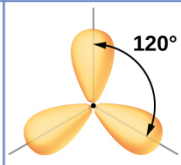
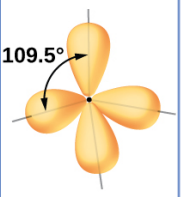
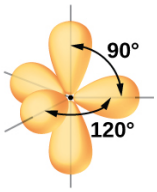
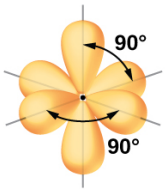
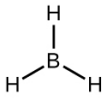
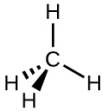
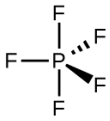
Number of regions	Two regions of high electron density (bonds and/or unshared pairs)	Three regions of high electron density (bonds and/or unshared pairs)	Four regions of high electron density (bonds and/or unshared pairs)	Five regions of high electron density (bonds and/or unshared pairs)	Six regions of high electron density (bonds and/or unshared pairs)
Spatial arrangement					
Line-dash-wedge notation					
Electron region geometry	Linear; 180° angle	Trigonal planar; all angles 120°	Tetrahedral; all angles 109.5°	Trigonal bipyramidal; angles of 90° or 120° An attached atom may be equatorial (in the plane of the triangle) or axial (above or below the plane of the triangle).	Octahedral; all angles 90° or 180°

Figure 5.7.3: The basic electron-pair geometries predicted by VSEPR theory maximize the space around any region of electron density (bonds or lone pairs).

Electron-pair Geometry versus Molecular Structure

It is important to note that electron-pair geometry around a central atom is *not* the same thing as its molecular structure. The electron-pair geometries shown in Figure 5.7.3 describe all regions where electrons are located, bonds as well as lone pairs. Molecular structure describes the location of the *atoms*, not the electrons.

We differentiate between these two situations by naming the geometry that includes *all* electron pairs the electron-pair geometry. The structure that includes only the placement of the atoms in the molecule is called the molecular structure. The electron-pair geometries will be the same as the molecular structures when there are no lone electron pairs around the central atom, but they will be different when there are lone pairs present on the central atom.

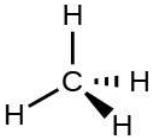


Figure 5.7.4: The molecular structure of the methane molecule, CH_4, is shown with a tetrahedral arrangement of the hydrogen atoms. VSEPR structures like this one are often drawn using the wedge and dash notation, in which solid lines represent bonds in the plane of the page, solid wedges represent bonds coming up out of the plane, and dashed lines represent bonds going down into the plane.

For example, the methane molecule, CH_4, which is the major component of natural gas, has four bonding pairs of electrons around the central carbon atom; the electron-pair geometry is tetrahedral, as is the molecular structure (Figure 5.7.4). On the other hand,

the ammonia molecule, NH_3, also has four electron pairs associated with the nitrogen atom, and thus has a tetrahedral electron-pair geometry. One of these regions, however, is a lone pair, which is not included in the molecular structure, and this lone pair influences the shape of the molecule (Figure 5.7.5).

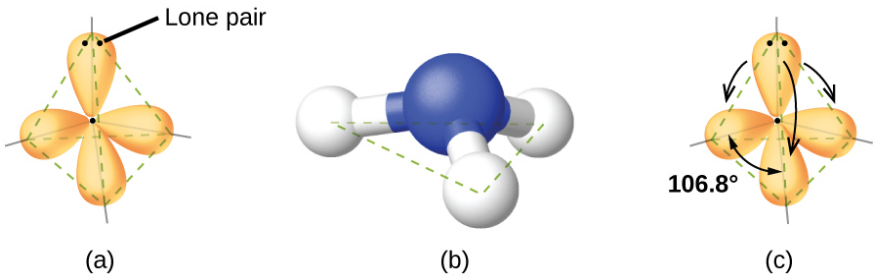


Figure 5.7.5: (a) The electron-pair geometry for the ammonia molecule is tetrahedral with one lone pair and three single bonds. (b) The trigonal pyramidal molecular structure is determined from the electron-pair geometry. (c) The actual bond angles deviate slightly from the idealized angles because the lone pair takes up a larger region of space than do the single bonds, causing the HNH angle to be slightly smaller than 109.5°.

As seen in Figure 5.7.5, small distortions from the ideal angles in Figure 5.7.3 can result from differences in repulsion between various regions of electron density. VSEPR theory predicts these distortions by establishing an order of repulsions and an order of the amount of space occupied by different kinds of electron pairs. The order of electron-pair repulsions from greatest to least repulsion is:

$$\text{lone pair-lone pair} > \text{lone pair-bonding pair} > \text{bonding pair-bonding pair}$$

This order of repulsions determines the amount of space occupied by different regions of electrons. A lone pair of electrons occupies a larger region of space than the electrons in a triple bond; in turn, electrons in a triple bond occupy more space than those in a double bond, and so on. The order of sizes from largest to smallest is:

$$\text{lone pair} > \text{triple bond} > \text{double bond} > \text{single bond}$$

Consider formaldehyde, H_2CO, which is used as a preservative for biological and anatomical specimens (Figure 5.7.1). This molecule has regions of high electron density that consist of two single bonds and one double bond. The basic geometry is trigonal planar with 120° bond angles, but we see that the double bond causes slightly larger angles (121°), and the angle between the single bonds is slightly smaller (118°).

In the ammonia molecule, the three hydrogen atoms attached to the central nitrogen are not arranged in a flat, trigonal planar molecular structure, but rather in a three-dimensional trigonal pyramid (Figure 5.7.5) with the nitrogen atom at the apex and the three hydrogen atoms forming the base. The ideal bond angles in a trigonal pyramid are based on the tetrahedral electron pair geometry. Again, there are slight deviations from the ideal because lone pairs occupy larger regions of space than do bonding electrons. The H–N–H bond angles in NH_3 are slightly smaller than the 109.5° angle in a regular tetrahedron (Figure 5.7.3) because the lone pair-bonding pair repulsion is greater than the bonding pair-bonding pair repulsion (Figure 5.7.5). Figure 5.7.6 illustrates the ideal molecular structures, which are predicted based on the electron-pair geometries for various combinations of lone pairs and bonding pairs.

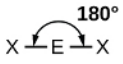
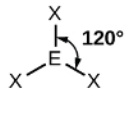
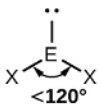
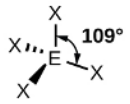
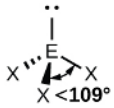
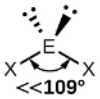
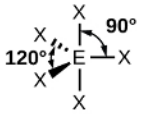
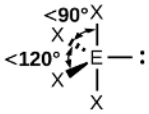
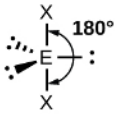
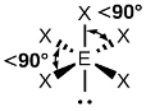
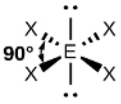
Number of electron regions	Electron region geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramid	 Square planar	 T-shape	 Linear

Figure 5.7.6: The molecular structures are identical to the electron-pair geometries when there are no lone pairs present (first column). For a particular number of electron pairs (row), the molecular structures for one or more lone pairs are determined based on modifications of the corresponding electron-pair geometry.

According to VSEPR theory, the terminal atom locations (Xs in Figure 5.7.6) are equivalent within the linear, trigonal planar, and tetrahedral electron-pair geometries (the first three rows of the table). It does not matter which X is replaced with a lone pair because the molecules can be rotated to convert positions. For trigonal bipyramidal electron-pair geometries, however, there are two distinct X positions, as shown in Figure 5.7.7 an axial position (if we hold a model of a trigonal bipyramid by the two axial positions, we have an axis around which we can rotate the model) and an equatorial position (three positions form an equator around the middle of the molecule). As shown in Figure 5.7.6, the axial position is surrounded by bond angles of 90°, whereas the equatorial position has more space available because of the 120° bond angles. In a trigonal bipyramidal electron-pair geometry, lone pairs always occupy equatorial positions because these more spacious positions can more easily accommodate the larger lone pairs.

Theoretically, we can come up with three possible arrangements for the three bonds and two lone pairs for the ClF_3 molecule (Figure 5.7.7). The stable structure is the one that puts the lone pairs in equatorial locations, giving a T-shaped molecular structure.

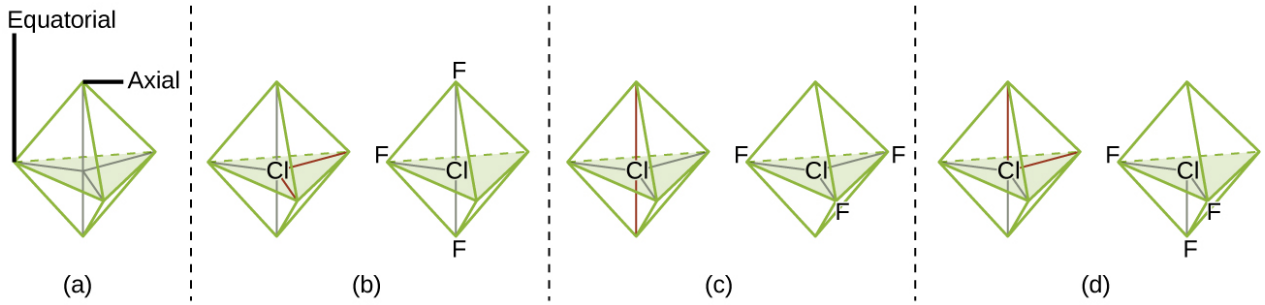


Figure 5.7.7: (a) In a trigonal bipyramid, the two axial positions are located directly across from one another, whereas the three equatorial positions are located in a triangular arrangement. (b–d) The two lone pairs (red lines) in ClF_3 have several possible arrangements, but the T-shaped molecular structure (b) is the one actually observed, consistent with the larger lone pairs both occupying equatorial positions.

When a central atom has two lone electron pairs and four bonding regions, we have an octahedral electron-pair geometry. The two lone pairs are on opposite sides of the octahedron (180° apart), giving a square planar molecular structure that minimizes lone pair-lone pair repulsions (Figure 5.7.6).

Predicting Electron Pair Geometry and Molecular Structure

The following procedure uses VSEPR theory to determine the electron pair geometries and the molecular structures:

1. Write the Lewis structure of the molecule or polyatomic ion.
2. Count the number of regions of electron density (lone pairs and bonds) around the central atom. A single, double, or triple bond counts as one region of electron density.
3. Identify the electron-pair geometry based on the number of regions of electron density: linear, trigonal planar, tetrahedral, trigonal bipyramidal, or octahedral (Figure 5.7.6, first column).
4. Use the number of lone pairs to determine the molecular structure (Figure 5.7.6). If more than one arrangement of lone pairs and chemical bonds is possible, choose the one that will minimize repulsions, remembering that lone pairs occupy more space than multiple bonds, which occupy more space than single bonds. In trigonal bipyramidal arrangements, repulsion is minimized when every lone pair is in an equatorial position. In an octahedral arrangement with two lone pairs, repulsion is minimized when the lone pairs are on opposite sides of the central atom.

The following examples illustrate the use of VSEPR theory to predict the molecular structure of molecules or ions that have no lone pairs of electrons. In this case, the molecular structure is identical to the electron pair geometry.

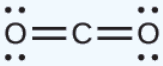
Example 5.7.1: Predicting Electron-pair Geometry and Molecular Structure: CO_2 and BCl_3

Predict the electron-pair geometry and molecular structure for each of the following:

a. carbon dioxide, CO_2, a molecule produced by the combustion of fossil fuels
b. boron trichloride, BCl_3, an important industrial chemical

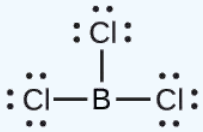
Solution

(a) We write the Lewis structure of CO_2 as:



This shows us two regions of high electron density around the carbon atom—each double bond counts as one region, and there are no lone pairs on the carbon atom. Using VSEPR theory, we predict that the two regions of electron density arrange themselves on opposite sides of the central atom with a bond angle of 180°. The electron-pair geometry and molecular structure are identical, and CO_2 molecules are linear.

(b) We write the Lewis structure of BCl_3 as:



Thus we see that BCl_3 contains three bonds, and there are no lone pairs of electrons on boron. The arrangement of three regions of high electron density gives a trigonal planar electron-pair geometry. The B–Cl bonds lie in a plane with 120° angles between them. BCl_3 also has a trigonal planar molecular structure (Figure 5.7.8).

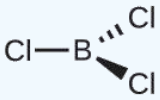


Figure 5.7.8:

The electron-pair geometry and molecular structure of BCl_3 are both trigonal planar. Note that the VSEPR geometry indicates the correct bond angles (120°), unlike the Lewis structure shown above.

? Exercise 5.7.1

Carbonate, is a common polyatomic ion found in various materials from eggshells to antacids. What are the electron-pair geometry and molecular structure of this polyatomic ion?

Answer

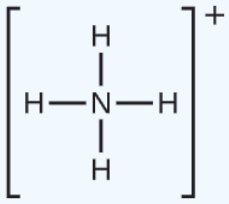
The electron-pair geometry is trigonal planar and the molecular structure is trigonal planar. Due to resonance, all three C–O bonds are identical. Whether they are single, double, or an average of the two, each bond counts as one region of electron density.

✓ Example 5.7.2: Predicting Electron-pair Geometry and Molecular Structure: Ammonium

Two of the top 50 chemicals produced in the United States, ammonium nitrate and ammonium sulfate, both used as fertilizers, contain the ammonium ion. Predict the electron-pair geometry and molecular structure of the cation.

Solution

We write the Lewis structure of as:



We can see that NH_4^+ contains four bonds from the nitrogen atom to hydrogen atoms and no lone pairs. We expect the four regions of high electron density to arrange themselves so that they point to the corners of a tetrahedron with the central nitrogen atom in the middle (Figure 5.7.6). Therefore, the electron pair geometry of Figure 5.7.9.

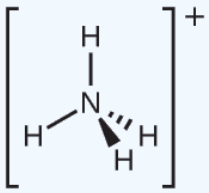


Figure 5.7.9: The ammonium ion displays a tetrahedral electron-pair geometry as well as a tetrahedral molecular structure.

Exercise 5.7.2

Identify a molecule with trigonal bipyramidal molecular structure.

Answer

Any molecule with five electron pairs around the central atoms including no lone pairs will be trigonal bipyramidal. PF_5 is a common example.

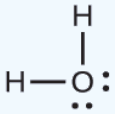
The next several examples illustrate the effect of lone pairs of electrons on molecular structure.

Example 5.7.3: Predicting Electron-pair Geometry and Molecular Structure: Lone Pairs on the Central Atom

Predict the electron-pair geometry and molecular structure of a water molecule.

Solution

The Lewis structure of H_2O indicates that there are four regions of high electron density around the oxygen atom: two lone pairs and two chemical bonds:



We predict that these four regions are arranged in a tetrahedral fashion (Figure 5.7.10), as indicated in Figure 5.7.6. Thus, the electron-pair geometry is tetrahedral and the molecular structure is bent with an angle slightly less than 109.5°. In fact, the bond angle is 104.5°.

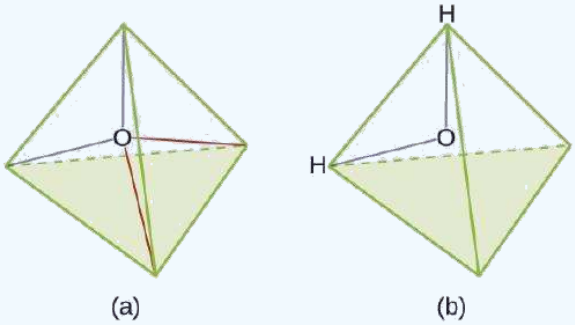


Figure 5.7.10: (a) H_2O has four regions of electron density around the central atom, so it has a tetrahedral electron-pair geometry. (b) Two of the electron regions are lone pairs, so the molecular structure is bent.

Exercise 5.7.3

The hydronium ion, H_3O^+, forms when acids are dissolved in water. Predict the electron-pair geometry and molecular structure of this cation.

Answer

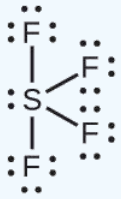
electron pair geometry: tetrahedral; molecular structure: trigonal pyramidal

Example 5.7.4: Predicting Electron-pair Geometry and Molecular Structure: SF_4

Sulfur tetrafluoride, SF_4, is extremely valuable for the preparation of fluorine-containing compounds used as herbicides (i.e., SF_4 is used as a fluorinating agent). Predict the electron-pair geometry and molecular structure of a SF_4 molecule.

Solution

The Lewis structure of SF_4 indicates five regions of electron density around the sulfur atom: one lone pair and four bonding pairs:



We expect these five regions to adopt a trigonal bipyramidal electron-pair geometry. To minimize lone pair repulsions, the lone pair occupies one of the equatorial positions. The molecular structure (Figure 5.7.11) is that of a seesaw (Figure 5.7.6).

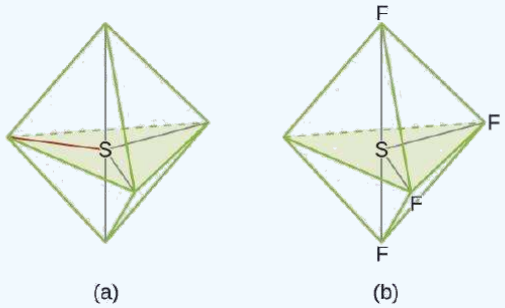


Figure 5.7.11: (a) SF4 has a trigonal bipyramidal arrangement of the five regions of electron density. (b) One of the regions is a lone pair, which results in a seesaw-shaped molecular structure.

Exercise 5.7.4

Predict the electron pair geometry and molecular structure for molecules of XeF_2.

Answer

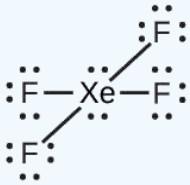
The electron-pair geometry is trigonal bipyramidal. The molecular structure is linear.

Example 5.7.5: Predicting Electron-pair Geometry and Molecular Structure: XeF_4

Of all the noble gases, xenon is the most reactive, frequently reacting with elements such as oxygen and fluorine. Predict the electron-pair geometry and molecular structure of the XeF_4 molecule.

Solution

The Lewis structure of XeF_4 indicates six regions of high electron density around the xenon atom: two lone pairs and four bonds:



These six regions adopt an octahedral arrangement (Figure 5.7.6), which is the electron-pair geometry. To minimize repulsions, the lone pairs should be on opposite sides of the central atom (Figure 5.7.12). The five atoms are all in the same plane and have a square planar molecular structure.

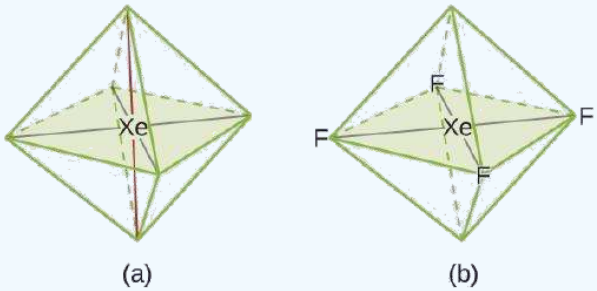


Figure 5.7.12: (a) XeF_4 adopts an octahedral arrangement with two lone pairs (red lines) and four bonds in the electron-pair geometry. (b) The molecular structure is square planar with the lone pairs directly across from one another.

Exercise 5.7.5

In a certain molecule, the central atom has three lone pairs and two bonds. What will the electron pair geometry and molecular structure be?

Answer

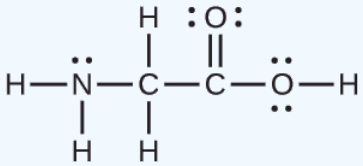
electron pair geometry: trigonal bipyramidal; molecular structure: linear

Molecular Structure for Multicenter Molecules

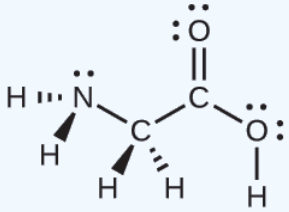
When a molecule or polyatomic ion has only one central atom, the molecular structure completely describes the shape of the molecule. Larger molecules do not have a single central atom, but are connected by a chain of interior atoms that each possess a "local" geometry. The way these local structures are oriented with respect to each other also influences the molecular shape, but such considerations are largely beyond the scope of this introductory discussion. For our purposes, we will only focus on determining the local structures.

Example 5.7.6: Predicting Structure in Multicenter Molecules

The Lewis structure for the simplest amino acid, glycine, $H_2NCH_2CO_2H$, is shown here. Predict the local geometry for the nitrogen atom, the two carbon atoms, and the oxygen atom with a hydrogen atom attached:



Solution



Consider each central atom independently. The electron-pair geometries:

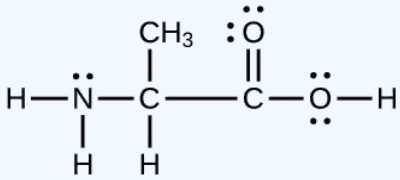
- nitrogen—four regions of electron density; tetrahedral
- carbon ($\underline{C}H_2$)—four regions of electron density; tetrahedral
- carbon ($\underline{C}O_2$)—three regions of electron density; trigonal planar
- oxygen ($\underline{O}H$)—four regions of electron density; tetrahedral

The local structures:

- nitrogen—three bonds, one lone pair; trigonal pyramidal
- carbon ($\underline{C}H_2$)—four bonds, no lone pairs; tetrahedral
- carbon ($\underline{C}O_2$)—three bonds (double bond counts as one bond), no lone pairs; trigonal planar
- oxygen ($\underline{O}H$)—two bonds, two lone pairs; bent (109°)

Exercise 5.7.6

Another amino acid is alanine, which has the Lewis structure shown here. Predict the electron-pair geometry and local structure of the nitrogen atom, the three carbon atoms, and the oxygen atom with hydrogen attached:



Answer

electron-pair geometries: nitrogen—tetrahedral; carbon ($\underline{C}H$)—tetrahedral; carbon ($\underline{C}H_3$)—tetrahedral; carbon ($\underline{C}O_2$)—trigonal planar; oxygen ($\underline{O}H$)—tetrahedral; local structures: nitrogen—trigonal pyramidal; carbon ($\underline{C}H$)—tetrahedral; carbon ($\underline{C}H_3$)—tetrahedral; carbon ($\underline{C}O_2$)—trigonal planar; oxygen ($\underline{O}H$)—bent (109°)

Link to Learning

The molecular shape simulator lets you build various molecules and practice naming their electron-pair geometries and molecular structures.

Example 5.7.7: Molecular Simulation

Using molecular shape simulator allows us to control whether bond angles and/or lone pairs are displayed by checking or unchecking the boxes under “Options” on the right. We can also use the “Name” checkboxes at bottom-left to display or hide the electron pair geometry (called “electron geometry” in the simulator) and/or molecular structure (called “molecular shape” in the simulator).

Build the molecule HCN in the simulator based on the following Lewis structure:

$$H{-}C{\equiv}N$$

Click on each bond type or lone pair at right to add that group to the central atom. Once you have the complete molecule, rotate it to examine the predicted molecular structure. What molecular structure is this?

Solution

The molecular structure is linear.

Exercise 5.7.7

Build a more complex molecule in the simulator. Identify the electron-group geometry, molecular structure, and bond angles. Then try to find a chemical formula that would match the structure you have drawn.

Answer

Answers will vary. For example, an atom with four single bonds, a double bond, and a lone pair has an octahedral electron-group geometry and a square pyramidal molecular structure. $XeOF_4$ is a molecule that adopts this structure.

Molecular Polarity and Dipole Moment

As discussed previously, polar covalent bonds connect two atoms with differing electronegativities, leaving one atom with a partial positive charge (δ+) and the other atom with a partial negative charge (δ–), as the electrons are pulled toward the more electronegative atom. This separation of charge gives rise to a bond dipole moment. The magnitude of a bond dipole moment is represented by the Greek letter mu (μ) and is given by the formula shown here, where Q is the magnitude of the partial charges (determined by the electronegativity difference) and r is the distance between the charges:

$$\mu = Qr$$

This bond moment can be represented as a vector, a quantity having both direction and magnitude (Figure 5.7.13). Dipole vectors are shown as arrows pointing along the bond from the less electronegative atom toward the more electronegative atom. A small plus sign is drawn on the less electronegative end to indicate the partially positive end of the bond. The length of the arrow is proportional to the magnitude of the electronegativity difference between the two atoms.

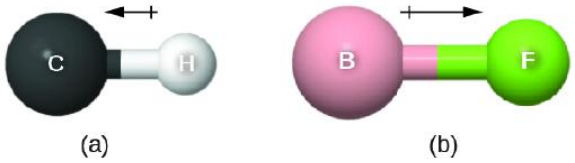


Figure 5.7.13: (a) There is a small difference in electronegativity between C and H, represented as a short vector. (b) The electronegativity difference between B and F is much larger, so the vector representing the bond moment is much longer.

A whole molecule may also have a separation of charge, depending on its molecular structure and the polarity of each of its bonds. If such a charge separation exists, the molecule is said to be a polar molecule (or dipole); otherwise the molecule is said to be nonpolar. The dipole moment measures the extent of net charge separation in the molecule as a whole. We determine the dipole moment by adding the bond moments in three-dimensional space, taking into account the molecular structure.

For diatomic molecules, there is only one bond, so its bond dipole moment determines the molecular polarity. Homonuclear diatomic molecules such as Br_2 and N_2 have no difference in electronegativity, so their dipole moment is zero. For heteronuclear molecules such as CO, there is a small dipole moment. For HF, there is a larger dipole moment because there is a larger difference in electronegativity.

When a molecule contains more than one bond, the geometry must be taken into account. If the bonds in a molecule are arranged such that their bond moments cancel (vector sum equals zero), then the molecule is nonpolar. This is the situation in CO_2 (Figure 5.7.14). Each of the bonds is polar, but the molecule as a whole is nonpolar. From the Lewis structure, and using VSEPR theory, we determine that the CO_2 molecule is linear with polar C=O bonds on opposite sides of the carbon atom. The bond moments cancel because they are pointed in opposite directions. In the case of the water molecule (Figure 5.7.14), the Lewis structure again shows that there are two bonds to a central atom, and the electronegativity difference again shows that each of these bonds has a nonzero bond moment. In this case, however, the molecular structure is bent because of the lone pairs on O, and the two bond moments do not cancel. Therefore, water does have a net dipole moment and is a polar molecule (dipole).

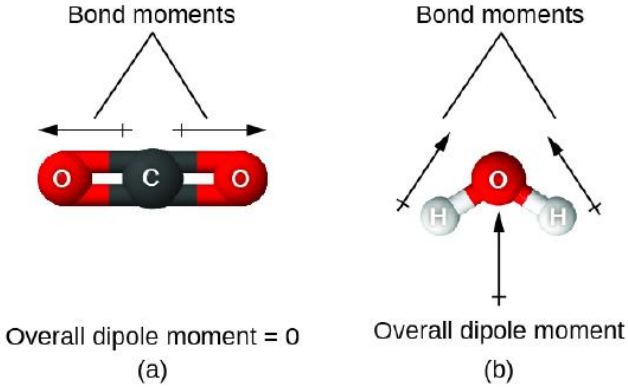
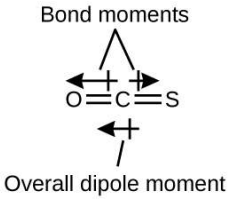


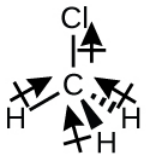
Figure 5.7.14: The overall dipole moment of a molecule depends on the individual bond dipole moments and how they are arranged. (a) Each CO bond has a bond dipole moment, but they point in opposite directions so that the net CO_2 molecule is nonpolar. (b) In contrast, water is polar because the OH bond moments do not cancel out.

The OCS molecule has a structure similar to CO_2, but a sulfur atom has replaced one of the oxygen atoms. To determine if this molecule is polar, we draw the molecular structure. VSEPR theory predicts a linear molecule:

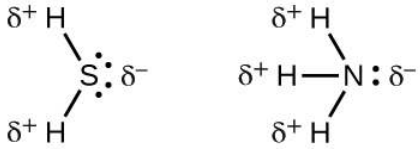


The C-O bond is considerably polar. Although C and S have very similar electronegativity values, S is slightly more electronegative than C, and so the C-S bond is just slightly polar. Because oxygen is more electronegative than sulfur, the oxygen end of the molecule is the negative end.

Chloromethane, CH_3Cl, is a tetrahedral molecule with three slightly polar C-H bonds and a more polar C-Cl bond. The relative electronegativities of the bonded atoms is H < C < Cl, and so the bond moments all point toward the Cl end of the molecule and sum to yield a considerable dipole moment (the molecules are relatively polar).



For molecules of high symmetry such as BF_3 (trigonal planar), CH_4 (tetrahedral), PF_5 (trigonal bipyramidal), and SF_6 (octahedral), all the bonds are of identical polarity (same bond moment) and they are oriented in geometries that yield nonpolar molecules (dipole moment is zero). Molecules of less geometric symmetry, however, may be polar even when all bond moments are identical. For these molecules, the directions of the equal bond moments are such that they sum to give a nonzero dipole moment and a polar molecule. Examples of such molecules include hydrogen sulfide, H_2S (nonlinear), and ammonia, NH_3 (trigonal pyramidal).



To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment does not cancel.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure 5.7.15). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

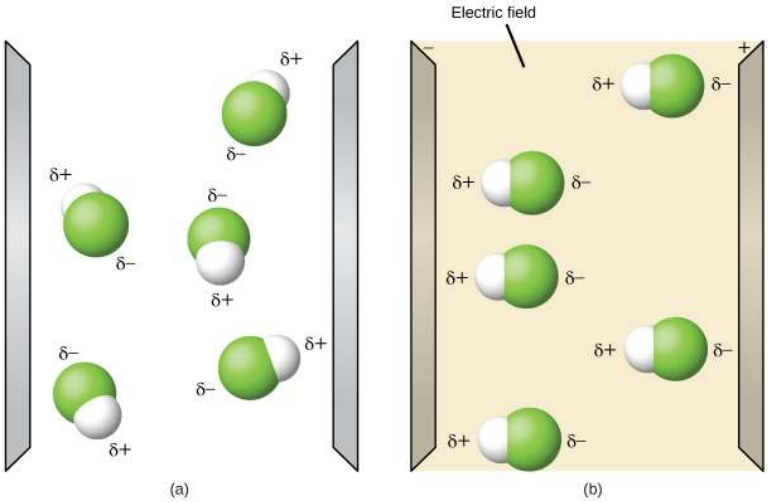


Figure 5.7.15: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction.

Link to Learning

The molecule polarity simulation provides many ways to explore dipole moments of bonds and molecules.

Example 5.7.8: Polarity Simulations

Open the molecule polarity simulation and select the "Three Atoms" tab at the top. This should display a molecule ABC with three electronegativity adjustors. You can display or hide the bond moments, molecular dipoles, and partial charges at the right. Turning on the Electric Field will show whether the molecule moves when exposed to a field, similar to Figure 5.7.15

Use the electronegativity controls to determine how the molecular dipole will look for the starting bent molecule if:

a. A and C are very electronegative and B is in the middle of the range.
b. A is very electronegative, and B and C are not.

Solution

a. Molecular dipole moment points immediately between A and C.
b. Molecular dipole moment points along the A–B bond, toward A.

Exercise 5.7.8

Determine the partial charges that will give the largest possible bond dipoles.

Answer

The largest bond moments will occur with the largest partial charges. The two solutions above represent how unevenly the electrons are shared in the bond. The bond moments will be maximized when the electronegativity difference is greatest. The controls for A and C should be set to one extreme, and B should be set to the opposite extreme. Although the magnitude of the bond moment will not change based on whether B is the most electronegative or the least, the direction of the bond moment will.

This page titled 5.7: Molecular Structure and Polarity is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **7.6: Molecular Structure and Polarity** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

5.8: Key Terms

axial position
location in a trigonal bipyramidal geometry in which there is another atom at a 180° angle and the equatorial positions are at a 90° angle

binary acid
compound that contains hydrogen and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions when dissolved in water)

binary compound
compound containing two different elements.

bond angle
angle between any two covalent bonds that share a common atom

bond dipole moment
separation of charge in a bond that depends on the difference in electronegativity and the bond distance represented by partial charges or a vector

bond distance
(also, bond length) distance between the nuclei of two bonded atoms

bond length
distance between the nuclei of two bonded atoms at which the lowest potential energy is achieved

covalent bond
bond formed when electrons are shared between atoms

dipole moment
property of a molecule that describes the separation of charge determined by the sum of the individual bond moments based on the molecular structure

double bond
covalent bond in which two pairs of electrons are shared between two atoms

electron-pair geometry
arrangement around a central atom of all regions of electron density (bonds, lone pairs, or unpaired electrons)

electronegativity
tendency of an atom to attract electrons in a bond to itself

equatorial position
one of the three positions in a trigonal bipyramidal geometry with 120° angles between them; the axial positions are located at a 90° angle

formal charge
charge that would result on an atom by taking the number of valence electrons on the neutral atom and subtracting the nonbonding electrons and the number of bonds (one-half of the bonding electrons)

free radical
molecule that contains an odd number of electrons

hypervalent molecule

molecule containing at least one main group element that has more than eight electrons in its valence shell

inert pair effect
tendency of heavy atoms to form ions in which their valence *s* electrons are not lost

ionic bond
strong electrostatic force of attraction between cations and anions in an ionic compound

Lewis structure
diagram showing lone pairs and bonding pairs of electrons in a molecule or an ion

Lewis symbol
symbol for an element or monatomic ion that uses a dot to represent each valence electron in the element or ion

linear
shape in which two outside groups are placed on opposite sides of a central atom

lone pair
two (a pair of) valence electrons that are not used to form a covalent bond

molecular structure
arrangement of atoms in a molecule or ion

molecular structure
structure that includes only the placement of the atoms in the molecule

nomenclature
system of rules for naming objects of interest

octahedral
shape in which six outside groups are placed around a central atom such that a three-dimensional shape is generated with four groups forming a square and the other two forming the apex of two pyramids, one above and one below the square plane

octet rule
guideline that states main group atoms will form structures in which eight valence electrons interact with each nucleus, counting bonding electrons as interacting with both atoms connected by the bond

oxyacid
compound that contains hydrogen, oxygen, and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions when dissolved in water)

polar covalent bond
covalent bond between atoms of different electronegativities; a covalent bond with a positive end and a negative end

polar molecule
(also, dipole) molecule with an overall dipole moment

pure covalent bond
(also, nonpolar covalent bond) covalent bond between atoms of identical electronegativities

resonance
situation in which one Lewis structure is insufficient to describe the bonding in a molecule and the average of multiple structures is observed

resonance forms

two or more Lewis structures that have the same arrangement of atoms but different arrangements of electrons

resonance hybrid
average of the resonance forms shown by the individual Lewis structures

single bond
bond in which a single pair of electrons is shared between two atoms

tetrahedral
shape in which four outside groups are placed around a central atom such that a three-dimensional shape is generated with four corners and 109.5° angles between each pair and the central atom

trigonal bipyramidal
shape in which five outside groups are placed around a central atom such that three form a flat triangle with 120° angles between each pair and the central atom, and the other two form the apex of two pyramids, one above and one below the triangular plane

trigonal planar
shape in which three outside groups are placed in a flat triangle around a central atom with 120° angles between each pair and the central atom

triple bond
bond in which three pairs of electrons are shared between two atoms

valence shell electron-pair repulsion theory (VSEPR)
theory used to predict the bond angles in a molecule based on positioning regions of high electron density as far apart as possible to minimize electrostatic repulsion

vector
quantity having magnitude and direction

This page titled 5.8: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

5.9: Key Equations

$$\text{formal charge} = \#\text{ valence shell electrons (free atom)} - \#\text{ lone pair electrons} - \tfrac{1}{2}\ \#\text{ bonding}$$

This page titled 5.9: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

5.10: Summary

Atoms gain or lose electrons to form ions with particularly stable electron configurations. The charges of cations formed by the representative metals may be determined readily because, with few exceptions, the electronic structures of these ions have either a noble gas configuration or a completely filled electron shell. The charges of anions formed by the nonmetals may also be readily determined because these ions form when nonmetal atoms gain enough electrons to fill their valence shells.

Covalent bonds form when electrons are shared between atoms and are attracted by the nuclei of both atoms. In pure covalent bonds, the electrons are shared equally. In polar covalent bonds, the electrons are shared unequally, as one atom exerts a stronger force of attraction on the electrons than the other. The ability of an atom to attract a pair of electrons in a chemical bond is called its electronegativity. The difference in electronegativity between two atoms determines how polar a bond will be. In a diatomic molecule with two identical atoms, there is no difference in electronegativity, so the bond is nonpolar or pure covalent. When the electronegativity difference is very large, as is the case between metals and nonmetals, the bonding is characterized as ionic.

Chemists use nomenclature rules to clearly name compounds. Ionic and molecular compounds are named using somewhat-different methods. Binary ionic compounds typically consist of a metal and a nonmetal. The name of the metal is written first, followed by the name of the nonmetal with its ending changed to *–ide*. For example, K_2O is called potassium oxide. If the metal can form ions with different charges, a Roman numeral in parentheses follows the name of the metal to specify its charge. Thus, $FeCl_2$ is iron(II) chloride and $FeCl_3$ is iron(III) chloride. Some compounds contain polyatomic ions; the names of common polyatomic ions should be memorized. Molecular compounds can form compounds with different ratios of their elements, so prefixes are used to specify the numbers of atoms of each element in a molecule of the compound. Examples include SF_6, sulfur hexafluoride, and N_2O_4, dinitrogen tetroxide. Acids are an important class of compounds containing hydrogen and having special nomenclature rules. Binary acids are named using the prefix *hydro-*, changing the *–ide* suffix to *–ic*, and adding "acid;" HCl is hydrochloric acid. Oxyacids are named by changing the ending of the anion (*–ate* to *–ic* and *–ite* to *–ous*), and adding "acid;" H_2CO_3 is carbonic acid.

Valence electronic structures can be visualized by drawing Lewis symbols (for atoms and monatomic ions) and Lewis structures (for molecules and polyatomic ions). Lone pairs, unpaired electrons, and single, double, or triple bonds are used to indicate where the valence electrons are located around each atom in a Lewis structure. Most structures—especially those containing second row elements—obey the octet rule, in which every atom (except H) is surrounded by eight electrons. Exceptions to the octet rule occur for odd-electron molecules (free radicals), electron-deficient molecules, and hypervalent molecules.

In a Lewis structure, formal charges can be assigned to each atom by treating each bond as if one-half of the electrons are assigned to each atom. These hypothetical formal charges are a guide to determining the most appropriate Lewis structure. A structure in which the formal charges are as close to zero as possible is preferred. Resonance occurs in cases where two or more Lewis structures with identical arrangements of atoms but different distributions of electrons can be written. The actual distribution of electrons (the resonance hybrid) is an average of the distribution indicated by the individual Lewis structures (the resonance forms).

VSEPR theory predicts the three-dimensional arrangement of atoms in a molecule. It states that valence electrons will assume an electron-pair geometry that minimizes repulsions between areas of high electron density (bonds and/or lone pairs). Molecular structure, which refers only to the placement of atoms in a molecule and not the electrons, is equivalent to electron-pair geometry only when there are no lone electron pairs around the central atom. A dipole moment measures a separation of charge. For one bond, the bond dipole moment is determined by the difference in electronegativity between the two atoms. For a molecule, the overall dipole moment is determined by both the individual bond moments and how these dipoles are arranged in the molecular structure. Polar molecules (those with an appreciable dipole moment) interact with electric fields, whereas nonpolar molecules do not.

This page titled 5.10: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

5.11: Exercises

1.

Does a cation gain protons to form a positive charge or does it lose electrons?

2.

Iron(III) sulfate [$Fe_2(SO_4)_3$] is composed of Fe^{3+} and $SO_4{}^{2-}$ ions. Explain why a sample of iron(III) sulfate is uncharged.

3.

Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: P, I, Mg, Cl, In, Cs, O, Pb, Co?

4.

Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: Br, Ca, Na, N, F, Al, Sn, S, Cd?

5.

Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

(a) P

(b) Mg

(c) Al

(d) O

(e) Cl

(f) Cs

6.

Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

(a) I

(b) Sr

(c) K

(d) N

(e) S

(f) In

7.

Write the electron configuration for each of the following ions:

(a) As^{3-}

(b) I^-

(c) Be^{2+}

(d) Cd^{2+}

(e) O^{2-}

(f) Ga^{3+}

(g) Li^{+}

(h) N^{3-}

(i) Sn^{2+}

(j) Co^{2+}

(k) Fe^{2+}

(l) As^{3+}

8.

Write the electron configuration for the monatomic ions formed from the following elements (which form the greatest concentration of monatomic ions in seawater):

(a) Cl

(b) Na

(c) Mg

(d) Ca

(e) K

(f) Br

(g) Sr

(h) F

9.

Write out the full electron configuration for each of the following atoms and for the monatomic ion found in binary ionic compounds containing the element:

(a) Al

(b) Br

(c) Sr

(d) Li

(e) As

(f) S

10.

From the labels of several commercial products, prepare a list of six ionic compounds in the products. For each compound, write the formula. (You may need to look up some formulas in a suitable reference.)

11.

Why is it incorrect to speak of a molecule of solid NaCl?

12.

What information can you use to predict whether a bond between two atoms is covalent or ionic?

13.

Predict which of the following compounds are ionic and which are covalent, based on the location of their constituent atoms in the periodic table:

(a) Cl_2CO

(b) MnO

(c) NCl_3

(d) $CoBr_2$

(e) K_2S

(f) CO

(g) CaF_2

(h) HI

(i) CaO

(j) IBr

(k) CO_2

14.

Explain the difference between a nonpolar covalent bond, a polar covalent bond, and an ionic bond.

15.

From its position in the periodic table, determine which atom in each pair is more electronegative:

(a) Br or Cl

(b) N or O

(c) S or O

(d) P or S

(e) Si or N

(f) Ba or P

(g) N or K

16.

From its position in the periodic table, determine which atom in each pair is more electronegative:

(a) N or P

(b) N or Ge

(c) S or F

(d) Cl or S

(e) H or C

(f) Se or P

(g) C or Si

17.

From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:

(a) C, F, H, N, O

(b) Br, Cl, F, H, I

(c) F, H, O, P, S

(d) Al, H, Na, O, P

(e) Ba, H, N, O, As

18.

From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:

(a) As, H, N, P, Sb

(b) Cl, H, P, S, Si

(c) Br, Cl, Ge, H, Sr

(d) Ca, H, K, N, Si

(e) Cl, Cs, Ge, H, Sr

19.

Which atoms can bond to sulfur so as to produce a positive partial charge on the sulfur atom?

20.

Which is the most polar bond?

(a) C–C

(b) C–H

(c) N–H

(d) O–H

(e) Se–H

21.

Identify the more polar bond in each of the following pairs of bonds:

(a) HF or HCl

(b) NO or CO

(c) SH or OH

(d) PCl or SCl

(e) CH or NH

(f) SO or PO

(g) CN or NN

22.

Which of the following molecules or ions contain polar bonds?

(a) O_3

(b) S_8

(c) O_2^{2-}

(d) NO_3^-

(e) CO_2

(f) H_2S

(g) BH_4^-

23.

Name the following compounds:

(a) CsCl

(b) BaO

(c) K_2S

(d) $BeCl_2$

(e) HBr

(f) AlF_3

24.

Name the following compounds:

(a) NaF

(b) Rb_2O

(c) BCl_3

(d) H_2Se

(e) P_4O_6

(f) ICl_3

25.

Write the formulas of the following compounds:

(a) rubidium bromide

(b) magnesium selenide

(c) sodium oxide

(d) calcium chloride

(e) hydrogen fluoride

(f) gallium phosphide

(g) aluminum bromide

(h) ammonium sulfate

26.

Write the formulas of the following compounds:

(a) lithium carbonate

(b) sodium perchlorate

(c) barium hydroxide

(d) ammonium carbonate

(e) sulfuric acid

(f) calcium acetate

(g) magnesium phosphate

(h) sodium sulfite

27.

Write the formulas of the following compounds:

(a) chlorine dioxide

(b) dinitrogen tetraoxide

(c) potassium phosphide

(d) silver(I) sulfide

(e) aluminum fluoride trihydrate

(f) silicon dioxide

28.

Write the formulas of the following compounds:

(a) barium chloride

(b) magnesium nitride

(c) sulfur dioxide

(d) nitrogen trichloride

(e) dinitrogen trioxide

(f) tin(IV) chloride

29.

Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

(a) Cr_2O_3

(b) $FeCl_2$

(c) CrO_3

(d) $TiCl_4$

(e) $CoCl_2 \cdot 6H_2O$

(f) MoS_2

30.

Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

(a) $NiCO_3$

(b) MoO_3

(c) $Co(NO_3)_2$

(d) V_2O_5

(e) MnO_2

(f) Fe_2O_3

31.

The following ionic compounds are found in common household products. Write the formulas for each compound:

(a) potassium phosphate

(b) copper(II) sulfate

(c) calcium chloride

(d) titanium(IV) oxide

(e) ammonium nitrate

(f) sodium bisulfate (the common name for sodium hydrogen sulfate)

32.

The following ionic compounds are found in common household products. Name each of the compounds:

(a) $Ca(H_2PO_4)_2$

(b) $FeSO_4$

(c) $CaCO_3$

(d) MgO

(e) $NaNO_2$

(f) KI

33.

What are the IUPAC names of the following compounds?

(a) manganese dioxide

(b) mercurous chloride (Hg_2Cl_2)

(c) ferric nitrate [$Fe(NO_3)_3$]

(d) titanium tetrachloride

(e) cupric bromide ($CuBr_2$)

34.

Write the Lewis symbols for each of the following ions:

(a) As^{3-}

(b) I^-

(c) Be^{2+}

(d) O^{2-}

(e) Ga^{3+}

(f) Li^+

(g) N^{3-}

35.

Many monatomic ions are found in seawater, including the ions formed from the following list of elements. Write the Lewis symbols for the monatomic ions formed from the following elements:

(a) Cl

(b) Na

(c) Mg

(d) Ca

(e) K

(f) Br

(g) Sr

(h) F

36.

Write the Lewis symbols of the ions in each of the following ionic compounds and the Lewis symbols of the atom from which they are formed:

(a) MgS

(b) Al_2O_3

(c) $GaCl_3$

(d) K_2O

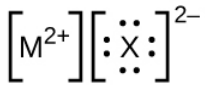
(e) Li_3N

(f) KF

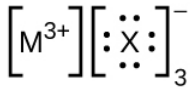
37.

In the Lewis structures listed here, M and X represent various elements in the third period of the periodic table. Write the formula of each compound using the chemical symbols of each element:

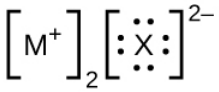
(a)



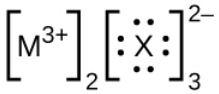
(b)



(c)



(d)



38.

Write the Lewis structure for the diatomic molecule P_2, an unstable form of phosphorus found in high-temperature phosphorus vapor.

39.

Write Lewis structures for the following:

(a) H_2

(b) HBr

(c) PCl_3

(d) SF_2

(e) H_2CCH_2

(f) HNNH

(g) H_2CNH

(h) NO^-

(i) N_2

(j) CO

(k) CN^-

40.

Write Lewis structures for the following:

(a) O_2

(b) H_2CO

(c) AsF_3

(d) ClNO

(e) $SiCl_4$

(f) H_3O^+

(g) $NH_4{}^+$

(h) $BF_4{}^-$

(i) HCCH

(j) ClCN

(k) $C_2{}^{2+}$

41.

Write Lewis structures for the following:

(a) ClF_3

(b) PCl_5

(c) BF_3

(d) $PF_6{}^-$

42.

Write Lewis structures for the following:

(a) SeF_6

(b) XeF_4

(c) $SeCl_3{}^+$

(d) Cl_2BBCl_2 (contains a B–B bond)

43.

Write Lewis structures for:

(a) PO_4^{3-}

(b) ICl_4^-

(c) SO_3^{2-}

(d) HONO

44.

Correct the following statement: “The bonds in solid $PbCl_2$ are ionic; the bond in a HCl molecule is covalent. Thus, all of the valence electrons in $PbCl_2$ are located on the Cl^- ions, and all of the valence electrons in a HCl molecule are shared between the H and Cl atoms.”

45.

Write Lewis structures for the following molecules or ions:

(a) SbH_3

(b) XeF_2

(c) Se_8 (a cyclic molecule with a ring of eight Se atoms)

46.

Methanol, H_3COH, is used as the fuel in some race cars. Ethanol, C_2H_5OH, is used extensively as motor fuel in Brazil. Both methanol and ethanol produce CO_2 and H_2O when they burn. Write the chemical equations for these combustion reactions using Lewis structures instead of chemical formulas.

47.

Many planets in our solar system contain organic chemicals including methane (CH_4) and traces of ethylene (C_2H_4), ethane (C_2H_6), propyne (H_3CCCH), and diacetylene (HCCCCH). Write the Lewis structures for each of these molecules.

48.

Carbon tetrachloride was formerly used in fire extinguishers for electrical fires. It is no longer used for this purpose because of the formation of the toxic gas phosgene, Cl_2CO. Write the Lewis structures for carbon tetrachloride and phosgene.

49.

Identify the atoms that correspond to each of the following electron configurations. Then, write the Lewis symbol for the common ion formed from each atom:

(a) $1s^2 2s^2 2p^5$

(b) $1s^2 2s^2 2p^6 3s^2$

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

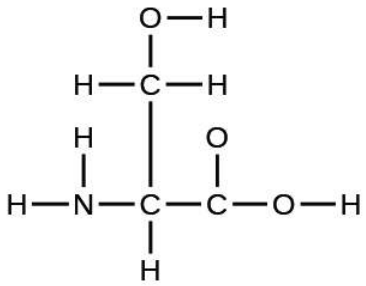
(d) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

(e) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$

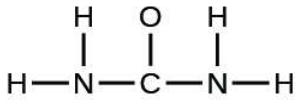
50.

The arrangement of atoms in several biologically important molecules is given here. Complete the Lewis structures of these molecules by adding multiple bonds and lone pairs. Do not add any more atoms.

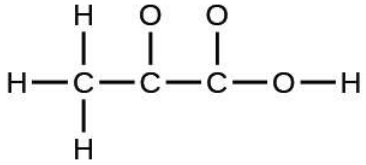
(a) the amino acid serine:



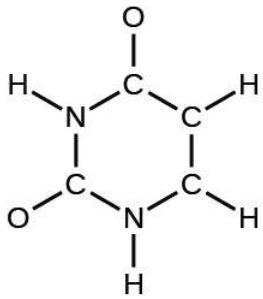
(b) urea:



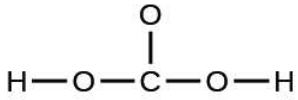
(c) pyruvic acid:



(d) uracil:



(e) carbonic acid:



51.

A compound with a molar mass of about 28 g/mol contains 85.7% carbon and 14.3% hydrogen by mass. Write the Lewis structure for a molecule of the compound.

52.

A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen by mass. Write the Lewis structure for a molecule of the compound.

53.

Two arrangements of atoms are possible for a compound with a molar mass of about 45 g/mol that contains 52.2% C, 13.1% H, and 34.7% O by mass. Write the Lewis structures for the two molecules.

54.

How are single, double, and triple bonds similar? How do they differ?

55.

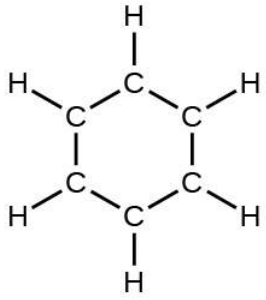
Write resonance forms that describe the distribution of electrons in each of these molecules or ions.

(a) selenium dioxide, OSeO

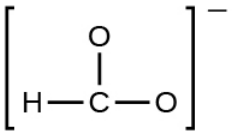
(b) nitrate ion, NO_3^-

(c) nitric acid, HNO_3 (N is bonded to an OH group and two O atoms)

(d) benzene, C_6H_6:



(e) the formate ion:



56.

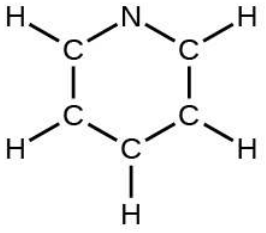
Write resonance forms that describe the distribution of electrons in each of these molecules or ions.

(a) sulfur dioxide, SO_2

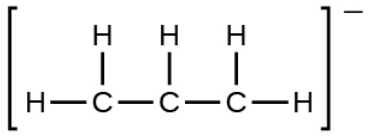
(b) carbonate ion, CO_3^{2-}

(c) hydrogen carbonate ion, HCO_3^- (C is bonded to an OH group and two O atoms)

(d) pyridine:



(e) the allyl ion:



57.

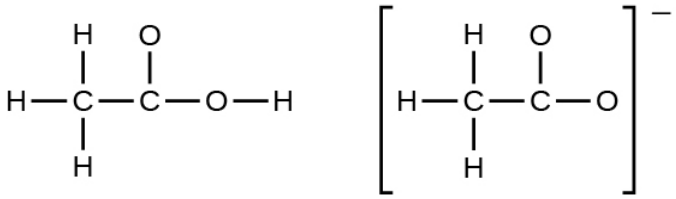
Write the resonance forms of ozone, O_3, the component of the upper atmosphere that protects the Earth from ultraviolet radiation.

58.

Sodium nitrite, which has been used to preserve bacon and other meats, is an ionic compound. Write the resonance forms of the nitrite ion, NO_2^-.

59.

In terms of the bonds present, explain why acetic acid, CH_3CO_2H, contains two distinct types of carbon-oxygen bonds, whereas the acetate ion, formed by loss of a hydrogen ion from acetic acid, only contains one type of carbon-oxygen bond. The skeleton structures of these species are shown:



60.

Write the Lewis structures for the following, and include resonance structures where appropriate. Indicate which has the strongest carbon-oxygen bond.

(a) CO_2

(b) CO

61.

Toothpastes containing sodium hydrogen carbonate (sodium bicarbonate) and hydrogen peroxide are widely used. Write Lewis structures for the hydrogen carbonate ion and hydrogen peroxide molecule, with resonance forms where appropriate.

62.

Determine the formal charge of each element in the following:

(a) HCl

(b) CF_4

(c) PCl_3

(d) PF_5

63.

Determine the formal charge of each element in the following:

(a) H_3O^+

(b) SO_4^{2-}

(c) NH_3

(d) O_2^{2-}

(e) H_2O_2

64.

Calculate the formal charge of chlorine in the molecules Cl_2, $BeCl_2$, and ClF_5.

65.

Calculate the formal charge of each element in the following compounds and ions:

(a) F_2CO

(b) NO^-

(c) BF_4^-

(d) $SnCl_3^-$

(e) H_2CCH_2

(f) ClF_3

(g) SeF_6

(h) PO_4^{3-}

66.

Draw all possible resonance structures for each of these compounds. Determine the formal charge on each atom in each of the resonance structures:

(a) O_3

(b) SO_2

(c) NO_2^-

(d) NO_3^-

67.

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in nitrosyl chloride: ClNO or ClON?

68.

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in hypochlorous acid: HOCl or OClH?

69.

Based on formal charge considerations, which of the following would likely be the correct arrangement of atoms in sulfur dioxide: OSO or SOO?

70.

Draw the structure of hydroxylamine, H_3NO, and assign formal charges; look up the structure. Is the actual structure consistent with the formal charges?

71.

Iodine forms a series of fluorides (listed here). Write Lewis structures for each of the four compounds and determine the formal charge of the iodine atom in each molecule:

(a) IF

(b) IF_3

(c) IF_5

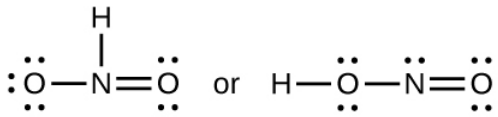
(d) IF_7

72.

Write the Lewis structure and chemical formula of the compound with a molar mass of about 70 g/mol that contains 19.7% nitrogen and 80.3% fluorine by mass, and determine the formal charge of the atoms in this compound.

73.

Which of the following structures would we expect for nitrous acid? Determine the formal charges:



74.

Sulfuric acid is the industrial chemical produced in greatest quantity worldwide. About 90 billion pounds are produced each year in the United States alone. Write the Lewis structure for sulfuric acid, H_2SO_4, which has two oxygen atoms and two OH groups bonded to the sulfur.

75.

Explain why the HOH molecule is bent, whereas the HBeH molecule is linear.

76.

What feature of a Lewis structure can be used to tell if a molecule's (or ion's) electron-pair geometry and molecular structure will be identical?

77.

Explain the difference between electron-pair geometry and molecular structure.

78.

Why is the H–N–H angle in NH_3 smaller than the H–C–H bond angle in CH_4? Why is the H–N–H angle in NH_4^+ identical to the H–C–H bond angle in CH_4?

79.

Explain how a molecule that contains polar bonds can be nonpolar.

80.

As a general rule, MX_n molecules (where M represents a central atom and X represents terminal atoms; n = 2 – 5) are polar if there is one or more lone pairs of electrons on M. NH_3 (M = N, X = H, n = 3) is an example. There are two molecular structures with lone pairs that are exceptions to this rule. What are they?

81.

Predict the electron pair geometry and the molecular structure of each of the following molecules or ions:

(a) SF_6

(b) PCl_5

(c) BeH_2

(d) CH_3^+

82.

Identify the electron pair geometry and the molecular structure of each of the following molecules or ions:

(a) IF_6^+

(b) CF_4

(c) BF_3

(d) SiF_5^-

(e) $BeCl_2$

83.

What are the electron-pair geometry and the molecular structure of each of the following molecules or ions?

(a) ClF_5

(b) ClO_2^-

(c) $TeCl_4^{2-}$

(d) PCl_3

(e) SeF_4

(f) PH_2^-

84.

Predict the electron pair geometry and the molecular structure of each of the following ions:

(a) H_3O^+

(b) PCl_4^-

(c) $SnCl_3^+$

(d) $BrCl_4^-$

(e) ICl_3

(f) XeF_4

(g) SF_2

85.

Identify the electron pair geometry and the molecular structure of each of the following molecules:

(a) ClNO (N is the central atom)

(b) CS_2

(c) Cl_2CO (C is the central atom)

(d) Cl_2SO (S is the central atom)

(e) SO_2F_2 (S is the central atom)

(f) XeO_2F_2 (Xe is the central atom)

(g) $ClOF_2^+$ (Cl is the central atom)

86.

Predict the electron pair geometry and the molecular structure of each of the following:

(a) IOF_5 (I is the central atom)

(b) $POCl_3$ (P is the central atom)

(c) Cl_2SeO (Se is the central atom)

(d) $ClSO^+$ (S is the central atom)

(e) F_2SO (S is the central atom)

(f) NO_2^-

(g) SiO_4^{4-}

87.

Which of the following molecules and ions contain polar bonds? Which of these molecules and ions have dipole moments?

(a) ClF_5

(b) ClO_2^-

(c) $TeCl_4^{2-}$

(d) PCl_3

(e) SeF_4

(f) PH_2^-

(g) XeF_2

88.

Which of these molecules and ions contain polar bonds? Which of these molecules and ions have dipole moments?

(a) H_3O^+

(b) PCl_4^-

(c) $SnCl_3^-$

(d) $BrCl_4^-$

(e) ICl_3

(f) XeF_4

(g) SF_2

89.

Which of the following molecules have dipole moments?

(a) CS_2

(b) SeS_2

(c) CCl_2F_2

(d) PCl_3 (P is the central atom)

(e) ClNO (N is the central atom)

90.

Identify the molecules with a dipole moment:

(a) SF_4

(b) CF_4

(c) Cl_2CCBr_2

(d) CH_3Cl

(e) H_2CO

91.

The molecule XF_3 has a dipole moment. Is X boron or phosphorus?

92.

The molecule XCl_2 has a dipole moment. Is X beryllium or sulfur?

93.

Is the Cl_2BBCl_2 molecule polar or nonpolar?

94.

There are three possible structures for PCl_2F_3 with phosphorus as the central atom. Draw them and discuss how measurements of dipole moments could help distinguish among them.

95.

Describe the molecular structure around the indicated atom or atoms:

(a) the sulfur atom in sulfuric acid, H_2SO_4 [$(HO)_2SO_2$]

(b) the chlorine atom in chloric acid, $HClO_3$ [$HOClO_2$]

(c) the oxygen atom in hydrogen peroxide, HOOH

(d) the nitrogen atom in nitric acid, HNO_3 [$HONO_2$]

(e) the oxygen atom in the OH group in nitric acid, HNO_3 [$HONO_2$]

(f) the central oxygen atom in the ozone molecule, O_3

(g) each of the carbon atoms in propyne, CH_3CCH

(h) the carbon atom in Freon, CCl_2F_2

(i) each of the carbon atoms in allene, H_2CCCH_2

96.

Draw the Lewis structures and predict the shape of each compound or ion:

(a) CO_2

(b) NO_2^-

(c) SO_3

(d) SO_3^{2-}

97.

A molecule with the formula AB_2, in which A and B represent different atoms, could have one of three different shapes. Sketch and name the three different shapes that this molecule might have. Give an example of a molecule or ion for each shape.

98.

A molecule with the formula AB_3, in which A and B represent different atoms, could have one of three different shapes. Sketch and name the three different shapes that this molecule might have. Give an example of a molecule or ion that has each shape.

99.

Draw the Lewis electron dot structures for these molecules, including resonance structures where appropriate:

(a) CS_3^{2-}

(b) CS_2

(c) CS

(d) predict the molecular shapes for CS_3^{2-} and CS_2 and explain how you arrived at your predictions

100.

What is the molecular structure of the stable form of FNO_2? (N is the central atom.)

101.

A compound with a molar mass of about 42 g/mol contains 85.7% carbon and 14.3% hydrogen. What is its molecular structure?

102.

Use the simulation to perform the following exercises for a two-atom molecule:

(a) Adjust the electronegativity value so the bond dipole is pointing toward B. Then determine what the electronegativity values must be to switch the dipole so that it points toward A.

(b) With a partial positive charge on A, turn on the electric field and describe what happens.

(c) With a small partial negative charge on A, turn on the electric field and describe what happens.

(d) Reset all, and then with a large partial negative charge on A, turn on the electric field and describe what happens.

103.

Use the simulation to perform the following exercises for a real molecule. You may need to rotate the molecules in three dimensions to see certain dipoles.

(a) Sketch the bond dipoles and molecular dipole (if any) for O_3. Explain your observations.

(b) Look at the bond dipoles for NH_3. Use these dipoles to predict whether N or H is more electronegative.

(c) Predict whether there should be a molecular dipole for NH_3 and, if so, in which direction it will point. Check the molecular dipole box to test your hypothesis.

104.

Use the Molecule Shape simulator to build a molecule. Starting with the central atom, click on the double bond to add one double bond. Then add one single bond and one lone pair. Rotate the molecule to observe the complete geometry. Name the electron group geometry and molecular structure and predict the bond angle. Then click the check boxes at the bottom and right of the simulator to check your answers.

105.

Use the Molecule Shape simulator to explore real molecules. On the Real Molecules tab, select H_2O. Switch between the “real” and “model” modes. Explain the difference observed.

106.

Use the Molecule Shape simulator to explore real molecules. On the Real Molecules tab, select “model” mode and S_2O. What is the model bond angle? Explain whether the “real” bond angle should be larger or smaller than the ideal model angle.

This page titled 5.11: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

6: Advanced Theories of Covalent Bonding

We have examined the basic ideas of bonding, showing that atoms share electrons to form molecules with stable Lewis structures and that we can predict the shapes of those molecules by valence shell electron pair repulsion (VSEPR) theory. These ideas provide an important starting point for understanding chemical bonding. But these models sometimes fall short in their abilities to predict the behavior of real substances. How can we reconcile the geometries of *s*, *p*, and *d* atomic orbitals with molecular shapes that show angles like 120° and 109.5°? Furthermore, we know that electrons and magnetic behavior are related through electromagnetic fields.

Yet oxygen demonstrates very different magnetic behavior than nitrogen. We can pour liquid nitrogen through a magnetic field with no visible interactions, while liquid oxygen (shown in Figure 6.1) is attracted to the magnet and floats in the magnetic field. We need to understand the additional concepts of valence bond theory, orbital hybridization, and molecular orbital theory to understand these observations.



Faculty Feature

This chapter on advanced theories of chemical bonding delves into the cutting-edge computational approaches and quantum mechanics that are redefining our understanding of atomic and molecule interactions. Dr. Artur Izmaylov is a prominent contributor to this discipline, and his research focuses on improving quantum chemistry approaches for accurately modelling complicated systems. His research focuses on electronic structure and dynamics, particularly in highly coupled systems, where traditional approaches may fall short. Dr. Izmaylov's study sheds further light on the creation, breaking, and modification of chemical bonds at the quantum level by developing sophisticated algorithms and upgrading non-adiabatic process models. This chapter will emphasize the significance of these sophisticated theories in comprehending the underlying basis of chemical reactions and interactions, stretching the frontiers of how we predict and manipulate bonding in both molecular and solid-state systems.



Read more about Dr. Artur Izmaylov and his research here: https://www.utsc.utoronto.ca/~aizmaylov/

6.1: Introduction
6.2: Valence Bond Theory
6.3: Hybrid Atomic Orbitals
6.4: Multiple Bonds
6.5: Molecular Orbital Theory
6.6: Key Terms
6.7: Key Equations
6.8: Summary
6.9: Exercises

This page titled 6: Advanced Theories of Covalent Bonding is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

6.1: Introduction

We have examined the basic ideas of bonding, showing that atoms share electrons to form molecules with stable Lewis structures and that we can predict the shapes of those molecules by valence shell electron pair repulsion (VSEPR) theory. These ideas provide an important starting point for understanding chemical bonding. But these models sometimes fall short in their abilities to predict the behavior of real substances. How can we reconcile the geometries of *s*, *p*, and *d* atomic orbitals with molecular shapes that show angles like 120° and 109.5°? Furthermore, we know that electrons and magnetic behavior are related through electromagnetic fields. Both N_2 and O_2 have fairly similar Lewis structures that contain lone pairs of electrons.

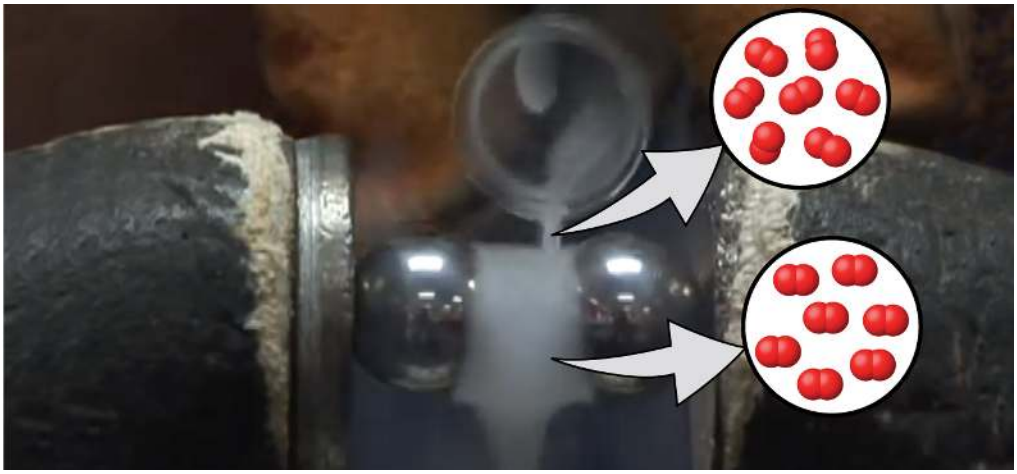


Figure 6.1.1: Oxygen molecules orient randomly most of the time, as shown in the top magnified view. However, when we pour liquid oxygen through a magnet, the molecules line up with the magnetic field, and the attraction allows them to stay suspended between the poles of the magnet where the magnetic field is strongest. Other diatomic molecules (like N_2) flow past the magnet. The detailed explanation of bonding described in this chapter allows us to understand this phenomenon. (credit: modification of work by Jefferson Lab)

Yet oxygen demonstrates very different magnetic behavior than nitrogen. We can pour liquid nitrogen through a magnetic field with no visible interactions, while liquid oxygen (shown in Figure 6.1.1) is attracted to the magnet and floats in the magnetic field. We need to understand the additional concepts of valence bond theory, orbital hybridization, and molecular orbital theory to understand these observations.



Faculty Feature

This chapter on advanced theories of chemical bonding delves into the cutting-edge computational approaches and quantum mechanics that are redefining our understanding of atomic and molecule interactions. Dr. Artur Izmaylov is a prominent contributor to this discipline, and his research focuses on improving quantum chemistry approaches for accurately modelling complicated systems. His research focuses on electronic structure and dynamics, particularly in highly coupled systems, where traditional approaches may fall short. Dr. Izmaylov's study sheds further light on the creation, breaking, and modification of chemical bonds at the quantum level by developing sophisticated algorithms and upgrading non-adiabatic process models. This chapter will emphasize the significance of these sophisticated theories in comprehending the underlying basis of chemical reactions and interactions, stretching the frontiers of how we predict and manipulate bonding in both molecular and solid-state systems.



Read more about Dr. Artur Izmaylov and his research here: https://www.utsc.utoronto.ca/~aizmaylov/

This page titled 6.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **8.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

6.2: Valence Bond Theory

Learning Objectives

By the end of this section, you will be able to:

- Describe the formation of covalent bonds in terms of atomic orbital overlap
- Define and give examples of σ and π bonds

As we know, a scientific theory is a strongly supported explanation for observed natural laws or large bodies of experimental data. For a theory to be accepted, it must explain experimental data and be able to predict behavior. For example, VSEPR theory has gained widespread acceptance because it predicts three-dimensional molecular shapes that are consistent with experimental data collected for thousands of different molecules. However, VSEPR theory does not provide an explanation of chemical bonding.

Atomic Orbital Overlap

There are successful theories that describe the electronic structure of atoms. We can use quantum mechanics to predict the specific regions around an atom where electrons are likely to be located: A spherical shape for an *s* orbital, a dumbbell shape for a *p* orbital, and so forth. However, these predictions only describe the orbitals around free atoms. When atoms bond to form molecules, atomic orbitals are not sufficient to describe the regions where electrons will be located in the molecule. A more complete understanding of electron distributions requires a model that can account for the electronic structure of molecules. One popular theory holds that a covalent bond forms when a pair of electrons is shared by two atoms and is simultaneously attracted by the nuclei of both atoms. In the following sections, we will discuss how such bonds are described by valence bond theory and hybridization.

Valence bond theory describes a covalent bond as the overlap of half-filled atomic orbitals (each containing a single electron) that yield a pair of electrons shared between the two bonded atoms. We say that orbitals on two different atoms overlap when a portion of one orbital and a portion of a second orbital occupy the same region of space. According to valence bond theory, a covalent bond results when two conditions are met: (1) an orbital on one atom overlaps an orbital on a second atom and (2) the single electrons in each orbital combine to form an electron pair. The mutual attraction between this negatively charged electron pair and the two atoms' positively charged nuclei serves to physically link the two atoms through a force we define as a covalent bond. The strength of a covalent bond depends on the extent of overlap of the orbitals involved. Orbitals that overlap extensively form bonds that are stronger than those that have less overlap.

The energy of the system depends on how much the orbitals overlap. Figure 5.2 illustrates how the sum of the energies of two hydrogen atoms (the colored curve) changes as they approach each other. When the atoms are far apart there is no overlap, and by convention we set the sum of the energies at zero. As the atoms move together, their orbitals begin to overlap. Each electron begins to feel the attraction of the nucleus in the other atom. In addition, the electrons begin to repel each other, as do the nuclei. While the atoms are still widely separated, the attractions are slightly stronger than the repulsions, and the energy of the system decreases. (A bond begins to form.) As the atoms move closer together, the overlap increases, so the attraction of the nuclei for the electrons continues to increase (as do the repulsions among electrons and between the nuclei). At some specific distance between the atoms, which varies depending on the atoms involved, the energy reaches its lowest (most stable) value. This optimum distance between the two bonded nuclei is the bond distance between the two atoms. The bond is stable because at this point, the attractive and repulsive forces combine to create the lowest possible energy configuration. If the distance between the nuclei were to decrease further, the repulsions between nuclei and the repulsions as electrons are confined in closer proximity to each other would become stronger than the attractive forces. The energy of the system would then rise (making the system destabilized), as shown at the far left of Figure 5.2.

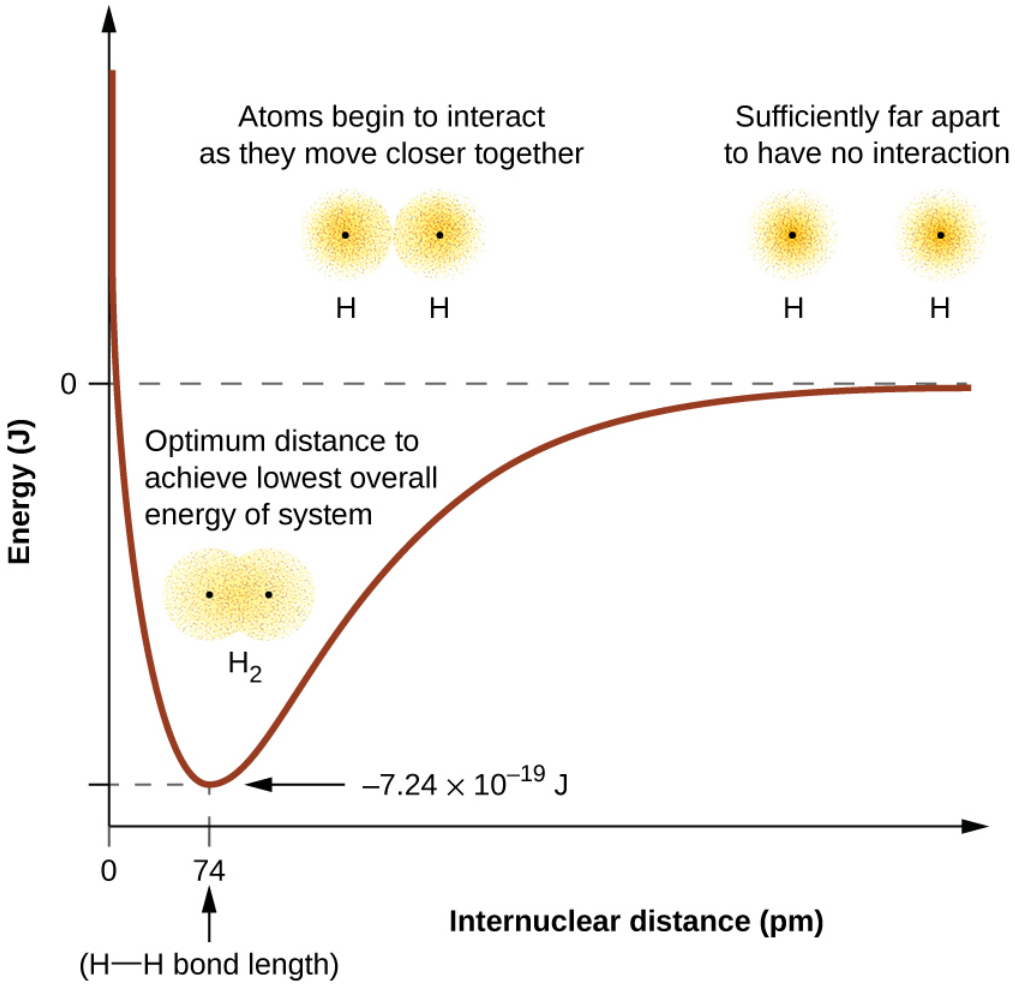


Figure 5.2 (a) The interaction of two hydrogen atoms changes as a function of distance. (b) The energy of the system changes as the atoms interact. The lowest (most stable) energy occurs at a distance of 74 pm, which is the bond length observed for the H_2 molecule.

In addition to the distance between two orbitals, the orientation of orbitals also affects their overlap (other than for two *s* orbitals, which are spherically symmetric). Greater overlap is possible when orbitals are oriented such that they overlap on a direct line between the two nuclei. Figure 5.3 illustrates this for two *p* orbitals from different atoms; the overlap is greater when the orbitals overlap end to end rather than at an angle.

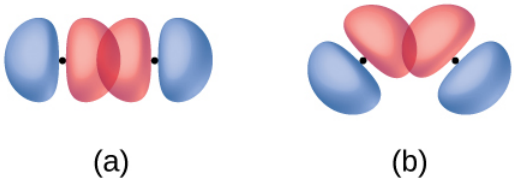


Figure 5.3 (a) The overlap of two *p* orbitals is greatest when the orbitals are directed end to end. (b) Any other arrangement results in less overlap. The dots indicate the locations of the nuclei.

The overlap of two *s* orbitals (as in H_2), the overlap of an *s* orbital and a *p* orbital (as in HCl), and the end-to-end overlap of two *p* orbitals (as in Cl_2) all produce sigma bonds (σ bonds), as illustrated in Figure 5.4. A σ bond is a covalent bond in which the electron density is concentrated in the region along the internuclear axis; that is, a line between the nuclei would pass through the center of the overlap region. Single bonds in Lewis structures are described as σ bonds in valence bond theory.

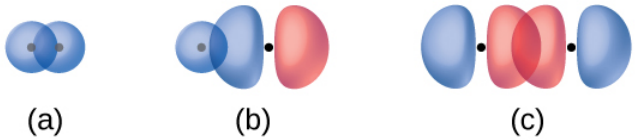


Figure 5.4 Sigma (σ) bonds form from the overlap of the following: (a) two *s* orbitals, (b) an *s* orbital and a *p* orbital, and (c) two *p* orbitals. The dots indicate the locations of the nuclei.

A pi bond (π bond) is a type of covalent bond that results from the side-by-side overlap of two *p* orbitals, as illustrated in Figure 5.5. In a π bond, the regions of orbital overlap lie on opposite sides of the internuclear axis. Along the axis itself, there is a node, that is, a plane with no probability of finding an electron.

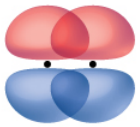
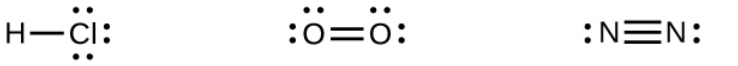


Figure 5.5 Pi (π) bonds form from the side-by-side overlap of two *p* orbitals. The dots indicate the location of the nuclei.

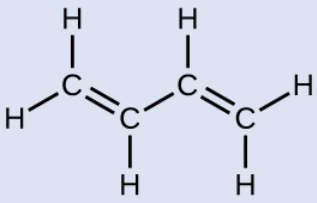
While all single bonds are σ bonds, multiple bonds consist of both σ and π bonds. As the Lewis structures below suggest, O_2 contains a double bond, and N_2 contains a triple bond. The double bond consists of one σ bond and one π bond, and the triple bond consists of one σ bond and two π bonds. Between any two atoms, the first bond formed will always be a σ bond, but there can only be one σ bond in any one location. In any multiple bond, there will be one σ bond, and the remaining one or two bonds will be π bonds. These bonds are described in more detail later in this chapter.



One σ bond No π bonds	One σ bond One π bond	One σ bond Two π bonds

Example 5.1

Counting σ and π Bonds



Butadiene, C_4H_6, is used to make synthetic rubber. Identify the number of σ and π bonds contained in this molecule.

Solution

There are six σ C–H bonds and one σ C–C bond, for a total of seven from the single bonds. There are two double bonds that each have a π bond in addition to the σ bond. This gives a total nine σ and two π bonds overall.

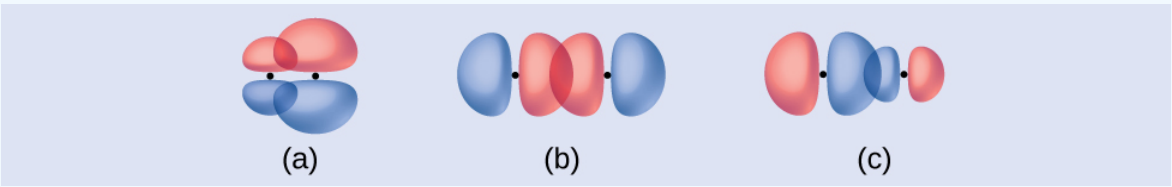
Check Your Learning

Identify each illustration as depicting a σ or π bond:

(a) side-by-side overlap of a 4*p* and a 2*p* orbital

(b) end-to-end overlap of a 4*p* and 4*p* orbital

(c) end-to-end overlap of a 4*p* and a 2*p* orbital



Answer:

(a) is a π bond with a node along the axis connecting the nuclei while (b) and (c) are σ bonds that overlap along the axis.

Dipole Moments and Ionic Character

Now that we have seen the importance of understanding the connection between the location of electrons in atoms and the properties of elements, we can expand our understanding of the connection between atoms. This will be an introduction to more advanced aspects of the chemical bond, which is the very heart of chemistry itself. With the sole exception of the noble gases, atoms by themselves do not possess the most stable possible electron configuration. That is where the concept of chemical bonding comes into its own: atoms can attain a stable configuration by exchanging electrons with another atom, resulting in the formation of ions.

Ions, in turn, can associate by charge – simple Coulombic attraction – resulting in the formation of compounds we call ionic compounds. We will look at the ionic nature of bonds first, from a simple positive-negative attraction standpoint. Just as important is that some atoms bond by sharing rather than exchanging electrons; the sharing of electrons gives rise to the covalent bond. To add just one more dimension, some chemical species are neither completely ionic nor completely covalent; these species possess a permanent dipole, and are classified as polar.

In your introductory physics course, you will likely discuss the concept of Coulombic interactions in much more rigorous detail than we will do here. We are interested primarily in the differences in properties between species that arise from their relative covalent, ionic, or polar nature – not in a rigorous model of those properties. We are concerned with the connection between potential energy and force and the relative separation (or lack of separation) between charges. We begin by defining the electric or Coulomb force as the product of the charges divided by the square of the distance between those charges:

$$F = \frac{Q_1 Q_2}{d^2}$$

Here, *Q* is taken to be the fundamental constant of electron charge: one electron has a charge of 1.60218×10^{-19} C. (We will work exclusively in the SI system, so distances will be measured in meters (m)).

And as you may recall, energy is force times distance, so

$$E = \frac{Q_1 Q_2}{d}$$

To illustrate the trend in attractive force, we will consider first the attractive force between two ions of single charge separated by a distance of 2*d*:

$$F = \frac{(1)(-1)}{(2d)^2} = -\frac{1}{4d^2}$$

And then the attractive force between two ions of double charge separated by a distance *d*:

$$F = \frac{(2)(-2)}{(d)^2} = -\frac{4}{d^2}$$

The force of attraction increases with the charge and decreases with increased distance. If all matter were composed of ions that would be the end of the story, but it clearly is not. There are molecules for which the charge – either positive or negative – is permanently concentrated more on one atom than the other. Notice we say atom, because these compounds are not composed of ions but rather of atoms sharing electrons through covalent bonds.

Bond Dipole Moments

The concept of a bond dipole moment allows us to examine the partial separation of charge between atoms. It is a simple model when applied to diatomic molecules, which will be more than sufficient for our purposes. The dipole moment of a bond is defined as the charge times the distance – charge once again being measured in multiples of the charge on an electron, or coulombs. The distance will always be in meters. Because we are considering very small charges and distances, and because it is the relative separation of charge rather than the actual value for it that we are interested in, we will introduce a new unit called the Debye, named after the physical chemist Peter Debye:

$$1 \text{ Debye (D)} = 3.336 \times 10^{-30} \text{ C-m} \qquad \mu = Q \times d$$

The usefulness of the Debye unit will be shown by example:
For HCl, the bond dipole moment is known to be 1.08 D
For HI, the bond dipole moment is known to be 0.44 D

Comparing the two, we can see that HI is less polar than HCl, which is what we would expect based on electronegativity values.

We have now made a transition between the concept of an ionic compound and a partially ionic one. Of course, the partially ionic compound must also by definition be partially covalent.

Partial Ionic Character

The concept of the bond dipole moment helps bridge the concepts of ionic and covalent bonding. Because there is a separation of charge that is less complete than it is in an ionic bond, we can refer to polar bonds as being partially ionic in nature. In contrast to sodium chloride, hydrogen chloride shows partial charges (indicated with a delta notation) on the hydrogen and chlorine. As you would expect from the electronegativity values, hydrogen carries a partial positive charge, while chlorine carries a partial negative charge. Where do these charges come from?

It is easy to come up with the partial charges by comparing the actual dipole moments (which can be obtained experimentally, using spectroscopy) with the dipole expected in the limiting case (that is, if we were to consider the molecule ionic). The actual dipole moment is 1.03 D.

Example 5.2

Finding the Partial Ionic Character

What are the partial charges of an HCl molecule, whose bond length is 0.127 nm?

Solution

The bond dipole moment is $\left(1.60218 \times 10^{-19}\text{C}\right)\left(0.127 \times 10^{-9}\text{m}\right)$or 2.03×10^{-29}C-m. Converted to D, this is $\left(2.03 \times 10^{-29}\text{C-m}\right)\left(\frac{1 \text{ Debye}}{3.336 \times 10^{-30}\text{C-m}}\right)$ or 6.09 D. Were HCl completely ionic, this would be its molecular dipole moment. To get the partial ionic character, we divide the experimentally measured bond moment by this limiting value: $\%\text{ionic character} = \frac{\mu_{\text{exp}}}{\mu_{\text{lim}}} \times 100\% = \frac{(1.03 \quad \text{D}}{(6.09 \quad \text{D})} \times 100\% = 17\%$. This means the bond is about 17% ionic – or, put another way, the positive charge in H is +0.17 and the partial negative charge on chlorine, –0.17.

Check Your Learning

Repeat the calculation for HI, which has a dipole moment of 0.42 D and a bond length of 0.161 nm.

Answer: Calculated 7.73, percent 5.43

What does the result suggest about the relative polarity of the HI bond vs. that of the HCl bond? Does the calculated dipole and percent ionic character reconcile with the difference in electronegativity between Cl and I?

The electron configuration of an atom or ion is key to understanding the chemical behavior of an element. The atoms that make up the element combine in various ways, ranging from the mostly ionic (NaCl) to the partially ionic (HCl) to what we will call purely covalent. At the most fundamental level, all chemical bonds involve electrons, and a significant percentage of chemical and physical properties can be explained by considering the location and separation of charge in a species. By understanding the structure of matter at the atomic level, we can begin to build an understanding of the behavior of matter at both the microscopic and macroscopic levels.

An understanding of dipoles and partial ionic character is fundamental to understanding the interactions between particles, which we will examine in the chapter on liquids and solids. These intermolecular forces become important in the liquid and solid states of matter.

This page titled 6.2: Valence Bond Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

6.3: Hybrid Atomic Orbitals

Learning Objectives

By the end of this section, you will be able to:

- Explain the concept of atomic orbital hybridization
- Determine the hybrid orbitals associated with various molecular geometries

Thinking in terms of overlapping atomic orbitals is one way for us to explain how chemical bonds form in diatomic molecules. However, to understand how molecules with more than two atoms form stable bonds, we require a more detailed model. As an example, let us consider the water molecule, in which we have one oxygen atom bonding to two hydrogen atoms. Oxygen has the electron configuration $1s^2 2s^2 2p^4$, with two unpaired electrons (one in each of the two 2*p* orbitals). Valence bond theory would predict that the two O–H bonds form from the overlap of these two 2*p* orbitals with the 1*s* orbitals of the hydrogen atoms. If this were the case, the bond angle would be 90°, as shown in Figure 6.3.1, because *p* orbitals are perpendicular to each other. Experimental evidence shows that the bond angle is 104.5°, not 90°. The prediction of the valence bond theory model does not match the real-world observations of a water molecule; a different model is needed.

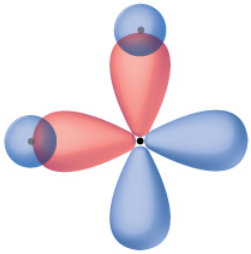


Figure 6.3.1: The hypothetical overlap of two of the 2*p* orbitals on an oxygen atom (red) with the 1*s* orbitals of two hydrogen atoms (blue) would produce a bond angle of 90°. This is not consistent with experimental evidence.[1]

Quantum-mechanical calculations suggest why the observed bond angles in H_2O differ from those predicted by the overlap of the 1*s* orbital of the hydrogen atoms with the 2*p* orbitals of the oxygen atom. The mathematical expression known as the wave function, ψ, contains information about each orbital and the wavelike properties of electrons in an isolated atom. When atoms are bound together in a molecule, the wave functions combine to produce new mathematical descriptions that have different shapes. This process of combining the wave functions for atomic orbitals is called **hybridization** and is mathematically accomplished by the *linear combination of atomic orbitals*, LCAO, (a technique that we will encounter again later). The new orbitals that result are called **hybrid orbitals**. The valence orbitals in an *isolated* oxygen atom are a 2*s* orbital and three 2*p* orbitals. The valence orbitals in an oxygen atom in a water molecule differ; they consist of four equivalent hybrid orbitals that point approximately toward the corners of a tetrahedron (Figure 6.3.2). Consequently, the overlap of the O and H orbitals should result in a tetrahedral bond angle (109.5°). The observed angle of 104.5° is experimental evidence for which quantum-mechanical calculations give a useful explanation: Valence bond theory must include a hybridization component to give accurate predictions.

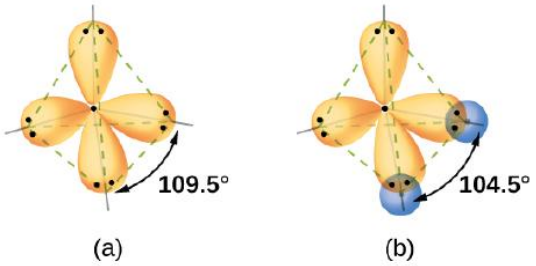


Figure 6.3.2: (a) A water molecule has four regions of electron density, so VSEPR theory predicts a tetrahedral arrangement of hybrid orbitals. (b) Two of the hybrid orbitals on oxygen contain lone pairs, and the other two overlap with the 1*s* orbitals of hydrogen atoms to form the O–H bonds in H_2O. This description is more consistent with the experimental structure.

The following ideas are important in understanding hybridization:

1. Hybrid orbitals do not exist in isolated atoms. They are formed only in covalently bonded atoms.
2. Hybrid orbitals have shapes and orientations that are very different from those of the atomic orbitals in isolated atoms.

3. A set of hybrid orbitals is generated by combining atomic orbitals. The number of hybrid orbitals in a set is equal to the number of atomic orbitals that were combined to produce the set.
4. All orbitals in a set of hybrid orbitals are equivalent in shape and energy.
5. The type of hybrid orbitals formed in a bonded atom depends on its electron-pair geometry as predicted by the VSEPR theory.
6. Hybrid orbitals overlap to form σ bonds. Unhybridized orbitals overlap to form π bonds.

In the following sections, we shall discuss the common types of hybrid orbitals.

sp Hybridization

The beryllium atom in a gaseous $BeCl_2$ molecule is an example of a central atom with no lone pairs of electrons in a linear arrangement of three atoms. There are two regions of valence electron density in the $BeCl_2$ molecule that correspond to the two covalent Be–Cl bonds. To accommodate these two electron domains, two of the Be atom's four valence orbitals will mix to yield two hybrid orbitals. This hybridization process involves mixing of the valence *s* orbital with one of the valence *p* orbitals to yield two equivalent ***sp* hybrid orbitals** that are oriented in a linear geometry (Figure 6.3.3). In this figure, the set of *sp* orbitals appears similar in shape to the original *p* orbital, but there is an important difference. The number of atomic orbitals combined always equals the number of hybrid orbitals formed. The *p* orbital is one orbital that can hold up to two electrons. The *sp* set is two equivalent orbitals that point 180° from each other. The two electrons that were originally in the *s* orbital are now distributed to the two *sp* orbitals, which are half filled. In gaseous $BeCl_2$, these half-filled hybrid orbitals will overlap with orbitals from the chlorine atoms to form two identical σ bonds.

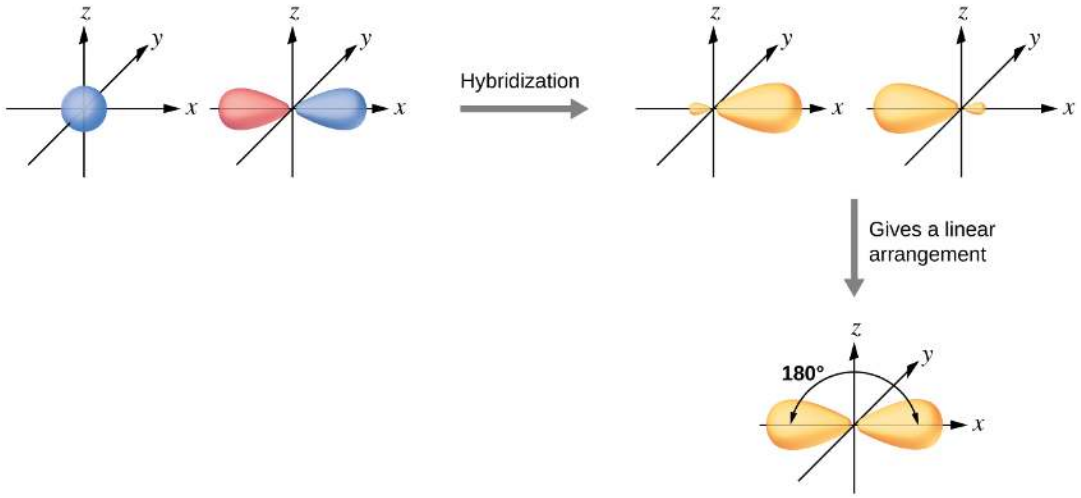


Figure 6.3.3: Hybridization of an *s* orbital (blue) and a *p* orbital (red) of the same atom produces two *sp* hybrid orbitals (yellow). Each hybrid orbital is oriented primarily in just one direction. Note that each *sp* orbital contains one lobe that is significantly larger than the other. The set of two *sp* orbitals are oriented at 180°, which is consistent with the geometry for two domains.

We illustrate the electronic differences in an isolated Be atom and in the bonded Be atom in the orbital energy-level diagram in Figure 6.3.4. These diagrams represent each orbital by a horizontal line (indicating its energy) and each electron by an arrow. Energy increases toward the top of the diagram. We use one upward arrow to indicate one electron in an orbital and two arrows (up and down) to indicate two electrons of opposite spin.

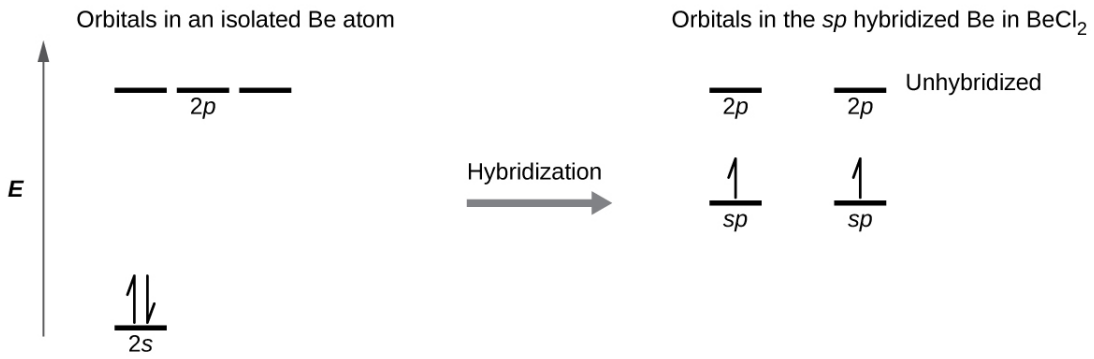


Figure 6.3.4: This orbital energy-level diagram shows the *sp* hybridized orbitals on Be in the linear $BeCl_2$ molecule. Each of the two *sp* hybrid orbitals holds one electron and is thus half filled and available for bonding via overlap with a Cl 3*p* orbital.

When atomic orbitals hybridize, the valence electrons occupy the newly created orbitals. The Be atom had two valence electrons, so each of the *sp* orbitals gets one of these electrons. Each of these electrons pairs up with the unpaired electron on a chlorine atom when a hybrid orbital and a chlorine orbital overlap during the formation of the Be–Cl bonds.

Any central atom surrounded by just two regions of valence electron density in a molecule will exhibit *sp* hybridization. Other examples include the mercury atom in the linear $HgCl_2$ molecule, the zinc atom in $Zn(CH_3)_2$, which contains a linear C–Zn–C arrangement, and the carbon atoms in HCCH and CO_2.

Link to Learning

Check out the University of Wisconsin-Oshkosh website to learn about visualizing hybrid orbitals in three dimensions.

sp^2 Hybridization

The valence orbitals of a central atom surrounded by three regions of electron density consist of a set of three ***sp*^{2} hybrid orbitals** and one unhybridized *p* orbital. This arrangement results from sp^2 hybridization, the mixing of one *s* orbital and two *p* orbitals to produce three identical hybrid orbitals oriented in a trigonal planar geometry (Figure 6.3.5).

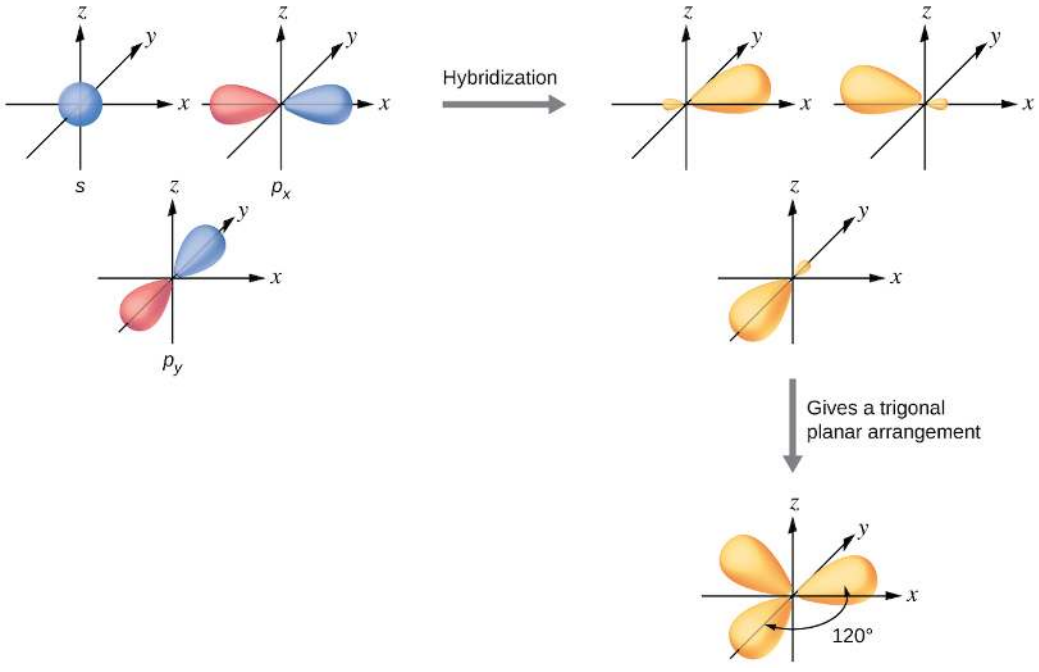


Figure 6.3.5: The hybridization of an *s* orbital (blue) and two *p* orbitals (red) produces three equivalent sp^2 hybridized orbitals (yellow) oriented at 120° with respect to each other. The remaining unhybridized *p* orbital is not shown here, but is located along the z axis.

Although quantum mechanics yields the "plump" orbital lobes as depicted in Figure 6.3.5, sometimes for clarity these orbitals are drawn thinner and without the minor lobes, as in Figure 6.3.6, to avoid obscuring other features of a given illustration. We will use these "thinner" representations whenever the true view is too crowded to easily visualize.

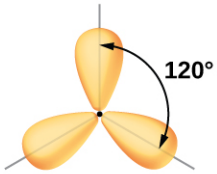


Figure 6.3.6: This alternate way of drawing the trigonal planar sp^2 hybrid orbitals is sometimes used in more crowded figures.

The observed structure of the borane molecule, BH_3, suggests sp^2 hybridization for boron in this compound. The molecule is trigonal planar, and the boron atom is involved in three bonds to hydrogen atoms (Figure 6.3.7). We can illustrate the comparison of orbitals and electron distribution in an isolated boron atom and in the bonded atom in BH_3 as shown in the orbital energy level diagram in Figure 6.3.8. We redistribute the three valence electrons of the boron atom in the three sp^2 hybrid orbitals, and each boron electron pairs with a hydrogen electron when B–H bonds form.

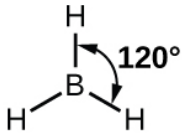


Figure 6.3.7: BH_3 is an electron-deficient molecule with a trigonal planar structure.

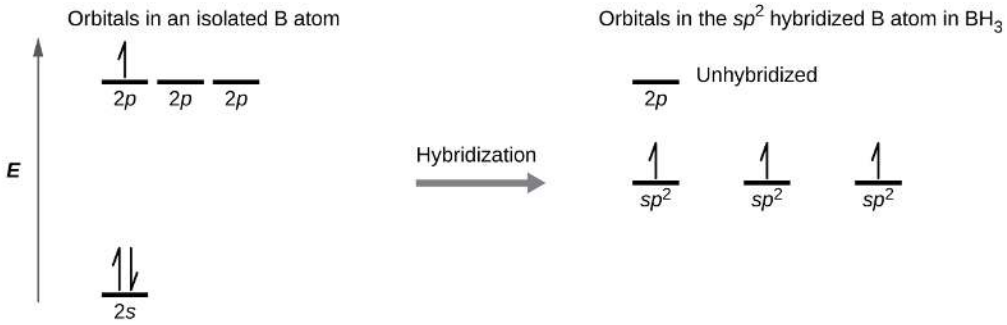


Figure 6.3.8: In an isolated B atom, there are one 2*s* and three 2*p* valence orbitals. When boron is in a molecule with three regions of electron density, three of the orbitals hybridize and create a set of three sp^2 orbitals and one unhybridized 2*p* orbital. The three half-filled hybrid orbitals each overlap with an orbital from a hydrogen atom to form three σ bonds in BH_3.

Any central atom surrounded by three regions of electron density will exhibit sp^2 hybridization. This includes molecules with a lone pair on the central atom, such as ClNO (Figure 6.3.9), or molecules with two single bonds and a double bond connected to the central atom, as in formaldehyde, CH_2O, and ethene, H_2CCH_2.

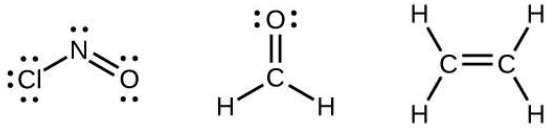


Figure 6.3.9: The central atom(s) in each of the structures shown contain three regions of electron density and are sp^2 hybridized. As we know from the discussion of VSEPR theory, a region of electron density contains all of the electrons that point in one direction. A lone pair, an unpaired electron, a single bond, or a multiple bond would each count as one region of electron density.

sp^3 Hybridization

The valence orbitals of an atom surrounded by a tetrahedral arrangement of bonding pairs and lone pairs consist of a set of four ***sp*^{3} hybrid orbitals**. The hybrids result from the mixing of one *s* orbital and all three *p* orbitals that produces four identical sp^3 hybrid orbitals (Figure 6.3.10). Each of these hybrid orbitals points toward a different corner of a tetrahedron.

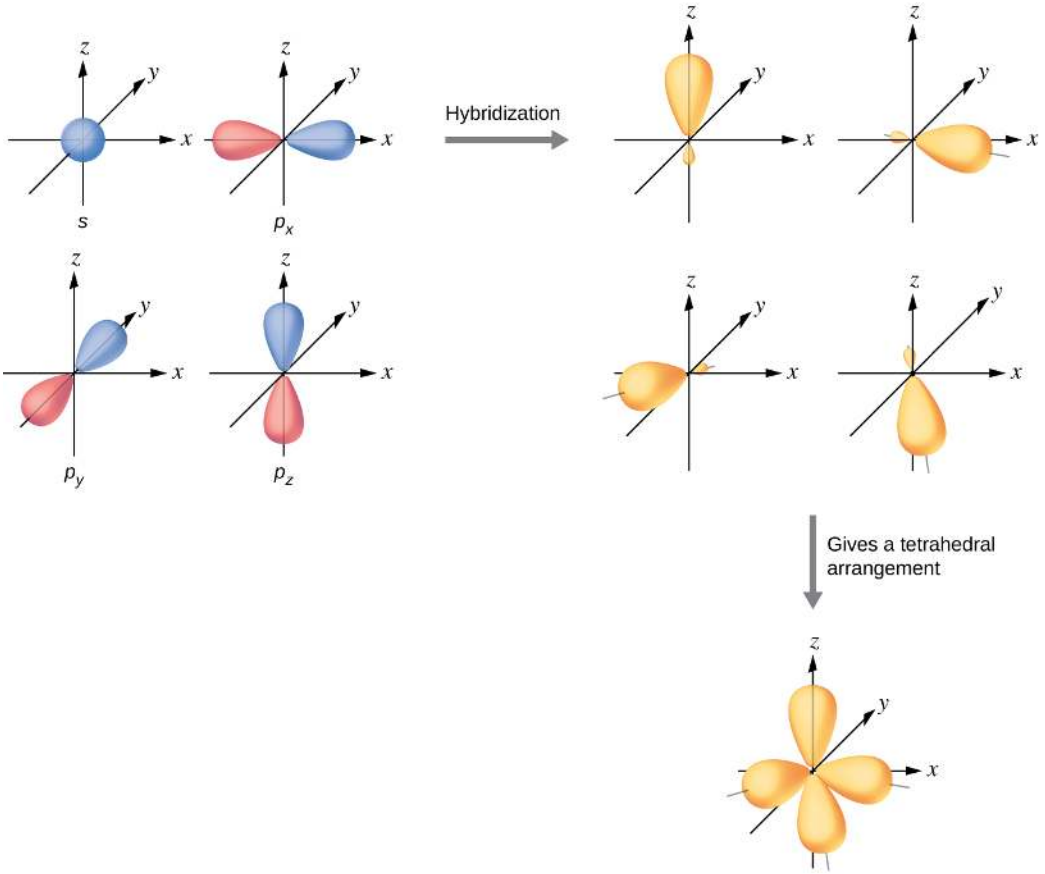


Figure 6.3.10: The hybridization of an *s* orbital (blue) and three *p* orbitals (red) produces four equivalent sp^3 hybridized orbitals (yellow) oriented at 109.5° with respect to each other.

A molecule of methane, CH_4, consists of a carbon atom surrounded by four hydrogen atoms at the corners of a tetrahedron. The carbon atom in methane exhibits sp^3 hybridization. We illustrate the orbitals and electron distribution in an isolated carbon atom and in the bonded atom in CH_4 in Figure 6.3.11. The four valence electrons of the carbon atom are distributed equally in the hybrid orbitals, and each carbon electron pairs with a hydrogen electron when the C–H bonds form.

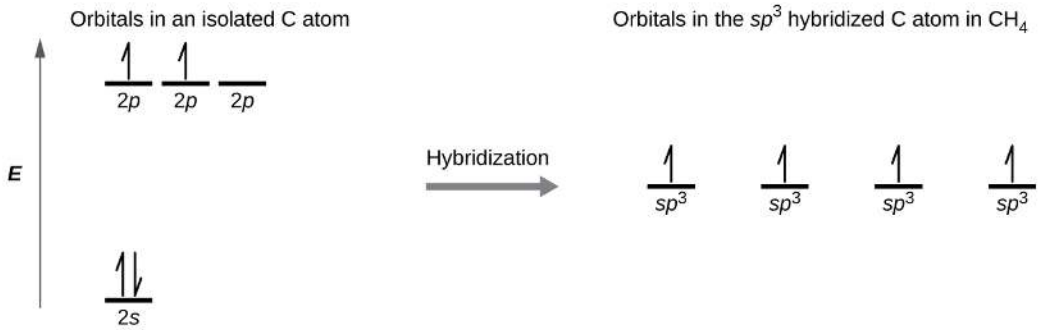


Figure 6.3.11: The four valence atomic orbitals from an isolated carbon atom all hybridize when the carbon bonds in a molecule like CH_4 with four regions of electron density. This creates four equivalent sp^3 hybridized orbitals. Overlap of each of the hybrid orbitals with a hydrogen orbital creates a C–H σ bond.

In a methane molecule, the 1*s* orbital of each of the four hydrogen atoms overlaps with one of the four sp^3 orbitals of the carbon atom to form a sigma (σ) bond. This results in the formation of four strong, equivalent covalent bonds between the carbon atom and each of the hydrogen atoms to produce the methane molecule, CH_4.

The structure of ethane, C_2H_6, is similar to that of methane in that each carbon in ethane has four neighboring atoms arranged at the corners of a tetrahedron—three hydrogen atoms and one carbon atom (Figure 6.3.12). However, in ethane an sp^3 orbital of one carbon atom overlaps end to end with an sp^3 orbital of a second carbon atom to form a σ bond between the two carbon atoms. Each

of the remaining sp^3 hybrid orbitals overlaps with an *s* orbital of a hydrogen atom to form carbon–hydrogen σ bonds. The structure and overall outline of the bonding orbitals of ethane are shown in Figure 6.3.12. The orientation of the two CH_3 groups is not fixed relative to each other. Experimental evidence shows that rotation around σ bonds occurs easily.

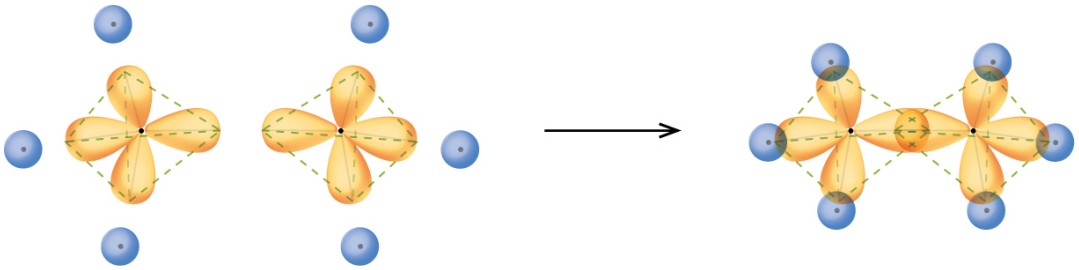


Figure 6.3.12: (a) In the ethane molecule, C_2H_6, each carbon has four sp^3 orbitals. (b) These four orbitals overlap to form seven σ bonds.

An sp^3 hybrid orbital can also hold a lone pair of electrons. For example, the nitrogen atom in ammonia is surrounded by three bonding pairs and a lone pair of electrons directed to the four corners of a tetrahedron. The nitrogen atom is sp^3 hybridized with one hybrid orbital occupied by the lone pair.

The molecular structure of water is consistent with a tetrahedral arrangement of two lone pairs and two bonding pairs of electrons. Thus we say that the oxygen atom is sp^3 hybridized, with two of the hybrid orbitals occupied by lone pairs and two by bonding pairs. Since lone pairs occupy more space than bonding pairs, structures that contain lone pairs have bond angles slightly distorted from the ideal. Perfect tetrahedra have angles of 109.5°, but the observed angles in ammonia (107.3°) and water (104.5°) are slightly smaller. Other examples of sp^3 hybridization include CCl_4, PCl_3, and NCl_3.

sp^3d and sp^3d^2 Hybridization

To describe the five bonding orbitals in a trigonal bipyramidal arrangement, we must use five of the valence shell atomic orbitals (the *s* orbital, the three *p* orbitals, and one of the *d* orbitals), which gives five ***sp*^{3}*d* hybrid orbitals**. With an octahedral arrangement of six hybrid orbitals, we must use six valence shell atomic orbitals (the *s* orbital, the three *p* orbitals, and two of the *d* orbitals in its valence shell), which gives six ***sp*^{3}*d*^{2} hybrid orbitals**. These hybridizations are only possible for atoms that have *d* orbitals in their valence subshells (that is, not those in the first or second period).

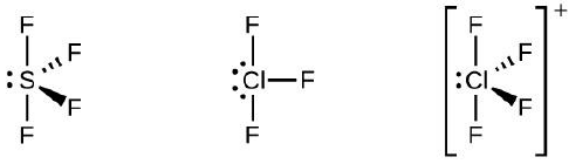


Figure 6.3.13: The three compounds pictured exhibit sp^3d hybridization in the central atom and a trigonal bipyramid form. SF_4 and have one lone pair of electrons on the central atom, and ClF_3 has two lone pairs giving it the T-shape shown.

In a molecule of phosphorus pentachloride, PCl_5, there are five P–Cl bonds (thus five pairs of valence electrons around the phosphorus atom) directed toward the corners of a trigonal bipyramid. We use the 3*s* orbital, the three 3*p* orbitals, and one of the 3*d* orbitals to form the set of five sp^3d hybrid orbitals (Figure 6.3.14) that are involved in the P–Cl bonds. Other atoms that exhibit sp^3d hybridization include the sulfur atom in SF_4 and the chlorine atoms in ClF_3 and in (The electrons on fluorine atoms are omitted for clarity.)

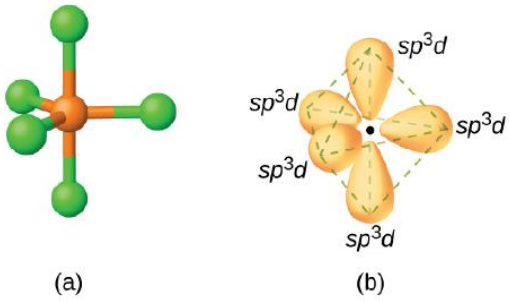


Figure 6.3.14: (a) The five regions of electron density around phosphorus in PCl_5 require five hybrid sp^3d orbitals. (b) These orbitals combine to form a trigonal bipyramidal structure with each large lobe of the hybrid orbital pointing at a vertex. As before, there are also small lobes pointing in the opposite direction for each orbital (not shown for clarity).

The sulfur atom in sulfur hexafluoride, SF_6, exhibits sp^3d^2 hybridization. A molecule of sulfur hexafluoride has six bonding pairs of electrons connecting six fluorine atoms to a single sulfur atom. There are no lone pairs of electrons on the central atom. To bond six fluorine atoms, the 3*s* orbital, the three 3*p* orbitals, and two of the 3*d* orbitals form six equivalent sp^3d^2 hybrid orbitals, each directed toward a different corner of an octahedron. Other atoms that exhibit sp^3d^2 hybridization include the phosphorus atom in the iodine atom in the interhalogens IF_5, and the xenon atom in XeF_4.

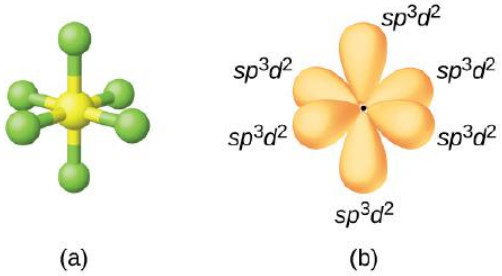


Figure 6.3.15: (a) Sulfur hexafluoride, SF_6, has an octahedral structure that requires sp^3d^2 hybridization. (b) The six sp^3d^2 orbitals form an octahedral structure around sulfur. Again, the minor lobe of each orbital is not shown for clarity.

Assignment of Hybrid Orbitals to Central Atoms

The hybridization of an atom is determined based on the number of regions of electron density that surround it. The geometrical arrangements characteristic of the various sets of hybrid orbitals are shown in Figure 6.3.16. These arrangements are identical to those of the electron-pair geometries predicted by VSEPR theory. VSEPR theory predicts the shapes of molecules, and hybrid orbital theory provides an explanation for how those shapes are formed. To find the hybridization of a central atom, we can use the following guidelines:

1. Determine the Lewis structure of the molecule.
2. Determine the number of regions of electron density around an atom using VSEPR theory, in which single bonds, multiple bonds, radicals, and lone pairs each count as one region.
3. Assign the set of hybridized orbitals from Figure 6.3.16 that corresponds to this geometry.

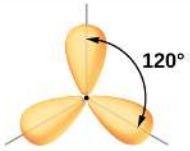
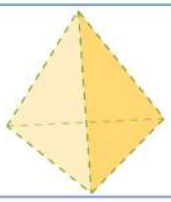
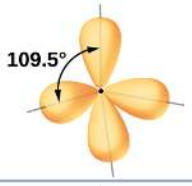
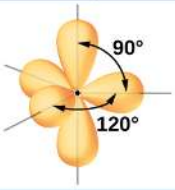
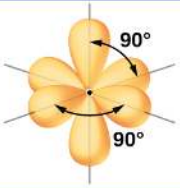
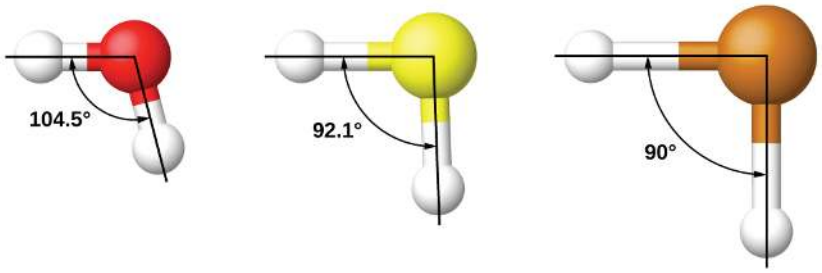
Regions of Electron Density	Arrangement		Hybridization	
2		linear	*sp*	
3		trigonal planar	sp^2	
4		tetrahedral	sp^3	
5		trigonal bipyramidal	sp^3d	
6		octahedral	sp^3d^2	

Figure 6.3.16: The shapes of hybridized orbital sets are consistent with the electron-pair geometries. For example, an atom surrounded by three regions of electron density is sp^2 hybridized, and the three sp^2 orbitals are arranged in a trigonal planar fashion.

It is important to remember that hybridization was devised to rationalize experimentally observed molecular geometries. The model works well for molecules containing small central atoms, in which the valence electron pairs are close together in space. However, for larger central atoms, the valence-shell electron pairs are farther from the nucleus, and there are fewer repulsions. Their compounds exhibit structures that are often not consistent with VSEPR theory, and hybridized orbitals are not necessary to explain the observed data. For example, we have discussed the H–O–H bond angle in H_2O, 104.5°, which is more consistent with sp^3 hybrid orbitals (109.5°) on the central atom than with 2*p* orbitals (90°). Sulfur is in the same group as oxygen, and H_2S has a similar Lewis structure. However, it has a much smaller bond angle (92.1°), which indicates much less hybridization on sulfur than oxygen. Continuing down the group, tellurium is even larger than sulfur, and for H_2Te, the observed bond angle (90°) is consistent with overlap of the 5*p* orbitals, without invoking hybridization. We invoke hybridization where it is necessary to explain the observed structures.

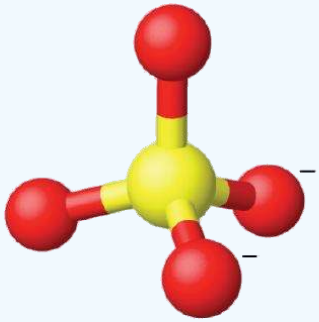


Example 6.3.1: Assigning Hybridization

Ammonium sulfate is important as a fertilizer. What is the hybridization of the sulfur atom in the sulfate ion,

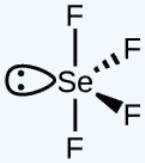
Solution

The Lewis structure of sulfate shows there are four regions of electron density. The hybridization is sp^3.



Exercise 6.3.1

What is the hybridization of the selenium atom in SeF_4?



Answer

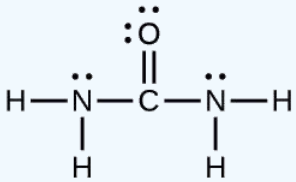
The selenium atom is sp^3d hybridized.

Example 6.3.2: Assigning Hybridization

Urea, $NH_2C(O)NH_2$, is sometimes used as a source of nitrogen in fertilizers. What is the hybridization of the carbon atom in urea?

Solution

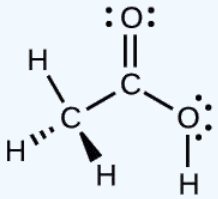
The Lewis structure of urea is



The carbon atom is surrounded by three regions of electron density, positioned in a trigonal planar arrangement. The hybridization in a trigonal planar electron pair geometry is sp^2 (Figure 6.3.16), which is the hybridization of the carbon atom in urea.

Exercise 6.3.1

Acetic acid, $H_3CC(O)OH$, is the molecule that gives vinegar its odor and sour taste. What is the hybridization of the two carbon atoms in acetic acid?



Answer

$H_3\underline{C}$, sp^3; $\underline{C}(O)OH$, sp^2

Footnotes

- 1Note that orbitals may sometimes be drawn in an elongated “balloon” shape rather than in a more realistic “plump” shape in order to make the geometry easier to visualize.

This page titled 6.3: Hybrid Atomic Orbitals is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **8.2: Hybrid Atomic Orbitals** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

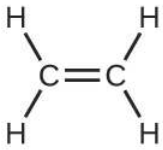
6.4: Multiple Bonds

Learning Objectives

By the end of this section, you will be able to:

- Describe multiple covalent bonding in terms of atomic orbital overlap
- Relate the concept of resonance to π-bonding and electron delocalization

The hybrid orbital model appears to account well for the geometry of molecules involving single covalent bonds. Is it also capable of describing molecules containing double and triple bonds? We have already discussed that multiple bonds consist of σ and π bonds. Next we can consider how we visualize these components and how they relate to hybrid orbitals. The Lewis structure of ethene, C_2H_4, shows us that each carbon atom is surrounded by one other carbon atom and two hydrogen atoms.



The three bonding regions form a trigonal planar electron-pair geometry. Thus we expect the σ bonds from each carbon atom are formed using a set of sp^2 hybrid orbitals that result from hybridization of two of the 2*p* orbitals and the 2*s* orbital (Figure 6.4.1).

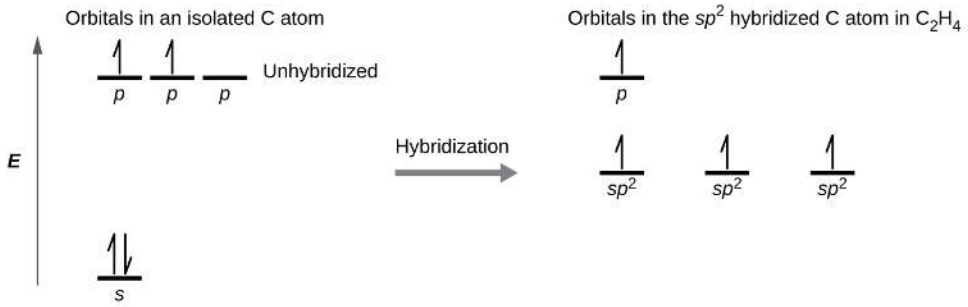


Figure 6.4.1: In ethene, each carbon atom is sp^2 hybridized, and the sp^2 orbitals and the *p* orbital are singly occupied. The hybrid orbitals overlap to form σ bonds, while the *p* orbitals on each carbon atom overlap to form a π bond.

These orbitals form the C–H single bonds and the σ bond in the Figure 6.4.2. The π bond in the C=C double bond results from the overlap of the third (remaining) 2p orbital on each carbon atom that is not involved in hybridization. This unhybridized p orbital (lobes shown in red and blue in Figure 6.4.2) is perpendicular to the plane of the sp2 hybrid orbitals. Thus the unhybridized 2p orbitals overlap in a side-by-side fashion, above and below the internuclear axis (Figure 6.4.2) and form a π bond.

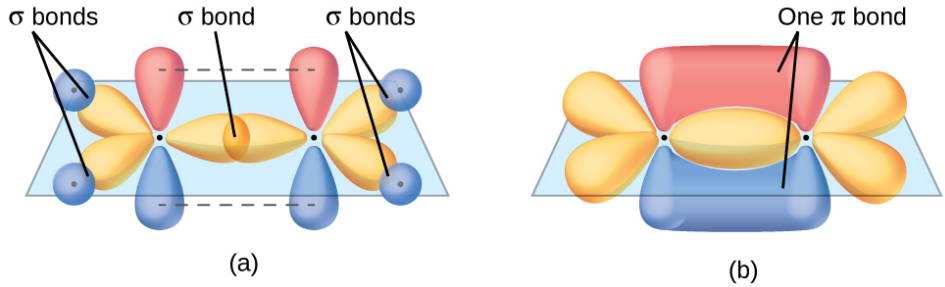


Figure 6.4.2: In the ethene molecule, C_2H_4, there are (a) five σ bonds. One C–C σ bond results from overlap of sp^2 hybrid orbitals on the carbon atom with one sp^2 hybrid orbital on the other carbon atom. Four C–H bonds result from the overlap between the C atoms' sp^2 orbitals with *s* orbitals on the hydrogen atoms. (b) The π bond is formed by the side-by-side overlap of the two unhybridized *p* orbitals in the two carbon atoms. The two lobes of the π bond are above and below the plane of the σ system.

In an ethene molecule, the four hydrogen atoms and the two carbon atoms are all in the same plane. If the two planes of sp^2 hybrid orbitals tilted relative to each other, the *p* orbitals would not be oriented to overlap efficiently to create the π bond. The planar configuration for the ethene molecule occurs because it is the most stable bonding arrangement. This is a significant difference

between σ and π bonds; rotation around single (σ) bonds occurs easily because the end-to-end orbital overlap does not depend on the relative orientation of the orbitals on each atom in the bond. In other words, rotation around the internuclear axis does not change the extent to which the σ bonding orbitals overlap because the bonding electron density is symmetric about the axis. Rotation about the internuclear axis is much more difficult for multiple bonds; however, this would drastically alter the off-axis overlap of the π bonding orbitals, essentially breaking the π bond.

In molecules with *sp* hybrid orbitals, two unhybridized *p* orbitals remain on the atom (Figure 6.4.3). We find this situation in acetylene, $\mathrm{H-C\equiv C-H}$ (Figure 6.4.4), The remaining sp orbitals form σ bonds with hydrogen atoms. The two unhybridized *p* orbitals per carbon are positioned such that they overlap side by side and, hence, form two π bonds. The two carbon atoms of acetylene are thus bound together by one σ bond and two π bonds, giving a triple bond.

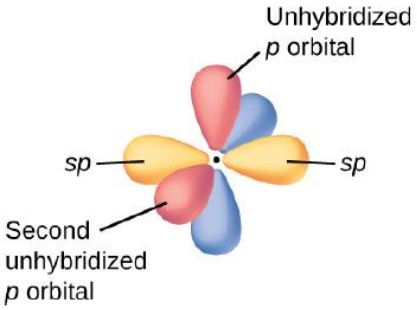


Figure 6.4.3: Diagram of the two linear *sp* hybrid orbitals of a carbon atom, which lie in a straight line, and the two unhybridized *p* orbitals at perpendicular angles.

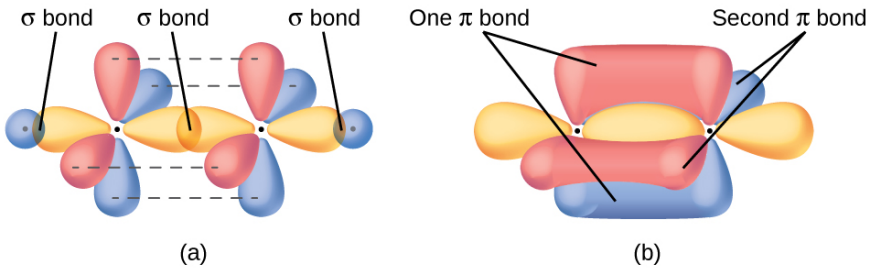


Figure 6.4.4: (a) In the acetylene molecule, C_2H_2, there are two C–H σ bonds and a triple bond involving one C–C σ bond and two C–C π bonds. The dashed lines, each connecting two lobes, indicate the side-by-side overlap of the four unhybridized *p* orbitals. (b) This shows the overall outline of the bonds in C_2H_2. The two lobes of each of the π bonds are positioned across from each other around the line of the C–C σ bond.

Hybridization involves only σ bonds, lone pairs of electrons, and single unpaired electrons (radicals). Structures that account for these features describe the correct hybridization of the atoms. However, many structures also include resonance forms. Remember that resonance forms occur when various arrangements of π bonds are possible. Since the arrangement of π bonds involves only the unhybridized orbitals, resonance does not influence the assignment of hybridization.

For example, molecule benzene has two resonance forms (Figure 6.4.5). We can use either of these forms to determine that each of the carbon atoms is bonded to three other atoms with no lone pairs, so the correct hybridization is sp^2. The electrons in the unhybridized *p* orbitals form π bonds. Neither resonance structure completely describes the electrons in the π bonds. They are not located in one position or the other, but in reality are delocalized throughout the ring. Valence bond theory does not easily address delocalization. Bonding in molecules with resonance forms is better described by molecular orbital theory. (See the next module.)

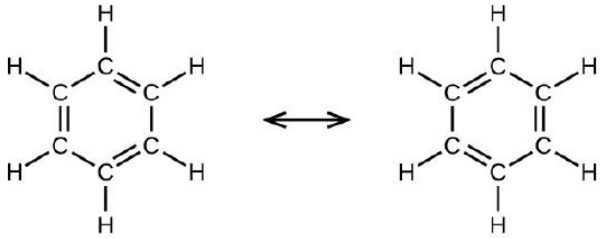


Figure 6.4.5: Each carbon atom in benzene, C_6H_6, is sp^2 hybridized, independently of which resonance form is considered. The electrons in the π bonds are not located in one set of *p* orbitals or the other, but rather delocalized throughout the molecule.

Example 6.4.1: Assignment of Hybridization Involving Resonance

Some acid rain results from the reaction of sulfur dioxide with atmospheric water vapor, followed by the formation of sulfuric acid. Sulfur dioxide, SO_2, is a major component of volcanic gases as well as a product of the combustion of sulfur-containing coal. What is the hybridization of the S atom in SO_2?

Solution

The resonance structures of SO_2 are



The sulfur atom is surrounded by two bonds and one lone pair of electrons in either resonance structure. Therefore, the electron-pair geometry is trigonal planar, and the hybridization of the sulfur atom is sp^2.

Exercise 6.4.1

Another acid in acid rain is nitric acid, HNO_3, which is produced by the reaction of nitrogen dioxide, NO_2, with atmospheric water vapor. What is the hybridization of the nitrogen atom in NO_2? (Note: the lone electron on nitrogen occupies a hybridized orbital just as a lone pair would.)

Answer

sp^2

This page titled 6.4: Multiple Bonds is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **8.3: Multiple Bonds** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

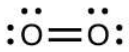
6.5: Molecular Orbital Theory

Learning Objectives

By the end of this section, you will be able to:

- Outline the basic quantum-mechanical approach to deriving molecular orbitals from atomic orbitals
- Describe traits of bonding and antibonding molecular orbitals
- Calculate bond orders based on molecular electron configurations
- Write molecular electron configurations for first- and second-row diatomic molecules
- Relate these electron configurations to the molecules' stabilities and magnetic properties

For almost every covalent molecule that exists, we can now draw the Lewis structure, predict the electron-pair geometry, predict the molecular geometry, and come close to predicting bond angles. However, one of the most important molecules we know, the oxygen molecule O_2, presents a problem with respect to its Lewis structure. We would write the following Lewis structure for O_2:



This electronic structure adheres to all the rules governing Lewis theory. There is an O=O double bond, and each oxygen atom has eight electrons around it. However, this picture is at odds with the magnetic behavior of oxygen. By itself, O_2 is not magnetic, but it is attracted to magnetic fields. Thus, when we pour liquid oxygen past a strong magnet, it collects between the poles of the magnet and defies gravity. Such attraction to a magnetic field is called **paramagnetism**, and it arises in molecules that have unpaired electrons. And yet, the Lewis structure of O_2 indicates that all electrons are paired. How do we account for this discrepancy?

Magnetic susceptibility measures the force experienced by a substance in a magnetic field. When we compare the weight of a sample to the weight measured in a magnetic field (Figure 6.5.1), paramagnetic samples that are attracted to the magnet will appear heavier because of the force exerted by the magnetic field. We can calculate the number of unpaired electrons based on the increase in weight.

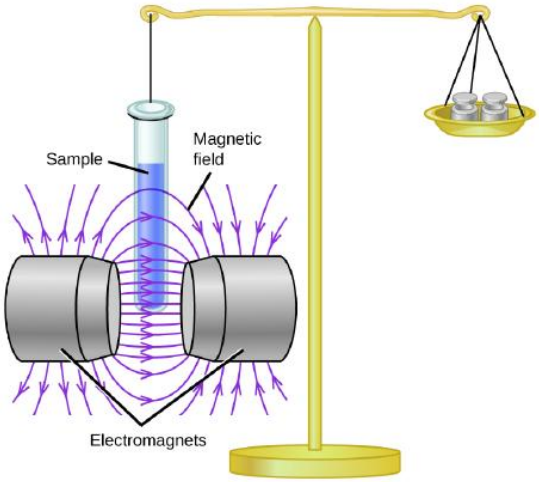


Figure 6.5.1: A Gouy balance compares the mass of a sample in the presence of a magnetic field with the mass with the electromagnet turned off to determine the number of unpaired electrons in a sample.

Experiments show that each O_2 molecule has two unpaired electrons. The Lewis-structure model does not predict the presence of these two unpaired electrons. Unlike oxygen, the apparent weight of most molecules decreases slightly in the presence of an inhomogeneous magnetic field. Materials in which all of the electrons are paired are diamagnetic and weakly repel a magnetic field. Paramagnetic and diamagnetic materials do not act as permanent magnets. Only in the presence of an applied magnetic field do they demonstrate attraction or repulsion.

Link to Learning

Water, like most molecules, contains all paired electrons. Living things contain a large percentage of water, so they demonstrate diamagnetic behavior. If you place a frog near a sufficiently large magnet, it will levitate. You can see videos of diamagnetic floating frogs, strawberries, and more.

Molecular orbital theory (MO theory) provides an explanation of chemical bonding that accounts for the paramagnetism of the oxygen molecule. It also explains the bonding in a number of other molecules, such as violations of the octet rule and more molecules with more complicated bonding (beyond the scope of this text) that are difficult to describe with Lewis structures. Additionally, it provides a model for describing the energies of electrons in a molecule and the probable location of these electrons. Unlike valence bond theory, which uses hybrid orbitals that are assigned to one specific atom, MO theory uses the combination of atomic orbitals to yield molecular orbitals that are *delocalized* over the entire molecule rather than being localized on its constituent atoms. MO theory also helps us understand why some substances are electrical conductors, others are semiconductors, and still others are insulators. Table 6.5.1 summarizes the main points of the two complementary bonding theories. Both theories provide different, useful ways of describing molecular structure.

Table 6.5.1: Comparison of Bonding Theories

Valence Bond Theory	Molecular Orbital Theory
considers bonds as localized between one pair of atoms	considers electrons delocalized throughout the entire molecule
creates bonds from overlap of atomic orbitals (*s*, *p*, *d*…) and hybrid orbitals (*sp*, sp^2, sp^3…)	combines atomic orbitals to form molecular orbitals (σ, σ*, π, π*)
forms σ or π bonds	creates bonding and antibonding interactions based on which orbitals are filled
predicts molecular shape based on the number of regions of electron density	predicts the arrangement of electrons in molecules
needs multiple structures to describe resonance	

Molecular orbital theory describes the distribution of electrons in molecules in much the same way that the distribution of electrons in atoms is described using atomic orbitals. Using quantum mechanics, the behavior of an electron in a molecule is still described by a wave function, Ψ, analogous to the behavior in an atom. Just like electrons around isolated atoms, electrons around atoms in molecules are limited to discrete (quantized) energies. The region of space in which a valence electron in a molecule is likely to be found is called a molecular orbital (Ψ^2). Like an atomic orbital, a molecular orbital is full when it contains two electrons with opposite spin.

We will consider the molecular orbitals in molecules composed of two identical atoms (H_2 or Cl_2, for example). Such molecules are called homonuclear diatomic molecules. In these diatomic molecules, several types of molecular orbitals occur.

The mathematical process of combining atomic orbitals to generate molecular orbitals is called the **linear combination of atomic orbitals** (LCAO). The wave function describes the wavelike properties of an electron. Molecular orbitals are combinations of atomic orbital wave functions. Combining waves can lead to constructive interference, in which peaks line up with peaks, or destructive interference, in which peaks line up with troughs (Figure 6.5.2). In orbitals, the waves are three dimensional, and they combine with in-phase waves producing regions with a higher probability of electron density and out-of-phase waves producing nodes, or regions of no electron density.

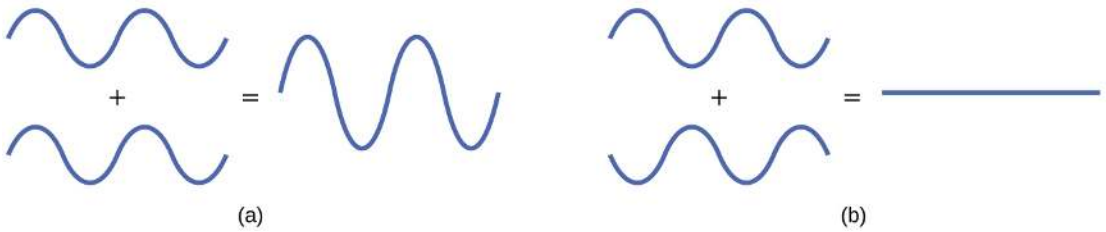


Figure 6.5.2: (a) When in-phase waves combine, constructive interference produces a wave with greater amplitude. (b) When out-of-phase waves combine, destructive interference produces a wave with less (or no) amplitude.

There are two types of molecular orbitals that can form from the overlap of two atomic *s* orbitals on adjacent atoms. The two types are illustrated in Figure 6.5.3. The in-phase combination produces a lower energy σ_s molecular orbital (read as "sigma-s") in which most of the electron density is directly between the nuclei. The out-of-phase addition (which can also be thought of as subtracting the wave functions) produces a higher energy molecular orbital (read as "sigma-s-star") molecular orbital in which there is a node between the nuclei. The asterisk signifies that the orbital is an antibonding orbital. Electrons in a σ_s orbital are attracted by both nuclei at the same time and are more stable (of lower energy) than they would be in the isolated atoms. Adding electrons to these orbitals creates a force that holds the two nuclei together, so we call these orbitals bonding orbitals. Electrons in the orbitals are located well away from the region between the two nuclei. The attractive force between the nuclei and these electrons pulls the two nuclei apart. Hence, these orbitals are called antibonding orbitals. Electrons fill the lower-energy bonding orbital before the higher-energy antibonding orbital, just as they fill lower-energy atomic orbitals before they fill higher-energy atomic orbitals.

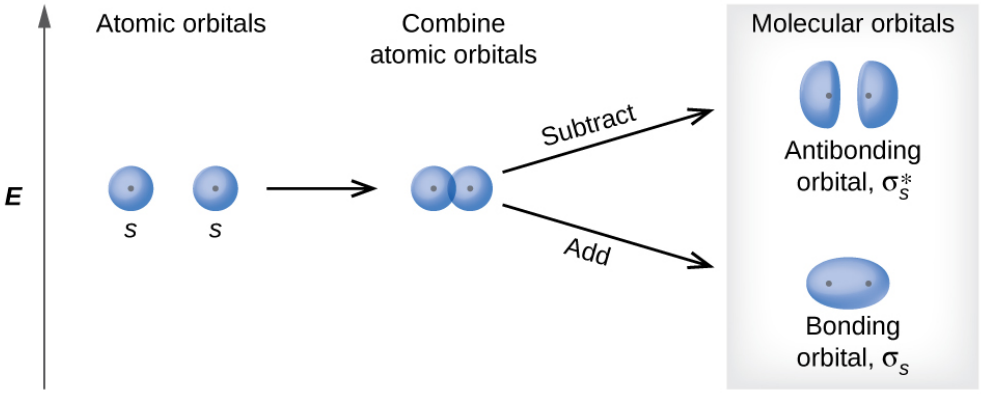


Figure 6.5.3: Sigma (σ) and sigma-star (σ*) molecular orbitals are formed by the combination of two *s* atomic orbitals. The dots (·) indicate the locations of nuclei.

Link to Learning

You can watch animations visualizing the calculated atomic orbitals combining to form various molecular orbitals at the Orbitron website.

In *p* orbitals, the wave function gives rise to two lobes with opposite phases, analogous to how a two-dimensional wave has both parts above and below the average. We indicate the phases by shading the orbital lobes different colors. When orbital lobes of the same phase overlap, constructive wave interference increases the electron density. When regions of opposite phase overlap, the destructive wave interference decreases electron density and creates nodes. When *p* orbitals overlap end to end, they create σ and σ* orbitals (Figure 6.5.4). If two atoms are located along the *x*-axis in a Cartesian coordinate system, the two p_x orbitals overlap end to end and form σ_{px} (bonding) and (antibonding) (read as "sigma-p-x" and "sigma-p-x star," respectively). Just as with *s*-orbital overlap, the asterisk indicates the orbital with a node between the nuclei, which is a higher-energy, antibonding orbital.

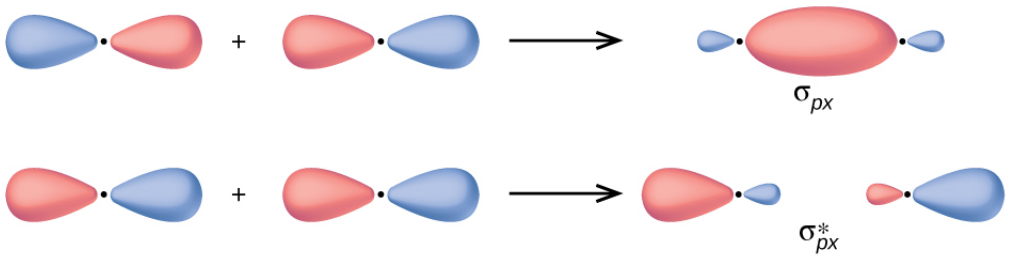


Figure 6.5.4: Combining wave functions of two *p* atomic orbitals along the internuclear axis creates two molecular orbitals, σ_p and σ_p^*.

The side-by-side overlap of two *p* orbitals gives rise to a **pi (π) bonding molecular orbital** and a **π* antibonding molecular orbital**, as shown in Figure 6.5.5. In valence bond theory, we describe π bonds as containing a nodal plane containing the internuclear axis and perpendicular to the lobes of the *p* orbitals, with electron density on either side of the node. In molecular orbital theory, we describe the π orbital by this same shape, and a π bond exists when this orbital contains electrons. Electrons in

this orbital interact with both nuclei and help hold the two atoms together, making it a bonding orbital. For the out-of-phase combination, there are two nodal planes created, one along the internuclear axis and a perpendicular one between the nuclei.

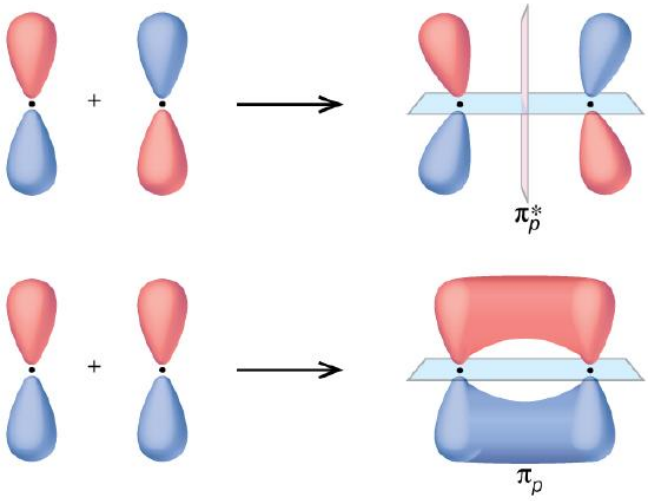
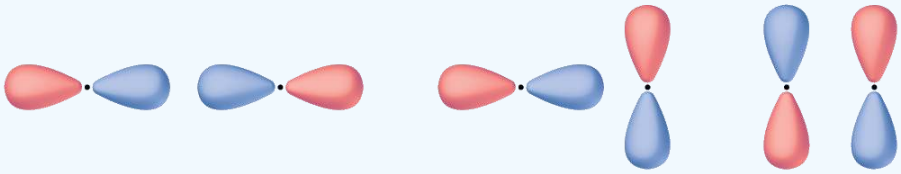


Figure 6.5.5: Side-by-side overlap of each two *p* orbitals results in the formation of two π molecular orbitals. Combining the out-of-phase orbitals results in an antibonding molecular orbital with two nodes. One contains the internuclear axis, and one is perpendicular to the axis. Combining the in-phase orbitals results in a bonding orbital. There is a node (blue) containing the internuclear axis with the two lobes of the orbital located above and below this node.

In the molecular orbitals of diatomic molecules, each atom also has two sets of *p* orbitals oriented side by side (p_y and p_z), so these four atomic orbitals combine pairwise to create two π orbitals and two π* orbitals. The π_{py} and orbitals are oriented at right angles to the π_{pz} and π^*_{pz} orbitals. Except for their orientation, the \(π_{py}\) and \(π_{pz}\) orbitals are identical and have the same energy; they are **degenerate orbitals**. The π^*_{py} and π^*_{pz} antibonding orbitals are also degenerate and identical except for their orientation. A total of six molecular orbitals results from the combination of the six atomic p orbitals in two atoms: σ_{px} and σ^*_{px}, π_{py} and π^*_{py}, π_{pz} and π^*_{pz}.

Example 6.5.1: Molecular Orbitals

Predict what type (if any) of molecular orbital would result from adding the wave functions so each pair of orbitals shown overlap. The orbitals are all similar in energy.



Solution

a. is an in-phase combination, resulting in a π_{3p} orbital
b. will not result in a new orbital because the in-phase component (bottom) and out-of-phase component (top) cancel out. Only orbitals with the correct alignment can combine.
c. is an out-of-phase combination, resulting in a π^*_{3p} orbital.

Exercise 6.5.1

Label the molecular orbital shown as σ or π, bonding or antibonding and indicate where the node occurs.



Answer

The orbital is located along the internuclear axis, so it is a σ orbital. There is a node bisecting the internuclear axis, so it is an antibonding orbital.



Portrait of a Chemist: Walter Kohn: Nobel Laureate

Walter Kohn (Figure 6.5.6) is a theoretical physicist who studies the electronic structure of solids. His work combines the principles of quantum mechanics with advanced mathematical techniques. This technique, called density functional theory, makes it possible to compute properties of molecular orbitals, including their shape and energies. Kohn and mathematician John Pople were awarded the Nobel Prize in Chemistry in 1998 for their contributions to our understanding of electronic structure. Kohn also made significant contributions to the physics of semiconductors.

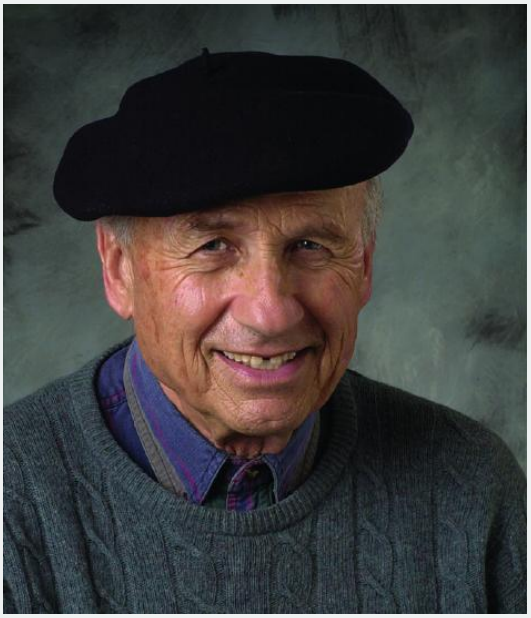


Figure 6.5.6: Walter Kohn developed methods to describe molecular orbitals. (credit: image courtesy of Walter Kohn)

Kohn's biography has been remarkable outside the realm of physical chemistry as well. He was born in Austria, and during World War II he was part of the Kindertransport program that rescued 10,000 children from the Nazi regime. His summer jobs included discovering gold deposits in Canada and helping Polaroid explain how its instant film worked. Dr. Kohn passed away in 2016 at the age of 93.

How Sciences Interconnect: Computational Chemistry in Drug Design

While the descriptions of bonding described in this chapter involve many theoretical concepts, they also have many practical, real-world applications. For example, drug design is an important field that uses our understanding of chemical bonding to develop pharmaceuticals. This interdisciplinary area of study uses biology (understanding diseases and how they operate) to identify specific targets, such as a binding site that is involved in a disease pathway. By modeling the structures of the binding site and potential drugs, computational chemists can predict which structures can fit together and how effectively they will bind (see Figure 6.5.7). Thousands of potential candidates can be narrowed down to a few of the most promising candidates. These candidate molecules are then carefully tested to determine side effects, how effectively they can be transported through

the body, and other factors. Dozens of important new pharmaceuticals have been discovered with the aid of computational chemistry, and new research projects are underway.

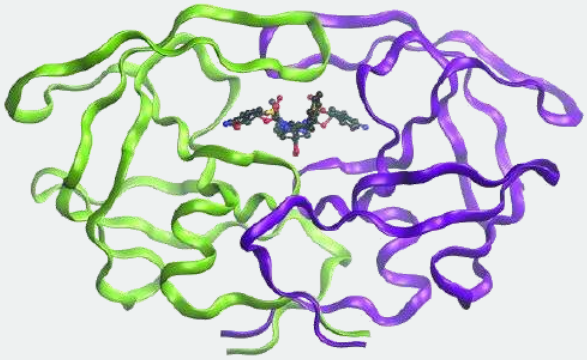


Figure 6.5.7: The molecule shown, HIV-1 protease, is an important target for pharmaceutical research. By designing molecules that bind to this protein, scientists are able to drastically inhibit the progress of the disease.

Molecular Orbital Energy Diagrams

The relative energy levels of atomic and molecular orbitals are typically shown in a molecular orbital diagram (Figure 6.5.8). For a diatomic molecule, the atomic orbitals of one atom are shown on the left, and those of the other atom are shown on the right. Each horizontal line represents one orbital that can hold two electrons. The molecular orbitals formed by the combination of the atomic orbitals are shown in the center. Dashed lines show which of the atomic orbitals combine to form the molecular orbitals. For each pair of atomic orbitals that combine, one lower-energy (bonding) molecular orbital and one higher-energy (antibonding) orbital result. Thus we can see that combining the six 2*p* atomic orbitals results in three bonding orbitals (one σ and two π) and three antibonding orbitals (one σ* and two π*).

We predict the distribution of electrons in these molecular orbitals by filling the orbitals in the same way that we fill atomic orbitals, by the Aufbau principle. Lower-energy orbitals fill first, electrons spread out among degenerate orbitals before pairing, and each orbital can hold a maximum of two electrons with opposite spins (Figure 6.5.8). Just as we write electron configurations for atoms, we can write the molecular electronic configuration by listing the orbitals with superscripts indicating the number of electrons present. For clarity, we place parentheses around molecular orbitals with the same energy. In this case, each orbital is at a different energy, so parentheses separate each orbital. Thus we would expect a diatomic molecule or ion containing seven electrons (such as Be_2^+) would have the molecular electron configuration $(\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^1$. It is common to omit the core electrons from molecular orbital diagrams and configurations and include only the valence electrons.

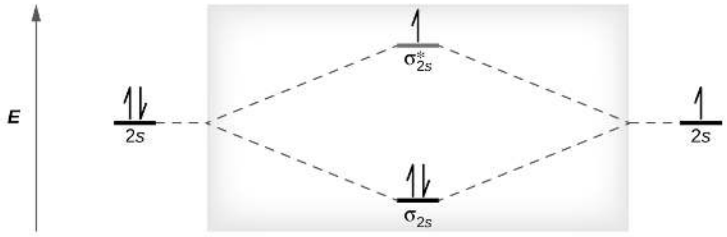


Figure 6.5.8: This is the molecular orbital diagram for the homonuclear diatomic showing the molecular orbitals of the valence shell only. The molecular orbitals are filled in the same manner as atomic orbitals, using the Aufbau principle and Hund's rule.

Bond Order

The filled molecular orbital diagram shows the number of electrons in both bonding and antibonding molecular orbitals. The net contribution of the electrons to the bond strength of a molecule is identified by determining the **bond order** that results from the filling of the molecular orbitals by electrons.

When using Lewis structures to describe the distribution of electrons in molecules, we define bond order as the number of bonding pairs of electrons between two atoms. Thus a single bond has a bond order of 1, a double bond has a bond order of 2, and a triple bond has a bond order of 3. We define bond order differently when we use the molecular orbital description of the distribution of electrons, but the resulting bond order is usually the same. The MO technique is more accurate and can handle cases when the Lewis structure method fails, but both methods describe the same phenomenon.

In the molecular orbital model, an electron contributes to a bonding interaction if it occupies a bonding orbital and it contributes to an antibonding interaction if it occupies an antibonding orbital. The bond order is calculated by subtracting the destabilizing (antibonding) electrons from the stabilizing (bonding) electrons. Since a bond consists of two electrons, we divide by two to get the bond order. We can determine bond order with the following equation:

$$\text{bond order } = \frac{(\text{number of bonding electrons }) - (\text{ number of antibonding electrons })}{2}$$

The order of a covalent bond is a guide to its strength; a bond between two given atoms becomes stronger as the bond order increases (Table 8.1). If the distribution of electrons in the molecular orbitals between two atoms is such that the resulting bond would have a bond order of zero, a stable bond does not form. We next look at some specific examples of MO diagrams and bond orders.

Bonding in Diatomic Molecules

A dihydrogen molecule (H_2) forms from two hydrogen atoms. When the atomic orbitals of the two atoms combine, the electrons occupy the molecular orbital of lowest energy, the σ_{1s} bonding orbital. A dihydrogen molecule, H_2, readily forms because the energy of a H_2 molecule is lower than that of two H atoms. The σ_{1s} orbital that contains both electrons is lower in energy than either of the two 1*s* atomic orbitals.

A molecular orbital can hold two electrons, so both electrons in the H_2 molecule are in the σ1s bonding orbital; the electron configuration is $(\sigma_{1s})^2$. We represent this configuration by a molecular orbital energy diagram (Figure 6.5.9) in which a single upward arrow indicates one electron in an orbital, and two (upward and downward) arrows indicate two electrons of opposite spin.

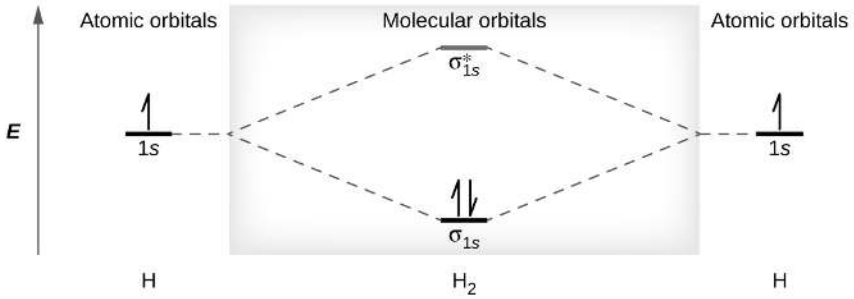


Figure 6.5.9: The molecular orbital energy diagram predicts that H_2 will be a stable molecule with lower energy than the separated atoms.

A dihydrogen molecule contains two bonding electrons and no antibonding electrons so we have

$$\text{bond order in } H_2 = \frac{(2-0)}{2} = 1$$

Because the bond order for the H–H bond is equal to 1, the bond is a single bond.

A helium atom has two electrons, both of which are in its 1*s* orbital. Two helium atoms do not combine to form a dihelium molecule, He_2, with four electrons, because the stabilizing effect of the two electrons in the lower-energy bonding orbital would be offset by the destabilizing effect of the two electrons in the higher-energy antibonding molecular orbital. We would write the hypothetical electron configuration of He_2 as

$$\text{bond order in } He_2 = \frac{(2-2)}{2} = 0$$

A bond order of zero indicates that no bond is formed between two atoms.

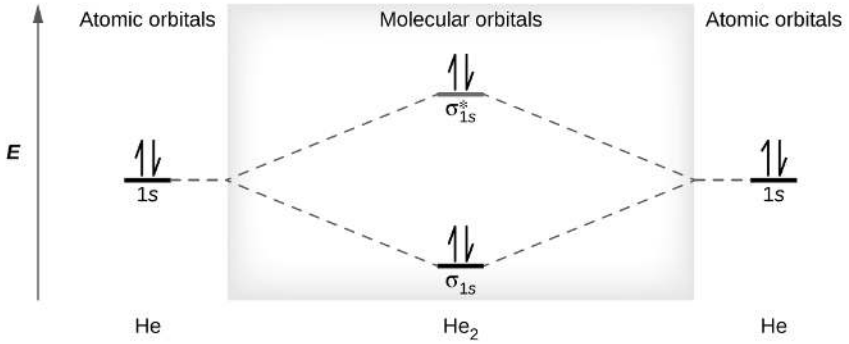


Figure 6.5.10: The molecular orbital energy diagram predicts that He_2 will not be a stable molecule, since it has equal numbers of bonding and antibonding electrons.

The Diatomic Molecules of the Second Period

Eight possible homonuclear diatomic molecules might be formed by the atoms of the second period of the periodic table: Li_2, Be_2, B_2, C_2, N_2, O_2, F_2, and Ne_2. However, we can predict that the Be_2 molecule and the Ne_2 molecule would not be stable. We can see this by a consideration of the molecular electron configurations (Table 6.5.2).

We predict valence molecular orbital electron configurations just as we predict electron configurations of atoms. Valence electrons are assigned to valence molecular orbitals with the lowest possible energies. Consistent with Hund's rule, whenever there are two or more degenerate molecular orbitals, electrons fill each orbital of that type singly before any pairing of electrons takes place.

As we saw in valence bond theory, σ bonds are generally more stable than π bonds formed from degenerate atomic orbitals. Similarly, in molecular orbital theory, σ orbitals are usually more stable than π orbitals. However, this is not always the case. The MOs for the valence orbitals of the second period are shown in Figure 6.5.11. Looking at Ne_2 molecular orbitals, we see that the order is consistent with the generic diagram shown in the previous section. However, for atoms with three or fewer electrons in the *p* orbitals (Li through N) we observe a different pattern, in which the σ_p orbital is higher in energy than the π_p set. Obtain the molecular orbital diagram for a homonuclear diatomic ion by adding or subtracting electrons from the diagram for the neutral molecule.

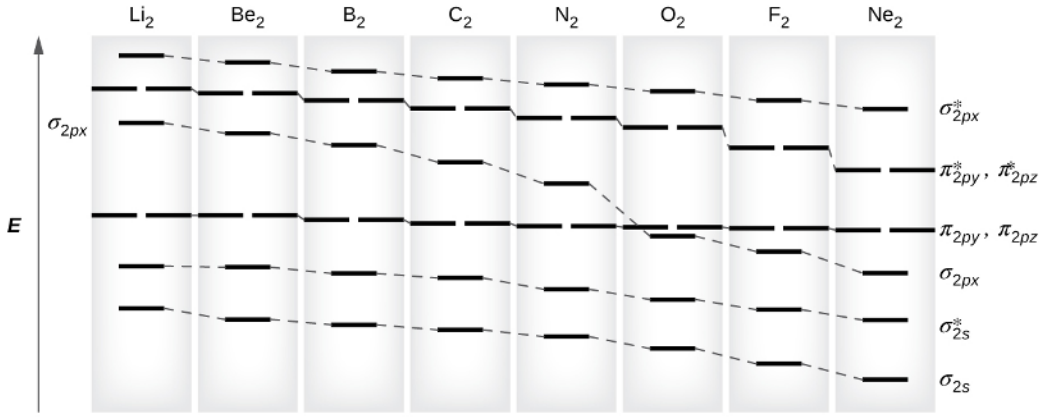


Figure 6.5.11: This shows the MO diagrams for each homonuclear diatomic molecule in the second period. The orbital energies decrease across the period as the effective nuclear charge increases and atomic radius decreases. Between N_2 and O_2, the order of the orbitals changes.

This switch in orbital ordering occurs because of a phenomenon called s-p mixing. s-p mixing does not create new orbitals; it merely influences the energies of the existing molecular orbitals. The σ_s wavefunction mathematically combines with the σ_p wavefunction, with the result that the σ_s orbital becomes more stable, and the σ_p orbital becomes less stable (Figure 6.5.12). Similarly, the antibonding orbitals also undergo s-p mixing, with the σ_{s*} becoming more stable and the σ_{p*} becoming less stable.

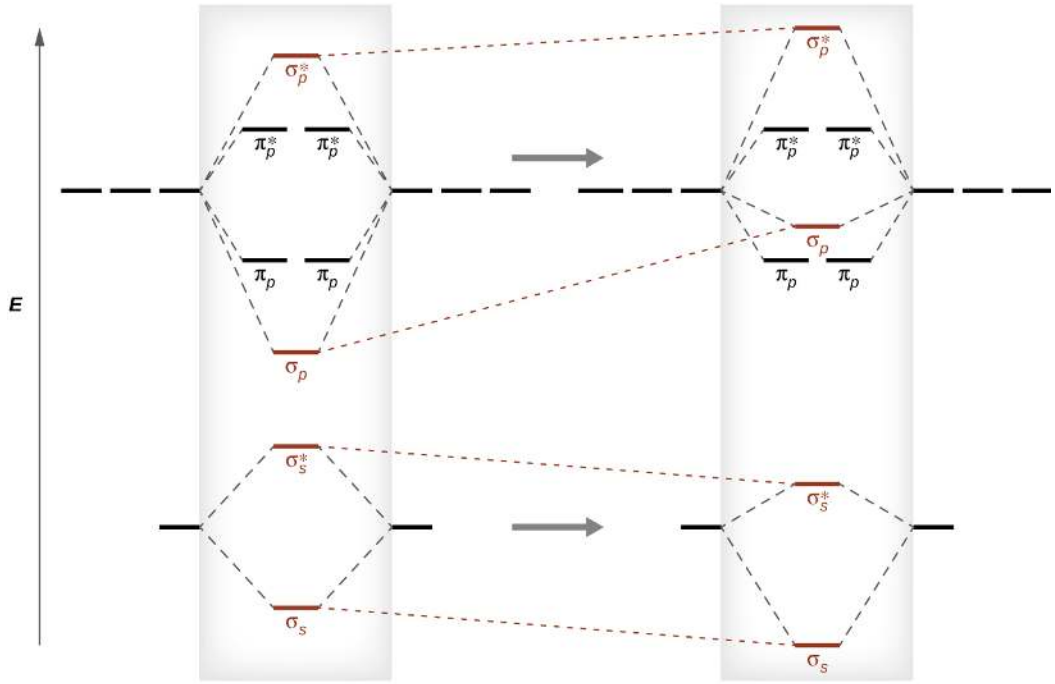


Figure 6.5.12: Without mixing, the MO pattern occurs as expected, with the σ_p orbital lower in energy than the π_p orbitals. When s-p mixing occurs, the orbitals shift as shown, with the σ_p orbital higher in energy than the π_p orbitals.

s-p mixing occurs when the *s* and *p* orbitals have similar energies. The energy difference between 2*s* and 2*p* orbitals in O, F, and Ne is greater than that in Li, Be, B, C, and N. Because of this, O_2, F_2, and Ne_2 exhibit negligible s-p mixing (not sufficient to change the energy ordering), and their MO diagrams follow the normal pattern, as shown in Figure 6.5.11. All of the other period 2 diatomic molecules do have s-p mixing, which leads to the pattern where the σ_p orbital is raised above the π_p set.

Using the MO diagrams shown in Figure 6.5.11: , we can add in the electrons and determine the molecular electron configuration and bond order for each of the diatomic molecules. As shown in Table 6.5.2, Be_2 and Ne_2 molecules would have a bond order of 0, and these molecules do not exist.

Table 6.5.2: Electron Configuration and Bond Order for Molecular Orbitals in Homonuclear Diatomic Molecules of Period Two Elements

Molecule	Electron Configuration	Bond Order
Li_2	$(\sigma_{2s})^2$	1
Be_2 (unstable)	$(\sigma_{2s})^2(\sigma_{2s}^*)^2$	0
B_2	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2py},\pi_{2pz})^2$	1
C_2	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2py},\pi_{2pz})^4$	2
N_2	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2py},\pi_{2pz})^4(\sigma_{2px})^2$	3
O_2	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2px})^2(\pi_{2py},\pi_{2pz})^4\left(\pi_{2py}^*,\pi_{2pz}^*\right)^2$	2
F_2	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2px})^2(\pi_{2py},\pi_{2pz})^4\left(\pi_{2py}^*,\pi_{2pz}^*\right)^4$	1
Ne_2 (unstable)	$(\sigma_{2s})^2(\sigma_{2s}^*)^2(\sigma_{2px})^2(\pi_{2py},\pi_{2pz})^4\left(\pi_{2py}^*,\pi_{2pz}^*\right)^4\left(\sigma_{2px}^*\right)^2$	0

The combination of two lithium atoms to form a lithium molecule, Li_2, is analogous to the formation of H_2, but the atomic orbitals involved are the valence 2*s* orbitals. Each of the two lithium atoms has one valence electron. Hence, we have two valence electrons available for the σ_{2s} bonding molecular orbital. Because both valence electrons would be in a bonding orbital, we would predict the

Li_2 molecule to be stable. The molecule is, in fact, present in appreciable concentration in lithium vapor at temperatures near the boiling point of the element. All of the other molecules in Table 6.5.2 with a bond order greater than zero are also known.

The O2 molecule has enough electrons to half fill the $\left(\pi^*_{2py}, \pi^*_{2pz}\right)$ level. We expect the two electrons that occupy these two degenerate orbitals to be unpaired, and this molecular electronic configuration for O2 is in accord with the fact that the oxygen molecule has two unpaired electrons (Figure 8.40). The presence of two unpaired electrons has proved to be difficult to explain using Lewis structures, but the molecular orbital theory explains it quite well. In fact, the unpaired electrons of the oxygen molecule provide a strong piece of support for the molecular orbital theory.

How Sciences Interconnect: Band Theory

When two identical atomic orbitals on different atoms combine, two molecular orbitals result (see Figure 6.5.3). The bonding orbital is lower in energy than the original atomic orbitals because the atomic orbitals are in-phase in the molecular orbital. The antibonding orbital is higher in energy than the original atomic orbitals because the atomic orbitals are out-of-phase.

In a solid, similar things happen, but on a much larger scale. Remember that even in a small sample there are a huge number of atoms (typically > 10^{23} atoms), and therefore a huge number of atomic orbitals that may be combined into molecular orbitals. When *N* valence atomic orbitals, all of the same energy and each containing one (1) electron, are combined, *N*/2 (filled) bonding orbitals and *N*/2 (empty) antibonding orbitals will result. Each bonding orbital will show an energy lowering as the atomic orbitals are *mostly* in-phase, but each of the bonding orbitals will be a little different and have slightly different energies. The antibonding orbitals will show an increase in energy as the atomic orbitals are *mostly* out-of-phase, but each of the antibonding orbitals will also be a little different and have slightly different energies. The allowed energy levels for all the bonding orbitals are so close together that they form a band, called the valence band. Likewise, all the antibonding orbitals are very close together and form a band, called the conduction band. Figure 6.5.13 shows the bands for three important classes of materials: insulators, semiconductors, and conductors.

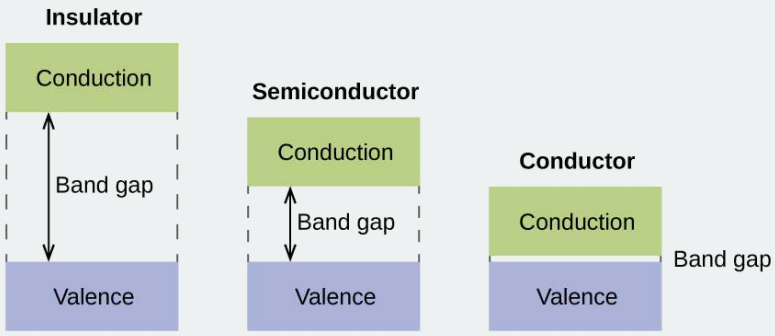


Figure 6.5.13: Molecular orbitals in solids are so closely spaced that they are described as bands. The valence band is lower in energy and the conduction band is higher in energy. The type of solid is determined by the size of the "band gap" between the valence and conduction bands. Only a very small amount of energy is required to move electrons from the valence band to the conduction band in a conductor, and so they conduct electricity well. In an insulator, the band gap is large, so that very few electrons move, and they are poor conductors of electricity. Semiconductors are in between: they conduct electricity better than insulators, but not as well as conductors.

In order to conduct electricity, electrons must move from the filled valence band to the empty conduction band where they can move throughout the solid. The size of the band gap, or the energy difference between the top of the valence band and the bottom of the conduction band, determines how easy it is to move electrons between the bands. Only a small amount of energy is required in a conductor because the band gap is very small. This small energy difference is "easy" to overcome, so they are good conductors of electricity. In an insulator, the band gap is so "large" that very few electrons move into the conduction band; as a result, insulators are poor conductors of electricity. Semiconductors conduct electricity when "moderate" amounts of energy are provided to move electrons out of the valence band and into the conduction band. Semiconductors, such as silicon, are found in many electronics.

Semiconductors are used in devices such as computers, smartphones, and solar cells. Solar cells produce electricity when light provides the energy to move electrons out of the valence band. The electricity that is generated may then be used to power a light or tool, or it can be stored for later use by charging a battery. As of December 2014, up to 46% of the energy in sunlight could be converted into electricity using solar cells.

Example 6.5.2: Molecular Orbital Diagrams, Bond Order, and Number of Unpaired Electrons

Draw the molecular orbital diagram for the oxygen molecule, O_2. From this diagram, calculate the bond order for O_2. How does this diagram account for the paramagnetism of O_2?

Solution

We draw a molecular orbital energy diagram similar to that shown in Figure 6.5.11. Each oxygen atom contributes six electrons, so the diagram appears as shown in Figure 6.5.14

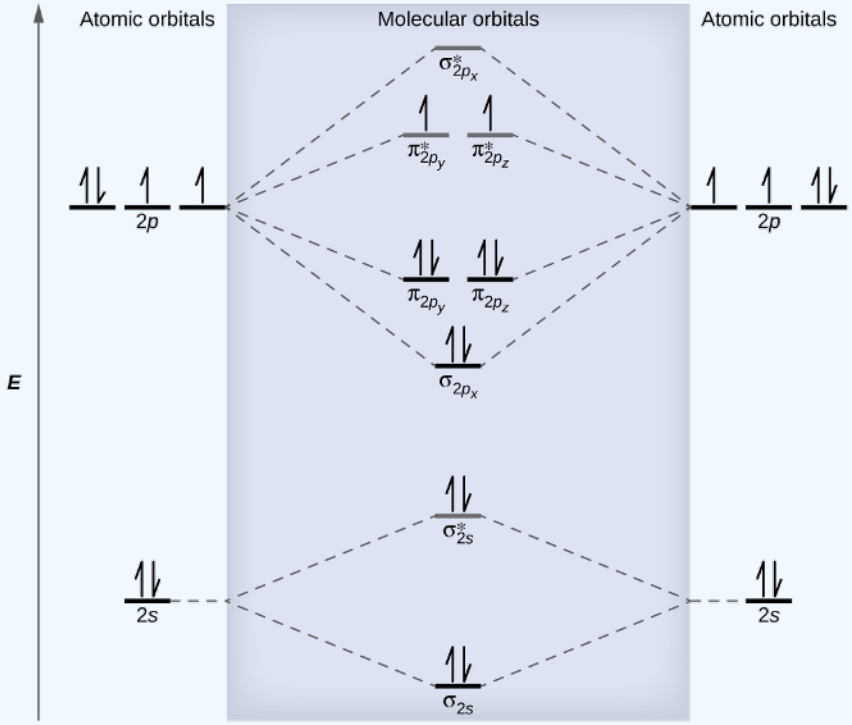


Figure 6.5.14: The molecular orbital energy diagram for O_2 predicts two unpaired electrons.

We calculate the bond order as

$$O_2 = \frac{(8-4)}{2} = 2$$

Oxygen's paramagnetism is explained by the presence of two unpaired electrons in the $(\pi_{2py}, \pi_{2pz})^*$ molecular orbitals.

Exercise 6.5.2

The main component of air is N_2. From the molecular orbital diagram of N_2, predict its bond order and whether it is diamagnetic or paramagnetic.

Answer

N_2 has a bond order of 3 and is diamagnetic.

Example 6.5.3: Ion Predictions with MO Diagrams

Give the molecular orbital configuration for the valence electrons in C_2^{2-}. Will this ion be stable?

Solution

Looking at the appropriate MO diagram, we see that the π orbitals are lower in energy than the σ_p orbital. The valence electron configuration for C_2 is $(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2py}, \pi_{2pz})^4$. Adding two more electrons to generate the C_2^{2-} anion will give a valence

electron configuration of $(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2py},\pi_{2pz})^4(\sigma_{2px})^2$. Since this has six more bonding electrons than antibonding, the bond order will be 3, and the ion should be stable.

Exercise 6.5.3

How many unpaired electrons would be present on a Be_2^{2-} ion? Would it be paramagnetic or diamagnetic?

Answer

two, paramagnetic

Link to Learning

Creating molecular orbital diagrams for molecules with more than two atoms relies on the same basic ideas as the diatomic examples presented here. However, with more atoms, computers are required to calculate how the atomic orbitals combine. See three-dimensional drawings of the molecular orbitals for C_6H_6.

This page titled 6.5: Molecular Orbital Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **8.4: Molecular Orbital Theory** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

6.6: Key Terms

antibonding orbital
molecular orbital located outside of the region between two nuclei; electrons in an antibonding orbital destabilize the molecule

bond order
number of pairs of electrons between two atoms; it can be found by the number of bonds in a Lewis structure or by the difference between the number of bonding and antibonding electrons divided by two

bonding orbital
molecular orbital located between two nuclei; electrons in a bonding orbital stabilize a molecule

degenerate orbitals
orbitals that have the same energy

diamagnetism
phenomenon in which a material is not magnetic itself but is repelled by a magnetic field; it occurs when there are only paired electrons present

homonuclear diatomic molecule
molecule consisting of two identical atoms

hybrid orbital
orbital created by combining atomic orbitals on a central atom

hybridization
model that describes the changes in the atomic orbitals of an atom when it forms a covalent compound

linear combination of atomic orbitals
technique for combining atomic orbitals to create molecular orbitals

molecular orbital
region of space in which an electron has a high probability of being found in a molecule

molecular orbital diagram
visual representation of the relative energy levels of molecular orbitals

molecular orbital theory
model that describes the behavior of electrons delocalized throughout a molecule in terms of the combination of atomic wave functions

node
plane separating different lobes of orbitals, where the probability of finding an electron is zero

overlap
coexistence of orbitals from two different atoms sharing the same region of space, leading to the formation of a covalent bond

paramagnetism
phenomenon in which a material is not magnetic itself but is attracted to a magnetic field; it occurs when there are unpaired electrons present

pi bond (π bond)
covalent bond formed by side-by-side overlap of atomic orbitals; the electron density is found on opposite sides of the internuclear axis

s-p mixing

change that causes σ_p orbitals to be less stable than π_p orbitals due to the mixing of *s* and *p*-based molecular orbitals of similar energies.

sigma bond (σ bond)

covalent bond formed by overlap of atomic orbitals along the internuclear axis

***sp* hybrid orbital**

one of a set of two orbitals with a linear arrangement that results from combining one *s* and one *p* orbital

***sp*^{2} hybrid orbital**

one of a set of three orbitals with a trigonal planar arrangement that results from combining one *s* and two *p* orbitals

***sp*^{3} hybrid orbital**

one of a set of four orbitals with a tetrahedral arrangement that results from combining one *s* and three *p* orbitals

***sp*^{3}*d* hybrid orbital**

one of a set of five orbitals with a trigonal bipyramidal arrangement that results from combining one *s*, three *p*, and one *d* orbital

***sp*^{3}*d*^{2} hybrid orbital**

one of a set of six orbitals with an octahedral arrangement that results from combining one *s*, three *p*, and two *d* orbitals

valence bond theory

description of bonding that involves atomic orbitals overlapping to form σ or π bonds, within which pairs of electrons are shared

π bonding orbital

molecular orbital formed by side-by-side overlap of atomic orbitals, in which the electron density is found on opposite sides of the internuclear axis

π* bonding orbital

antibonding molecular orbital formed by out of phase side-by-side overlap of atomic orbitals, in which the electron density is found on both sides of the internuclear axis, and there is a node between the nuclei

σ bonding orbital

molecular orbital in which the electron density is found along the axis of the bond

σ* bonding orbital

antibonding molecular orbital formed by out-of-phase overlap of atomic orbital along the axis of the bond, generating a node between the nuclei

---

This page titled 6.6: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

6.7: Key Equations

$$\text{bond order} = \frac{(\text{number of bonding electron}) - (\text{number of antibonding electrons})}{}$$

This page titled 6.7: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

6.8: Summary

Valence bond theory describes bonding as a consequence of the overlap of two separate atomic orbitals on different atoms that creates a region with one pair of electrons shared between the two atoms. When the orbitals overlap along an axis containing the nuclei, they form a σ bond. When they overlap in a fashion that creates a node along this axis, they form a π bond. Dipole moments can be used to determine partial separations of charges between atoms.

We can use hybrid orbitals, which are mathematical combinations of some or all of the valence atomic orbitals, to describe the electron density around covalently bonded atoms. These hybrid orbitals either form sigma (σ) bonds directed toward other atoms of the molecule or contain lone pairs of electrons. We can determine the type of hybridization around a central atom from the geometry of the regions of electron density about it. Two such regions imply *sp* hybridization; three, sp^2 hybridization; four, sp^3 hybridization; five, sp^3d hybridization; and six, sp^3d^2 hybridization. Pi (π) bonds are formed from unhybridized atomic orbitals (*p* or *d* orbitals).

Multiple bonds consist of a σ bond located along the axis between two atoms and one or two π bonds. The σ bonds are usually formed by the overlap of hybridized atomic orbitals, while the π bonds are formed by the side-by-side overlap of unhybridized orbitals. Resonance occurs when there are multiple unhybridized orbitals with the appropriate alignment to overlap, so the placement of π bonds can vary.

Molecular orbital (MO) theory describes the behavior of electrons in a molecule in terms of combinations of the atomic wave functions. The resulting molecular orbitals may extend over all the atoms in the molecule. Bonding molecular orbitals are formed by in-phase combinations of atomic wave functions, and electrons in these orbitals stabilize a molecule. Antibonding molecular orbitals result from out-of-phase combinations of atomic wave functions and electrons in these orbitals make a molecule less stable. Molecular orbitals located along an internuclear axis are called σ MOs. They can be formed from *s* orbitals or from *p* orbitals oriented in an end-to-end fashion. Molecular orbitals formed from *p* orbitals oriented in a side-by-side fashion have electron density on opposite sides of the internuclear axis and are called π orbitals.

We can describe the electronic structure of diatomic molecules by applying molecular orbital theory to the valence electrons of the atoms. Electrons fill molecular orbitals following the same rules that apply to filling atomic orbitals; Hund's rule and the Aufbau principle tell us that lower-energy orbitals will fill first, electrons will spread out before they pair up, and each orbital can hold a maximum of two electrons with opposite spins. Materials with unpaired electrons are paramagnetic and attracted to a magnetic field, while those with all-paired electrons are diamagnetic and repelled by a magnetic field. Correctly predicting the magnetic properties of molecules is in advantage of molecular orbital theory over Lewis structures and valence bond theory.

This page titled 6.8: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

6.9: Exercises

1. Explain how σ and π bonds are similar and how they are different.

2.

Use valence bond theory to explain the bonding in F_2, HF, and ClBr. Sketch the overlap of the atomic orbitals involved in the bonds.

3.

Use valence bond theory to explain the bonding in O_2. Sketch the overlap of the atomic orbitals involved in the bonds in O_2.

4.

How many σ and π bonds are present in the molecule HCN?

5.

A friend tells you N_2 has three π bonds due to overlap of the three *p*-orbitals on each N atom. Do you agree?

6.

Draw the Lewis structures for CO_2 and CO, and predict the number of σ and π bonds for each molecule.

(a) CO_2

(b) CO

7.

Why is the concept of hybridization required in valence bond theory?

8.

Give the shape that describes each hybrid orbital set:

(a) sp^2

(b) sp^3d

(c) sp

(d) sp^3d^2

9.

Explain why a carbon atom cannot form five bonds using sp^3d hybrid orbitals.

10.

What is the hybridization of the central atom in each of the following?

(a) BeH_2

(b) SF_6

(c) PO_4^{3-}

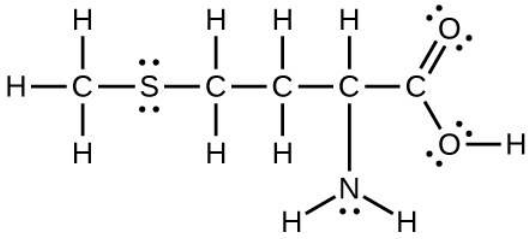
(d) PCl_5

11.

A molecule with the formula AB_3 could have one of four different shapes. Give the shape and the hybridization of the central A atom for each.

12.

Methionine, $CH_3SCH_2CH_2CH(NH_2)CO_2H$, is an amino acid found in proteins. The Lewis structure of this compound is shown below. What is the hybridization type of each carbon, oxygen, the nitrogen, and the sulfur?



13.

Sulfuric acid is manufactured by a series of reactions represented by the following equations:

$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$
$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$
$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(l)$

Draw a Lewis structure, predict the molecular geometry by VSEPR, and determine the hybridization of sulfur for the following:

(a) circular S_8 molecule

(b) SO_2 molecule

(c) SO_3 molecule

(d) H_2SO_4 molecule (the hydrogen atoms are bonded to oxygen atoms)

14.

Two important industrial chemicals, ethene, C_2H_4, and propene, C_3H_6, are produced by the steam (or thermal) cracking process:

$$2C_3H_8(g) \longrightarrow C_2H_4(g) + C_3H_6(g) + CH_4(g) + H_2(g)$$

For each of the four carbon compounds, do the following:

(a) Draw a Lewis structure.

(b) Predict the geometry about the carbon atom.

(c) Determine the hybridization of each type of carbon atom.

15.

Analysis of a compound indicates that it contains 77.55% Xe and 22.45% F by mass.

(a) What is the empirical formula for this compound? *(Assume this is also the molecular formula in responding to the remaining parts of this exercise).*

(b) Write a Lewis structure for the compound.

(c) Predict the shape of the molecules of the compound.

(d) What hybridization is consistent with the shape you predicted?

16.

Consider nitrous acid, HNO_2 (HONO).

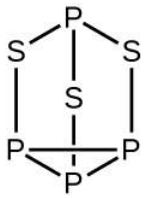
(a) Write a Lewis structure.

(b) What are the electron pair and molecular geometries of the internal oxygen and nitrogen atoms in the HNO_2 molecule?

(c) What is the hybridization on the internal oxygen and nitrogen atoms in HNO_2?

17.

Strike-anywhere matches contain a layer of $KClO_3$ and a layer of P_4S_3. The heat produced by the friction of striking the match causes these two compounds to react vigorously, which sets fire to the wooden stem of the match. $KClO_3$ contains the ClO_3^- ion. P_4S_3 is an unusual molecule with the skeletal structure.



(a) Write Lewis structures for P_4S_3 and the ClO_3^- ion.

(b) Describe the geometry about the P atoms, the S atom, and the Cl atom in these species.

(c) Assign a hybridization to the P atoms, the S atom, and the Cl atom in these species.

(d) Determine the oxidation states and formal charge of the atoms in P_4S_3 and the ClO_3^- ion.

18.

Identify the hybridization of each carbon atom in the following molecule. (The arrangement of atoms is given; you need to determine how many bonds connect each pair of atoms.)



19.

Write Lewis structures for NF_3 and PF_5. On the basis of hybrid orbitals, explain the fact that NF_3, PF_3, and PF_5 are stable molecules, but NF_5 does not exist.

20.

In addition to NF_3, two other fluoro derivatives of nitrogen are known: N_2F_4 and N_2F_2. What shapes do you predict for these two molecules? What is the hybridization for the nitrogen in each molecule?

21.

The bond energy of a C–C single bond averages 347 kJ mol^{-1}; that of a $C \equiv C$ triple bond averages 839 kJ mol^{-1}. Explain why the triple bond is not three times as strong as a single bond.

22.

For the carbonate ion, CO_3^{2-}, draw all of the resonance structures. Identify which orbitals overlap to create each bond.

23.

A useful solvent that will dissolve salts as well as organic compounds is the compound acetonitrile, H_3CCN. It is present in paint strippers.

(a) Write the Lewis structure for acetonitrile, and indicate the direction of the dipole moment in the molecule.

(b) Identify the hybrid orbitals used by the carbon atoms in the molecule to form σ bonds.

(c) Describe the atomic orbitals that form the π bonds in the molecule. Note that it is not necessary to hybridize the nitrogen atom.

24.

For the molecule allene, $H_2C = C = CH_2$, give the hybridization of each carbon atom. Will the hydrogen atoms be in the same plane or perpendicular planes?

25.

Identify the hybridization of the central atom in each of the following molecules and ions that contain multiple bonds:

(a) ClNO (N is the central atom)

(b) CS_2

(c) Cl_2CO (C is the central atom)

(d) Cl_2SO (S is the central atom)

(e) SO_2F_2 (S is the central atom)

(f) XeO_2F_2 (Xe is the central atom)

(g) $ClOF_2^+$ (Cl is the central atom)

26.

Describe the molecular geometry and hybridization of the N, P, or S atoms in each of the following compounds.

(a) H_3PO_4, phosphoric acid, used in cola soft drinks

(b) NH_4NO_3, ammonium nitrate, a fertilizer and explosive

(c) S_2Cl_2, disulfur dichloride, used in vulcanizing rubber

(d) $K_4[O_3POPO_3]$, potassium pyrophosphate, an ingredient in some toothpastes

27.

For each of the following molecules, indicate the hybridization requested and whether or not the electrons will be delocalized:

(a) ozone (O_3) central O hybridization

(b) carbon dioxide (CO_2) central C hybridization

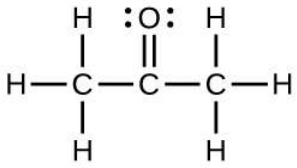
(c) nitrogen dioxide (NO_2) central N hybridization

(d) phosphate ion (PO_4^{3-}) central P hybridization

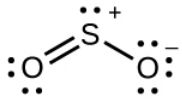
28.

For each of the following structures, determine the hybridization requested and whether the electrons will be delocalized:

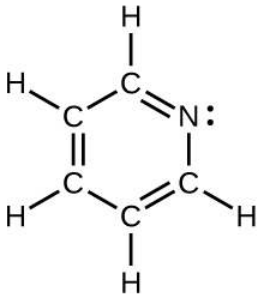
(a) Hybridization of each carbon



(b) Hybridization of sulfur



(c) All atoms



29.

Draw the orbital diagram for carbon in CO_2 showing how many carbon atom electrons are in each orbital.

30.

Sketch the distribution of electron density in the bonding and antibonding molecular orbitals formed from two *s* orbitals and from two *p* orbitals.

31.

How are the following similar, and how do they differ?

(a) σ molecular orbitals and π molecular orbitals

(b) ψ for an atomic orbital and ψ for a molecular orbital

(c) bonding orbitals and antibonding orbitals

32.

If molecular orbitals are created by combining five atomic orbitals from atom A and five atomic orbitals from atom B combine, how many molecular orbitals will result?

33.

Can a molecule with an odd number of electrons ever be diamagnetic? Explain why or why not.

34.

Can a molecule with an even number of electrons ever be paramagnetic? Explain why or why not.

35.

Why are bonding molecular orbitals lower in energy than the parent atomic orbitals?

36.

Calculate the bond order for an ion with this configuration:

$$(\sigma_{2s})^2(\sigma^*_{2s})^2(\sigma_{2px})^2(\pi_{2py},\ \pi_{2pz})^4(\pi^*_{2py},\ \pi^*_{2pz})^3$$

37.

Explain why an electron in the bonding molecular orbital in the H_2 molecule has a lower energy than an electron in the 1*s* atomic orbital of either of the separated hydrogen atoms.

38.

Predict the valence electron molecular orbital configurations for the following, and state whether they will be stable or unstable ions.

(a) $Na_2{}^{2}+$

(b) $Mg_2{}^{2}+$

(c) $Al_2{}^{2}+$

(d) $Si_2{}^{2}+$

(e) $P_2{}^{2}+$

(f) $S_2{}^{2}+$

(g) $F_2{}^{2}+$

(h) $Ar_2{}^{2}+$

39.

Determine the bond order of each member of the following groups, and determine which member of each group is predicted by the molecular orbital model to have the strongest bond.

(a) H_2, $H_2{}^{+}$, $H_2{}^{-}$

(b) O_2, $O_2{}^{2}+$, $O_2{}^{2}-$

(c) Li_2, $Be_2{}^{+}$, Be_2

(d) F_2, $F_2{}^{+}$, $F_2{}^{-}$

(e) N_2, $N_2{}^{+}$, $N_2{}^{-}$

40.

For the first ionization energy for an N_2 molecule, what molecular orbital is the electron removed from?

41.

Compare the atomic and molecular orbital diagrams to identify the member of each of the following pairs that has the highest first ionization energy (the most tightly bound electron) in the gas phase:

(a) H and H_2

(b) N and N_2

(c) O and O_2

(d) C and C_2

(e) B and B_2

42.

Which of the period 2 homonuclear diatomic molecules are predicted to be paramagnetic?

43.

A friend tells you that the 2*s* orbital for fluorine starts off at a much lower energy than the 2*s* orbital for lithium, so the resulting σ_{2s} molecular orbital in F_2 is more stable than in Li_2. Do you agree?

44.

True or false: Boron contains $2s^2 2p^1$ valence electrons, so only one *p* orbital is needed to form molecular orbitals.

45.

What charge would be needed on F_2 to generate an ion with a bond order of 2?

46.

Predict whether the MO diagram for S_2 would show s-p mixing or not.

47.

Explain why ${N_2}^{2+}$ is diamagnetic, while ${O_2}^{4+}$, which has the same number of valence electrons, is paramagnetic.

48.

Using the MO diagrams, predict the bond order for the stronger bond in each pair:

(a) B_2 or ${B_2}^+$

(b) F_2 or ${F_2}^+$

(c) O_2 or ${O_2}^{2+}$

(d) ${C_2}^+$ or ${C_2}^-$

This page titled 6.9: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

7: Thermochemistry

Useful forms of energy are also available from a variety of chemical reactions other than combustion. For example, the energy produced by the batteries in a cell phone, car, or flashlight results from chemical reactions. This chapter introduces many of the basic ideas necessary to explore the relationships between chemical changes and energy, with a focus on thermal energy.

7.1: Introduction
7.2: Energy Basics
7.3: Calorimetry
7.4: Enthalpy
7.5: Strengths of Ionic and Covalent Bonds
7.6: Key Terms
7.7: Key Equations
7.8: Summary
7.9: Exercises

This page titled 7: Thermochemistry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

7.1: Introduction

Chemical reactions, such as those that occur when you light a match, involve changes in energy as well as matter. Societies at all levels of development could not function without the energy released by chemical reactions. In 2012, about 85% of US energy consumption came from the combustion of petroleum products, coal, wood, and garbage. We use this energy to produce electricity (38%); to transport food, raw materials, manufactured goods, and people (27%); for industrial production (21%); and to heat and power our homes and businesses (10%).[1] While these combustion reactions help us meet our essential energy needs, they are also recognized by the majority of the scientific community as a major contributor to global climate change.



Figure 7.1.1: Sliding a match head along a rough surface initiates a combustion reaction that produces energy in the form of heat and light. (credit: modification of work by Laszlo Ilyes)

Useful forms of energy are also available from a variety of chemical reactions other than combustion. For example, the energy produced by the batteries in a cell phone, car, or flashlight results from chemical reactions. This chapter introduces many of the basic ideas necessary to explore the relationships between chemical changes and energy, with a focus on thermal energy.

The chemical bond is simply another form of energy, and its strength is indicated in exactly the same units as any other process involving energy. That is, in joules (or kilojoules) – or, when taken on a molar basis, in J/mol or kJ/mol. All of the principles involving energy we have examined up to now apply to bond energies exactly as they do to all other forms of energy: that is, energy may change form or be absorbed or released, but it cannot be created or destroyed. The energy of a chemical bond is indicated by the bond enthalpy, taken by convention to be the energy required to break a chemical bond. The same energy is released when a chemical bond is formed. Given that a chemical reaction involves breaking some bonds and making others, the enthalpy change for a chemical reaction can be estimated by analyzing the bonds broken and formed during the reaction. While this procedure gives an estimate of the overall enthalpy, it should be noted that a far more precise method involves using enthalpies of formation; the fact that there can be significant differences between, for example, an O–H bond in water and that in acetic acid accounts for the discrepancy in calculating ΔH using bond enthalpies versus the use of enthalpies of formation.

Footnotes

- 1US Energy Information Administration, *Primary Energy Consumption by Source and Sector, 2012*, http://www.eia.gov/totalenergy/data/...012_energy.pdf. Data derived from US Energy Information Administration, *Monthly Energy Review* (January 2014).

This page titled 7.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

7.2: Energy Basics

Learning Objectives

By the end of this section, you will be able to:

- Define energy, distinguish types of energy, and describe the nature of energy changes that accompany chemical and physical changes
- Distinguish the related properties of heat, thermal energy, and temperature
- Define and distinguish specific heat and heat capacity, and describe the physical implications of both
- Perform calculations involving heat, specific heat, and temperature change

Chemical changes and their accompanying changes in energy are important parts of our everyday world (Figure 7.2.1). The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating, and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.



(a) (b) (c)

Figure 7.2.1: The energy involved in chemical changes is important to our daily lives: (a) A cheeseburger for lunch provides the energy you need to get through the rest of the day; (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school; and (c) coke, a processed form of coal, provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily. (credit a: modification of work by "Pink Sherbet Photography"/Flickr; credit b: modification of work by Jeffery Turner)

Over 90% of the energy we use comes originally from the sun. Every day, the sun provides the earth with almost 10,000 times the amount of energy necessary to meet all of the world's energy needs for that day. Our challenge is to find ways to convert and store incoming solar energy so that it can be used in reactions or chemical processes that are both convenient and nonpolluting. Plants and many bacteria capture solar energy through photosynthesis. We release the energy stored in plants when we burn wood or plant products such as ethanol. We also use this energy to fuel our bodies by eating food that comes directly from plants or from animals that got their energy by eating plants. Burning coal and petroleum also releases stored solar energy: These fuels are fossilized plant and animal matter.

This chapter will introduce the basic ideas of an important area of science concerned with the amount of heat absorbed or released during chemical and physical changes—an area called thermochemistry. The concepts introduced in this chapter are widely used in almost all scientific and technical fields. Food scientists use them to determine the energy content of foods. Biologists study the energetics of living organisms, such as the metabolic combustion of sugar into carbon dioxide and water. The oil, gas, and transportation industries, renewable energy providers, and many others endeavor to find better methods to produce energy for our commercial and personal needs. Engineers strive to improve energy efficiency, find better ways to heat and cool our homes, refrigerate our food and drinks, and meet the energy and cooling needs of computers and electronics, among other applications. Understanding thermochemical principles is essential for chemists, physicists, biologists, geologists, every type of engineer, and just about anyone who studies or does any kind of science.

Energy

Energy can be defined as the capacity to supply heat or do work. One type of work (w) is the process of causing matter to move against an opposing force. For example, we do work when we inflate a bicycle tire—we move matter (the air in the pump) against the opposing force of the air already in the tire.

Like matter, energy comes in different types. One scheme classifies energy into two types: potential energy, the energy an object has because of its relative position, composition, or condition, and kinetic energy, the energy that an object possesses because of its motion. Water at the top of a waterfall or dam has potential energy because of its position; when it flows downward through generators, it has kinetic energy that can be used to do work and produce electricity in a hydroelectric plant (Figure 7.2.2). A battery has potential energy because the chemicals within it can produce electricity that can do work.



(a)



(b)

Figure 7.2.2: (a) Water at a higher elevation, for example, at the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As the water falls, some of its potential energy is converted into kinetic energy. (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy. (credit a: modification of work by Steve Jurvetson; credit b: modification of work by "curimedia"/Wikimedia commons)

Energy can be converted from one form into another, but all of the energy present before a change occurs always exists in some form after the change is completed. This observation is expressed in the law of conservation of energy: during a chemical or physical change, energy can be neither created nor destroyed, although it can be changed in form. (This is also one version of the first law of thermodynamics, as you will learn later.)

When one substance is converted into another, there is always an associated conversion of one form of energy into another. Heat is usually released or absorbed, but sometimes the conversion involves light, electrical energy, or some other form of energy. For example, chemical energy (a type of potential energy) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car's engine, the rapidly expanding gaseous products of this chemical reaction generate mechanical energy (a type of kinetic energy) when they move the cylinders' pistons.

According to the law of conservation of matter (seen in an earlier chapter), there is no detectable change in the total amount of matter during a chemical change. When chemical reactions occur, the energy changes are relatively modest and the mass changes are too small to measure, so the laws of conservation of matter and energy hold well. However, in nuclear reactions, the energy changes are much larger (by factors of a million or so), the mass changes are measurable, and matter-energy conversions are significant. This will be examined in more detail in a later chapter on nuclear chemistry.

Thermal Energy, Temperature, and Heat

Thermal energy is kinetic energy associated with the random motion of atoms and molecules. Temperature is a quantitative measure of "hot" or "cold." When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE), and we say that the object is "hot." When the atoms and molecules are moving slowly, they have lower average KE, and we say that the object is "cold" (Figure 7.2.3). Assuming that no chemical reaction or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. And, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

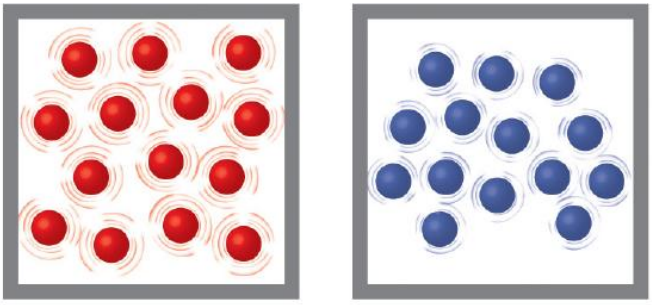


Figure 7.2.3: (a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.

Link to Learning

Click on this interactive simulation to view the effects of temperature on molecular motion.

Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure 7.2.4. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

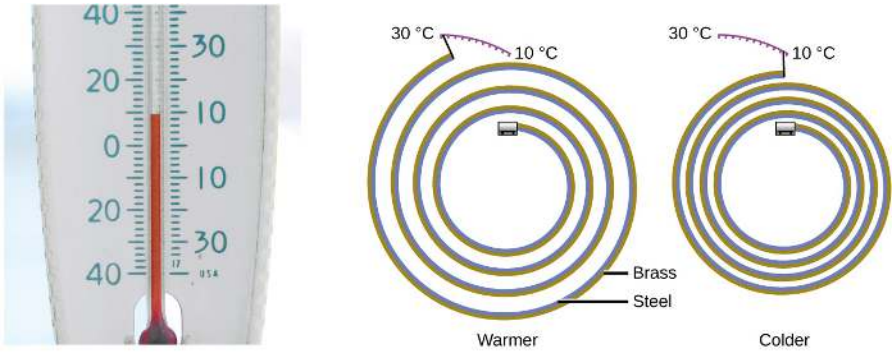


Figure 7.2.4: (a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature. (credit a: modification of work by “dwstucke”/Flickr)

Link to Learning

The following demonstration allows one to view the effects of heating and cooling a coiled bimetallic strip.

Heat (q) is the transfer of thermal energy between two bodies at different temperatures. Heat flow (a redundant term, but one commonly used) increases the thermal energy of one body and decreases the thermal energy of the other. Suppose we initially have a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature (Figure 7.2.5).

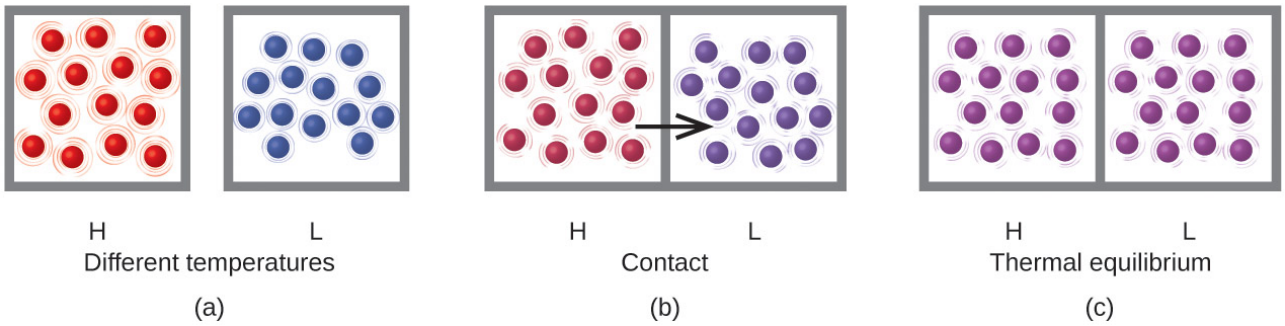


Figure 7.2.5: (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they contact each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach "thermal equilibrium" when both substances are at the same temperature and their molecules have the same average kinetic energy.

Link to Learning

Click on the PhET simulation to explore energy forms and changes. Visit the Energy Systems tab to create combinations of energy sources, transformation methods, and outputs. Click on Energy Symbols to visualize the transfer of energy.

Matter undergoing chemical reactions and physical changes can release or absorb heat. A change that releases heat is called an exothermic process. For example, the combustion reaction that occurs when using an oxyacetylene torch is an exothermic process—this process also releases energy in the form of light as evidenced by the torch's flame (Figure 7.2.6). A reaction or change that absorbs heat is an endothermic process. A cold pack used to treat muscle strains provides an example of an endothermic process. When the substances in the cold pack (water and a salt like ammonium nitrate) are brought together, the resulting process absorbs heat, leading to the sensation of cold.



Figure 7.2.6: (a) An oxyacetylene torch produces heat by the combustion of acetylene in oxygen. The energy released by this exothermic reaction heats and then melts the metal being cut. The sparks are tiny bits of the molten metal flying away. (b) A cold pack uses an endothermic process to create the sensation of cold. (credit a: modification of work by "Skatebiker"/Wikimedia commons)

Historically, energy was measured in units of calories (cal). A calorie is the amount of energy required to raise one gram of water by 1 degree C (1 kelvin). However, this quantity depends on the atmospheric pressure and the starting temperature of the water. The ease of measurement of energy changes in calories has meant that the calorie is still frequently used. The Calorie (with a capital C), or large calorie, commonly used in quantifying food energy content, is a kilocalorie. The SI unit of heat, work, and energy is the joule. A joule (J) is defined as the amount of energy used when a force of 1 newton moves an object 1 meter. It is named in honor of the English physicist James Prescott Joule. One joule is equivalent to 1 kg m^2/s^2, which is also called 1 newton–meter. A kilojoule (kJ) is 1000 joules. To standardize its definition, 1 calorie has been set to equal 4.184 joules.

We now introduce two concepts useful in describing heat flow and temperature change. The heat capacity (C) of a body of matter is the quantity of heat (q) it absorbs or releases when it experiences a temperature change (ΔT) of 1 degree Celsius (or equivalently, 1 kelvin):

$$C = \frac{q}{\Delta T}$$

Heat capacity is determined by both the type and amount of substance that absorbs or releases heat. It is therefore an extensive property—its value is proportional to the amount of the substance.

For example, consider the heat capacities of two cast iron frying pans. The heat capacity of the large pan is five times greater than that of the small pan because, although both are made of the same material, the mass of the large pan is five times greater than the mass of the small pan. More mass means more atoms are present in the larger pan, so it takes more energy to make all of those atoms vibrate faster. The heat capacity of the small cast iron frying pan is found by observing that it takes 18,150 J of energy to raise the temperature of the pan by 50.0 °C:

$$C_{\text{small pan}} = \frac{18,140J}{50.0^\circ C} = 363J/^\circ C$$

The larger cast iron frying pan, while made of the same substance, requires 90,700 J of energy to raise its temperature by 50.0 °C. The larger pan has a (proportionally) larger heat capacity because the larger amount of material requires a (proportionally) larger amount of energy to yield the same temperature change:

$$C_{\text{large pan}} = \frac{90,700J}{50.0^\circ C} = 1814J/^\circ C$$

The specific heat capacity (*c*) of a substance, commonly called its "specific heat," is the quantity of heat required to raise the temperature of 1 gram of a substance by 1 degree Celsius (or 1 kelvin):

$$c = \frac{q}{m\Delta T}$$

Specific heat capacity depends only on the kind of substance absorbing or releasing heat. It is an intensive property—the type, but not the amount, of the substance is all that matters. For example, the small cast iron frying pan has a mass of 808 g. The specific heat of iron (the material used to make the pan) is therefore:

$$c_{\text{iron}} = \frac{18,140J}{(808g)\,(50.0^\circ C)} = 0.449J/g^\circ C$$

The large frying pan has a mass of 4040 g. Using the data for this pan, we can also calculate the specific heat of iron:

$$c_{\text{iron}} = \frac{90,700J}{(4040g)\,(50.0^\circ C)} = 0.449J/g^\circ C$$

Although the large pan is more massive than the small pan, since both are made of the same material, they both yield the same value for specific heat (for the material of construction, iron). Note that specific heat is measured in units of energy per temperature per mass and is an intensive property, being derived from a ratio of two extensive properties (heat and mass). The molar heat capacity, also an intensive property, is the heat capacity per mole of a particular substance and has units of J/mol °C (Figure 7.2.7).



Figure 7.2.7: Because of its larger mass, a large frying pan has a larger heat capacity than a small frying pan. Because they are made of the same material, both frying pans have the same specific heat. (credit: Mark Blaser)

Water has a relatively high specific heat (about 4.2 J/g °C for the liquid and 2.09 J/g °C for the solid); most metals have much lower specific heats (usually less than 1 J/g °C). The specific heat of a substance varies somewhat with temperature. However, this variation is usually small enough that we will treat specific heat as constant over the range of temperatures that will be considered in this chapter. Specific heats of some common substances are listed in Table 7.2.1.

Table 7.2.1: Specific Heats of Common Substances at 25 °C and 1 bar

Substance	Symbol (*state*)	Specific Heat (J/g °C)
helium	He(g)	5.193
water	$H_2O(l)$	4.184
ethanol	$C_2H_6O(l)$	2.376
ice	$H_2O(s)$	2.093 (at −10 °C)
water vapor	$H_2O(g)$	1.864
nitrogen	$N_2(g)$	1.040
air		1.007
oxygen	$O_2(g)$	0.918
aluminum	Al(s)	0.897
carbon dioxide	$CO_2(g)$	0.853
argon	Ar(g)	0.522
iron	Fe(s)	0.449
copper	Cu(s)	0.385
lead	Pb(s)	0.130
gold	Au(s)	0.129
silicon	Si(s)	0.712

If we know the mass of a substance and its specific heat, we can determine the amount of heat, q, entering or leaving the substance by measuring the temperature change before and after the heat is gained or lost:

$$\begin{aligned} q &= (\text{ specific heat }) \times (\text{ mass of substance }) \times (\text{ temperature change }) \\ &= c \times m \times \Delta T \\ &= c \times m \times (T_{\text{final}} - T_{\text{initial}}) \end{aligned} \tag{7.2.1}$$

In this equation, c is the specific heat of the substance, m is its mass, and ΔT (which is read “delta T”) is the temperature change, $T_{\text{final}} - T_{\text{initial}}$. If a substance gains thermal energy, its temperature increases, its final temperature is higher than its initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a positive value, and the value of q is positive. If a substance loses thermal energy, its temperature decreases, the final temperature is lower than the initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a negative value, and the value of q is negative.

Example 7.2.1: Measuring Heat

A flask containing 8.0 10^2 g of water is heated, and the temperature of the water increases from 21 °C to 85 °C. How much heat did the water absorb?

Solution

To answer this question, consider these factors:

- the specific heat of the substance being heated (in this case, water)
- the amount of substance being heated (in this case, 8.0 × 10^2 g)
- the magnitude of the temperature change (in this case, from 21 °C to 85 °C).

The specific heat of water is 4.184 J/g °C, so to heat 1 g of water by 1 °C requires 4.184 J. We note that since 4.184 J is required to heat 1 g of water by 1 °C, we will need *800 times as much* to heat 8.0 × 10^2 g of water by 1 °C. Finally, we observe that since 4.184 J are required to heat 1 g of water by 1 °C, we will need *64 times as much* to heat it by 64 °C (that is, from 21 °C to 85 °C).

This can be summarized using Equation 7.2.1:

$$\begin{aligned} q &= c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}}) \\ &= \left(4.184 J/\frac{g}{\circ}{}^{\circ}C\right) \times \left(8.0 \times 10^2 g\right) \times (85 - 21)^{\circ}C \\ &= (4.184 J/g^{\circ}C) \times \left(8.0 \times 10^2 g\right) \times (64)^{\circ}C \\ &= 210,000 J \left(= 2.1 \times 10^2 kJ\right) \end{aligned}$$

Because the temperature increased, the water absorbed heat and q is positive.

Exercise 7.2.1

How much heat, in joules, must be added to a 502 g iron skillet to increase its temperature from 25 °C to 250 °C? The specific heat of iron is 0.449 J/g °C.

Answer

5.07 10^4 J

Note that the relationship between heat, specific heat, mass, and temperature change can be used to determine any of these quantities (not just heat) if the other three are known or can be deduced.

Example 7.2.2: Determining Other Quantities

A piece of unknown metal weighs 348 g. When the metal piece absorbs 6.64 kJ of heat, its temperature increases from 22.4 °C to 43.6 °C. Determine the specific heat of this metal (which might provide a clue to its identity).

Solution

Since mass, heat, and temperature change are known for this metal, we can determine its specific heat using the relationship in Equation 7.2.1:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

Substituting the known values:

$$6640 J = c \times (348 g) \times (43.6 - 22.4)^{\circ}C$$

Solving:

$$c = \frac{6640 J}{(348 g) \times (21.2^{\circ}C)} = 0.900 J/g^{\circ}C$$

Comparing this value with the values in Table 7.2.1, this value matches the specific heat of aluminum, which suggests that the unknown metal may be aluminum.

Exercise 7.2.2

A piece of unknown metal weighs 217 g. When the metal piece absorbs 1.43 kJ of heat, its temperature increases from 24.5 °C to 39.1 °C. Determine the specific heat of this metal, and predict its identity.

Answer

c = 0.451 J/g °C; the metal is likely to be iron

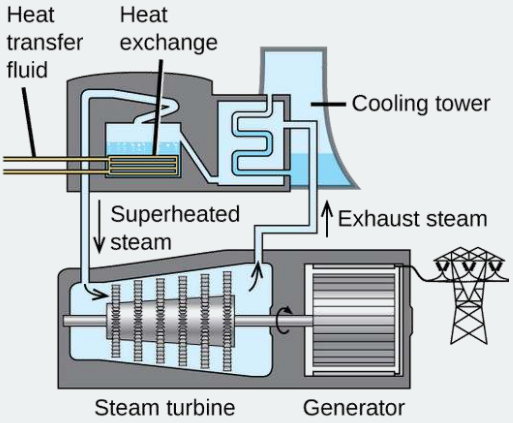
Chemistry in Everyday Life: Solar Thermal Energy Power Plants

The sunlight that reaches the earth contains thousands of times more energy than we presently capture. Solar thermal systems provide one possible solution to the problem of converting energy from the sun into energy we can use. Large-scale solar thermal plants have different design specifics, but all concentrate sunlight to heat some substance; the heat “stored” in that substance is then converted into electricity.

The Solana Generating Station in Arizona’s Sonora Desert produces 280 megawatts of electrical power. It uses parabolic mirrors that focus sunlight on pipes filled with a heat transfer fluid (HTF) (Figure 7.2.8). The HTF then does two things: It turns water into steam, which spins turbines, which in turn produces electricity, and it melts and heats a mixture of salts, which functions as a thermal energy storage system. After the sun goes down, the molten salt mixture can then release enough of its stored heat to produce steam to run the turbines for 6 hours. Molten salts are used because they possess a number of beneficial properties, including high heat capacities and thermal conductivities.



(a)



(b)

Figure 7.2.8: This solar thermal plant uses parabolic trough mirrors to concentrate sunlight. (credit a: modification of work by Bureau of Land Management)

The 377-megawatt Ivanpah Solar Generating System, located in the Mojave Desert in California, is the largest solar thermal power plant in the world (Figure 7.2.9). Its 170,000 mirrors focus huge amounts of sunlight on three water-filled towers, producing steam at over 538 °C that drives electricity-producing turbines. It produces enough energy to power 140,000 homes. Water is used as the working fluid because of its large heat capacity and heat of vaporization.



(a)



(b)

Figure 7.2.9: (a) The Ivanpah solar thermal plant uses 170,000 mirrors to concentrate sunlight on water-filled towers. (b) It covers 4000 acres of public land near the Mojave Desert and the California-Nevada border. (credit a: modification of work by Craig Dietrich; credit b: modification of work by “USFWS Pacific Southwest Region”/Flickr)

This page titled 7.2: Energy Basics is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **5.1: Energy Basics** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

7.3: Calorimetry

Learning Objectives

By the end of this section, you will be able to:

- Explain the technique of calorimetry
- Calculate and interpret heat and related properties using typical calorimetry data

One technique we can use to measure the amount of heat involved in a chemical or physical process is known as calorimetry. Calorimetry is used to measure amounts of heat transferred to or from a substance. To do so, the heat is exchanged with a calibrated object (calorimeter). The temperature change measured by the calorimeter is used to derive the amount of heat transferred by the process under study. The measurement of heat transfer using this approach requires the definition of a system (the substance or substances undergoing the chemical or physical change) and its surroundings (all other matter, including components of the measurement apparatus, that serve to either provide heat to the system or absorb heat from the system).

A calorimeter is a device used to measure the amount of heat involved in a chemical or physical process. For example, when an exothermic reaction occurs in solution in a calorimeter, the heat produced by the reaction is absorbed by the solution, which increases its temperature. When an endothermic reaction occurs, the heat required is absorbed from the thermal energy of the solution, which decreases its temperature (Figure 7.3.1). The temperature change, along with the specific heat and mass of the solution, can then be used to calculate the amount of heat involved in either case.

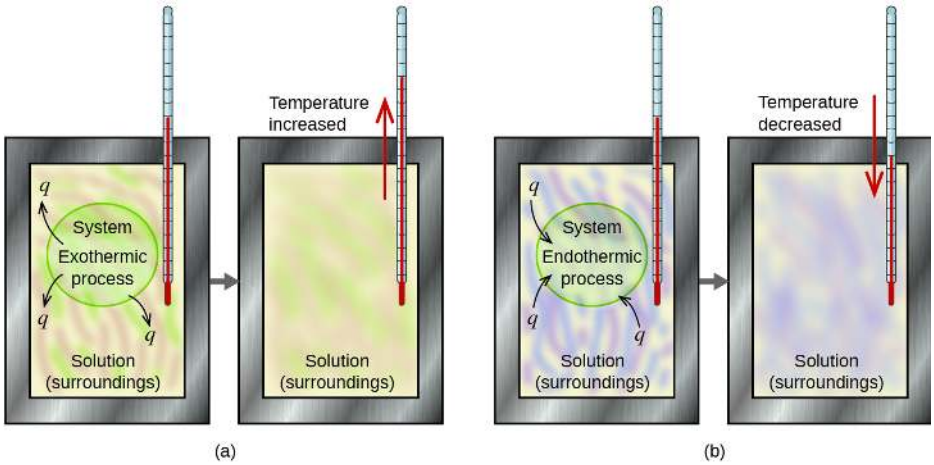


Figure 7.3.1: In a calorimetric determination, either (a) an exothermic process occurs and heat, q, is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q, is positive, indicating that thermal energy is transferred from the surroundings to the system.

Calorimetry measurements are important in understanding the heat transferred in reactions involving everything from microscopic proteins to massive machines. During her time at the National Bureau of Standards, research chemist Reatha Clark King performed calorimetric experiments to understand the precise heats of various flourine compounds. Her work was important to NASA in their quest for better rocket fuels.

Scientists use well-insulated calorimeters that all but prevent the transfer of heat between the calorimeter and its environment, which effectively limits the "surroundings" to the nonsystem components with the calorimeter (and the calorimeter itself). This enables the accurate determination of the heat involved in chemical processes, the energy content of foods, and so on. General chemistry students often use simple calorimeters constructed from polystyrene cups (Figure 7.3.2). These easy-to-use "coffee cup" calorimeters allow more heat exchange with the outside environment, and therefore produce less accurate energy values.

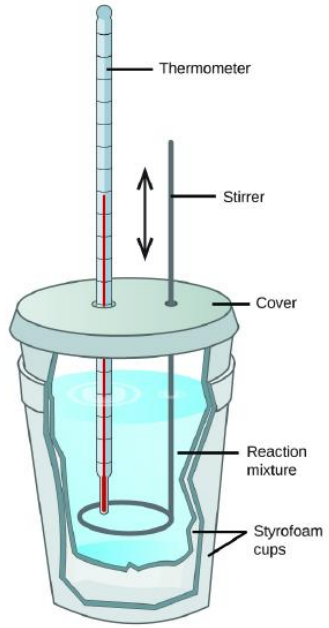


Figure 7.3.2: A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.

Commercial solution calorimeters are also available. Relatively inexpensive calorimeters often consist of two thin-walled cups that are nested in a way that minimizes thermal contact during use, along with an insulated cover, handheld stirrer, and simple thermometer. More expensive calorimeters used for industry and research typically have a well-insulated, fully enclosed reaction vessel, motorized stirring mechanism, and a more accurate temperature sensor (Figure 7.3.3).

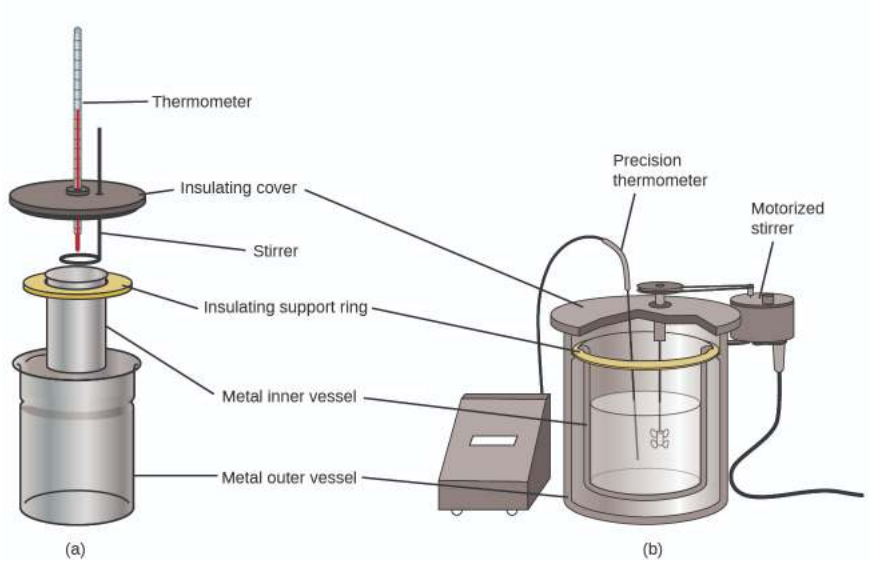


Figure 7.3.3: Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.

Before discussing the calorimetry of chemical reactions, consider a simpler example that illustrates the core idea behind calorimetry. Suppose we initially have a high-temperature substance, such as a hot piece of metal (M), and a low-temperature substance, such as cool water (W). If we place the metal in the water, heat will flow from M to W. The temperature of M will decrease, and the temperature of W will increase, until the two substances have the same temperature—that is, when they reach thermal equilibrium (Figure 7.3.4). If this occurs in a calorimeter, ideally all of this heat transfer occurs between the two substances, with no heat gained or lost by either its external environment. Under these ideal circumstances, the net heat change is zero:

$$q_{\text{substance}}\, M + q_{\text{substance}}\, W = 0$$

This relationship can be rearranged to show that the heat gained by substance M is equal to the heat lost by substance W:

$$q_{\text{substance M}} = -q_{\text{substance W}}$$

The magnitude of the heat (change) is therefore the same for both substances, and the negative sign merely shows that $q_{\text{substance M}}$ and $q_{\text{substance W}}$ are opposite in direction of heat flow (gain or loss) but does not indicate the arithmetic sign of either q value (that is determined by whether the matter in question gains or loses heat, per definition). In the specific situation described, $q_{\text{substance M}}$ is a negative value and $q_{\text{substance W}}$ is positive, since heat is transferred from M to W.

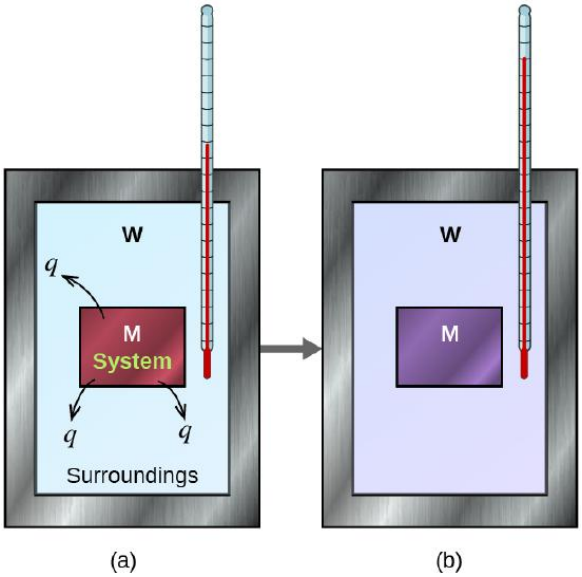


Figure 7.3.4: In a simple calorimetry process, (a) heat, q, is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.

Example 7.3.1: Heat Transfer between Substances at Different Temperatures

A 360.0-g piece of rebar (a steel rod used for reinforcing concrete) is dropped into 425 mL of water at 24.0 °C. The final temperature of the water was measured as 42.7 °C. Calculate the initial temperature of the piece of rebar. Assume the specific heat of steel is approximately the same as that for iron (Table 5.1), and that all heat transfer occurs between the rebar and the water (there is no heat exchange with the surroundings).

Solution

The temperature of the water increases from 24.0 °C to 42.7 °C, so the water absorbs heat. That heat came from the piece of rebar, which initially was at a higher temperature. Assuming that all heat transfer was between the rebar and the water, with no heat "lost" to the outside environment, then *heat given off by rebar = −heat taken in by water*, or:

$$q_{\text{rebar}} = -q_{\text{water}}$$

Since we know how heat is related to other measurable quantities, we have:

$$(c \times m \times \Delta T)_{\text{rebar}} = -(c \times m \times \Delta T)_{\text{water}}$$

Letting f = final and i = initial, in expanded form, this becomes:

$$c_{\text{rebar}} \times m_{\text{rebar}} \times (T_{f,\text{ rebar}} - T_{i,\text{ rebar}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{f,\text{water}} - T_{i,\text{ water}})$$

The density of water is 1.0 g/mL, so 425 mL of water = 425 g. Noting that the final temperature of both the rebar and water is 42.7 °C, substituting known values yields:

$$(0.449 J/g^{\circ}C)\,(360.0 g)\,(42.7^{\circ}C - T_{i,\text{ rebar}}) = -(4.184 J/g^{\circ}C)\,(425 g)\,(42.7^{\circ}C - 24.0^{\circ}C)$$

$$T_{i,\text{ rebar}} = \frac{(4.184 J/g^{\circ}C)\,(425 g)\,(42.7^{\circ}C - 24.0^{\circ}C)}{(0.449 J/g^{\circ}C)\,(360.0 g)} + 42.7^{\circ}C$$

Solving this gives $T_{i,\text{ rebar}}$ = 248 °C, so the initial temperature of the rebar was 248 °C.

Exercise 7.3.1A

A 248-g piece of copper is dropped into 390 mL of water at 22.6 °C. The final temperature of the water was measured as 39.9 °C. Calculate the initial temperature of the piece of copper. Assume that all heat transfer occurs between the copper and the water.

Answer

The initial temperature of the copper was 335.6 °C.

Exercise 7.3.1B

A 248-g piece of copper initially at 314 °C is dropped into 390 mL of water initially at 22.6 °C. Assuming that all heat transfer occurs between the copper and the water, calculate the final temperature.

Answer

The final temperature (reached by both copper and water) is 38.7 °C.

This method can also be used to determine other quantities, such as the specific heat of an unknown metal.

Example 7.3.2: Identifying a Metal by Measuring Specific Heat

A 59.7 g piece of metal that had been submerged in boiling water was quickly transferred into 60.0 mL of water initially at 22.0 °C. The final temperature is 28.5 °C. Use these data to determine the specific heat of the metal. Use this result to identify the metal.

Solution

Assuming perfect heat transfer, *heat given off by metal* = −*heat taken in by water*, or:

$$q_{\text{metal}} = -q_{\text{water}}$$

In expanded form, this is:

$$c_{\text{metal}} \times m_{\text{metal}} \times (T_{f,\text{ metal}} - T_{i,\text{ metal}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{f,\text{water}} - T_{i,\text{ water}})$$

Noting that since the metal was submerged in boiling water, its initial temperature was 100.0 °C; and that for water, 60.0 mL = 60.0 g; we have:

$$(c_{\text{metal}})\,(59.7g)\,(28.5^\circ C - 100.0^\circ C) = -(4.18 J/g^\circ C)\,(60.0g)\,(28.5^\circ C - 22.0^\circ C)$$

Solving this:

$$c_{\text{metal}} = \frac{-(4.184 J/g^\circ C)\,(60.0g)\,(6.5^\circ C)}{(59.7g)\,(-71.5^\circ C)} = 0.38 J/g^\circ C$$

Comparing this with values in Table 5.1, our experimental specific heat is closest to the value for copper (0.39 J/g °C), so we identify the metal as copper.

Exercise 7.3.2

A 92.9-g piece of a silver/gray metal is heated to 178.0 °C, and then quickly transferred into 75.0 mL of water initially at 24.0 °C. After 5 minutes, both the metal and the water have reached the same temperature: 29.7 °C. Determine the specific heat and the identity of the metal. (Note: You should find that the specific heat is close to that of two different metals. Explain how you can confidently determine the identity of the metal).

Answer

c_{metal}= 0.13 J/g °C

This specific heat is close to that of either gold or lead. It would be difficult to determine which metal this was based solely on the numerical values. However, the observation that the metal is silver/gray in addition to the value for the specific heat indicates that the metal is lead.

When we use calorimetry to determine the heat involved in a chemical reaction, the same principles we have been discussing apply. The amount of heat absorbed by the calorimeter is often small enough that we can neglect it (though not for highly accurate measurements, as discussed later), and the calorimeter minimizes energy exchange with the outside environment. Because energy is neither created nor destroyed during a chemical reaction, the heat produced or consumed in the reaction (the "system"), \(q_{\text{reaction}}\), plus the heat absorbed or lost by the solution (the "surroundings"), $q_{solution}$, must add up to zero:

$$q_{reaction} + q_{solution} = 0$$

This means that the amount of heat produced or consumed in the reaction equals the amount of heat absorbed or lost by the solution:

$$q_{reaction} = -q_{solution}$$

This concept lies at the heart of all calorimetry problems and calculations.

Example 7.3.3: Heat Produced by an Exothermic Reaction

When 50.0 mL of 1.00 M HCl(*aq*) and 50.0 mL of 1.00 M NaOH(*aq*), both at 22.0 °C, are added to a coffee cup calorimeter, the temperature of the mixture reaches a maximum of 28.9 °C. What is the approximate amount of heat produced by this reaction?

$$\mathrm{HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)}$$

Solution

To visualize what is going on, imagine that you could combine the two solutions so quickly that no reaction took place while they mixed; then after mixing, the reaction took place. At the instant of mixing, you have 100.0 mL of a mixture of HCl and NaOH at 22.0 °C. The HCl and NaOH then react until the solution temperature reaches 28.9 °C.

The heat given off by the reaction is equal to that taken in by the solution. Therefore:

$$q_{reaction} = -q_{solution}$$

(It is important to remember that this relationship only holds if the calorimeter does not absorb any heat from the reaction, and there is no heat exchange between the calorimeter and the outside environment.)

Next, we know that the heat absorbed by the solution depends on its specific heat, mass, and temperature change:

$$q_{solution} = (c \times m \times \Delta T)_{solution}$$

To proceed with this calculation, we need to make a few more reasonable assumptions or approximations. Since the solution is aqueous, we can proceed as if it were water in terms of its specific heat and mass values. The density of water is approximately 1.0 g/mL, so 100.0 mL has a mass of about 1.0 10^2 g (two significant figures). The specific heat of water is approximately 4.184 J/g °C, so we use that for the specific heat of the solution. Substituting these values gives:

$$q_{solution} = (4.184 J/g^{\circ}C)\left(1.0 \times 10^2 g\right)(28.9^{\circ}C - 22.0^{\circ}C) = 2.9 \times 10^3 J$$

Finally, since we are trying to find the heat of the reaction, we have:

$$q_{reaction} = -q_{solution} = -2.9 \times 10^3 J$$

The negative sign indicates that the reaction is exothermic. It produces 2.9 kJ of heat.

Exercise 7.3.3

When 100 mL of 0.200 M NaCl(*aq*) and 100 mL of 0.200 M $AgNO_3$(*aq*), both at 21.9 °C, are mixed in a coffee cup calorimeter, the temperature increases to 23.5 °C as solid AgCl forms. How much heat is produced by this precipitation reaction? What assumptions did you make to determine your value?

Answer

1.34 1.3 kJ; assume no heat is absorbed by the calorimeter, no heat is exchanged between the calorimeter and its surroundings, and that the specific heat and mass of the solution are the same as those for water

Chemistry in Everyday Life: Thermochemistry of Hand Warmers

When working or playing outdoors on a cold day, you might use a hand warmer to warm your hands (Figure 7.3.5). A common reusable hand warmer contains a supersaturated solution of $NaC_2H_3O_2$ (sodium acetate) and a metal disc. Bending the disk creates nucleation sites around which the metastable $NaC_2H_3O_2$ quickly crystallizes (a later chapter on solutions will investigate saturation and supersaturation in more detail).

The process

$$NaC_2H_3O_2(aq) \longrightarrow NaC_2H_3O_2(s)$$

is exothermic, and the heat produced by this process is absorbed by your hands, thereby warming them (at least for a while). If the hand warmer is reheated, the $NaC_2H_3O_2$ redissolves and can be reused.



Figure 7.3.5: Chemical hand warmers produce heat that warms your hand on a cold day. In this one, you can see the metal disc that initiates the exothermic precipitation reaction. (credit: modification of work by Science Buddies TV/YouTube)

Another common hand warmer produces heat when it is ripped open, exposing iron and water in the hand warmer to oxygen in the air. One simplified version of this exothermic reaction is

$$2\,Fe(s) + \tfrac{3}{2}O_2(g) \longrightarrow Fe_2O_3(s)$$

Salt in the hand warmer catalyzes the reaction, so it produces heat more rapidly; cellulose, vermiculite, and activated carbon help distribute the heat evenly. Other types of hand warmers use lighter fluid (a platinum catalyst helps lighter fluid oxidize exothermically), charcoal (charcoal oxidizes in a special case), or electrical units that produce heat by passing an electrical current from a battery through resistive wires.

Link to Learning

This link shows the precipitation reaction that occurs when the disk in a chemical hand warmer is flexed.

Example 7.3.4: Heat Flow in an Instant Ice Pack

When solid ammonium nitrate dissolves in water, the solution becomes cold. This is the basis for an “instant ice pack” (Figure 7.3.6). When 3.21 g of solid NH_4NO_3 dissolves in 50.0 g of water at 24.9 °C in a calorimeter, the temperature decreases to 20.3 °C.

Calculate the value of q for this reaction and explain the meaning of its arithmetic sign. State any assumptions that you made.



Figure 7.3.6: An instant cold pack consists of a bag containing solid ammonium nitrate and a second bag of water. When the bag of water is broken, the pack becomes cold because the dissolution of ammonium nitrate is an endothermic process that removes thermal energy from the water. The cold pack then removes thermal energy from your body.

Solution

We assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself), in which case:

$$q_{rxn} = -q_{soln}$$

with "rxn" and "soln" used as shorthand for "reaction" and "solution," respectively.

Assuming also that the specific heat of the solution is the same as that for water, we have:

$$\begin{aligned} q_{rxn} &= -q_{soln} = -(c \times m \times \Delta T)_{soln} \\ &= -\left[(4.184 J/g^{\circ}C) \times (53.2 g) \times (20.3^{\circ}C - 24.9^{\circ}C)\right] \\ &= -\left[(4.184 J/g^{\circ}C) \times (53.2 g) \times (-4.6^{\circ}C)\right] \\ &+1.0 \times 10^{3} J = +1.0 kJ \end{aligned}$$

The positive sign for q indicates that the dissolution is an endothermic process.

? Exercise 7.3.4

When a 3.00-g sample of KCl was added to 3.00 10^2 g of water in a coffee cup calorimeter, the temperature decreased by 1.05 °C. How much heat is involved in the dissolution of the KCl? What assumptions did you make?

Answer

1.33 kJ; assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself) and that the specific heat of the solution is the same as that for water

If the amount of heat absorbed by a calorimeter is too large to neglect or if we require more accurate results, then we must take into account the heat absorbed both by the solution and by the calorimeter.

The calorimeters described are designed to operate at constant (atmospheric) pressure and are convenient to measure heat flow accompanying processes that occur in solution. A different type of calorimeter that operates at constant volume, colloquially known as a bomb calorimeter, is used to measure the energy produced by reactions that yield large amounts of heat and gaseous products, such as combustion reactions. (The term "bomb" comes from the observation that these reactions can be vigorous enough

to resemble explosions that would damage other calorimeters.) This type of calorimeter consists of a robust steel container (the "bomb") that contains the reactants and is itself submerged in water (Figure 7.3.7). The sample is placed in the bomb, which is then filled with oxygen at high pressure. A small electrical spark is used to ignite the sample. The energy produced by the reaction is absorbed by the steel bomb and the surrounding water. The temperature increase is measured and, along with the known heat capacity of the calorimeter, is used to calculate the energy produced by the reaction. Bomb calorimeters require calibration to determine the heat capacity of the calorimeter and ensure accurate results. The calibration is accomplished using a reaction with a known *q*, such as a measured quantity of benzoic acid ignited by a spark from a nickel fuse wire that is weighed before and after the reaction. The temperature change produced by the known reaction is used to determine the heat capacity of the calorimeter. The calibration is generally performed each time before the calorimeter is used to gather research data.

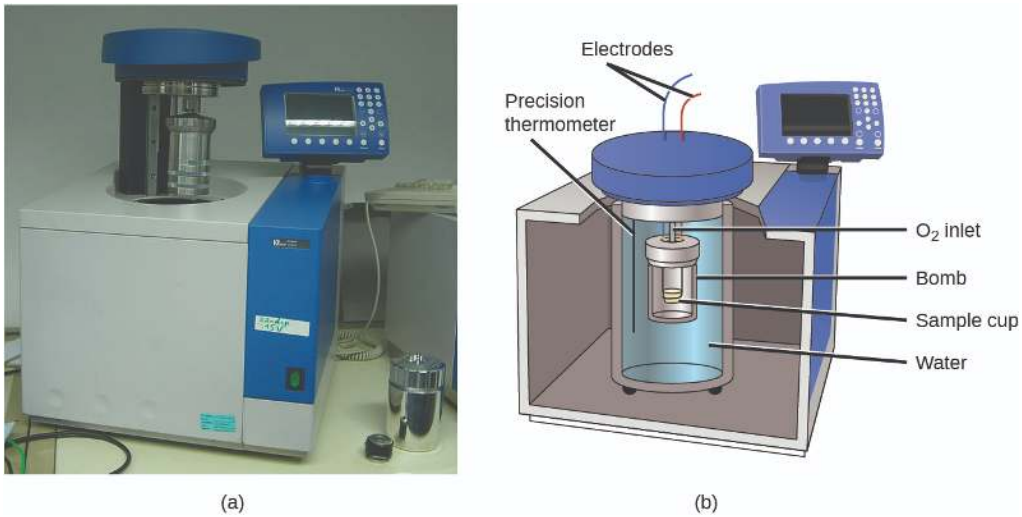


Figure 7.3.7: (a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight "bomb," which is submerged in water and surrounded by insulating materials. (credit a: modification of work by "Harbor1"/Wikimedia commons)

Link to Learning

Click on this link to view how a bomb calorimeter is prepared for action. This site shows calorimetric calculations using sample data.

Example 7.3.5: Bomb Calorimetry

When 3.12 g of glucose, $C_6H_{12}O_6$, is burned in a bomb calorimeter, the temperature of the calorimeter increases from 23.8 °C to 35.6 °C. The calorimeter contains 775 g of water, and the bomb itself has a heat capacity of 893 J/°C. How much heat was produced by the combustion of the glucose sample?

Solution

The combustion produces heat that is primarily absorbed by the water and the bomb. (The amounts of heat absorbed by the reaction products and the unreacted excess oxygen are relatively small and dealing with them is beyond the scope of this text. We will neglect them in our calculations.)

The heat produced by the reaction is absorbed by the water and the bomb:

$$\begin{aligned} q_{rxn} &= -\left(q_{\text{water}} + q_{\text{bomb}}\right) \\ &= -\left[(4.184 J/g^{\circ}C) \times (775 g) \times (35.6^{\circ}C - 23.8^{\circ}C) + 893 J/^{\circ}C \times (35.6^{\circ}C - 23.8^{\circ}C)\right] \\ &= -(38,300 J + 10,500 J) \\ &= -48,800 J = -48.8 kJ \end{aligned}$$

This reaction released 48.7 kJ of heat when 3.12 g of glucose was burned.

Exercise 7.3.5

When 0.963 g of benzene, C_6H_6, is burned in a bomb calorimeter, the temperature of the calorimeter increases by 8.39 °C. The bomb has a heat capacity of 784 J/°C and is submerged in 925 mL of water. How much heat was produced by the combustion of the benzene sample?

Answer

q_{rx} = –39.0 kJ (the reaction produced 39.0 kJ of heat)

Since the first one was constructed in 1899, 35 calorimeters have been built to measure the heat produced by a living person.[2] These whole-body calorimeters of various designs are large enough to hold an individual human being. More recently, whole-room calorimeters allow for relatively normal activities to be performed, and these calorimeters generate data that more closely reflect the real world. These calorimeters are used to measure the metabolism of individuals under different environmental conditions, different dietary regimes, and with different health conditions, such as diabetes.

For example Carla Prado's team at University of Alberta undertook whole-body calorimetry to understand the energy expenditures of women who had recently given birth. Studies like this help develop better recommendations and regimens for nutrition, exercise, and general wellbeing during this period of significant physiological change. In humans, metabolism is typically measured in Calories per day. A nutritional calorie (Calorie) is the energy unit used to quantify the amount of energy derived from the metabolism of foods; one Calorie is equal to 1000 calories (1 kcal), the amount of energy needed to heat 1 kg of water by 1 °C.

Chemistry in Everyday Life: Measuring Nutritional Calories

In your day-to-day life, you may be more familiar with energy being given in Calories, or nutritional calories, which are used to quantify the amount of energy in foods. One calorie (cal) = exactly 4.184 joules, and one Calorie (note the capitalization) = 1000 cal, or 1 kcal. (This is approximately the amount of energy needed to heat 1 kg of water by 1 °C.)

The macronutrients in food are proteins, carbohydrates, and fats or oils. Proteins provide about 4 Calories per gram, carbohydrates also provide about 4 Calories per gram, and fats and oils provide about 9 Calories/g. Nutritional labels on food packages show the caloric content of one serving of the food, as well as the breakdown into Calories from each of the three macronutrients (Figure 7.3.8).

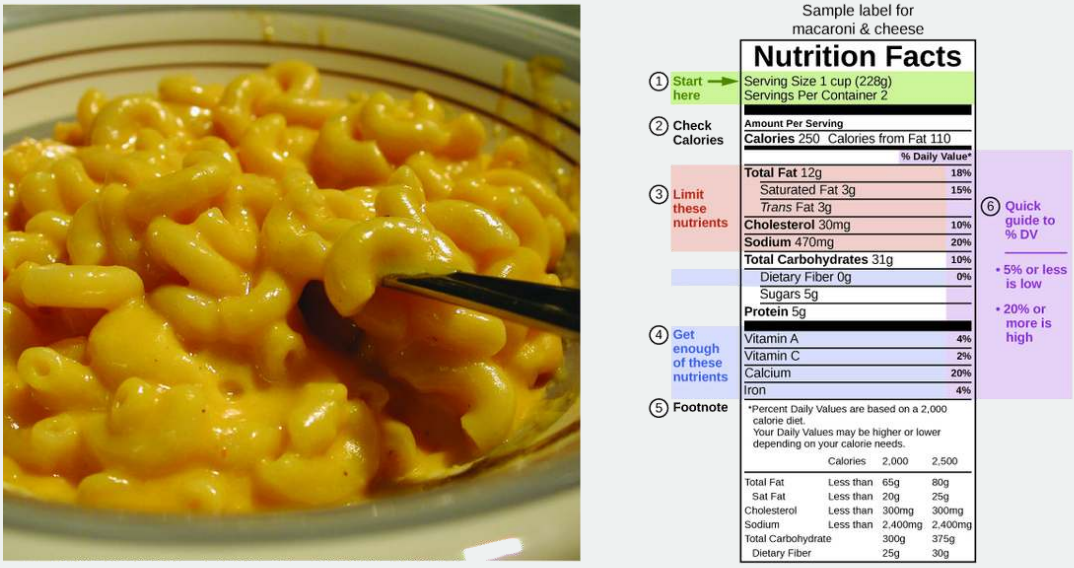


Figure 7.3.8: (a) Macaroni and cheese contain energy in the form of the macronutrients in the food. (b) The food's nutritional information is shown on the package label. In the US, the energy content is given in Calories (per serving); the rest of the world usually uses kilojoules. (credit a: modification of work by "Rex Roof"/Flickr)

For the example shown in (b), the total energy per 228-g portion is calculated by:

$$(5g\,\text{protein} \times 4\,\text{Calories}/g) + (31g\,\text{carb} \times 4\,\text{Calories}/g) + (12g\,\text{fat} \times 9\,\text{Calories}/g) = 252\,\text{Calories}$$

So, you can use food labels to count your Calories. But where do the values come from? And how accurate are they? The caloric content of foods can be determined by using bomb calorimetry; that is, by burning the food and measuring the energy it contains. A sample of food is weighed, mixed in a blender, freeze-dried, ground into powder, and formed into a pellet. The pellet is burned inside a bomb calorimeter, and the measured temperature change is converted into energy per gram of food.

Today, the caloric content on food labels is derived using a method called the Atwater system that uses the average caloric content of the different chemical constituents of food, protein, carbohydrate, and fats. The average amounts are those given in the equation and are derived from the various results given by bomb calorimetry of whole foods. The carbohydrate amount is discounted a certain amount for the fiber content, which is indigestible carbohydrate. To determine the energy content of a food, the quantities of carbohydrate, protein, and fat are each multiplied by the average Calories per gram for each and the products summed to obtain the total energy.

Link to Learning

Click on this link to access the US Department of Agriculture (USDA) National Nutrient Database, containing nutritional information on over 8000 foods.

Footnotes

- 2Francis D. Reardon et al. “The Snellen human calorimeter revisited, re-engineered and upgraded: Design and performance characteristics.” *Medical and Biological Engineering and Computing* 8 (2006)721–28, http://link.springer.com/article/10....517-006-0086-5.

This page titled 7.3: Calorimetry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **5.2: Calorimetry** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

7.4: Enthalpy

Learning Objectives

By the end of this section, you will be able to:

- State the first law of thermodynamics
- Define enthalpy and explain its classification as a state function
- Write and balance thermochemical equations
- Calculate enthalpy changes for various chemical reactions
- Explain Hess's law and use it to compute reaction enthalpies

Thermochemistry is a branch of chemical thermodynamics, the science that deals with the relationships between heat, work, and other forms of energy in the context of chemical and physical processes. As we concentrate on thermochemistry in this chapter, we need to consider some widely used concepts of thermodynamics.

Substances act as reservoirs of energy, meaning that energy can be added to them or removed from them. Energy is stored in a substance when the kinetic energy of its atoms or molecules is raised. The greater kinetic energy may be in the form of increased translations (travel or straight-line motions), vibrations, or rotations of the atoms or molecules. When thermal energy is lost, the intensities of these motions decrease and the kinetic energy falls. The total of all possible kinds of energy present in a substance is called the internal energy (U), sometimes symbolized as E.

As a system undergoes a change, its internal energy can change, and energy can be transferred from the system to the surroundings, or from the surroundings to the system. Energy is transferred into a system when it absorbs heat (q) from the surroundings or when the surroundings do work (w) on the system. For example, energy is transferred into room-temperature metal wire if it is immersed in hot water (the wire absorbs heat from the water), or if you rapidly bend the wire back and forth (the wire becomes warmer because of the work done on it). Both processes increase the internal energy of the wire, which is reflected in an increase in the wire's temperature. Conversely, energy is transferred out of a system when heat is lost from the system, or when the system does work on the surroundings.

The relationship between internal energy, heat, and work can be represented by the equation:

$$\Delta U = q + w$$

as shown in Figure 7.4.1. This is one version of the first law of thermodynamics, and it shows that the internal energy of a system changes through heat flow into or out of the system (positive q is heat flow in; negative q is heat flow out) or work done on or by the system. The work, w, is positive if it is done on the system and negative if it is done by the system.

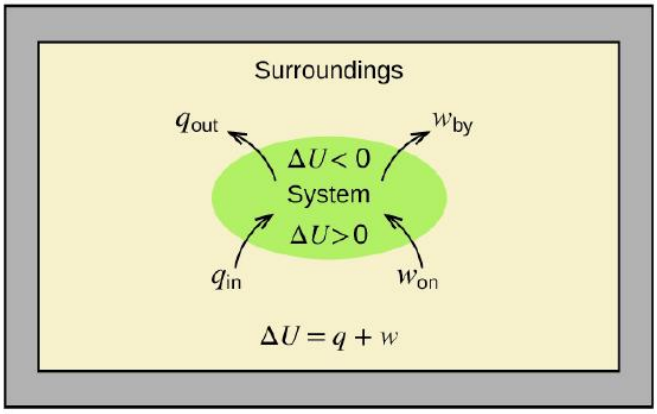


Figure 7.4.1: The internal energy, U, of a system can be changed by heat flow and work. If heat flows into the system, q_{in}, or work is done on the system, w_{on}, its internal energy increases, $\Delta U > 0$. If heat flows out of the system, q_{out}, or work is done by the system, w_{by}, its internal energy decreases, $\Delta U < 0$.

A type of work called expansion work (or pressure-volume work) occurs when a system pushes back the surroundings against a restraining pressure, or when the surroundings compress the system. An example of this occurs during the operation of an internal

combustion engine. The reaction of gasoline and oxygen is exothermic. Some of this energy is given off as heat, and some does work pushing the piston in the cylinder. The substances involved in the reaction are the system, and the engine and the rest of the universe are the surroundings. The system loses energy by both heating and doing work on the surroundings, and its internal energy decreases. (The engine is able to keep the car moving because this process is repeated many times per second while the engine is running.) We will consider how to determine the amount of work involved in a chemical or physical change in the chapter on thermodynamics.

Link to Learning

This view of an internal combustion engine illustrates the conversion of energy produced by the exothermic combustion reaction of a fuel such as gasoline into energy of motion.

As discussed, the relationship between internal energy, heat, and work can be represented as $\Delta U = q + w$. Internal energy is an example of a state function (or state variable), whereas heat and work are not state functions. The value of a state function depends only on the state that a system is in, and not on how that state is reached. If a quantity is not a state function, then its value *does* depend on how the state is reached. An example of a state function is altitude or elevation. If you stand on the summit of Mt. Kilimanjaro, you are at an altitude of 5895 m, and it does not matter whether you hiked there or parachuted there. The distance you traveled to the top of Kilimanjaro, however, is not a state function. You could climb to the summit by a direct route or by a more roundabout, circuitous path (Figure 7.4.2). The distances traveled would differ (distance is not a state function) but the elevation reached would be the same (altitude is a state function).

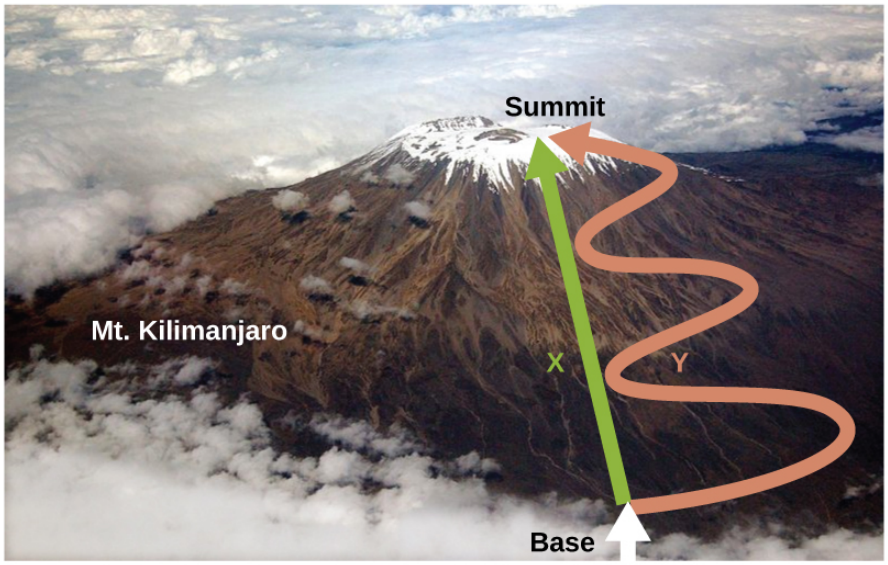


Figure 7.4.2: Paths X and Y represent two different routes to the summit of Mt. Kilimanjaro. Both have the same change in elevation (altitude or elevation on a mountain is a state function; it does not depend on path), but they have very different distances traveled (distance walked is not a state function; it depends on the path). (credit: modification of work by Paul Shaffner)

Chemists ordinarily use a property known as enthalpy (H) to describe the thermodynamics of chemical and physical processes. Enthalpy is defined as the sum of a system's internal energy (U) and the mathematical product of its pressure (P) and volume (V):

$$H = U + PV$$

Enthalpy is also a state function. Enthalpy values for specific substances cannot be measured directly; only enthalpy *changes* for chemical or physical processes can be determined. For processes that take place at constant pressure (a common condition for many chemical and physical changes), the enthalpy change (ΔH) is:

$$\Delta H = \Delta U + P\Delta V$$

The mathematical product $P\Delta V$ represents work (w), namely, expansion or pressure-volume work as noted. By their definitions, the arithmetic signs of ΔV and w will **always** be opposite:

$$P\Delta V = -w$$

Substituting this equation and the definition of internal energy into the enthalpy-change equation yields:

$$\begin{aligned}\Delta H &= \Delta U + P\Delta V \\ &= q_p + w - w \\ &= q_p\end{aligned}$$

where q_p is the **heat of reaction** under conditions of constant pressure.

And so, if a chemical or physical process is carried out at constant pressure with the only work done caused by expansion or contraction, then the heat flow (q_p) and enthalpy change (ΔH) for the process are equal.

The heat given off when you operate a Bunsen burner is equal to the enthalpy change of the methane combustion reaction that takes place, since it occurs at the essentially constant pressure of the atmosphere. On the other hand, the heat produced by a reaction measured in a bomb calorimeter is not equal to ΔH because the closed, constant-volume metal container prevents the pressure from remaining constant (it may increase or decrease if the reaction yields increased or decreased amounts of gaseous species). Chemists usually perform experiments under normal atmospheric conditions, at constant external pressure with $q = \Delta H$, which makes enthalpy the most convenient choice for determining heat changes for chemical reactions.

The following conventions apply when using ΔH:

- A negative value of an enthalpy change, $\Delta H < 0$, indicates an exothermic reaction; a positive value, $\Delta H > 0$, indicates an endothermic reaction. If the direction of a chemical equation is reversed, the arithmetic sign of its ΔH is changed (a process that is endothermic in one direction is exothermic in the opposite direction).
- Chemists use a thermochemical equation to represent the changes in both matter and energy. In a thermochemical equation, the enthalpy change of a reaction is shown as a ΔH value following the equation for the reaction. This ΔH value indicates the amount of heat associated with the reaction involving the number of moles of reactants and products *as shown in the chemical equation*. For example, consider this equation:

  $$\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)} \quad \Delta H = -286 kJ$$

  This equation indicates that when 1 mole of hydrogen gas and mole of oxygen gas at some temperature and pressure change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released to the surroundings. If the coefficients of the chemical equation are multiplied by some factor, the enthalpy change must be multiplied by that same factor (ΔH is an extensive property):

  - two-fold increase in amounts

    $$\mathrm{2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)} \quad \Delta H = 2 \times (-286 kJ) = -572 kJ$$

  - two-fold decrease in amounts

    $$\mathrm{\tfrac{1}{2}H_2(g) + \tfrac{1}{4}O_2(g) \longrightarrow \tfrac{1}{2}H_2O(l)} \quad \Delta H = \frac{1}{2} \times (-286 kJ) = -143 kJ$$

- The enthalpy change of a reaction depends on the physical states of the reactants and products, so these must be shown. For example, when 1 mole of hydrogen gas and mole of oxygen gas change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released. If gaseous water forms, only 242 kJ of heat are released.

  $$\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(g)} \quad \Delta H = -242 kJ$$

Example 7.4.1: Writing Thermochemical Equations

When 0.0500 mol of HCl(*aq*) reacts with 0.0500 mol of NaOH(*aq*) to form 0.0500 mol of NaCl(*aq*), 2.9 kJ of heat are produced. Write a balanced thermochemical equation for the reaction of one mole of HCl.

$$\mathrm{HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)}$$

Solution

For the reaction of 0.0500 mol acid (HCl), q = −2.9 kJ. The reactants are provided in stoichiometric amounts (same molar ratio as in the balanced equation), and so the amount of acid may be used to calculate a molar enthalpy change. Since ΔH is an extensive property, it is proportional to the amount of acid neutralized:

$$\Delta H = 1 \text{ mol HCl} \times \frac{-2.9kJ}{0.0500 \text{ mol HCl}} = -58kJ$$

The thermochemical equation is then

$$\text{HCl(aq)} + \text{NaOH(aq)} \longrightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \quad \Delta H = -58kJ$$

? Exercise 7.4.1

When 1.34 g Zn(*s*) reacts with 60.0 mL of 0.750 M HCl(*aq*), 3.14 kJ of heat are produced. Determine the enthalpy change per mole of zinc reacting for the reaction:

$$\text{Zn(s)} + 2\,\text{HCl(aq)} \longrightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

Answer

ΔH = −153 kJ

Be sure to take both stoichiometry and limiting reactants into account when determining the ΔH for a chemical reaction.

✓ Example 7.4.2: Writing Thermochemical Equations

A gummy bear contains 2.67 g sucrose, $C_{12}H_{22}O_{11}$. When it reacts with 7.19 g potassium chlorate, $KClO_3$, 43.7 kJ of heat are produced. Write a thermochemical equation for the reaction of one mole of sucrose:

$$\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(aq)} + 8\,\text{KClO}_3\text{(aq)} \longrightarrow 12\,\text{CO}_2\text{(g)} + 11\,\text{H}_2\text{O(l)} + 8\,\text{KCl(aq)}$$

Solution

Unlike the previous example exercise, this one does not involve the reaction of stoichiometric amounts of reactants, and so the *limiting reactant* must be identified (it limits the yield of the reaction and the amount of thermal energy produced or consumed).

The provided amounts of the two reactants are

$$(2.67\,\text{g})\left(\frac{1\,\text{mol}}{342.3\,\text{g}}\right) = 0.00780\,\text{mol}\,\text{C}_{12}\text{H}_{22}\text{O}_{11}$$

$$(7.19\,\text{g})\left(\frac{1\,\text{mol}}{122.5\,\text{g}}\right) = 0.0587\,\text{mol}\,\text{KCIO}$$

The provided molar ratio of perchlorate-to-sucrose is then

$$\frac{0.0587\,\text{mol}\,\text{KClO}_3}{0.00780\,\text{mol}\,\text{C}_{12}\text{H}_{22}\text{O}_{11}} = 7.52$$

The balanced equation indicates 8 mol $KClO_3$ are required for reaction with 1 mol $C_{12}H_{22}O_{11}$. Since the provided amount of $KClO_3$ is less than the stoichiometric amount, it is the limiting reactant and may be used to compute the enthalpy change:

$$\Delta H = \frac{-43.7\,\text{kJ}}{0.0587\,\text{molKClO}_3} = \frac{744\,\text{kJ}}{\text{mol}\,\text{KClO}_3}$$

Because the equation, as written, represents the reaction of 8 mol $KClO_3$, the enthalpy change is

$$\left(\frac{744\,\text{kJ}}{1\ \text{mol}\,\text{KClO}_3}\right)\left(8\ \text{mol}\,\text{KClO}_3\right) = 5960\,\text{kJ}$$

The enthalpy change for this reaction is −5960 kJ, and the thermochemical equation is:

$$\text{C}_{12}\text{H}_{22}\text{O}_{11} + 8\,\text{KClO}_3 \longrightarrow 12\,\text{CO}_2 + 11\,\text{H}_2\text{O} + 8\,\text{KCl} \quad \Delta H = -5960\,\text{kJ}$$

Exercise 7.4.1

When 1.42 g of iron reacts with 1.80 g of chlorine, 3.22 g of $\mathrm{FeCl_2(s)}$ and 8.60 kJ of heat is produced. What is the enthalpy change for the reaction when 1 mole of $\mathrm{FeCl_2(s)}$ is produced?

Answer

$\Delta H = -338$ kJ

Enthalpy changes are typically tabulated for reactions in which both the reactants and products are at the same conditions. A standard state is a commonly accepted set of conditions used as a reference point for the determination of properties under other different conditions. For chemists, the IUPAC standard state refers to materials under a pressure of 1 bar and solutions at 1 M, and does not specify a temperature. Many thermochemical tables list values with a standard state of 1 atm. Because the ΔH of a reaction changes very little with such small changes in pressure (1 bar = 0.987 atm), ΔH values (except for the most precisely measured values) are essentially the same under both sets of standard conditions. We will include a superscripted "o" in the enthalpy change symbol to designate standard state. Since the usual (but not technically standard) temperature is 298.15 K, this temperature will be assumed unless some other temperature is specified. Thus, the symbol is used to indicate an enthalpy change for a process occurring under these conditions. (The symbol ΔH is used to indicate an enthalpy change for a reaction occurring under nonstandard conditions.)

The enthalpy changes for many types of chemical and physical processes are available in the reference literature, including those for combustion reactions, phase transitions, and formation reactions. As we discuss these quantities, it is important to pay attention to the *extensive* nature of enthalpy and enthalpy changes. Since the enthalpy change for a given reaction is proportional to the amounts of substances involved, it may be reported on that basis (i.e., as the ΔH for specific amounts of reactants). However, we often find it more useful to divide one extensive property (ΔH) by another (amount of substance), and report a per-amount *intensive* value of ΔH, often "normalized" to a per-mole basis. (Note that this is similar to determining the intensive property specific heat from the extensive property heat capacity, as seen previously.)

Standard Enthalpy of Combustion

Standard enthalpy of combustion is the enthalpy change when 1 mole of a substance burns (combines vigorously with oxygen) under standard state conditions; it is sometimes called "heat of combustion." For example, the enthalpy of combustion of ethanol, −1366.8 kJ/mol, is the amount of heat produced when one mole of ethanol undergoes complete combustion at 25 °C and 1 atmosphere pressure, yielding products also at 25 °C and 1 atm.

$$\mathrm{C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2 + 3\,H_2O(l)} \quad \Delta H^\circ = -1366.8\,\mathrm{kJ}$$

Enthalpies of combustion for many substances have been measured; a few of these are listed in Table 7.4.1. Many readily available substances with large enthalpies of combustion are used as fuels, including hydrogen, carbon (as coal or charcoal), and hydrocarbons (compounds containing only hydrogen and carbon), such as methane, propane, and the major components of gasoline.

Table 7.4.1: Standard Molar Enthalpies of Combustion

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH_c° (kJ/mol) at 25 °C
carbon	$C(s) + O_2(g) \longrightarrow CO_2(g)$	−393.5
hydrogen	$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$	−285.8
magnesium	$Mg(s) + \frac{1}{2}O_2(g) \longrightarrow MgO(s)$	−601.6
sulfur	$S(s) + O_2(g) \longrightarrow SO_2(g)$	−296.8
carbon monoxide	$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$	−283.0
methane	$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$	−890.8
acetylene	$C_2H_2(g) + \frac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(l)$	−1301.1

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH_c° (kJ/mol) at 25 °C
ethanol	$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$	−1366.8
methanol	$CH_3OH(l) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$	−726.1
isooctane	$C_8H_{18}(l) + \frac{25}{2}O_2(g) \longrightarrow 8CO_2(g) + 9H_2O(l)$	−5461

Example 7.4.3: Using Enthalpy of Combustion

As Figure 7.4.3 suggests, the combustion of gasoline is a highly exothermic process. Let us determine the approximate amount of heat produced by burning 1.00 L of gasoline, assuming the enthalpy of combustion of gasoline is the same as that of isooctane, a common component of gasoline. The density of isooctane is 0.692 g/mL.



Figure 7.4.3: The combustion of gasoline is very exothermic. (credit: modification of work by “AlexEagle”/Flickr)

Solution

Starting with a known amount (1.00 L of isooctane), we can perform conversions between units until we arrive at the desired amount of heat or energy. The enthalpy of combustion of isooctane provides one of the necessary conversions. Table 7.4.1 gives this value as −5460 kJ per 1 mole of isooctane (C_8H_{18}).

Using these data,

$$1.00\ \text{L}\,\cancel{\text{C}_8\text{H}_{18}} \times \left(\frac{1000\ \text{mL}\,\cancel{\text{C}_8\text{H}_{18}}}{1\ \text{L}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \left(\frac{0.692\ \text{g}\,\cancel{\text{C}_8\text{H}_{18}}}{1\ \text{mL}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \left(\frac{1\ \text{mol}\,\cancel{\text{C}_8\text{H}_{18}}}{114\ \text{g}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \frac{-5460\ \text{kJ}}{1\ \text{mol}\,\cancel{\text{C}_8\text{H}_{18}}} = -3.31 \times 10^4\ \text{kJ}$$

The combustion of 1.00 L of isooctane produces 33,100 kJ of heat. (This amount of energy is enough to melt 99.2 kg, or about 218 lbs, of ice.)

Note: If you do this calculation one step at a time, you would find:

$$1.00\,\text{L}\,\text{C}_8\text{H}_{18} \longrightarrow 1.00 \times 10^3\ \text{mL}\text{C}_8\text{H}_{18}$$
$$1.00 \times 10^3\ \text{mL}\,\text{C}_8\text{H}_{18} \longrightarrow 692\,\text{g}\,\text{C}_8\text{H}_{18}$$
$$692\,\text{g}\,\text{C}_8\text{H}_{18} \longrightarrow 6.07\,\text{mol}\,\text{C}_8\text{H}_{18}$$
$$6.07\,\text{mol}\,\text{C}_8\text{H}_{18} \longrightarrow -3.31 \times 10^4 kJ$$

Exercise 7.4.3

How much heat is produced by the combustion of 125 g of acetylene?

Answer

6.25 10^3 kJ

https://chem.libretexts.org/@go/page/481002

Chemistry in Everyday Life: Emerging Algae-Based Energy Technologies (Biofuels)

As reserves of fossil fuels diminish and become more costly to extract, the search is ongoing for replacement fuel sources for the future. Among the most promising biofuels are those derived from algae (Figure 5.22). The species of algae used are nontoxic, biodegradable, and among the world's fastest growing organisms. About 50% of algal weight is oil, which can be readily converted into fuel such as biodiesel. Algae can yield 26,000 gallons of biofuel per hectare—much more energy per acre than other crops. Some strains of algae can flourish in brackish water that is not usable for growing other crops. Algae can produce biodiesel, biogasoline, ethanol, butanol, methane, and even jet fuel.



(a) (b) (c)

Figure 5.22 (a) Tiny algal organisms can be (b) grown in large quantities and eventually (c) turned into a useful fuel such as biodiesel. (credit a: modification of work by Micah Sittig; credit b: modification of work by Robert Kerton; credit c: modification of work by John F. Williams)

According to the US Department of Energy, only 39,000 square kilometers (about 0.4% of the land mass of the US or less than of the area used to grow corn) can produce enough algal fuel to replace all the petroleum-based fuel used in the US. The cost of algal fuels is becoming more competitive—for instance, the US Air Force is producing jet fuel from algae at a total cost of under $5 per gallon.[3] The process used to produce algal fuel is as follows: grow the algae (which use sunlight as their energy source and CO_2 as a raw material); harvest the algae; extract the fuel compounds (or precursor compounds); process as necessary (e.g., perform a transesterification reaction to make biodiesel); purify; and distribute (Figure 7.4.4).



Figure 7.4.4: Algae convert sunlight and carbon dioxide into oil that is harvested, extracted, purified, and transformed into a variety of renewable fuels.

Link to Learning

Click here to learn more about the process of creating algae biofuel.

Standard Enthalpy of Formation

A standard enthalpy of formation is an enthalpy change for a reaction in which exactly 1 mole of a pure substance is formed from free elements in their most stable states under standard state conditions. These values are especially useful for computing or predicting enthalpy changes for chemical reactions that are impractical or dangerous to carry out, or for processes for which it is difficult to make measurements. If we have values for the appropriate standard enthalpies of formation, we can determine the enthalpy change for any reaction, which we will practice in the next section on Hess's law.

The standard enthalpy of formation of $CO_2(g)$ is −393.5 kJ/mol. This is the enthalpy change for the exothermic reaction:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H_f^\circ = \Delta H^\circ = -393.5\,kJ}$$

starting with the reactants at a pressure of 1 atm and 25 °C (with the carbon present as graphite, the most stable form of carbon under these conditions) and ending with one mole of CO_2, also at 1 atm and 25 °C. For nitrogen dioxide, $NO_2(g)$, is 33.2 kJ/mol. This is the enthalpy change for the reaction:

$$\tfrac{1}{2}\mathrm{N_2(g)} + \mathrm{O_2(g)} \longrightarrow \mathrm{NO_2(g)} \quad \Delta H_f^\circ = \Delta H^\circ = +33.2\,\mathrm{kJ}$$

A reaction equation with mole of $\mathrm{N_2}$ and 1 mole of $\mathrm{O_2}$ is correct in this case because the standard enthalpy of formation always refers to 1 mole of product, $\mathrm{NO_2(g)}$.

You will find a table of standard enthalpies of formation of many common substances in Appendix G. These values indicate that formation reactions range from highly exothermic (such as −2984 kJ/mol for the formation of $\mathrm{P_4O_{10}}$) to strongly endothermic (such as +226.7 kJ/mol for the formation of acetylene, $\mathrm{C_2H_2}$). By definition, the standard enthalpy of formation of an element in its most stable form is equal to zero under standard conditions, which is 1 atm for gases and 1 M for solutions.

Example 7.4.4: Evaluating an Enthalpy of Formation

Ozone, $\mathrm{O_3(g)}$, forms from oxygen, $\mathrm{O_2(g)}$, by an endothermic process. Ultraviolet radiation is the source of the energy that drives this reaction in the upper atmosphere. Assuming that both the reactants and products of the reaction are in their standard states, determine the standard enthalpy of formation, of ozone from the following information:

$$3\,\mathrm{O_2(g)} \longrightarrow 2\,\mathrm{O_3(g)} \quad \Delta H^\circ = +286\,\mathrm{kJ}$$

Solution

ΔH_f° is the enthalpy change for the formation of one mole of a substance in its standard state from the elements in their standard states. Thus, for $\mathrm{O_3(g)}$ is the enthalpy change for the reaction:

$$\tfrac{3}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{O_3(g)}$$

For the formation of 2 mol of $\mathrm{O_3(g)}$, $\Delta H^\circ = +286\,\mathrm{kJ}$. This ratio, can be used as a conversion factor to find the heat produced when 1 mole of $O_3(g)$ is formed, which is the enthalpy of formation for $\mathrm{O_3(g)}$:

$$\Delta H^\circ \text{ for 1 mole of } \mathrm{O_3}(g) = 1\ \mathrm{mol\ \cancel{O_3}} \times \frac{286\,\mathrm{kJ}}{2\,\mathrm{mol\ O_3}} = 143\,\mathrm{kJ}$$

Therefore,

$$\Delta H_f^\circ\,[O_3(g)] = +143\,\mathrm{kJ/mol}.$$

Exercise 7.4.4

Hydrogen gas, $\mathrm{H_2}$, reacts explosively with gaseous chlorine, $\mathrm{Cl_2}$, to form hydrogen chloride, $\mathrm{HCl(g)}$. What is the enthalpy change for the reaction of 1 mole of $\mathrm{H_2(g)}$ with 1 mole of $\mathrm{Cl_2(g)}$ if both the reactants and products are at standard state conditions? The standard enthalpy of formation of $\mathrm{HCl(g)}$ is −92.3 kJ/mol.

Answer

For the reaction

$$\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl(g)} \quad \Delta H^\circ = -184.6\,\mathrm{kJ}$$

Example 7.4.5: Writing Reaction Equations for

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OH(l)}$
b. $\mathrm{Ca_3(PO_4)_2(s)}$

Solution

Remembering that reaction equations are for forming 1 mole of the compound from its constituent elements under standard conditions, we have:

a. $$2\,\mathrm{C(s,\,graphite)} + 3\,\mathrm{H_2(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{C_2H_5OH(l)}$$

b. $$3\,\mathrm{Ca(s)} + \tfrac{1}{2}\mathrm{P_4(s)} + 4\,\mathrm{O_2(g)} \longrightarrow \mathrm{Ca_3(PO_4)_2(s)}$$

Note: The standard state of carbon is graphite, and phosphorus exists as $\mathrm{P_4}$.

Exercise 7.4.5

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OC_2H_5(l)}$
b. $\mathrm{Na_2CO_3(s)}$

Answer

a. $$4\,\mathrm{C(s,\,graphite)} + 5\,\mathrm{H_2(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{C_2H_5OC_2H_5(l)}$$

b. $$2\,\mathrm{Na(s)} + \mathrm{C(s,\,graphite)} + \tfrac{3}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{Na_2CO_3(s)}$$

Hess's Law

There are two ways to determine the amount of heat involved in a chemical change: measure it experimentally, or calculate it from other experimentally determined enthalpy changes. Some reactions are difficult, if not impossible, to investigate and make accurate measurements for experimentally. And even when a reaction is not hard to perform or measure, it is convenient to be able to determine the heat involved in a reaction without having to perform an experiment.

This type of calculation usually involves the use of Hess's law, which states: *If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps.* Hess's law is valid because enthalpy is a state function: Enthalpy changes depend only on where a chemical process starts and ends, but not on the path it takes from start to finish. For example, we can think of the reaction of carbon with oxygen to form carbon dioxide as occurring either directly or by a two-step process. The direct process is written:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)} \quad \Delta H^\circ = -394kJ$$

In the two-step process, first carbon monoxide is formed:

$$\mathrm{C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g)} \quad \Delta H^\circ = -111kJ$$

Then, carbon monoxide reacts further to form carbon dioxide:

$$\mathrm{CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g)} \quad \Delta H^\circ = -283kJ$$

The equation describing the overall reaction is the sum of these two chemical changes:

step 1	$\mathrm{C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g)}$
step 2	$\mathrm{CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g)}$
sum	$\mathrm{C(s) + \cancel{CO(g)} + O_2(g) \longrightarrow \cancel{CO(g)} + CO_2(g)}$

Because the $\mathrm{CO(g)}$ produced in Step 1 is consumed in Step 2, the net change is:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)}$$

According to Hess's law, the enthalpy change of the reaction will equal the sum of the enthalpy changes of the steps.

step 1	$\mathrm{C(s) + \frac{1}{2}O_2(g) \longrightarrow \cancel{CO(g)}}$	$\Delta H^\circ = -111\ \mathrm{kJ}$
step 2	$\mathrm{\cancel{CO(g)} + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)}$	$\Delta H^\circ = -283\ \mathrm{kJ}$
sum	$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)}$	$\Delta H^\circ = -394\ \mathrm{kJ}$

The result is shown in Figure 7.4.5. We see that ΔH of the overall reaction is the same whether it occurs in one step or two. This finding (overall ΔH for the reaction = sum of ΔH values for reaction "steps" in the overall reaction) is true in general for chemical and physical processes.

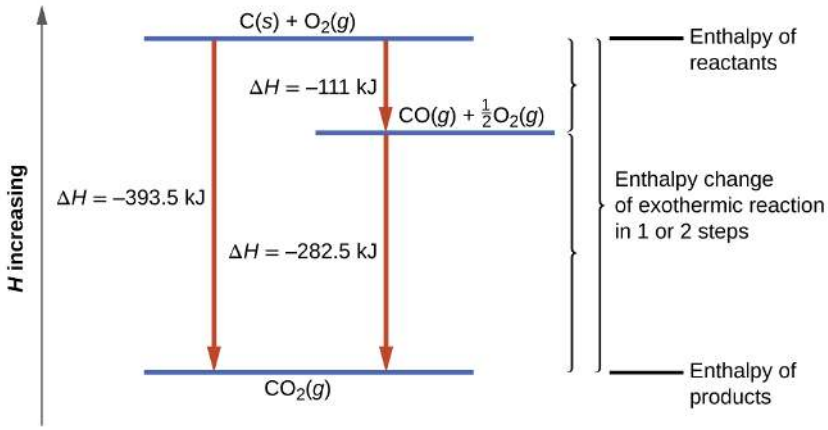


Figure 7.4.5: The formation of $\mathrm{CO_2(g)}$ from its elements can be thought of as occurring in two steps, which sum to the overall reaction, as described by Hess's law. The horizontal blue lines represent enthalpies. For an exothermic process, the products are at lower enthalpy than are the reactants.

Before we further practice using Hess's law, let us recall two important features of ΔH.

1. ΔH is directly proportional to the quantities of reactants or products. For example, the enthalpy change for the reaction forming 1 mole of $NO_2(g)$ is +33.2 kJ:

$$\mathrm{\frac{1}{2}N_2(g) + O_2(g) \longrightarrow NO_2(g)} \quad \Delta H = +33.2\ \mathrm{kJ}$$

When 2 moles of $\mathrm{NO_2}$ (twice as much) are formed, the ΔH will be twice as large:

$$\mathrm{N_2(g) + 2\,O_2(g) \longrightarrow 2\,NO_2(g)} \quad \Delta H = +66.4\ \mathrm{kJ}$$

In general, if we multiply or divide an equation by a number, then the enthalpy change should also be multiplied or divided by the same number.

2. ΔH for a reaction in one direction is equal in magnitude and opposite in sign to ΔH for the reaction in the reverse direction. For example, given that:

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)} \quad \Delta H = -184.6\ \mathrm{kJ}$$

Then, for the "reverse" reaction, the enthalpy change is also "reversed":

Example 7.4.6: Stepwise Calculation of Heat of Formation Using Hess's Law

Determine the enthalpy of formation, of $\mathrm{FeCl_3(s)}$ from the enthalpy changes of the following two-step process that occurs under standard state conditions:

$$\mathrm{Fe(s) + Cl_2(g) \longrightarrow FeCl_2(s)} \quad \Delta H^\circ = -341.8\ \mathrm{kJ}$$

$$\mathrm{FeCl_2(s) + \frac{1}{2}Cl_2(g) \longrightarrow FeCl_3(s)} \quad \Delta H^\circ = -57.7\ \mathrm{kJ}$$

Solution

We are trying to find the standard enthalpy of formation of $\mathrm{FeCl_3(s)}$, which is equal to ΔH° for the reaction:

$$\mathrm{Fe(s) + \frac{3}{2}Cl_2(g) \longrightarrow FeCl_3(s)} \quad \Delta H_f^\circ = ?$$

Looking at the reactions, we see that the reaction for which we want to find ΔH° is the sum of the two reactions with known ΔH values, so we must sum their ΔH values:

step 1	$\mathrm{Fe(s) + Cl_2(g) \longrightarrow \cancel{FeCl_2(s)}}$	$\Delta H^\circ = -341.8\,\mathrm{kJ}$
step 2	$\mathrm{\cancel{FeCl_2(s)} + \frac{1}{2}Cl_2(g) \longrightarrow FeCl_3(s)}$	$\Delta H^\circ = -57.7\,\mathrm{kJ}$
sum	$\mathrm{Fe(s) + \frac{3}{2}Cl_2(g) \longrightarrow FeCl_3(s)}$	$\Delta H^\circ = -399.5\,\mathrm{kJ}$

The enthalpy of formation, ΔH_f° of $\mathrm{FeCl_3(s)}$ is −399.5 kJ/mol.

Exercise 7.4.6

Calculate ΔH for the process:

$$\mathrm{N_2(g) + 2\,O_2(g) \longrightarrow 2\,NO_2(g)}$$

from the following information:

$$\mathrm{N_2(g) + O_2(g) \longrightarrow 2\,NO(g)} \quad \Delta H = 180.5\,\mathrm{kJ}$$

$$\mathrm{NO(g) + \tfrac{1}{2}O_2(g) \longrightarrow NO_2(g)} \quad \Delta H = -57.06\,\mathrm{kJ}$$

Answer

66.4 kJ

Here is a less straightforward example that illustrates the thought process involved in solving many Hess's law problems. It shows how we can find many standard enthalpies of formation (and other values of ΔH) if they are difficult to determine experimentally.

Example 7.4.7: A More Challenging Problem Using Hess's Law

Chlorine monofluoride can react with fluorine to form chlorine trifluoride:

i. $$\mathrm{ClF(g) + F_2(g) \longrightarrow ClF_3(g)} \quad \Delta H^\circ = ?$$

Use the reactions here to determine the ΔH^o for reaction (i):

ii. $$\mathrm{2\,OF_2(g) \longrightarrow O_2(g) + 2\,F_2(g)} \quad \Delta H^\circ_{(ii)} = -49.4\,\mathrm{kJ}$$

iii. $$\mathrm{2\,ClF(g) + O_2(g) \longrightarrow Cl_2O(g) + OF_2(g)} \quad \Delta H^\circ_{(iii)} = +214.0\,\mathrm{kJ}$$

iv. $$\mathrm{ClF_3(g) + O_2(g) \longrightarrow \tfrac{1}{2}Cl_2O(g) + \tfrac{3}{2}OF_2(g)} \quad \Delta H^\circ_{(in)} = +236.2\,\mathrm{kJ}$$

Solution

Our goal is to manipulate and combine reactions (ii), (iii), and (iv) such that they add up to reaction (i). Going from left to right in (i), we first see that ClF(g) is needed as a reactant. This can be obtained by multiplying reaction (iii) by $\frac{1}{2}$ which means that the ΔH^o change is also multiplied by $\frac{1}{2}$:

$$\mathrm{ClF(g) + \tfrac{1}{2}O_2(g) \longrightarrow \tfrac{1}{2}Cl_2O(g) + \tfrac{1}{2}OF_2(g)} \quad \Delta H^\circ = \frac{1}{2}(214.0) = +107.0\,\mathrm{kJ}$$

Next, we see that F_2 is also needed as a reactant. To get this, reverse and halve reaction *(ii)*, which means that the ΔH^o changes sign and is halved:

$$\mathrm{\tfrac{1}{2}O_2(g) + F_2(g) \longrightarrow OF_2(g)} \quad \Delta H^\circ = +24.7\,\mathrm{kJ}$$

To get $\mathrm{ClF_3}$ as a product, reverse (iv), changing the sign of ΔH^o:

$$\mathrm{\tfrac{1}{2}Cl_2O(g) + \tfrac{3}{2}OF_2(g) \longrightarrow ClF_3(g) + O_2(g)} \quad \Delta H^\circ = -236.2kJ$$

Now check to make sure that these reactions add up to the reaction we want:

step 1	$\mathrm{ClF(g)} + \cancel{\frac{1}{2}\mathrm{O_2(g)}} \longrightarrow \cancel{\frac{1}{2}\mathrm{Cl_2O(g)}} + \cancel{\frac{1}{2}\mathrm{OF_2(g)}}$	$\Delta H^\circ = +107.0\,\mathrm{kJ}$
step 2	$\cancel{\frac{1}{2}\mathrm{O_2(g)}} + \mathrm{F_2(g)} \longrightarrow \cancel{\mathrm{OF_2(g)}}$	$\Delta H^\circ = +24.7\,\mathrm{kJ}$
step 3	$\cancel{\frac{1}{2}\mathrm{Cl_2O(g)}} + \cancel{\frac{3}{2}\mathrm{OF_2(g)}} \longrightarrow \mathrm{ClF_3(g)} + \cancel{\mathrm{O_2(g)}}$	$\Delta H^\circ = -236.2\,\mathrm{kJ}$
sum	$\mathrm{ClF(g)} + \mathrm{F_2(g)} \longrightarrow \mathrm{ClF_3(g)}$	$\Delta H^\circ = -104.5\,\mathrm{kJ}$

Reactants $\frac{1}{2}\mathrm{O_2}$ and $\frac{1}{2}\mathrm{O_2}$ cancel out the product $\mathrm{O_2}$; the $\frac{1}{2}\mathrm{Cl_2O}$ product cancels $\frac{1}{2}\mathrm{Cl_2O}$ reactant and $\frac{3}{2}\mathrm{Cl_2O}$ reactant is cancelled by products $\frac{1}{2}\mathrm{OF_2}$ and $\mathrm{OF_2}$. This leaves only reactants $\mathrm{ClF(g)}$ and $\mathrm{F_2(g)}$ and product $\mathrm{ClF_3(g)}$, which are what we want. Since summing these three modified reactions yields the reaction of interest, summing the three modified ΔH^o values will give the desired ΔH^o:

$$\Delta H^\circ = (+107.0kJ) + (24.7kJ) + (-236.2kJ) = -104.5\,\mathrm{kJ}$$

Exercise 7.4.7

Aluminum chloride can be formed from its elements:

i. $$2\,\mathrm{Al(s)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{AlCl_3(s)} \quad \Delta H^\circ = ?$$

Use the reactions here to determine the ΔH^o for reaction *(i)*:

ii. $$\mathrm{HCl(g)} \longrightarrow \mathrm{HCl(aq)} \quad \Delta H^\circ_{(ii)} = -74.8\,\mathrm{kJ}$$

iii. $$\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl(g)} \quad \Delta H^\circ_{(iii)} = -185\,\mathrm{kJ}$$

iv. $$\mathrm{AlCl_3(aq)} \longrightarrow \mathrm{AlCl_3(s)} \quad \Delta H^\circ_{(iv)} = +323\,\mathrm{kJ}$$

v. $$2\,\mathrm{Al(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{AlCl_3(aq)} + 3\,\mathrm{H_2(g)} \quad \Delta H^\circ_{(v)} = -1049\,\mathrm{kJ}$$

Answer

−1407 kJ

We also can use Hess's law to determine the enthalpy change of any reaction if the corresponding enthalpies of formation of the reactants and products are available. The stepwise reactions we consider are:

i. decompositions of the reactants into their component elements (for which the enthalpy changes are proportional to the negative of the enthalpies of formation of the reactants), followed by
ii. re-combinations of the elements to give the products (with the enthalpy changes proportional to the enthalpies of formation of the products).

The standard enthalpy change of the overall reaction is therefore equal to: (ii) the sum of the standard enthalpies of formation of all the products plus (i) the sum of the negatives of the standard enthalpies of formation of the reactants. This is usually rearranged slightly to be written as follows, with ∑ representing "the sum of" and *n* standing for the stoichiometric coefficients:

$$\Delta H^\circ_{\text{reaction}} = \sum n \times \Delta H^\circ_f(\text{ products }) - \sum n \times \Delta H^\circ_f(\text{ reactants })$$

The following example shows in detail why this equation is valid, and how to use it to calculate the enthalpy change for a reaction of interest.

Example 7.4.8: Using Hess's Law

What is the standard enthalpy change for the reaction:

$$3\,\mathrm{NO_2(g)} + \mathrm{H_2O(l)} \longrightarrow 2\,\mathrm{HNO_3(aq)} + \mathrm{NO(g)} \quad \Delta H^\circ = ?$$

Solution: Using the Equation

Use the special form of Hess's law given previously, and values from Appendix G:

$$\begin{aligned}\Delta H^\circ_{\text{reaction}} &= \sum n \times \Delta H^\circ_f(\text{ products }) - \sum n \times \Delta H^\circ_f(\text{ reactants }) \\ &= \left[2\ \text{mol}\,\cancel{\text{HNO}_3\text{(aq)}} \times \frac{-207.4\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{HNO}_3\text{(aq)}}} + 1\ \text{mol}\,\cancel{\text{NO(g)}} \times \frac{+90.2\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{NO(g)}}}\right] \\ &\quad - \left[3\ \text{mol}\,\cancel{\text{NO}_2\text{(g)}} \times \frac{+33.2\,\text{kJ}}{1\ \text{mole}\,\cancel{\text{NO}_2\text{(g)}}} + 1\ \text{mol}\,\cancel{\text{H}_2\text{O(l)}} \times \frac{-285.8\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{H}_2\text{O(l)}}}\right] \\ &= [2 \times (-206.64) + 90.25] - [3 \times 33.2 + -(-285.83)] \\ &= -323.03 + 186.23 \\ &= -136.80\,\text{kJ}\end{aligned}$$

Solution: Supporting Why the General Equation is Valid

Alternatively, we can write this reaction as the sum of the decompositions of $3\,\text{NO}_2\text{(g)}$ and $H2O(l)$ into their constituent elements, and the formation of $2\,\text{HNO}_3\text{(aq)}$ and $1\,\text{NO(g)}$ from their constituent elements. Writing out these reactions, and noting their relationships to the ΔH° values for these compounds (from Appendix G), we have:

$$\begin{aligned} 3\,\text{NO}_2\text{(g)} &\longrightarrow \tfrac{3}{2}\text{N}_2\text{(g)} + 3\,\text{O}_2\text{(g)} && \Delta H^\circ_1 = -99.6\,\text{kJ} \\ \text{H}_2\text{O(l)} &\longrightarrow \text{H}_2\text{(g)} + \tfrac{1}{2}\text{O}_2\text{(g)} && \Delta H^\circ_2 = +285.8\,\text{kJ}\left[-1 \times \Delta H^\circ_f(\text{H}_2\text{O})\right] \\ \text{H}_2\text{(g)} + \text{N}_2\text{(g)} + 3\,\text{O}_2\text{(g)} &\longrightarrow 2\,\text{HNO}_3\text{(aq)} && \Delta H^\circ_3 = -414.8\,\text{kJ}\left[2 \times \Delta H^\circ_f(\text{HNO}_3)\right] \\ \tfrac{1}{2}\text{N}_2\text{(g)} + \tfrac{1}{2}\text{O}_2\text{(g)} &\longrightarrow \text{NO(g)} && \Delta H^\circ_4 = +90.2\,\text{kJ}\left[1 \times \Delta H^\circ_f(\text{NO})\right] \end{aligned}$$

Summing these reaction equations gives the reaction we are interested in:

$$3\,\text{NO}_2\text{(g)} + \text{H}_2\text{O(l)} \longrightarrow 2\,\text{HNO}_3\text{(aq)} + \text{NO(g)}$$

Summing their enthalpy changes gives the value we want to determine:

$$\begin{aligned}\Delta H^\circ_{rxn} &= \Delta H^\circ_1 + \Delta H^\circ_2 + \Delta H^\circ_3 + \Delta H^\circ_4 \\ &= (-99.6\,\text{kJ}) + (+285.8\,\text{kJ}) + (-414.8\,\text{kJ}) + (+90.2\,\text{kJ}) \\ &= -138.4\,\text{kJ}\end{aligned}$$

So the standard enthalpy change for this reaction is ΔH° = −138.4 kJ.

Note that this result was obtained by:

a. multiplying the of each product by its stoichiometric coefficient and summing those values,
b. multiplying the of each reactant by its stoichiometric coefficient and summing those values, and then
c. subtracting the result found in (2) from the result found in (1).

This is also the procedure in using the general equation, as shown.

? Exercise 7.4.8

Calculate the heat of combustion of 1 mole of ethanol, $\text{C}_2\text{H}_5\text{OH(l)}$, when $\text{H}_2\text{O(l)}$ and $\text{CO}_2\text{(g)}$ are formed. Use the following enthalpies of formation: $\text{C}_2\text{H}_5\text{OH(l)}$, −278 kJ/mol; $\text{H}_2\text{O(l)}$, −286 kJ/mol; and $\text{CO}_2\text{(g)}$, −394 kJ/mol.

Answer

−1368 kJ/mol

Footnotes

- 3For more on algal fuel, see http://www.theguardian.com/environme...n-fuel-problem.

This page titled 7.4: Enthalpy is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **5.3: Enthalpy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

7.5: Strengths of Ionic and Covalent Bonds

Learning Objectives

By the end of this section, you will be able to:

- Describe the energetics of covalent and ionic bond formation and breakage
- Use the Born-Haber cycle to compute lattice energies for ionic compounds
- Use average covalent bond energies to estimate enthalpies of reaction

A bond's strength describes how strongly each atom is joined to another atom, and therefore how much energy is required to break the bond between the two atoms. In this section, you will learn about the bond strength of covalent bonds, and then compare that to the strength of ionic bonds, which is related to the lattice energy of a compound.

Bond Strength: Covalent Bonds

Stable molecules exist because covalent bonds hold the atoms together. We measure the strength of a covalent bond by the energy required to break it, that is, the energy necessary to separate the bonded atoms. Separating any pair of bonded atoms requires energy. The stronger a bond, the greater the energy required to break it.

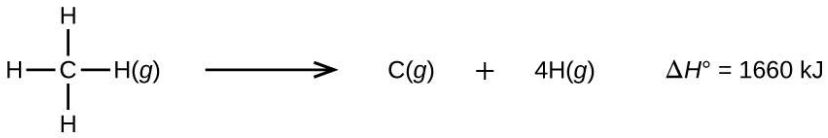
The energy required to break a specific covalent bond in one mole of gaseous molecules is called the **bond energy** or the **bond dissociation energy**. The bond energy for a diatomic molecule, $\mathrm{D_{X–Y}}$, is defined as the **standard enthalpy change** for the endothermic reaction:

$$\mathrm{XY(g) \longrightarrow X(g) + Y(g)} \qquad \mathrm{D_{\text{X}–\text{Y}}} = \Delta H^\circ$$

For example, the bond energy of the pure covalent $\mathrm{H–H}$ bond, $\mathrm{D_{H–H}}$, is 436 kJ per mole of $\mathrm{H–H}$ bonds broken:

$$\mathrm{H_2(g) \longrightarrow 2\,H(g)} \qquad D_{\mathrm{H–H}} = \Delta H^\circ = 436\ \mathrm{kJ}$$

Molecules with three or more atoms have two or more bonds. The sum of all bond energies in such a molecule is equal to the standard enthalpy change for the endothermic reaction that breaks all the bonds in the molecule. For example, the sum of the four $\mathrm{C–H}$ bond energies in $\mathrm{CH_4}$, 1660 kJ, is equal to the standard enthalpy change of the reaction:



The average $\mathrm{C–H}$ bond energy, $\mathrm{D_{C–H}}$, is 1660/4 = 415 kJ/mol because there are four moles of $\mathrm{C–H}$ bonds broken per mole of the reaction. Although the four $\mathrm{C–H}$ bonds are equivalent in the original molecule, they do not each require the same energy to break; once the first bond is broken (which requires 439 kJ/mol), the remaining bonds are easier to break. The 415 kJ/mol value is the average, not the exact value required to break any one bond.

The strength of a bond between two atoms increases as the number of electron pairs in the bond increases. Generally, as the bond strength increases, the bond length decreases. Thus, we find that triple bonds are stronger and shorter than double bonds between the same two atoms; likewise, double bonds are stronger and shorter than single bonds between the same two atoms. Average bond energies for some common bonds appear in Table 7.5.1, and a comparison of bond lengths and bond strengths for some common bonds appears in Table 7.5.2. When one atom bonds to various atoms in a group, the bond strength typically decreases as we move down the group. For example, $\mathrm{C–F}$ is 439 kJ/mol, $\mathrm{C–Cl}$ is 330 kJ/mol, and $\mathrm{C–Br}$ is 275 kJ/mol.

Table 7.5.1: Bond Energies (kJ/mol)

Bond	Bond Energy	Bond	Bond Energy	Bond	Bond Energy
H–H	436	C–S	260	F–Cl	255
H–C	415	C–Cl	330	F–Br	235
H–N	390	C–Br	275	Si–Si	230

Bond	Bond Energy	Bond	Bond Energy	Bond	Bond Energy
H–O	464	C–I	240	Si–P	215
H–F	569	N–N	160	Si–S	225
H–Si	395	N=N	418	Si–Cl	359
H–P	320	N≡N	946	Si–Br	290
H–S	340	N–O	200	Si–I	215
H–Cl	432	N–F	270	P–P	215
H–Br	370	N–P	210	P–S	230
H–I	295	N–Cl	200	P–Cl	330
C–C	345	N–Br	245	P–Br	270
C=C	611	O–O	140	P–I	215
C≡C	837	O=O	498	S–S	215
C–N	290	O–F	160	S–Cl	250
C=N	615	O–Si	370	S–Br	215
C≡N	891	O–P	350	Cl–Cl	243
C–O	350	O–Cl	205	Cl–Br	220
C=O	741	O–I	200	Cl–I	210
C≡O	1080	F–F	160	Br–Br	190
C–F	439	F–Si	540	Br–I	180
C–Si	360	F–P	489	I–I	150
C–P	265	F–S	285		

Table 7.5.2: Average Bond Lengths and Bond Energies for Some Common Bonds

Bond	Bond Length (Å)	Bond Energy (kJ/mol)
C–C	1.54	345
C=C	1.34	611
C≡C	1.20	837
C–N	1.43	290
C=N	1.38	615
C≡N	1.16	891
C–O	1.43	350
C=O	1.23	741
C≡O	1.13	1080

We can use bond energies to calculate approximate enthalpy changes for reactions where enthalpies of formation are not available. Calculations of this type will also tell us whether a reaction is exothermic or endothermic. An exothermic reaction (ΔH negative, heat produced) results when the bonds in the products are stronger than the bonds in the reactants. An endothermic reaction (ΔH positive, heat absorbed) results when the bonds in the products are weaker than those in the reactants.

The enthalpy change, ΔH, for a chemical reaction is approximately equal to the sum of the energy required to break all bonds in the reactants (energy "in", positive sign) plus the energy released when all bonds are formed in the products (energy "out," negative sign). This can be expressed mathematically in the following way:

$$\Delta H = \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}}$$

In this expression, the symbol $\sum$ means "the sum of" and D represents the bond energy in kilojoules per mole, which is always a positive number. The bond energy is obtained from a table (like Table 7.5.2) and will depend on whether the particular bond is a single, double, or triple bond. Thus, in calculating enthalpies in this manner, it is important that we consider the bonding in all reactants and products. Because D values are typically averages for one type of bond in many different molecules, this calculation provides a rough estimate, not an exact value, for the enthalpy of reaction.

Consider the following reaction:

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)}$$

or

$$\mathrm{H{-}H(g) + Cl{-}Cl(g) \longrightarrow 2\,H{-}Cl(g)}$$

To form two moles of HCl, one mole of H– H bonds and one mole of Cl– Cl bonds must be broken. The energy required to break these bonds is the sum of the bond energy of the H– H bond (436 kJ/mol) and the Cl– Cl bond (243 kJ/mol). During the reaction, two moles of H– Cl bonds are formed (bond energy = $432\ \text{kJ/mol}$), releasing $2 \times 432\ \text{kJ}$; or 864 kJ. Because the bonds in the products are stronger than those in the reactants, the reaction releases more energy than it consumes:

$$\begin{aligned}\Delta H &= \sum D_{\text{bonds broken}} - \sum D_{\text{bonds formed}} \\ &= [D_{\mathrm{H{-}H}} + D_{\mathrm{Cl{-}Cl}}] - 2D_{\mathrm{H{-}Cl}} \\ &= [436 + 243] - 2(432) \\ &= -185\ \text{kJ}\end{aligned}$$

This excess energy is released as heat, so the reaction is exothermic. Appendix G gives a value for the standard molar enthalpy of formation of HCl(g), ΔH_f^o of –92.307 kJ/mol. Twice that value is –184.6 kJ, which agrees well with the answer obtained earlier for the formation of two moles of HCl.

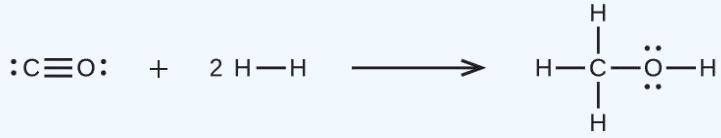
Example 7.5.1: Using Bond Energies to Calculate Approximate Enthalpy Changes

Methanol, CH_3OH, may be an excellent alternative fuel. The high-temperature reaction of steam and carbon produces a mixture of the gases carbon monoxide, CO, and hydrogen, H_2, from which methanol can be produced. Using the bond energies in Table 7.5.2, calculate the approximate enthalpy change, ΔH, for the reaction here:

$$\mathrm{CO(g) + 2\,H_2(g) \longrightarrow CH_3OH(g)}$$

Solution

First, we need to write the Lewis structures of the reactants and the products:



From this, we see that ΔH for this reaction involves the energy required to break a C–O triple bond and two H–H single bonds, as well as the energy produced by the formation of three C–H single bonds, a C–O single bond, and an O–H single bond. We can express this as follows:

$$\begin{aligned}\Delta H &= \Sigma D_{\text{bonds broken}} - \Sigma D_{\text{bonds formed}} \\ &= [D_{C\equiv O} + 2\,(D_{H-H})] - [3\,(D_{C-H}) + D_{C-O} + D_{O-H}]\end{aligned}$$

Using the bond energy values in Table 7.5.2, we obtain:

$$\begin{aligned}\Delta H &= [1080 + 2(436)] - [3(415) + 350 + 464] \\ &= -107 kJ\end{aligned}$$

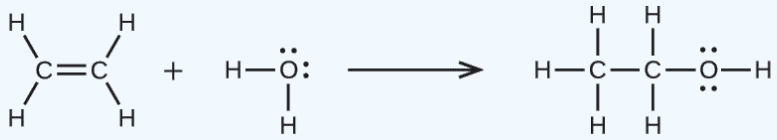
We can compare this value to the value calculated based on data from Appendix G:

$$\begin{aligned}\Delta H &= \left[\Delta H_f^\circ CH_3OH(g)\right] - \left[\Delta H_f^\circ CO(g) + 2 \times \Delta H_f^\circ H_2\right] \\ &= [-201.0] - [-110.52 + 2 \times 0] \\ &= -90.5 kJ\end{aligned}$$

Note that there is a fairly significant gap between the values calculated using the two different methods. This occurs because D values are the *average* of different bond strengths; therefore, they often give only rough agreement with other data.

? Exercise 7.5.1

Ethyl alcohol, CH_3CH_2OH, was one of the first organic chemicals deliberately synthesized by humans. It has many uses in industry, and it is the alcohol contained in alcoholic beverages. It can be obtained by the fermentation of sugar or synthesized by the hydration of ethylene in the following reaction:



Using the bond energies in Table 7.5.2, calculate an approximate enthalpy change, ΔH, for this reaction.

Answer

–35 kJ

Ionic Bond Strength and Lattice Energy

An ionic compound is stable because of the electrostatic attraction between its positive and negative ions. The lattice energy of a compound is a measure of the strength of this attraction. The **lattice energy** ($\Delta H_{lattice}$) of an ionic compound is defined as the energy required to separate one mole of the solid into its component gaseous ions. For the ionic solid MX, the lattice energy is the enthalpy change of the process:

$$MX(s) \longrightarrow M^{n+}(g) + X^{n-}(g) \quad \Delta H_{\text{lattice}}$$

Note that we are using the convention where the ionic solid is separated into ions, so our lattice energies will be *endothermic* (positive values). Some texts use the equivalent but opposite convention, defining lattice energy as the energy released when separate ions combine to form a lattice and giving negative (exothermic) values. Thus, if you are looking up lattice energies in another reference, be certain to check which definition is being used. In both cases, a larger magnitude for lattice energy indicates a more stable ionic compound. For sodium chloride, $\Delta H_{\text{lattice}}$ = 769 kJ. Thus, it requires 769 kJ to separate one mole of solid NaCl into gaseous Na^+ and Cl^- ions. When one mole each of gaseous Na^+ and Cl^- ions form solid NaCl, 769 kJ of heat is released.

The lattice energy $\Delta H_{\text{lattice}}$ of an ionic crystal can be expressed by the following equation (derived from Coulomb's law, governing the forces between electric charges):

$$\Delta H_{\text{lattice}} = \frac{C\,(Z^+)\,(Z^-)}{R_0}$$

in which C is a constant that depends on the type of crystal structure; Z^+ and Z^- are the charges on the ions; and R_o is the interionic distance (the sum of the radii of the positive and negative ions). Thus, the lattice energy of an ionic crystal increases rapidly as the charges of the ions increase and the sizes of the ions decrease. When all other parameters are kept constant, doubling the charge of both the cation and anion quadruples the lattice energy. For example, the lattice energy of LiF (Z^+ and Z^- = 1) is 1023 kJ/mol, whereas that of MgO (Z^+ and Z^- = 2) is 3900 kJ/mol (R_o is nearly the same—about 200 pm for both compounds).

Different interatomic distances produce different lattice energies. For example, we can compare the lattice energy of MgF_2 (2957 kJ/mol) to that of MgI_2 (2327 kJ/mol) to observe the effect on lattice energy of the smaller ionic size of F^- as compared to I^-.

Example 7.5.2: Lattice Energy Comparisons

The precious gem ruby is aluminum oxide, Al_2O_3, containing traces of Cr^{3+}. The compound Al_2Se_3 is used in the fabrication of some semiconductor devices. Which has the larger lattice energy, Al_2O_3 or Al_2Se_3?

Solution

In these two ionic compounds, the charges Z^+ and Z^- are the same, so the difference in lattice energy will depend upon R_o. The O^{2-} ion is smaller than the Se^{2-} ion. Thus, Al_2O_3 would have a shorter interionic distance than Al_2Se_3, and Al_2O_3 would have the larger lattice energy.

Exercise 7.5.1

Zinc oxide, ZnO, is a very effective sunscreen. How would the lattice energy of ZnO compare to that of NaCl?

Answer

ZnO would have the larger lattice energy because the Z values of both the cation and the anion in ZnO are greater, and the interionic distance of ZnO is smaller than that of NaCl.

The Born-Haber Cycle

It is not possible to measure lattice energies directly. However, the lattice energy can be calculated using the equation given in the previous section or by using a thermochemical cycle. The **Born-Haber cycle** is an application of Hess's law that breaks down the formation of an ionic solid into a series of individual steps:

- ΔH_f° the standard enthalpy of formation of the compound
- IE, the ionization energy of the metal
- EA, the electron affinity of the nonmetal
- ΔH_s° the enthalpy of sublimation of the metal
- D, the bond dissociation energy of the nonmetal
- $\Delta H_{lattice}$, the lattice energy of the compound

Figure 7.5.1: diagrams the Born-Haber cycle for the formation of solid cesium fluoride.

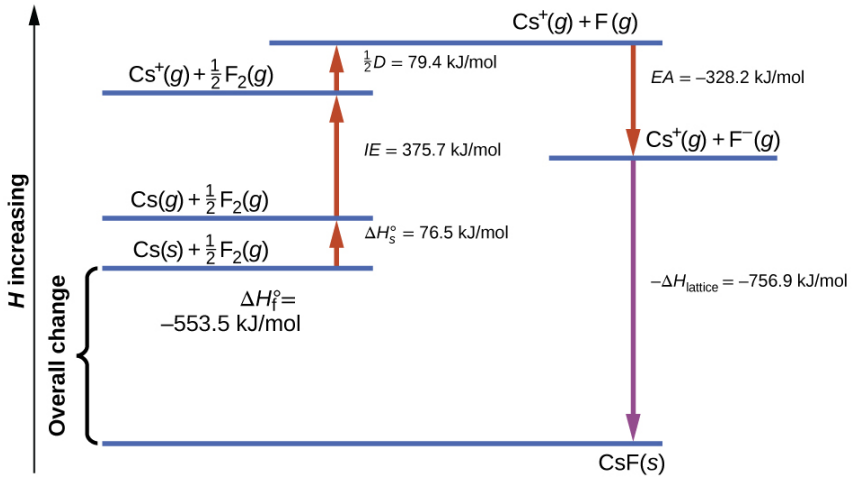


Figure 7.5.1: The Born-Haber cycle shows the relative energies of each step involved in the formation of an ionic solid from the necessary elements in their reference states.

We begin with the elements in their most common states, $Cs(s)$ and $F_2(g)$. The represents the conversion of solid cesium into a gas, and then the ionization energy converts the gaseous cesium atoms into cations. In the next step, we account for the energy required to break the F–F bond to produce fluorine atoms. Converting one mole of fluorine atoms into fluoride ions is an

exothermic process, so this step gives off energy (the electron affinity) and is shown as decreasing along the *y*-axis. We now have one mole of Cs cations and one mole of F anions. These ions combine to produce solid cesium fluoride. The enthalpy change in this step is the negative of the lattice energy, so it is also an exothermic quantity. The total energy involved in this conversion is equal to the experimentally determined enthalpy of formation, ΔH_f° of the compound from its elements. In this case, the overall change is exothermic.

Hess's law can also be used to show the relationship between the enthalpies of the individual steps and the enthalpy of formation. Table $7.5.3$ shows this for fluoride, CsF.

Table 7.5.3

Enthalpy of sublimation of $\mathrm{Cs(s)}$	$\mathrm{Cs(s)} \longrightarrow \mathrm{Cs(g)}$	$\Delta H = \Delta H_s^\circ = 76.5\,\mathrm{kJ/mol}$
One-half of the bond energy of $\mathrm{F_2}$	$\frac{1}{2}\mathrm{F_2(g)} \longrightarrow \mathrm{F(g)}$	$\Delta H = \frac{1}{2}D = 79.4\,\mathrm{kJ/mol}$
Ionization energy of $\mathrm{Cs(g)}$	$\mathrm{Cs(g)} \longrightarrow \mathrm{Cs^+(g) + e^-}$	$\Delta H = IE = 375.7\,\mathrm{kJ/mol}$
Electron affinity of F	$\mathrm{F(g) + e^-} \longrightarrow \mathrm{F^-(g)}$	$\Delta H = EA = -328.2\,\mathrm{kJ/mol}$
Negative of the lattice energy of $\mathrm{CsF(s)}$	$\mathrm{Cs^+(g) + F^-(g)} \longrightarrow \mathrm{CsF(s)}$	$\Delta H = -\Delta H_{\text{lattice}} = ?$
Enthalpy of formation of $\mathrm{CsF(s)}$, add steps 1–5	$\Delta H = \Delta H_f^\circ = \Delta H_s^\circ + \frac{1}{2}D + IE + (EA) + (-\Delta H_{\text{lattice}})$ $\mathrm{Cs(s)} + \frac{1}{2}\mathrm{F_2(g)} \longrightarrow \mathrm{CsF(s)}$	$\Delta H = -553.5\,\mathrm{kJ/mol}$

Thus, the lattice energy can be calculated from other values. For cesium fluoride, using this data, the lattice energy is:

$$\Delta H_{\text{lattice}} = 76.5 + 79.4 + 375.7 + (-328.2) - (-553.5) = 756.9\,\mathrm{kJ/mol}$$

The Born-Haber cycle may also be used to calculate any one of the other quantities in the equation for lattice energy, provided that the remainder is known. For example, if the relevant enthalpy of sublimation ΔH_s°, ionization energy (IE), bond dissociation enthalpy (D), lattice energy $\Delta H_{\text{lattice}}$, and standard enthalpy of formation are known, the Born-Haber cycle can be used to determine the electron affinity of an atom.

Lattice energies calculated for ionic compounds are typically much higher than bond dissociation energies measured for covalent bonds. Whereas lattice energies typically fall in the range of 600–4000 kJ/mol (some even higher), covalent bond dissociation energies are typically between 150–400 kJ/mol for single bonds. Keep in mind, however, that these are not directly comparable values. For ionic compounds, lattice energies are associated with many interactions, as cations and anions pack together in an extended lattice. For covalent bonds, the bond dissociation energy is associated with the interaction of just two atoms.

This page titled 7.5: Strengths of Ionic and Covalent Bonds is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **7.5: Strengths of Ionic and Covalent Bonds** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

7.6: Key Terms

bomb calorimeter
device designed to measure the energy change for processes occurring under conditions of constant volume; commonly used for reactions involving solid and gaseous reactants or products

bond energy
(also, bond dissociation energy) energy required to break a covalent bond in a gaseous substance

Born-Haber cycle
thermochemical cycle relating the various energetic steps involved in the formation of an ionic solid from the relevant elements

calorie (cal)
unit of heat or other energy; the amount of energy required to raise 1 gram of water by 1 degree Celsius; 1 cal is defined as 4.184 J

calorimeter
device used to measure the amount of heat absorbed or released in a chemical or physical process

calorimetry
process of measuring the amount of heat involved in a chemical or physical process

chemical thermodynamics
area of science that deals with the relationships between heat, work, and all forms of energy associated with chemical and physical processes

endothermic process
chemical reaction or physical change that absorbs heat

energy
capacity to supply heat or do work

enthalpy (*H*)
sum of a system’s internal energy and the mathematical product of its pressure and volume

enthalpy change (ΔH)
heat released or absorbed by a system under constant pressure during a chemical or physical process

exothermic process
chemical reaction or physical change that releases heat

expansion work (pressure-volume work)
work done as a system expands or contracts against external pressure

first law of thermodynamics
internal energy of a system changes due to heat flow in or out of the system or work done on or by the system

heat (*q*)
transfer of thermal energy between two bodies

heat capacity (*C*)
extensive property of a body of matter that represents the quantity of heat required to increase its temperature by 1 degree Celsius (or 1 kelvin)

Hess’s law

if a process can be represented as the sum of several steps, the enthalpy change of the process equals the sum of the enthalpy changes of the steps

hydrocarbon

compound composed only of hydrogen and carbon; the major component of fossil fuels

internal energy (*U*)

total of all possible kinds of energy present in a substance or substances

joule (J)

SI unit of energy; 1 joule is the kinetic energy of an object with a mass of 2 kilograms moving with a velocity of 1 meter per second, 1 J = 1 kg m^2/s and 4.184 J = 1 cal

kinetic energy

energy of a moving body, in joules, equal to $\frac{1}{2}mv^2$ (where m = mass and v = velocity)

lattice energy ($\Delta H_{\text{lattice}}$)

energy required to separate one mole of an ionic solid into its component gaseous ions

nutritional calorie (Calorie)

unit used for quantifying energy provided by digestion of foods, defined as 1000 cal or 1 kcal

potential energy

energy of a particle or system of particles derived from relative position, composition, or condition

specific heat capacity (*c*)

intensive property of a substance that represents the quantity of heat required to raise the temperature of 1 gram of the substance by 1 degree Celsius (or 1 kelvin)

standard enthalpy of combustion (ΔH_c°)

heat released when one mole of a compound undergoes complete combustion under standard conditions

standard enthalpy of formation (ΔH_f°)

enthalpy change of a chemical reaction in which 1 mole of a pure substance is formed from its elements in their most stable states under standard state conditions

standard state

set of physical conditions as accepted as common reference conditions for reporting thermodynamic properties; 1 bar of pressure, and solutions at 1 molar concentrations, usually at a temperature of 298.15 K

state function

property depending only on the state of a system, and not the path taken to reach that state

surroundings

all matter other than the system being studied

system

portion of matter undergoing a chemical or physical change being studied

temperature

intensive property of matter that is a quantitative measure of “hotness” and “coldness”

thermal energy

kinetic energy associated with the random motion of atoms and molecules

thermochemistry

study of measuring the amount of heat absorbed or released during a chemical reaction or a physical change

work (*w*)

energy transfer due to changes in external, macroscopic variables such as pressure and volume; or causing matter to move against an opposing force

---

This page titled 7.6: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

7.7: Key Equations

$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$

$\Delta U = q +$
$\Delta H_{\text{reaction}}{}^{\circ} = \sum n \times \Delta H_{\text{f}}{}^{\circ}(\text{products}) - \sum n \times \Delta H_{\text{f}}{}^{\circ}(\text{reactants})$

Bond energy for a diatomic molecule: $\text{XY}(g) \longrightarrow \text{X}(g) + \text{Y}(g)$	$\text{D}_{\text{X–Y}} = \Delta H^{\circ}$
Enthalpy change: $\Delta H = \Sigma \text{D}_{\text{bonds broken}} - \Sigma \text{D}_{\text{bonds formed}}$	
Lattice energy for a solid MX: $\text{MX}(s) \longrightarrow \text{M}^{n}+(g) + \text{X}^{n-}(g)$	
Lattice energy for an ionic crystal: $\Delta H_{\text{lattice}} = \frac{\text{C}(\text{Z}^{+})(\text{Z}^{-})}{}$	

This page titled 7.7: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

7.8: Summary

Energy is the capacity to supply heat or do work (applying a force to move matter). Kinetic energy (KE) is the energy of motion; potential energy is energy due to relative position, composition, or condition. When energy is converted from one form into another, energy is neither created nor destroyed (law of conservation of energy or first law of thermodynamics).

The thermal energy of matter is due to the kinetic energies of its constituent atoms or molecules. Temperature is an intensive property of matter reflecting hotness or coldness that increases as the average kinetic energy increases. Heat is the transfer of thermal energy between objects at different temperatures. Chemical and physical processes can absorb heat (endothermic) or release heat (exothermic). The SI unit of energy, heat, and work is the joule (J).

Specific heat and heat capacity are measures of the energy needed to change the temperature of a substance or object. The amount of heat absorbed or released by a substance depends directly on the type of substance, its mass, and the temperature change it undergoes.

Calorimetry is used to measure the amount of thermal energy transferred in a chemical or physical process. This requires careful measurement of the temperature change that occurs during the process and the masses of the system and surroundings. These measured quantities are then used to compute the amount of heat produced or consumed in the process using known mathematical relations.

Calorimeters are designed to minimize energy exchange between their contents and the external environment. They range from simple coffee cup calorimeters used by introductory chemistry students to sophisticated bomb calorimeters used to determine the energy content of food.

If a chemical change is carried out at constant pressure and the only work done is caused by expansion or contraction, q for the change is called the enthalpy change with the symbol ΔH, or ΔH° for reactions occurring under standard state conditions at 298 K. The value of ΔH for a reaction in one direction is equal in magnitude, but opposite in sign, to ΔH for the reaction in the opposite direction, and ΔH is directly proportional to the quantity of reactants and products. The standard enthalpy of formation, $\Delta H_\mathrm{f}^\circ$, is the enthalpy change accompanying the formation of 1 mole of a substance from the elements in their most stable states at 1 bar and 298.15 K. If the enthalpies of formation are available for the reactants and products of a reaction, the enthalpy change can be calculated using Hess's law: If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps.

The strength of a covalent bond is measured by its bond dissociation energy, that is, the amount of energy required to break that particular bond in a mole of molecules. Multiple bonds are stronger than single bonds between the same atoms. The enthalpy of a reaction can be estimated based on the energy input required to break bonds and the energy released when new bonds are formed. For ionic bonds, the lattice energy is the energy required to separate one mole of a compound into its gas phase ions. Lattice energy increases for ions with higher charges and shorter distances between ions. Lattice energies are often calculated using the Born-Haber cycle, a thermochemical cycle including all of the energetic steps involved in converting elements into an ionic compound.

This page titled 7.8: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

7.9: Exercises

6.1: Energy Basics

Q6.1.1

A burning match and a bonfire may have the same temperature, yet you would not sit around a burning match on a fall evening to stay warm. Why not?

S6.1.1

The temperature of 1 gram of burning wood is approximately the same for both a match and a bonfire. This is an intensive property and depends on the material (wood). However, the overall amount of produced heat depends on the amount of material; this is an extensive property. The amount of wood in a bonfire is much greater than that in a match; the total amount of produced heat is also much greater, which is why we can sit around a bonfire to stay warm, but a match would not provide enough heat to keep us from getting cold.

Q6.1.2

Prepare a table identifying several energy transitions that take place during the typical operation of an automobile.

Q6.1.3

Explain the difference between heat capacity and specific heat of a substance.

S6.1.3

Heat capacity refers to the heat required to raise the temperature of the mass of the substance 1 degree; specific heat refers to the heat required to raise the temperature of 1 gram of the substance 1 degree. Thus, heat capacity is an extensive property, and specific heat is an intensive one.

Q6.1.4

Calculate the heat capacity, in joules and in calories per degree, of the following:

a. 28.4 g of water
b. 1.00 oz of lead

Q6.1.5

Calculate the heat capacity, in joules and in calories per degree, of the following:

a. 45.8 g of nitrogen gas
b. 1.00 pound of aluminum metal

S5.1.5

a. 47.6 J/°C; 11.38 cal $°C^{-1}$;
b. 407 J/°C; 97.3 cal $°C^{-1}$

Q5.1.6

How much heat, in joules and in calories, must be added to a 75.0–g iron block with a specific heat of 0.449 J/g °C to increase its temperature from 25 °C to its melting temperature of 1535 °C?

S6.1.6

$q = mC\Delta^{\circ}T$

$q = (75.0g) \times (\frac{0.449\ J}{g\ ^{\circ}C}) \times (1,510^{\circ}K) = 50,800J$

50,800J ; 12,200cal

Q6.1.7

How much heat, in joules and in calories, is required to heat a 28.4-g (1-oz) ice cube from −23.0 °C to −1.0 °C?

S6.1.7

1310 J; 313 cal

Q6.1.8

How much would the temperature of 275 g of water increase if 36.5 kJ of heat were added?

S6.1.8

ΔT° = 31.7° C

Q6.1.9

If 14.5 kJ of heat were added to 485 g of liquid water, how much would its temperature increase?

S5.1.9

7.15 °C

Q5.1.10

A piece of unknown substance weighs 44.7 g and requires 2110 J to increase its temperature from 23.2 °C to 89.6 °C.

a. What is the specific heat of the substance?
b. If it is one of the substances found in Table, what is its likely identity?

S5.1.10

a.) Solve for the specific heat C and compare the values with the chart

$q = mC\Delta^{\circ}T$

$2110J = (44.7\ g)(C)(66.4^{\circ}C)$

$$C = \frac{2110\ J}{2970\ g^{\circ}C}$$

$$C = \frac{0.711\ J}{g^{\circ}C}$$

b.) Silicon

Q5.1.11

A piece of unknown solid substance weighs 437.2 g, and requires 8460 J to increase its temperature from 19.3 °C to 68.9 °C.

a. What is the specific heat of the substance?
b. If it is one of the substances found in Table, what is its likely identity?

S5.1.11

a. 0.390 J/g °C;
b. Copper is a likely candidate.

Q5.1.12

An aluminum kettle weighs 1.05 kg.

a. What is the heat capacity of the kettle?
b. How much heat is required to increase the temperature of this kettle from 23.0 °C to 99.0 °C?
c. How much heat is required to heat this kettle from 23.0 °C to 99.0 °C if it contains 1.25 L of water (density of 0.997 g/mL and a specific heat of 4.184 J/g °C)?

Q5.1.13

Most people find waterbeds uncomfortable unless the water temperature is maintained at about 85 °F. Unless it is heated, a waterbed that contains 892 L of water cools from 85 °F to 72 °F in 24 hours. Estimate the amount of electrical energy required over 24 hours, in kWh, to keep the bed from cooling. Note that 1 kilowatt-hour (kWh) = 3.6×10^6 J, and assume that the density of water is 1.0 g/mL (independent of temperature). What other assumptions did you make? How did they affect your calculated result (i.e., were they likely to yield “positive” or “negative” errors)?

S5.1.13

We assume that the density of water is 1.0 g/cm^3(1 g/mL) and that it takes as much energy to keep the water at 85 °F as to heat it from 72 °F to 85 °F. We also assume that only the water is going to be heated. Energy required = 7.47 kWh

5.2: Calorimetry

Q5.2.1

A 500-mL bottle of water at room temperature and a 2-L bottle of water at the same temperature were placed in a refrigerator. After 30 minutes, the 500-mL bottle of water had cooled to the temperature of the refrigerator. An hour later, the 2-L of water had cooled to the same temperature. When asked which sample of water lost the most heat, one student replied that both bottles lost the same amount of heat because they started at the same temperature and finished at the same temperature. A second student thought that the 2-L bottle of water lost more heat because there was more water. A third student believed that the 500-mL bottle of water lost more heat because it cooled more quickly. A fourth student thought that it was not possible to tell because we do not know the initial temperature and the final temperature of the water. Indicate which of these answers is correct and describe the error in each of the other answers.

Q5.2.2

Would the amount of heat measured for the reaction in Example be greater, lesser, or remain the same if we used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

S5.2.2

lesser; more heat would be lost to the coffee cup and the environment and so ΔT for the water would be lesser and the calculated q would be lesser

Q5.2.3

Would the amount of heat absorbed by the dissolution in Example appear greater, lesser, or remain the same if the experimenter used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

Q5.2.4

Would the amount of heat absorbed by the dissolution in Example appear greater, lesser, or remain the same if the heat capacity of the calorimeter were taken into account? Explain your answer.

S5.2.4

greater, since taking the calorimeter’s heat capacity into account will compensate for the thermal energy transferred to the solution from the calorimeter; this approach includes the calorimeter itself, along with the solution, as “surroundings”: $q_{rxn} = -(q_{solution} + q_{calorimeter})$; since both $q_{solution}$ and $q_{calorimeter}$ are negative, including the latter term (q_{rxn}) will yield a greater value for the heat of the dissolution

Q5.2.5

How many milliliters of water at 23 °C with a density of 1.00 g/mL must be mixed with 180 mL (about 6 oz) of coffee at 95 °C so that the resulting combination will have a temperature of 60 °C? Assume that coffee and water have the same density and the same specific heat.

Q5.2.6

How much will the temperature of a cup (180 g) of coffee at 95 °C be reduced when a 45 g silver spoon (specific heat 0.24 J/g °C) at 25 °C is placed in the coffee and the two are allowed to reach the same temperature? Assume that the coffee has the same density

and specific heat as water.

S5.2.6

The temperature of the coffee will drop 1 degree.

Q5.2.7

A 45-g aluminum spoon (specific heat 0.88 J/g °C) at 24 °C is placed in 180 mL (180 g) of coffee at 85 °C and the temperature of the two become equal.

a. What is the final temperature when the two become equal? Assume that coffee has the same specific heat as water.
b. The first time a student solved this problem she got an answer of 88 °C. Explain why this is clearly an incorrect answer.

Q5.2.8

The temperature of the cooling water as it leaves the hot engine of an automobile is 240 °F. After it passes through the radiator it has a temperature of 175 °F. Calculate the amount of heat transferred from the engine to the surroundings by one gallon of water with a specific heat of 4.184 J/g °C.

S5.2.8

$5.7 \times 10^2\ kJ$

Q5.2.8

A 70.0-g piece of metal at 80.0 °C is placed in 100 g of water at 22.0 °C contained in a calorimeter like that shown in Figure. The metal and water come to the same temperature at 24.6 °C. How much heat did the metal give up to the water? What is the specific heat of the metal?

Q5.2.9

If a reaction produces 1.506 kJ of heat, which is trapped in 30.0 g of water initially at 26.5 °C in a calorimeter like that in Figure, what is the resulting temperature of the water?

S5.2.9

38.5 °C

Q5.2.10

A 0.500-g sample of KCl is added to 50.0 g of water in a calorimeter (Figure). If the temperature decreases by 1.05 °C, what is the approximate amount of heat involved in the dissolution of the KCl, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

Q5.2.11

Dissolving 3.0 g of $CaCl_2(s)$ in 150.0 g of water in a calorimeter (Figure) at 22.4 °C causes the temperature to rise to 25.8 °C. What is the approximate amount of heat involved in the dissolution, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

S5.2.11

2.2 kJ; The heat produced shows that the reaction is exothermic.

Q5.2.12

When 50.0 g of 0.200 M NaCl(*aq*) at 24.1 °C is added to 100.0 g of 0.100 M $AgNO_3(aq)$ at 24.1 °C in a calorimeter, the temperature increases to 25.2 °C as AgCl(*s*) forms. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat in joules produced.

Q5.2.13

The addition of 3.15 g of $Ba(OH)_2 \cdot 8H_2O$ to a solution of 1.52 g of NH_4SCN in 100 g of water in a calorimeter caused the temperature to fall by 3.1 °C. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat absorbed by the reaction, which can be represented by the following equation:

$$Ba(OH)_2 \cdot 8H_2O_{(s)} + 2NH_4SCN_{(aq)} \rightarrow Ba(SCN)_{2(aq)} + 2NH_{3(aq)} + 10H_2O_{(l)} \quad (7.9.1)$$

S5.2.13

1.4 kJ

Q5.2.14

The reaction of 50 mL of acid and 50 mL of base described in Example increased the temperature of the solution by 6.9 degrees. How much would the temperature have increased if 100 mL of acid and 100 mL of base had been used in the same calorimeter starting at the same temperature of 22.0 °C? Explain your answer.

Q5.2.15

If the 3.21 g of NH_4NO_3 in Example were dissolved in 100.0 g of water under the same conditions, how much would the temperature change? Explain your answer.

S5.2.15

22.6. Since the mass and the heat capacity of the solution is approximately equal to that of the water, the two-fold increase in the amount of water leads to a two-fold decrease of the temperature change.

Q5.2.16

When 1.0 g of fructose, $C_6H_{12}O_6(s)$, a sugar commonly found in fruits, is burned in oxygen in a bomb calorimeter, the temperature of the calorimeter increases by 1.58 °C. If the heat capacity of the calorimeter and its contents is 9.90 kJ/°C, what is q for this combustion?

Q5.2.17

When a 0.740-g sample of trinitrotoluene (TNT), $C_7H_5N_2O_6$, is burned in a bomb calorimeter, the temperature increases from 23.4 °C to 26.9 °C. The heat capacity of the calorimeter is 534 J/°C, and it contains 675 mL of water. How much heat was produced by the combustion of the TNT sample?

S5.2.17

11.7 kJ

Q5.2.18

One method of generating electricity is by burning coal to heat water, which produces steam that drives an electric generator. To determine the rate at which coal is to be fed into the burner in this type of plant, the heat of combustion per ton of coal must be determined using a bomb calorimeter. When 1.00 g of coal is burned in a bomb calorimeter, the temperature increases by 1.48 °C. If the heat capacity of the calorimeter is 21.6 kJ/°C, determine the heat produced by combustion of a ton of coal (2.000×10^3 pounds).

Q5.2.19

The amount of fat recommended for someone with a daily diet of 2000 Calories is 65 g. What percent of the calories in this diet would be supplied by this amount of fat if the average number of Calories for fat is 9.1 Calories/g?

S5.2.19

30%

Q5.2.20

A teaspoon of the carbohydrate sucrose (common sugar) contains 16 Calories (16 kcal). What is the mass of one teaspoon of sucrose if the average number of Calories for carbohydrates is 4.1 Calories/g?

Q5.2.21

What is the maximum mass of carbohydrate in a 6-oz serving of diet soda that contains less than 1 Calorie per can if the average number of Calories for carbohydrates is 4.1 Calories/g?

S5.2.21

0.24 g

Q5.2.22

A pint of premium ice cream can contain 1100 Calories. What mass of fat, in grams and pounds, must be produced in the body to store an extra 1.1×10^3 Calories if the average number of Calories for fat is 9.1 Calories/g?

Q5.2.23

A serving of a breakfast cereal contains 3 g of protein, 18 g of carbohydrates, and 6 g of fat. What is the Calorie content of a serving of this cereal if the average number of Calories for fat is 9.1 Calories/g, for carbohydrates is 4.1 Calories/g, and for protein is 4.1 Calories/g?

S5.2.23

1.4×10^2 Calories

Q5.2.24

Which is the least expensive source of energy in kilojoules per dollar: a box of breakfast cereal that weighs 32 ounces and costs $4.23, or a liter of isooctane (density, 0.6919 g/mL) that costs $0.45? Compare the nutritional value of the cereal with the heat produced by combustion of the isooctane under standard conditions. A 1.0-ounce serving of the cereal provides 130 Calories.

5.3: Enthalpy

Q5.3.1

Explain how the heat measured in [link] differs from the enthalpy change for the exothermic reaction described by the following equation:

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l) \tag{7.9.2}$$

The enthalpy change of the indicated reaction is for exactly 1 mol HCL and 1 mol NaOH; the heat in the example is produced by 0.0500 mol HCl and 0.0500 mol NaOH. Using the data in the check your learning section of [link], calculate ΔH in kJ/mol of $AgNO_3(aq)$ for the reaction:

$$\mathrm{NaCl}(aq) + \mathrm{AgNO_3}(aq) \longrightarrow \mathrm{AgCl}(s) + \mathrm{NaNO_3}(aq) \tag{7.9.3}$$

Q5.3.2

Calculate the enthalpy of solution (ΔH for the dissolution) per mole of NH_4NO_3 under the conditions described in [link].

S5.3.2

25 kJ mol^{-1}

Q5.3.3

Calculate ΔH for the reaction described by the equation.

$$\mathrm{Ba(OH)_2 \cdot 8\,H_2O}(s) + 2\,\mathrm{NH_4SCN}(aq) \longrightarrow \mathrm{Ba(SCN)_2}(aq) + 2\,\mathrm{NH_3}(aq) + 10\,\mathrm{H_2O}(l)$$

Q5.3.4

Calculate the enthalpy of solution (ΔH for the dissolution) per mole of $CaCl_2$.

S5.3.4

81 kJ mol^{-1}

Q5.3.5

Although the gas used in an oxyacetylene torch is essentially pure acetylene, the heat produced by combustion of one mole of acetylene in such a torch is likely not equal to the enthalpy of combustion of acetylene listed in Table. Considering the conditions for which the tabulated data are reported, suggest an explanation.

Q5.3.6

How much heat is produced by burning 4.00 moles of acetylene under standard state conditions?

S5.3.6

5204.4 kJ

Q5.3.7

How much heat is produced by combustion of 125 g of methanol under standard state conditions?

Q5.3.8

How many moles of isooctane must be burned to produce 100 kJ of heat under standard state conditions?

S5.3.8

1.83×10^{-2} mol

Q5.3.9

What mass of carbon monoxide must be burned to produce 175 kJ of heat under standard state conditions?

Q5.3.10

When 2.50 g of methane burns in oxygen, 125 kJ of heat is produced. What is the enthalpy of combustion per mole of methane under these conditions?

S5.3.10

802 kJ mol^{-1}

Q5.3.11

How much heat is produced when 100 mL of 0.250 M HCl (density, 1.00 g/mL) and 200 mL of 0.150 M NaOH (density, 1.00 g/mL) are mixed?

$$\text{HCl}(aq) + \text{NaOH}(aq) \longrightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \quad \Delta H^\circ_{298} = -58 \text{ kJ} \tag{7.9.4}$$

If both solutions are at the same temperature and the heat capacity of the products is 4.19 J/g °C, how much will the temperature increase? What assumption did you make in your calculation?

Q5.3.12

A sample of 0.562 g of carbon is burned in oxygen in a bomb calorimeter, producing carbon dioxide. Assume both the reactants and products are under standard state conditions, and that the heat released is directly proportional to the enthalpy of combustion of graphite. The temperature of the calorimeter increases from 26.74 °C to 27.93 °C. What is the heat capacity of the calorimeter and its contents?

S5.3.12

15.5 kJ/ºC

Q5.3.13

Before the introduction of chlorofluorocarbons, sulfur dioxide (enthalpy of vaporization, 6.00 kcal/mol) was used in household refrigerators. What mass of SO_2 must be evaporated to remove as much heat as evaporation of 1.00 kg of CCl_2F_2 (enthalpy of vaporization is 17.4 kJ/mol)?

The vaporization reactions for SO_2 and CCl_2F_2 are

$\text{SO}_2(l) \longrightarrow \text{SO}_2(g)$ and $\text{CCl}_2\text{F}(l) \longrightarrow \text{CCl}_2\text{F}_2(g)$, respectively.

Q5.3.14

Homes may be heated by pumping hot water through radiators. What mass of water will provide the same amount of heat when cooled from 95.0 to 35.0 °C, as the heat provided when 100 g of steam is cooled from 110 °C to 100 °C.

S5.3.14

7.43 g

Q5.3.15

Which of the enthalpies of combustion in Table the table are also standard enthalpies of formation?

Q5.3.16

Does the standard enthalpy of formation of $H_2O(g)$ differ from $\Delta H°$ for the reaction $2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$?

S5.3.16

No.

Q5.3.17

Joseph Priestly prepared oxygen in 1774 by heating red mercury(II) oxide with sunlight focused through a lens. How much heat is required to decompose exactly 1 mole of red HgO(*s*) to Hg(*l*) and $O_2(g)$ under standard conditions?

Q5.3.18

How many kilojoules of heat will be released when exactly 1 mole of manganese, Mn, is burned to form $Mn_3O_4(s)$ at standard state conditions?

S5.3.18

459.6 kJ

Q5.3.19

How many kilojoules of heat will be released when exactly 1 mole of iron, Fe, is burned to form $Fe_2O_3(s)$ at standard state conditions?

Q5.3.20

The following sequence of reactions occurs in the commercial production of aqueous nitric acid:

$$4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(l) \quad \Delta H = -907\text{ kJ}$$

$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g) \quad \Delta H = -113\text{ kJ}$$

$$3\,NO_2 + H_2O(l) \longrightarrow 2\,HNO_2(aq) + NO(g) \quad \Delta H = -139\text{ kJ}$$

Determine the total energy change for the production of one mole of aqueous nitric acid by this process.

S5.3.20

495 kJ/mol

Q5.3.21

Both graphite and diamond burn.

$$C(s,\text{ diamond}) + O_2(g) \longrightarrow CO_2(g)$$

For the conversion of graphite to diamond:

$$C(s,\text{ graphite}) \longrightarrow C(s,\text{ diamond}) \quad \Delta H^\circ_{298} = 1.90\text{ kJ}$$

Which produces more heat, the combustion of graphite or the combustion of diamond?

Q5.3.22

From the molar heats of formation in Appendix G, determine how much heat is required to evaporate one mole of water: $H_2O(l) \longrightarrow H_2O(g)$

S5.3.22

44.01 kJ/mol

Q5.3.23

Which produces more heat?

$$\mathrm{Os}(s) \longrightarrow 2\,\mathrm{O}_2(g) \longrightarrow \mathrm{OsO}_4(s)$$

or

$$\mathrm{Os}(s) \longrightarrow 2\,\mathrm{O}_2(g) \longrightarrow \mathrm{OsO}_4(g)$$

for the phase change $\mathrm{OsO}_4(s) \longrightarrow \mathrm{OsO}_4(g) \quad \Delta H = 56.4\ \mathrm{kJ}$

Q5.3.24

Calculate ΔH°_{298} for the process

$$\mathrm{Sb}(s) + \frac{5}{2}\mathrm{Cl}_2(g) \longrightarrow \mathrm{SbCl}_5(g)$$

from the following information:

$$\mathrm{Sb}(s) + \frac{3}{2}\mathrm{Cl}_2(g) \longrightarrow \mathrm{SbCl}_3(g) \quad \Delta H^\circ_{298} = -314\ \mathrm{kJ}$$

$$\mathrm{SbCl}_3(s) + \mathrm{Cl}_2(g) \longrightarrow \mathrm{SbCl}_5(g) \quad \Delta H^\circ_{298} = -80\ \mathrm{kJ}$$

S5.3.24

394 kJ

Q5.3.25

Calculate ΔH°_{298} for the process $\mathrm{Zn}(s) + \mathrm{S}(s) + 2\,\mathrm{O}_2(g) \longrightarrow \mathrm{ZnSO}_4(s)$

from the following information:

$$\mathrm{Zn}(s) + \mathrm{S}(s) \longrightarrow \mathrm{ZnS}(s) \quad \Delta H^\circ_{298} = -206.0\ \mathrm{kJ}$$

$$\mathrm{ZnS}(s) + 2\,\mathrm{O}_2(g) \longrightarrow \mathrm{ZnSO}_4(s) \quad \Delta H^\circ_{298} = -776.8\ \mathrm{kJ}$$

Q5.3.26

Calculate ΔH for the process

$$\mathrm{Hg}_2\mathrm{Cl}_2(s) \longrightarrow 2\,\mathrm{Hg}(l) + \mathrm{Cl}_2(g)$$

from the following information:

$\mathrm{Hg}(l) + \mathrm{Cl}_2(g) \longrightarrow \mathrm{HgCl}_2(s) \quad \Delta H = -224\ \mathrm{kJ}$

$\mathrm{Hg}(l) + \mathrm{HgCl}_2(s) \longrightarrow \mathrm{Hg}_2\mathrm{Cl}_2(s) \quad \Delta H = -41.2\ \mathrm{kJ}$

S5.3.26

265 kJ

Q5.3.27

Calculate ΔH°_{298} for the process

$$\mathrm{Co}_3\mathrm{O}_4(s) \longrightarrow 3\,\mathrm{Co}(s) + 2\,\mathrm{O}_2(g)$$

from the following information:

$$\mathrm{Co}(s) + \frac{1}{2}\mathrm{O}_2(g) \longrightarrow \mathrm{CoO}(s) \quad \Delta H^\circ_{298} = -237.9\ \mathrm{kJ}$$

$$3\,\mathrm{Co}(s) + \mathrm{O}_2(g) \longrightarrow \mathrm{Co}_3\mathrm{O}_4(s) \quad \Delta H^\circ_{298} = -177.5\ \mathrm{kJ}$$

Q5.3.28

Calculate the standard molar enthalpy of formation of NO(g) from the following data:

$$\mathrm{N}_2(g) + 2\,\mathrm{O}_2 \longrightarrow 2\,\mathrm{NO}_2(g) \quad \Delta H^\circ_{298} = 66.4\ \mathrm{kJ}$$

$$2\,NO(g) + O_2 \longrightarrow 2\,NO_2(g) \quad \Delta H^\circ_{298} = -114.1\ kJ$$

S5.3.28

90.3 mol^{-1} of NO

Q5.3.29

Using the data in Appendix G, calculate the standard enthalpy change for each of the following reactions:

a. $N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$
b. $Si(s) + 2\,Cl_2(g) \longrightarrow SiCl_4(g)$
c. $Fe_2O_3(s) + 3\,H_2(g) \longrightarrow 2\,Fe(s) + 3\,H_2O(l)$
d. $2\,LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(g)$

Q5.3.30

Using the data in Appendix G, calculate the standard enthalpy change for each of the following reactions:

a. $Si(s) + 2\,F_2(g) \longrightarrow SiF_4(g)$
b. $2\,C(s) + 2\,H_2(g) + O_2(g) \longrightarrow CH_3CO_2H(l)$
c. $CH_4(g) + N_2(g) \longrightarrow HCN(g) + NH_3(g)$;
d. $CS_2(g) + 3\,Cl_2(g) \longrightarrow CCl_4(g) + S_2Cl_2(g)$

S5.3.30

a. −1615.0 kJ mol^{-1};
b. −484.3 kJ mol^{-1};
c. 164.2 kJ;
d. −232.1 kJ

Q5.3.31

The following reactions can be used to prepare samples of metals. Determine the enthalpy change under standard state conditions for each.

a. $2\,Ag_2O(s) \longrightarrow 4\,Ag(s) + O_2(g)$
b. $SnO(s) + CO(g) \longrightarrow Sn(s) + CO_2(g)$
c. $Cr_2O_3(s) + 3\,H_2(g) \longrightarrow 2\,Cr(s) + 3\,H_2O(l)$
d. $2\,Al(s) + Fe_2O_3(s) \longrightarrow Al_2O_3(s) + 2\,Fe(s)$

Q5.3.32

The decomposition of hydrogen peroxide, H_2O_2, has been used to provide thrust in the control jets of various space vehicles. Using the data in Appendix G, determine how much heat is produced by the decomposition of exactly 1 mole of H_2O_2 under standard conditions.

$$2\,H_2O_2(l) \longrightarrow 2\,H_2O(g) + O_2(g)$$

S5.3.32

−54.04 kJ mol^{-1}

Q5.3.33

Calculate the enthalpy of combustion of propane, $C_3H_8(g)$, for the formation of $H_2O(g)$ and $CO_2(g)$. The enthalpy of formation of propane is −104 kJ/mol.

Q5.3.34

Calculate the enthalpy of combustion of butane, $C_4H_{10}(g)$ for the formation of $H_2O(g)$ and $CO_2(g)$. The enthalpy of formation of butane is −126 kJ/mol.

S5.3.34

2660 kJ mol^{-1}

Q5.3.35

Both propane and butane are used as gaseous fuels. Which compound produces more heat per gram when burned?

Q5.3.36

The white pigment TiO_2 is prepared by the reaction of titanium tetrachloride, $TiCl_4$, with water vapor in the gas phase:

$$\mathrm{TiCl_4}(g) + 2\,\mathrm{H_2O}(g) \longrightarrow \mathrm{TiO_2}(s) + 4\,\mathrm{HCl}(g) \quad .$$

How much heat is evolved in the production of exactly 1 mole of $TiO_2(s)$ under standard state conditions?

S5.3.36

67.1 kJ

Q5.3.37

Water gas, a mixture of H_2 and CO, is an important industrial fuel produced by the reaction of steam with red hot coke, essentially pure carbon:

$$\mathrm{C}(s) + \mathrm{H_2O}(g) \longrightarrow \mathrm{CO}(g) + \mathrm{H_2}(g) \quad .$$

a. Assuming that coke has the same enthalpy of formation as graphite, calculate ΔH°_{298} for this reaction.
b. Methanol, a liquid fuel that could possibly replace gasoline, can be prepared from water gas and additional hydrogen at high temperature and pressure in the presence of a suitable catalyst:

$$2\,\mathrm{H_2}(g) + \mathrm{CO}(g) \longrightarrow \mathrm{CH_3OH}(g). \tag{7.9.5}$$

Under the conditions of the reaction, methanol forms as a gas. Calculate ΔH°_{298} for this reaction and for the condensation of gaseous methanol to liquid methanol.
c. Calculate the heat of combustion of 1 mole of liquid methanol to $H_2O(g)$ and $CO_2(g)$.

Q5.3.38

In the early days of automobiles, illumination at night was provided by burning acetylene, C_2H_2. Though no longer used as auto headlamps, acetylene is still used as a source of light by some cave explorers. The acetylene is (was) prepared in the lamp by the reaction of water with calcium carbide, CaC_2:

$$\mathrm{CaC_2}(s) + \mathrm{H_2O}(l) \longrightarrow \mathrm{Ca(OH)_2}(s) + \mathrm{C_2H_2}(g) \quad .$$

Calculate the standard enthalpy of the reaction. The ΔH_f° of CaC_2 is −15.14 kcal/mol.

S5.3.38

122.8 kJ

Q5.3.39

From the data in Table, determine which of the following fuels produces the greatest amount of heat per gram when burned under standard conditions: $CO(g)$, $CH_4(g)$, or $C_2H_2(g)$.

Q5.3.40

The enthalpy of combustion of hard coal averages −35 kJ/g, that of gasoline, 1.28×10^5 kJ/gal. How many kilograms of hard coal provide the same amount of heat as is available from 1.0 gallon of gasoline? Assume that the density of gasoline is 0.692 g/mL (the same as the density of isooctane).

S5.3.40

3.7 kg

Q5.3.41

Ethanol, C_2H_5OH, is used as a fuel for motor vehicles, particularly in Brazil.

a. Write the balanced equation for the combustion of ethanol to $CO_2(g)$ and $H_2O(g)$, and, using the data in Appendix G, calculate the enthalpy of combustion of 1 mole of ethanol.
b. The density of ethanol is 0.7893 g/mL. Calculate the enthalpy of combustion of exactly 1 L of ethanol.
c. Assuming that an automobile's mileage is directly proportional to the heat of combustion of the fuel, calculate how much farther an automobile could be expected to travel on 1 L of gasoline than on 1 L of ethanol. Assume that gasoline has the heat of combustion and the density of n–octane, C_8H_{18} ($\Delta H_f^\circ = -208.4\ \text{kJ/mol}$; density = 0.7025 g/mL).

Q5.3.42

Among the substances that react with oxygen and that have been considered as potential rocket fuels are diborane [B_2H_6, produces $B_2O_3(s)$ and $H_2O(g)$], methane [CH_4, produces $CO_2(g)$ and $H_2O(g)$], and hydrazine [N_2H_4, produces $N_2(g)$ and $H_2O(g)$]. On the basis of the heat released by 1.00 g of each substance in its reaction with oxygen, which of these compounds offers the best possibility as a rocket fuel? The ΔH_f° of $B_2H_6(g)$, $CH_4(g)$, and $N_2H_4(l)$ may be found in Appendix G.

S5.3.42

On the assumption that the best rocket fuel is the one that gives off the most heat, B_2H_6 is the prime candidate.

Q5.3.43

How much heat is produced when 1.25 g of chromium metal reacts with oxygen gas under standard conditions?

Ethylene, C_2H_2, a byproduct from the fractional distillation of petroleum, is fourth among the 50 chemical compounds produced commercially in the largest quantities. About 80% of synthetic ethanol is manufactured from ethylene by its reaction with water in the presence of a suitable catalyst.

$$\mathrm{C_2H_4}(g) + \mathrm{H_2O}(g) \longrightarrow \mathrm{C_2H_5OH}(l)$$

Using the data in the table in Appendix G, calculate $\Delta H°$ for the reaction.

S5.3.43

88.2 kJ

Q5.3.44

The oxidation of the sugar glucose, $C_6H_{12}O_6$, is described by the following equation:

$$\mathrm{C_6H_{12}O_6}(s) + 6\ \mathrm{O_2}(g) \longrightarrow 6\ \mathrm{CO_2}(g) + 6\ \mathrm{H_2O}(l) \quad \Delta H = -2816\ \text{kJ}$$

The metabolism of glucose gives the same products, although the glucose reacts with oxygen in a series of steps in the body.

a. How much heat in kilojoules can be produced by the metabolism of 1.0 g of glucose?
b. How many Calories can be produced by the metabolism of 1.0 g of glucose?

Q5.3.45

Propane, C_3H_8, is a hydrocarbon that is commonly used as a fuel.

a. Write a balanced equation for the complete combustion of propane gas.
b. Calculate the volume of air at 25 °C and 1.00 atmosphere that is needed to completely combust 25.0 grams of propane. Assume that air is 21.0 percent O_2 by volume. (Hint: we will see how to do this calculation in a later chapter on gases—for now use the information that 1.00 L of air at 25 °C and 1.00 atm contains 0.275 g of O_2 per liter.)
c. The heat of combustion of propane is −2,219.2 kJ/mol. Calculate the heat of formation, ΔH_f° of propane given that ΔH_f° of $H_2O(l)$ = −285.8 kJ/mol and ΔH_f° of $CO_2(g)$ = −393.5 kJ/mol.
d. Assuming that all of the heat released in burning 25.0 grams of propane is transferred to 4.00 kilograms of water, calculate the increase in temperature of the water.

S5.3.45

a. $\mathrm{C_3H_8}(g) + 5\ \mathrm{O_2}(g) \longrightarrow 3\ \mathrm{CO_2}(g) + 4\ \mathrm{H_2O}(l)$;
b. 330 L;

c. $-104.5\ \text{kJ mol}^{-1}$;
d. 75.4 °C

Q5.3.46

During a recent winter month in Sheboygan, Wisconsin, it was necessary to obtain 3500 kWh of heat provided by a natural gas furnace with 89% efficiency to keep a small house warm (the efficiency of a gas furnace is the percent of the heat produced by combustion that is transferred into the house).

a. Assume that natural gas is pure methane and determine the volume of natural gas in cubic feet that was required to heat the house. The average temperature of the natural gas was 56 °F; at this temperature and a pressure of 1 atm, natural gas has a density of 0.681 g/L.
b. How many gallons of LPG (liquefied petroleum gas) would be required to replace the natural gas used? Assume the LPG is liquid propane [C_3H_8: density, 0.5318 g/mL; enthalpy of combustion, 2219 kJ/mol for the formation of $CO_2(g)$ and $H_2O(l)$] and the furnace used to burn the LPG has the same efficiency as the gas furnace.
c. What mass of carbon dioxide is produced by combustion of the methane used to heat the house?
d. What mass of water is produced by combustion of the methane used to heat the house?
e. What volume of air is required to provide the oxygen for the combustion of the methane used to heat the house? Air contains 23% oxygen by mass. The average density of air during the month was 1.22 g/L.
f. How many kilowatt–hours (1 kWh = 3.6×10^6 J) of electricity would be required to provide the heat necessary to heat the house? Note electricity is 100% efficient in producing heat inside a house.
g. Although electricity is 100% efficient in producing heat inside a house, production and distribution of electricity is not 100% efficient. The efficiency of production and distribution of electricity produced in a coal-fired power plant is about 40%. A certain type of coal provides 2.26 kWh per pound upon combustion. What mass of this coal in kilograms will be required to produce the electrical energy necessary to heat the house if the efficiency of generation and distribution is 40%?

This page titled 7.9: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

8: Liquids and Solids and Phase Changes

The great distances between atoms and molecules in a gaseous phase, and the corresponding absence of any significant interactions between them, allows for simple descriptions of many physical properties that are the same for all gases, regardless of their chemical identities. As described in the final module of the chapter on gases, this situation changes at high pressures and low temperatures—conditions that permit the atoms and molecules to interact to a much greater extent. In the liquid and solid states, these interactions are of considerable strength and play an important role in determining a number of physical properties that *do* depend on the chemical identity of the substance. In this chapter, the nature of these interactions and their effects on various physical properties of liquid and solid phases will be examined.

8.1: Introduction
8.2: Intermolecular Forces
8.3: Properties of Liquids
8.4: Phase Transitions
8.5: Phase Diagrams
8.6: The Solid State of Matter
8.7: Key Terms
8.8: Key Equations
8.9: Summary
8.10: Exercises

This page titled 8: Liquids and Solids and Phase Changes is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

8.1: Introduction

Leprosy has been a devastating disease throughout much of human history. Aside from the symptoms and complications of the illness, its social stigma led sufferers to be cast out of communities and isolated in colonies; in some regions this practice lasted well into the twentieth century. At that time, the best potential treatment for leprosy was oil from the chaulmoogra tree, but the oil was extremely thick, causing blisters and making usage painful and ineffective. Healthcare professionals seeking a better application contacted Alice Ball, a young chemist at the University of Hawaii, who had focused her masters thesis on a similar plant. Ball initiated a sequence of procedures (repeated acidification and purification to change the characteristics of the oil and isolate the active substances (esters, discussed later in this text). The "Ball Method" as it later came to be called, became the standard treatment for leprosy for decades.



Figure 8.1.1: Solid carbon dioxide (“dry ice”, left) sublimes vigorously when placed in a liquid (right), cooling the liquid and generating a dense mist of water above the cylinder. (credit: modification of work by Paul Flowers)

In the liquid and solid states, atomic and molecular interactions are of considerable strength and play an important role in determining a number of physical properties of the substance. For example, the thickness, or viscosity, of the chaulmoogra oil was due to its intermolecular forces. In this chapter, the nature of these interactions and their effects on various physical properties of liquid and solid phases will be examined.

This page titled 8.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **10.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.2: Intermolecular Forces

Learning Objectives

By the end of this section, you will be able to:

- Describe the types of intermolecular forces possible between atoms or molecules in condensed phases (dispersion forces, dipole-dipole attractions, and hydrogen bonding)
- Identify the types of intermolecular forces experienced by specific molecules based on their structures
- Explain the relation between the intermolecular forces present within a substance and the temperatures associated with changes in its physical state

As was the case for gaseous substances, the kinetic molecular theory may be used to explain the behavior of solids and liquids. In the following description, the term *particle* will be used to refer to an atom, molecule, or ion. Note that we will use the popular phrase "intermolecular attraction" to refer to attractive forces between the particles of a substance, regardless of whether these particles are molecules, atoms, or ions.

Consider these two aspects of the molecular-level environments in solid, liquid, and gaseous matter:

- Particles in a solid are tightly packed together and often arranged in a regular pattern; in a liquid, they are close together with no regular arrangement; in a gas, they are far apart with no regular arrangement.
- Particles in a solid vibrate about fixed positions and do not generally move in relation to one another; in a liquid, they move past each other but remain in essentially constant contact; in a gas, they move independently of one another except when they collide.

The differences in the properties of a solid, liquid, or gas reflect the strengths of the attractive forces between the atoms, molecules, or ions that make up each phase. The phase in which a substance exists depends on the relative extents of its **intermolecular forces** (IMFs) and the kinetic energies (KE) of its molecules. IMFs are the various forces of attraction that may exist between the atoms and molecules of a substance due to electrostatic phenomena, as will be detailed in this module. These forces serve to hold particles close together, whereas the particles' KE provides the energy required to overcome the attractive forces and thus increase the distance between particles. Figure 8.2.1 illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

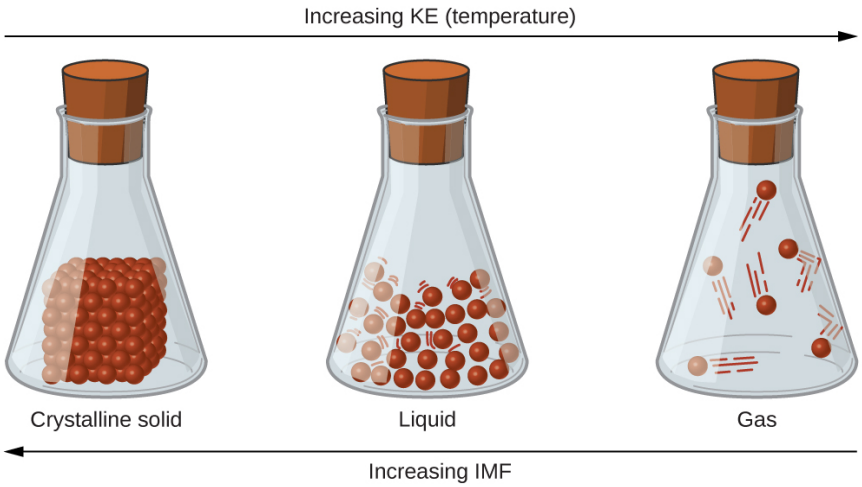


Figure 8.2.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

As an example of the processes depicted in this figure, consider a sample of water. When gaseous water is cooled sufficiently, the attractions between H_2O molecules will be capable of holding them together when they come into contact with each other; the gas condenses, forming liquid H_2O. For example, liquid water forms on the outside of a cold glass as the water vapor in the air is cooled by the cold glass, as seen in Figure 8.2.2.



Figure 8.2.2: Condensation forms when water vapor in the air is cooled enough to form liquid water, such as (a) on the outside of a cold beverage glass or (b) in the form of fog. (credit a: modification of work by Jenny Downing; credit b: modification of work by Cory Zanker)

We can also liquefy many gases by compressing them, if the temperature is not too high. The increased pressure brings the molecules of a gas closer together, such that the attractions between the molecules become strong relative to their KE. Consequently, they form liquids. Butane, C_4H_{10}, is the fuel used in disposable lighters and is a gas at standard temperature and pressure. Inside the lighter's fuel compartment, the butane is compressed to a pressure that results in its condensation to the liquid state, as shown in Figure 8.2.3.



Figure 8.2.3: Gaseous butane is compressed within the storage compartment of a disposable lighter, resulting in its condensation to the liquid state. (credit: modification of work by "Sam-Cat"/Flickr)

Finally, if the temperature of a liquid becomes sufficiently low, or the pressure on the liquid becomes sufficiently high, the molecules of the liquid no longer have enough KE to overcome the IMF between them, and a solid forms. A more thorough discussion of these and other changes of state, or phase transitions, is provided in a later module of this chapter.

Link to Learning

Access this interactive simulation on states of matter, phase transitions, and intermolecular forces. This simulation is useful for visualizing concepts introduced throughout this chapter.

Forces between Molecules

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intra*molecular forces. *Intra*molecular forces are those *within* the molecule that keep the molecule together, for example, the bonds between the atoms. *Inter*molecular forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 8.2.4: illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

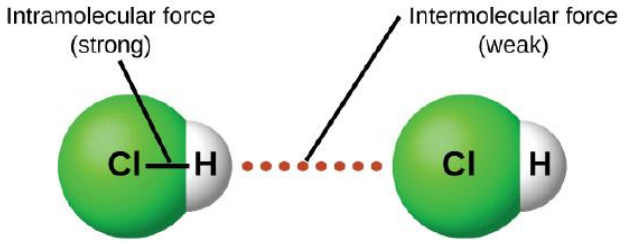


Figure 8.2.4: *Intra*molecular forces keep a molecule intact. *Inter*molecular forces hold multiple molecules together and determine many of a substance's properties.

All of the attractive forces between neutral atoms and molecules are known as van der Waals forces, although they are usually referred to more informally as intermolecular attraction. We will consider the various types of IMFs in the next three sections of this module.

Dispersion Forces

One of the three van der Waals forces is present in all condensed phases, regardless of the nature of the atoms or molecules composing the substance. This attractive force is called the London dispersion force in honor of German-born American physicist Fritz London who, in 1928, first explained it. This force is often referred to as simply the **dispersion force**. Because the electrons of an atom or molecule are in constant motion (or, alternatively, the electron's location is subject to quantum-mechanical variability), at any moment in time, an atom or molecule can develop a temporary, **instantaneous dipole** if its electrons are distributed asymmetrically. The presence of this dipole can, in turn, distort the electrons of a neighboring atom or molecule, producing an **induced dipole**. These two rapidly fluctuating, temporary dipoles thus result in a relatively weak electrostatic attraction between the species—a so-called dispersion force like that illustrated in Figure 8.2.5.

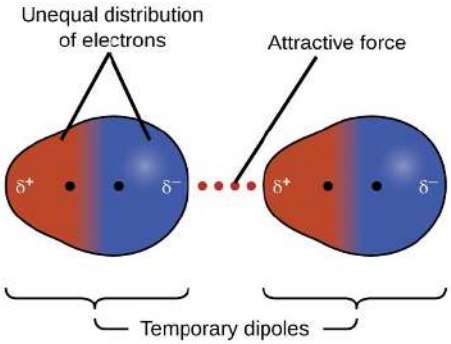


Figure 8.2.5: Dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules.

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. Larger and heavier atoms and molecules exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 8.2.1.

Table 8.2.1: Melting and Boiling Points of the Halogens

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as polarizability. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

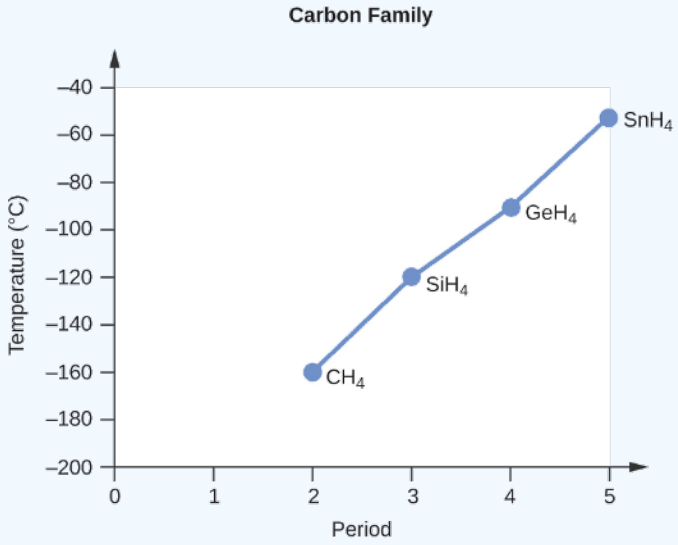
Example 8.2.1: London Forces and Their Effects

Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4, SiH_4, GeH_4, and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be $CH_4 < SiH_4 < GeH_4 < SnH_4$.

A graph of the actual boiling points of these compounds versus the period of the group 14 element shows this prediction to be correct:



Exercise 8.2.1

Order the following hydrocarbons from lowest to highest boiling point: C_2H_6, C_3H_8, and C_4H_{10}.

Answer

$C_2H_6 < C_3H_8 < C_4H_{10}$. All of these compounds are nonpolar and only have London dispersion forces: the larger the molecule, the larger the dispersion forces and the higher the boiling point. The ordering from lowest to highest boiling point is therefore $C_2H_6 < C_3H_8 < C_4H_{10}$.

The shapes of molecules also affect the magnitudes of the dispersion forces between them. For example, boiling points for the isomers *n*-pentane, isopentane, and neopentane (shown in Figure 8.2.6) are 36 °C, 27 °C, and 9.5 °C, respectively. Even though these compounds are composed of molecules with the same chemical formula, C_5H_{12}, the difference in boiling points suggests that dispersion forces in the liquid phase are different, being greatest for *n*-pentane and least for neopentane. The elongated shape of *n*-

pentane provides a greater surface area available for contact between molecules, resulting in correspondingly stronger dispersion forces. The more compact shape of isopentane offers a smaller surface area available for intermolecular contact and, therefore, weaker dispersion forces. Neopentane molecules are the most compact of the three, offering the least available surface area for intermolecular contact and, hence, the weakest dispersion forces. This behavior is analogous to the connections that may be formed between strips of VELCRO brand fasteners: the greater the area of the strip's contact, the stronger the connection.

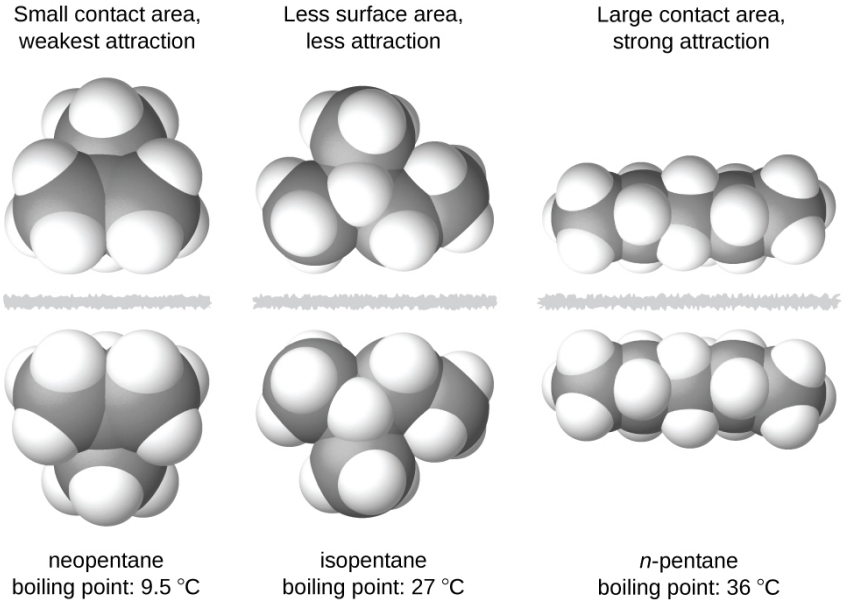


Figure 8.2.6: The strength of the dispersion forces increases with the contact area between molecules, as demonstrated by the boiling points of these pentane isomers.

Chemistry in Everyday Life: Geckos and Intermolecular Forces

Geckos have an amazing ability to adhere to most surfaces. They can quickly run up smooth walls and across ceilings that have no toe-holds, and they do this without having suction cups or a sticky substance on their toes. And while a gecko can lift its feet easily as it walks along a surface, if you attempt to pick it up, it sticks to the surface. How are geckos (as well as spiders and some other insects) able to do this? Although this phenomenon has been investigated for hundreds of years, scientists only recently uncovered the details of the process that allows geckos' feet to behave this way.

Geckos' toes are covered with hundreds of thousands of tiny hairs known as *setae*, with each seta, in turn, branching into hundreds of tiny, flat, triangular tips called *spatulae*. The huge numbers of spatulae on its setae provide a gecko, shown in Figure 8.2.7, with a large total surface area for sticking to a surface. In 2000, Kellar Autumn, who leads a multi-institutional gecko research team, found that geckos adhered equally well to both polar silicon dioxide and nonpolar gallium arsenide. This proved that geckos stick to surfaces because of dispersion forces—weak intermolecular attractions arising from temporary, synchronized charge distributions between adjacent molecules. Although dispersion forces are very weak, the total attraction over millions of spatulae is large enough to support many times the gecko's weight.

In 2014, two scientists developed a model to explain how geckos can rapidly transition from "sticky" to "non-sticky." Alex Greaney and Congcong Hu at Oregon State University described how geckos can achieve this by changing the angle between their spatulae and the surface. Geckos' feet, which are normally nonsticky, become sticky when a small shear force is applied. By curling and uncurling their toes, geckos can alternate between sticking and unsticking from a surface, and thus easily move across it. Later research led by Alyssa Stark at University of Akron showed that geckos can maintain their hold on hydrophobic surfaces (similar to the leaves in their habitats) equally well whether the surfaces were wet or dry. Stark's experiment used a ribbon to gently pull the geckos until they slipped, so that the researchers could determine the geckos' ability to hold various surfaces under wet and dry conditions. Further investigations may eventually lead to the development of better adhesives and other applications.

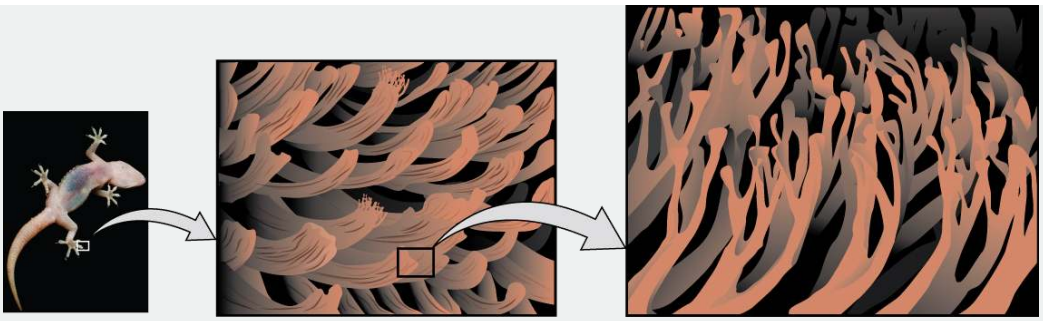


Figure 8.2.7: Geckos' toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko's millions of spatulae. By changing how the spatulae contact the surface, geckos can turn their stickiness "on" and "off." (credit photo: modification of work by "JC*+A!"/Flickr)

Link to Learning

Watch this video to learn more about Kellar Autumn's research that determined that van der Waals forces are responsible for a gecko's ability to cling and climb.

Dipole-Dipole Attractions

Recall from the chapter on chemical bonding and molecular geometry that *polar* molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a *dipole*. Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a dipole-dipole attraction—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in Figure 8.2.8.

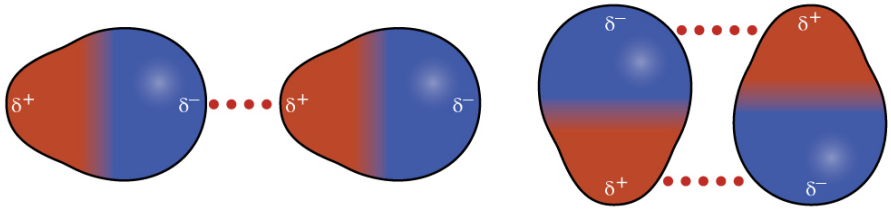


Figure 8.2.8: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F_2 molecules. Both HCl and F_2 consist of the same number of atoms and have approximately the same molecular mass. At a temperature of 150 K, molecules of both substances would have the same average KE. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to "stick together" to form a liquid, whereas the relatively weaker dispersion forces between nonpolar F_2 molecules are not, and so this substance is gaseous at this temperature. The higher normal boiling point of HCl (188 K) compared to F_2 (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F_2 molecules. We will often use values such as boiling or freezing points, or enthalpies of vaporization or fusion, as indicators of the relative strengths of IMFs of attraction present within different substances.

Example 8.2.2: Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

Solution

CO and N_2 are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N_2 is nonpolar, its molecules cannot exhibit

dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N_2 molecules, so CO is expected to have the higher boiling point.

? Exercise 8.2.2

Predict which will have the higher boiling point: ICl or Br_2. Explain your reasoning.

Answer

ICl. ICl and Br_2 have similar masses (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br_2 is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point.

Hydrogen Bonding

Nitrosyl fluoride (ONF, molecular mass 49 amu) is a gas at room temperature. Water (H_2O, molecular mass 18 amu) is a liquid, even though it has a lower molecular mass. We clearly cannot attribute this difference between the two compounds to dispersion forces. Both molecules have about the same shape and ONF is the heavier and larger molecule. It is, therefore, expected to experience more significant dispersion forces. Additionally, we cannot attribute this difference in boiling points to differences in the dipole moments of the molecules. Both molecules are polar and exhibit comparable dipole moments. The large difference between the boiling points is due to a particularly strong dipole-dipole attraction that may occur when a molecule contains a hydrogen atom bonded to a fluorine, oxygen, or nitrogen atom (the three most electronegative elements). The very large difference in electronegativity between the H atom (2.1) and the atom to which it is bonded (4.0 for an F atom, 3.5 for an O atom, or 3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to *highly concentrated partial charges* with these atoms. Molecules with F-H, O-H, or N-H moieties are very strongly attracted to similar moieties in nearby molecules, a particularly strong type of dipole-dipole attraction called hydrogen bonding. Examples of hydrogen bonds include HF···HF, H_2O···HOH, and H_3N···HNH_2, in which the hydrogen bonds are denoted by dots. Figure 8.2.9 illustrates hydrogen bonding between water molecules.

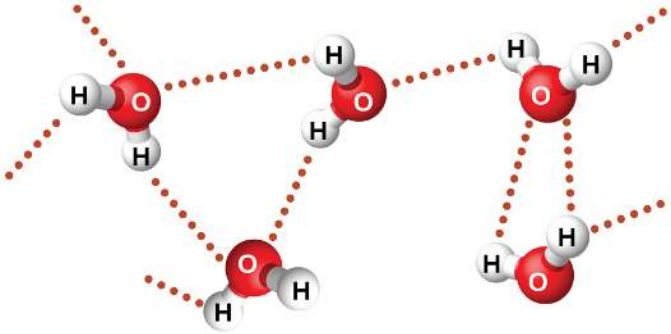


Figure 8.2.9: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

Despite use of the word “bond,” keep in mind that hydrogen bonds are *intermolecular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, only about 5 to 10% as strong, but are generally much stronger than other dipole-dipole attractions and dispersion forces.

Hydrogen bonds have a pronounced effect on the properties of condensed phases (liquids and solids). For example, consider the trends in boiling points for the binary hydrides of group 15 (NH_3, PH_3, AsH_3, and SbH_3), group 16 hydrides (H_2O, H_2S, H_2Se, and H_2Te), and group 17 hydrides (HF, HCl, HBr, and HI). The boiling points of the heaviest three hydrides for each group are plotted in Figure 8.2.10. As we progress down any of these groups, the polarities of the molecules decrease slightly, whereas the sizes of the molecules increase substantially. The effect of increasingly stronger dispersion forces dominates that of increasingly weaker dipole-dipole attractions, and the boiling points are observed to increase steadily.

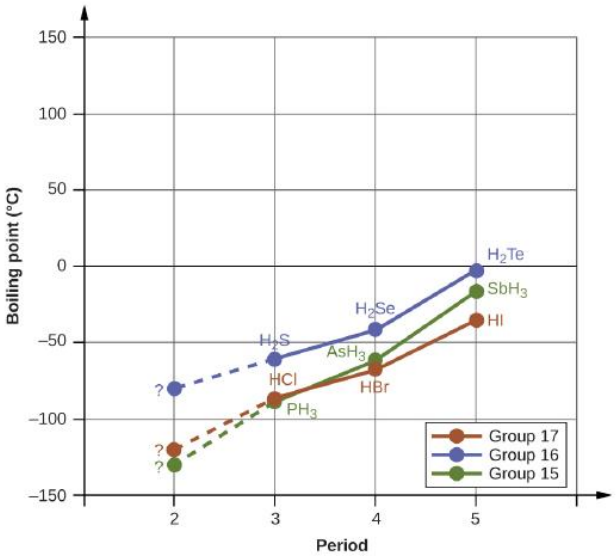


Figure 8.2.10: For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5.

If we use this trend to predict the boiling points for the lightest hydride for each group, we would expect NH_3 to boil at about −120 °C, H_2O to boil at about −80 °C, and HF to boil at about −110 °C. However, when we measure the boiling points for these compounds, we find that they are dramatically higher than the trends would predict, as shown in Figure 8.2.11. The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

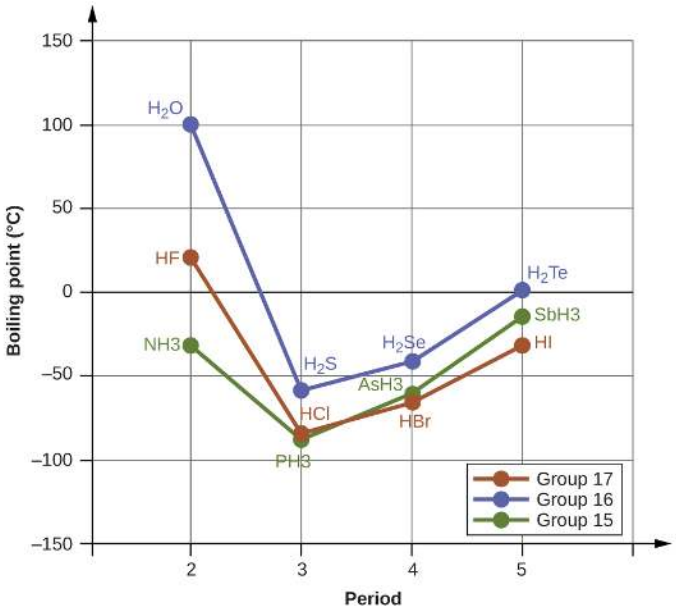


Figure 8.2.11: In comparison to periods 3−5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

✔ Example 8.2.3: Effect of Hydrogen Bonding on Boiling Points

Consider the compounds dimethylether (CH_3OCH_3), ethanol (CH_3CH_2OH), and propane ($CH_3CH_2CH_3$). Their boiling points, not necessarily in order, are −42.1 °C, −24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

Solution

The VSEPR-predicted shapes of CH_3OCH_3, CH_3CH_2OH, and $CH_3CH_2CH_3$ are similar, as are their molar masses (46 g/mol, 46 g/mol, and 44 g/mol, respectively), so they will exhibit similar dispersion forces. Since $CH_3CH_2CH_3$ is nonpolar, it may

exhibit *only* dispersion forces. Because CH_3OCH_3 is polar, it will also experience dipole-dipole attractions. Finally, CH_3CH_2OH has an −OH group, and so it will experience the uniquely strong dipole-dipole attraction known as hydrogen bonding. So the ordering in terms of strength of IMFs, and thus boiling points, is $CH_3CH_2CH_3 < CH_3OCH_3 < CH_3CH_2OH$. The boiling point of propane is −42.1 °C, the boiling point of dimethylether is −24.8 °C, and the boiling point of ethanol is 78.5 °C.

Exercise 8.2.3

Ethane (CH_3CH_3) has a melting point of −183 °C and a boiling point of −89 °C. Predict the melting and boiling points for methylamine (CH_3NH_2). Explain your reasoning.

Answer

The melting point and boiling point for methylamine are predicted to be significantly greater than those of ethane. CH_3CH_3 and CH_3NH_2 are similar in size and mass, but methylamine possesses an −NH group and therefore may exhibit hydrogen bonding. This greatly increases its IMFs, and therefore its melting and boiling points. It is difficult to predict values, but the known values are a melting point of −93 °C and a boiling point of −6 °C.

Scientist Focus: Dorothy June Sutor and Hydrogen Bonding



We often associate hydrogen bonding to only molecules that contain of H-F, H-N or H-O. As scientists, it is very important to challenge general assumptions, especially when presented with reproducible experimental evidence.

One such scientist that challenged the chemical theory of her time was Doroth June Sutor (1929-1990), who was a New Zealand-born crystallographer who worked during the 1950s and 1960s. She noticied that a number of the complexes she was studying exhibited "C-H----O" bonds, in which the distance between the O and H atoms was smaller than expected from the van der Waals radius. This was a bold observation as the electronegativity difference between C and H is quite small and hence the general consensus among chemists at the time was that no hydrogen bonding should exists with a C-H bond. Dr. Sutor worked with computers that required paper punch cards to program them and all of her work was completed using paper punch programming cards.

Because she was a woman working in a time period where sexism was common place, her work was widely disputed by her peers, who were more prominent in the crystallographic community and also male. Owing to these prejudicies, her work was mostly forgotten. In the 1980s, when more advanced crystallographic techniques were developed (neutron diffraction), her work was revisited and it was found she had been right all along! It is now standard practice to identify hydrogen bonding between the hydrogen atoms in C-H bonds oxygen atoms in close proximity, especially in DNA.

If you would like to read more, please see this page: https://en.Wikipedia.org/wiki/June_Sutor

How Sciences Interconnect: Hydrogen Bonding and DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism's characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism's offspring. A DNA molecule consists of two (anti-)parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in Figure 8.2.12

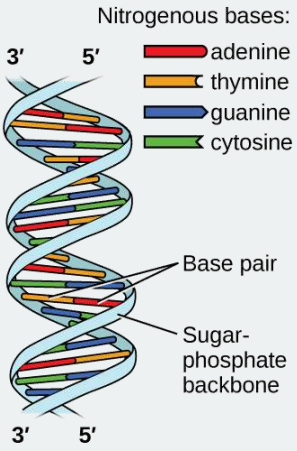


Figure 8.2.12: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure Figure 8.2.13

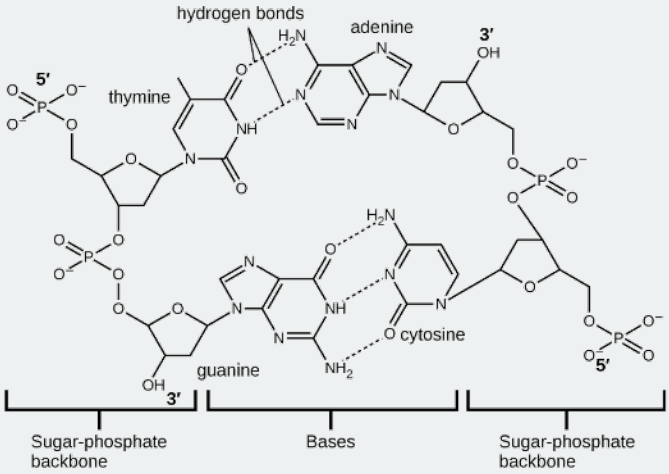


Figure 8.2.13: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called “complementary base pairs.”

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily “unzip” down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

This page titled 8.2: Intermolecular Forces is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **10.1: Intermolecular Forces** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.3: Properties of Liquids

Learning Objectives

By the end of this section, you will be able to:

- Distinguish between adhesive and cohesive forces
- Define viscosity, surface tension, and capillary rise
- Describe the roles of intermolecular attractive forces in each of these properties/phenomena

When you pour a glass of water, or fill a car with gasoline, you observe that water and gasoline flow freely. But when you pour syrup on pancakes or add oil to a car engine, you note that syrup and motor oil do not flow as readily. The **viscosity** of a liquid is a measure of its resistance to flow. Water, gasoline, and other liquids that flow freely have a low viscosity. Honey, syrup, motor oil, and other liquids that do not flow freely, like those shown in Figure 8.3.1, have higher viscosities. We can measure viscosity by measuring the rate at which a metal ball falls through a liquid (the ball falls more slowly through a more viscous liquid) or by measuring the rate at which a liquid flows through a narrow tube (more viscous liquids flow more slowly).



Figure 8.3.1: (a) Honey and (b) motor oil are examples of liquids with high viscosities; they flow slowly. (credit a: modification of work by Scott Bauer; credit b: modification of work by David Nagy)

The IMFs between the molecules of a liquid, the size and shape of the molecules, and the temperature determine how easily a liquid flows. As Table 8.3.1 shows, the more structurally complex are the molecules in a liquid and the stronger the IMFs between them, the more difficult it is for them to move past each other and the greater is the viscosity of the liquid. As the temperature increases, the molecules move more rapidly and their kinetic energies are better able to overcome the forces that hold them together; thus, the viscosity of the liquid decreases.

Table 8.3.1: Viscosities of Common Substances at 25 °C

Substance	Formula	Viscosity (mPa·s)
water	H_2O	0.890
mercury	Hg	1.526
ethanol	C_2H_5OH	1.074
octane	C_8H_{18}	0.508
ethylene glycol	$CH_2(OH)CH_2(OH)$	16.1
honey	variable	~2,000–10,000
motor oil	variable	~50–500

The various IMFs between identical molecules of a substance are examples of cohesive forces. The molecules within a liquid are surrounded by other molecules and are attracted equally in all directions by the cohesive forces within the liquid. However, the molecules on the surface of a liquid are attracted only by about one-half as many molecules. Because of the unbalanced molecular attractions on the surface molecules, liquids contract to form a shape that minimizes the number of molecules on the surface—that

is, the shape with the minimum surface area. A small drop of liquid tends to assume a spherical shape, as shown in Figure 8.3.2, because in a sphere, the ratio of surface area to volume is at a minimum. Larger drops are more greatly affected by gravity, air resistance, surface interactions, and so on, and as a result, are less spherical.

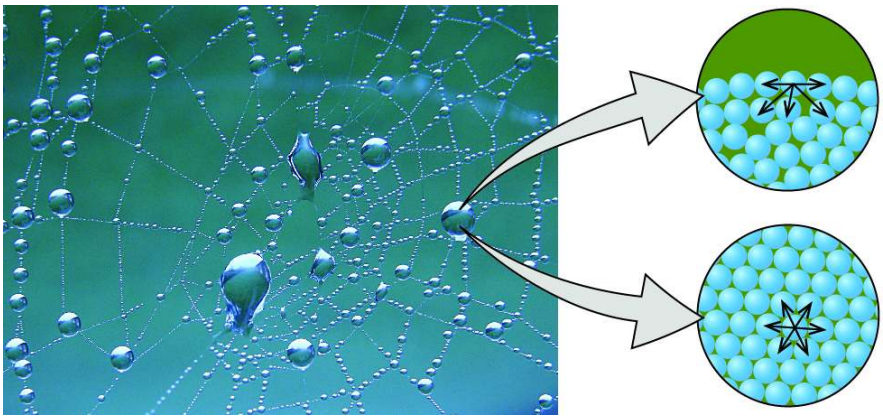


Figure 8.3.2: Attractive forces result in a spherical water drop that minimizes surface area; cohesive forces hold the sphere together; adhesive forces keep the drop attached to the web. (credit photo: modification of work by “OliBac”/Flickr)

Surface tension is defined as the energy required to increase the surface area of a liquid, or the force required to increase the length of a liquid surface by a given amount. This property results from the cohesive forces between molecules at the surface of a liquid, and it causes the surface of a liquid to behave like a stretched rubber membrane. Surface tensions of several liquids are presented in Table 8.3.2. Among common liquids, water exhibits a distinctly high surface tension due to strong hydrogen bonding between its molecules. As a result of this high surface tension, the surface of water represents a relatively “tough skin” that can withstand considerable force without breaking. A steel needle carefully placed on water will float. Some insects, like the one shown in Figure 8.3.3, even though they are denser than water, move on its surface because they are supported by the surface tension.

Table 8.3.2: Surface Tensions of Common Substances at 25 °C

Substance	Formula	Surface Tension (mN/m)
water	H_2O	71.99
mercury	Hg	458.48
ethanol	C_2H_5OH	21.97
octane	C_8H_{18}	21.14
ethylene glycol	$CH_2(OH)CH_2(OH)$	47.99

Surface tension is affected by a variety of variables, including the introduction of additional substances on the surface. In the late 1800s, Agnes Pockels, who was initially blocked from pursuing a scientific career but studied on her own, began investigating the impact and characteristics of soapy and greasy films in water. Using homemade materials, she developed an instrument known as a trough for measuring surface contaminants and their effects. With the support of renowned scientist Lord Rayleigh, her 1891 paper showed that surface contamination significantly reduces surface tension, and also that changing the characteristics of the surface (compressing or expanding it) also affects surface tension. Decades later, Irving Langmuir and Katharine Blodgett built on Pockels' work in their own trough and important advances in surface chemistry. Langmuir pioneered methods for producing single-molecule layers of film; Blodgett applied these to the development of non-reflective glass (critical for film-making and other applications), and also studied methods related to cleaning surfaces, which are important in semiconductor fabrication.

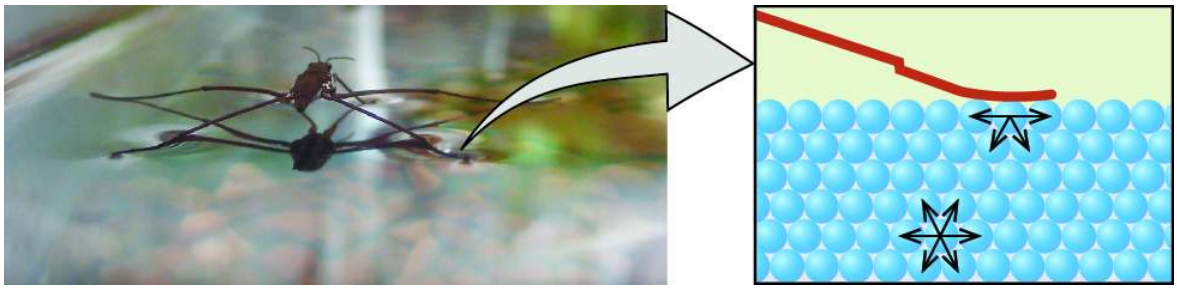


Figure 8.3.3: Surface tension (right) prevents this insect, a “water strider,” from sinking into the water.

The IMFs of attraction between two *different* molecules are called **adhesive forces**. Consider what happens when water comes into contact with some surface. If the adhesive forces between water molecules and the molecules of the surface are weak compared to the cohesive forces between the water molecules, the water does not “wet” the surface. For example, water does not wet waxed surfaces or many plastics such as polyethylene. Water forms drops on these surfaces because the cohesive forces within the drops are greater than the adhesive forces between the water and the plastic. Water spreads out on glass because the adhesive force between water and glass is greater than the cohesive forces within the water. When water is confined in a glass tube, its meniscus (surface) has a concave shape because the water wets the glass and creeps up the side of the tube. On the other hand, the cohesive forces between mercury atoms are much greater than the adhesive forces between mercury and glass. Mercury therefore does not wet glass, and it forms a convex meniscus when confined in a tube because the cohesive forces within the mercury tend to draw it into a drop (Figure 8.3.4).

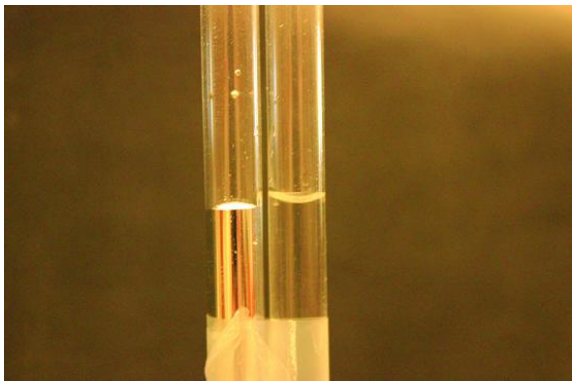


Figure 8.3.4: Differences in the relative strengths of cohesive and adhesive forces result in different meniscus shapes for mercury (left) and water (right) in glass tubes. (credit: Mark Ott)

If you place one end of a paper towel in spilled wine, as shown in Figure 8.3.5, the liquid wicks up the paper towel. A similar process occurs in a cloth towel when you use it to dry off after a shower. These are examples of **capillary action**—when a liquid flows within a porous material due to the attraction of the liquid molecules to the surface of the material and to other liquid molecules. The adhesive forces between the liquid and the porous material, combined with the cohesive forces within the liquid, may be strong enough to move the liquid upward against gravity.

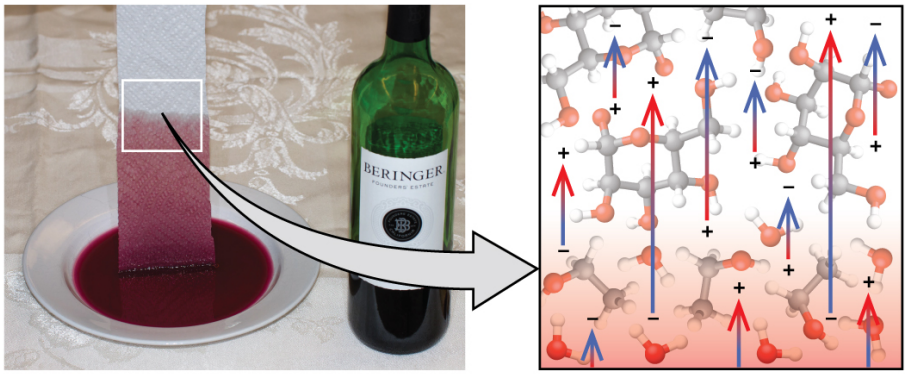


Figure 8.3.5: Wine wicks up a paper towel (left) because of the strong attractions of water (and ethanol) molecules to the −OH groups on the towel's cellulose fibers and the strong attractions of water molecules to other water (and ethanol) molecules (right). (credit photo: modification of work by Mark Blaser)

Towels soak up liquids like water because the fibers of a towel are made of molecules that are attracted to water molecules. Most cloth towels are made of cotton, and paper towels are generally made from paper pulp. Both consist of long molecules of cellulose that contain many −OH groups. Water molecules are attracted to these −OH groups and form hydrogen bonds with them, which draws the H_2O molecules up the cellulose molecules. The water molecules are also attracted to each other, so large amounts of water are drawn up the cellulose fibers.

Capillary action can also occur when one end of a small diameter tube is immersed in a liquid, as illustrated in Figure 8.3.6. If the liquid molecules are strongly attracted to the tube molecules, the liquid creeps up the inside of the tube until the weight of the liquid and the adhesive forces are in balance. The smaller the diameter of the tube is, the higher the liquid climbs. It is partly by capillary action occurring in plant cells called xylem that water and dissolved nutrients are brought from the soil up through the roots and into a plant. Capillary action is the basis for thin layer chromatography, a laboratory technique commonly used to separate small quantities of mixtures. You depend on a constant supply of tears to keep your eyes lubricated and on capillary action to pump tear fluid away.

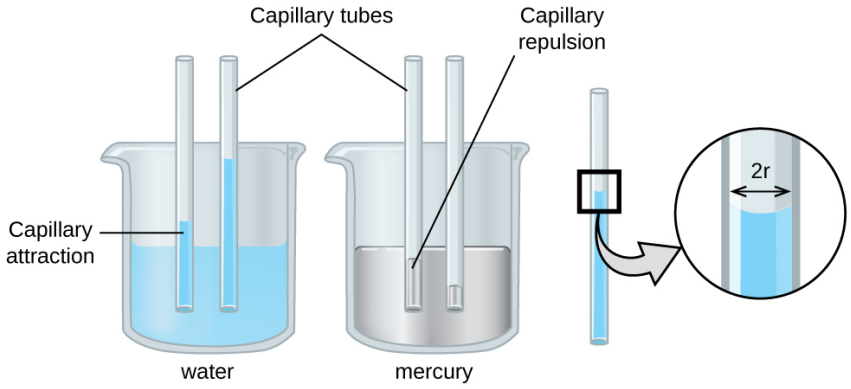


Figure 8.3.6: Depending upon the relative strengths of adhesive and cohesive forces, a liquid may rise (such as water) or fall (such as mercury) in a glass capillary tube. The extent of the rise (or fall) is directly proportional to the surface tension of the liquid and inversely proportional to the density of the liquid and the radius of the tube.

The height to which a liquid will rise in a capillary tube is determined by several factors as shown in the following equation:

$$h = \frac{2T\cos\theta}{r\rho g} \tag{8.3.1}$$

In this equation, h is the height of the liquid inside the capillary tube relative to the surface of the liquid outside the tube, T is the surface tension of the liquid, θ is the contact angle between the liquid and the tube, r is the radius of the tube, ρ is the density of the liquid, and g is the acceleration due to gravity, 9.8 m/s^2. When the tube is made of a material to which the liquid molecules are

strongly attracted, they will spread out completely on the surface, which corresponds to a contact angle of 0°. This is the situation for water rising in a glass tube.

Example 8.3.1: Capillary Rise

At 25 °C, how high will water rise in a glass capillary tube with an inner diameter of 0.25 mm?

For water, T = 71.99 mN/m and ρ = 1.0 g/cm^3.

Solution

The liquid will rise to a height h given by Equation 8.3.1:

$$h = \frac{2T\cos\theta}{r\rho g}$$

The Newton is defined as a kg m/s^2, and so the provided surface tension is equivalent to 0.07199 kg/s^2. The provided density must be converted into units that will cancel appropriately: ρ = 1000 kg/m^3. The diameter of the tube in meters is 0.00025 m, so the radius is 0.000125 m. For a glass tube immersed in water, the contact angle is θ = 0°, so cos θ = 1. Finally, acceleration due to gravity on the earth is g = 9.8 m/s^2. Substituting these values into the equation, and cancelling units, we have:

$$h = \frac{2\left(0.07199 kg/s^2\right)}{\left(0.000125 m\right)\left(1000 kg/m^3\right)\left(9.8 m/s^2\right)} = 0.12 m = 12 cm$$

Exercise 8.3.1

Water rises in a glass capillary tube to a height of 8.4 cm. What is the diameter of the capillary tube?

Answer

diameter = 0.36 mm

Chemistry in Everyday Life: Biomedical Applications of Capillary Action

Many medical tests require drawing a small amount of blood, for example to determine the amount of glucose in someone with diabetes or the hematocrit level in an athlete. This procedure can be easily done because of capillary action, the ability of a liquid to flow up a small tube against gravity, as shown in Figure 8.3.7. When your finger is pricked, a drop of blood forms and holds together due to surface tension—the unbalanced intermolecular attractions at the surface of the drop. Then, when the open end of a narrow-diameter glass tube touches the drop of blood, the adhesive forces between the molecules in the blood and those at the glass surface draw the blood up the tube. How far the blood goes up the tube depends on the diameter of the tube (and the type of fluid). A small tube has a relatively large surface area for a given volume of blood, which results in larger (relative) attractive forces, allowing the blood to be drawn farther up the tube. The liquid itself is held together by its own cohesive forces. When the weight of the liquid in the tube generates a downward force equal to the upward force associated with capillary action, the liquid stops rising.



Figure 8.3.7: Blood is collected for medical analysis by capillary action, which draws blood into a small diameter glass tube. (credit: modification of work by Centers for Disease Control and Prevention)

This page titled 8.3: Properties of Liquids is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **10.2: Properties of Liquids** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.4: Phase Transitions

Learning Objectives

By the end of this section, you will be able to:

- Define phase transitions and phase transition temperatures
- Explain the relation between phase transition temperatures and intermolecular attractive forces
- Describe the processes represented by typical heating and cooling curves, and compute heat flows and enthalpy changes accompanying these processes

We witness and utilize changes of physical state, or phase transitions, in a great number of ways. As one example of global significance, consider the evaporation, condensation, freezing, and melting of water. These changes of state are essential aspects of our earth's water cycle as well as many other natural phenomena and technological processes of central importance to our lives. In this module, the essential aspects of phase transitions are explored.

Vaporization and Condensation

When a liquid vaporizes in a closed container, gas molecules cannot escape. As these gas phase molecules move randomly about, they will occasionally collide with the surface of the condensed phase, and in some cases, these collisions will result in the molecules re-entering the condensed phase. The change from the gas phase to the liquid is called **condensation**. When the rate of condensation becomes equal to the rate of **vaporization**, neither the amount of the liquid nor the amount of the vapor in the container changes. The vapor in the container is then said to be *in equilibrium* with the liquid. Keep in mind that this is not a static situation, as molecules are continually exchanged between the condensed and gaseous phases. Such is an example of a **dynamic equilibrium**, the status of a system in which reciprocal processes (for example, vaporization and condensation) occur at equal rates. The pressure exerted by the vapor in equilibrium with a liquid in a closed container at a given temperature is called the liquid's **vapor pressure** (or equilibrium vapor pressure). The area of the surface of the liquid in contact with a vapor and the size of the vessel have no effect on the vapor pressure, although they do affect the time required for the equilibrium to be reached. We can measure the vapor pressure of a liquid by placing a sample in a closed container, like that illustrated in Figure 8.4.1, and using a manometer to measure the increase in pressure that is due to the vapor in equilibrium with the condensed phase.

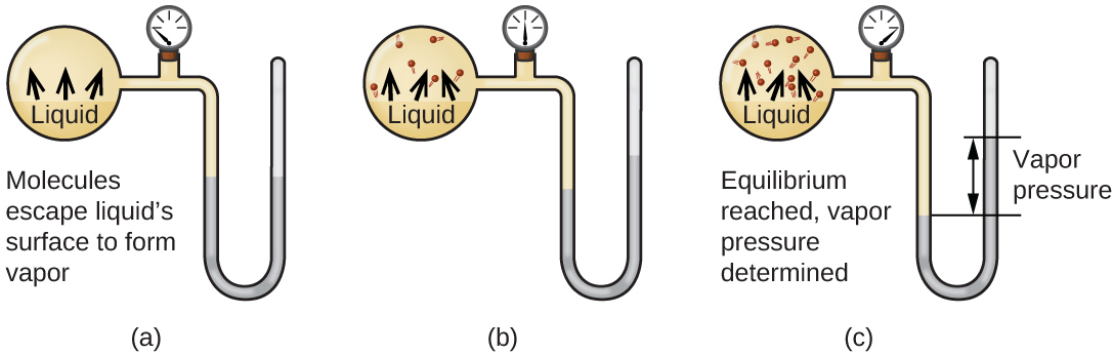
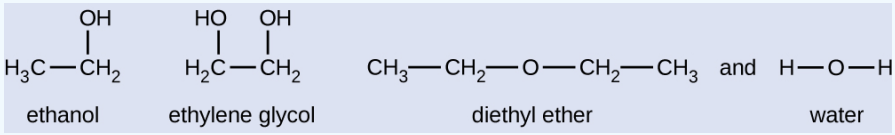


Figure 8.4.1: In a closed container, dynamic equilibrium is reached when (a) the rate of molecules escaping from the liquid to become the gas (b) increases and eventually (c) equals the rate of gas molecules entering the liquid. When this equilibrium is reached, the vapor pressure of the gas is constant, although the vaporization and condensation processes continue.

The chemical identities of the molecules in a liquid determine the types (and strengths) of intermolecular attractions possible; consequently, different substances will exhibit different equilibrium vapor pressures. Relatively strong intermolecular attractive forces will serve to impede vaporization as well as favoring "recapture" of gas-phase molecules when they collide with the liquid surface, resulting in a relatively low vapor pressure. Weak intermolecular attractions present less of a barrier to vaporization, and a reduced likelihood of gas recapture, yielding relatively high vapor pressures. The following example illustrates this dependence of vapor pressure on intermolecular attractive forces.

Example 8.4.1: Explaining Vapor Pressure in Terms of IMFs

Given the shown structural formulas for these four compounds, explain their relative vapor pressures in terms of types and extents of IMFs:



Solution

Diethyl ether has a very small dipole and most of its intermolecular attractions are London forces. Although this molecule is the largest of the four under consideration, its IMFs are the weakest and, as a result, its molecules most readily escape from the liquid. It also has the highest vapor pressure. Due to its smaller size, ethanol exhibits weaker dispersion forces than diethyl ether. However, ethanol is capable of hydrogen bonding and, therefore, exhibits stronger overall IMFs, which means that fewer molecules escape from the liquid at any given temperature, and so ethanol has a lower vapor pressure than diethyl ether. Water is much smaller than either of the previous substances and exhibits weaker dispersion forces, but its extensive hydrogen bonding provides stronger intermolecular attractions, fewer molecules escaping the liquid, and a lower vapor pressure than for either diethyl ether or ethanol. Ethylene glycol has two −OH groups, so, like water, it exhibits extensive hydrogen bonding. It is much larger than water and thus experiences larger London forces. Its overall IMFs are the largest of these four substances, which means its vaporization rate will be the slowest and, consequently, its vapor pressure the lowest.

Exercise 8.4.1

At 20 °C, the vapor pressures of several alcohols are given in this table. Explain these vapor pressures in terms of types and extents of IMFs for these alcohols:

Compound	methanol CH_3OH	ethanol C_2H_5OH	propanol C_3H_7OH	butanol C_4H_9OH
Vapor Pressure at 20 °C	11.9 kPa	5.95 kPa	2.67 kPa	0.56 kPa

Answer

All these compounds exhibit hydrogen bonding; these strong IMFs are difficult for the molecules to overcome, so the vapor pressures are relatively low. As the size of molecule increases from methanol to butanol, dispersion forces increase, which means that the vapor pressures decrease as observed:

$$P_{methanol} > P_{ethanol} > P_{propanol} > P_{butanol}.$$

As temperature increases, the vapor pressure of a liquid also increases due to the increased average KE of its molecules. Recall that at any given temperature, the molecules of a substance experience a range of kinetic energies, with a certain fraction of molecules having a sufficient energy to overcome IMF and escape the liquid (vaporize). At a higher temperature, a greater fraction of molecules have enough energy to escape from the liquid, as shown in Figure 8.4.2. The escape of more molecules per unit of time and the greater average speed of the molecules that escape both contribute to the higher vapor pressure.

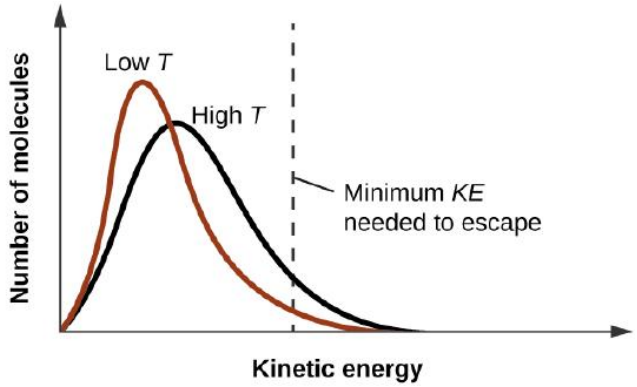


Figure 8.4.2: Temperature affects the distribution of kinetic energies for the molecules in a liquid. At the higher temperature, more molecules have the necessary kinetic energy, KE, to escape from the liquid into the gas phase.

Boiling Points

When the vapor pressure increases enough to equal the external atmospheric pressure, the liquid reaches its **boiling point**. The boiling point of a liquid is the temperature at which its equilibrium vapor pressure is equal to the pressure exerted on the liquid by its gaseous surroundings. For liquids in open containers, this pressure is that due to the earth's atmosphere. The **normal boiling point** of a liquid is defined as its boiling point when surrounding pressure is equal to 1 atm (101.3 kPa). Figure 8.4.3 shows the variation in vapor pressure with temperature for several different substances. Considering the definition of boiling point, these curves may be seen as depicting the dependence of a liquid's boiling point on surrounding pressure.

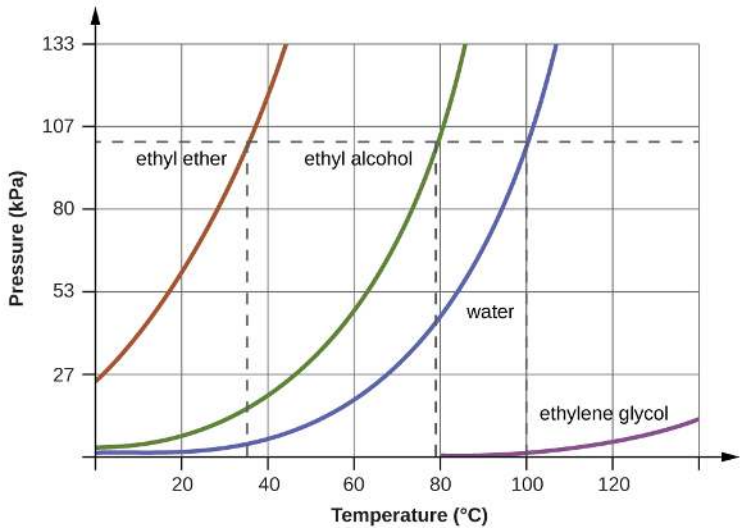


Figure 8.4.3: The boiling points of liquids are the temperatures at which their equilibrium vapor pressures equal the pressure of the surrounding atmosphere. Normal boiling points are those corresponding to a pressure of 1 atm (101.3 kPa.)

Example 8.4.2: A Boiling Point at Reduced Pressure

A typical atmospheric pressure in Leadville, Colorado (elevation 10,200 feet) is 68 kPa. Use the graph in Figure 8.4.3 to determine the boiling point of water at this elevation.

Solution

The graph of the vapor pressure of water versus temperature in Figure 8.4.3 indicates that the vapor pressure of water is 68 kPa at about 90 °C. Thus, at about 90 °C, the vapor pressure of water will equal the atmospheric pressure in Leadville, and water will boil.

Exercise 8.4.2

The boiling point of ethyl ether was measured to be 10 °C at a base camp on the slopes of Mount Everest. Use Figure 8.4.3 to determine the approximate atmospheric pressure at the camp.

Answer

Approximately 40 kPa (0.4 atm)

The quantitative relation between a substance's vapor pressure and its temperature is described by the Clausius-Clapeyron equation:

$$P = Ae^{-\Delta H_{vap}/RT}$$

where ΔH_{vap} is the enthalpy of vaporization for the liquid, R is the gas constant, and A is a constant whose value depends on the chemical identity of the substance. Temperature T must be in Kelvin in this equation. This equation is often rearranged into logarithmic form to yield the linear equation:

$$\ln P = -\frac{\Delta H_{vap}}{RT} + \ln A$$

This linear equation may be expressed in a two-point format that is convenient for use in various computations, as demonstrated in the example exercises that follow. If at temperature T_1, the vapor pressure is P_1, and at temperature T_2, the vapor pressure is P_2, the corresponding linear equations are:

$$\ln P_1 = -\frac{\Delta H_{vap}}{RT_1} + \ln A$$

and

$$\ln P_2 = -\frac{\Delta H_{vap}}{RT_2} + \ln A$$

Since the constant, A, is the same, these two equations may be rearranged to isolate $\ln A$ and then set them equal to one another:

$$\ln P_1 + \frac{\Delta H_{\text{vap}}}{RT_1} = \ln P_2 + \frac{\Delta H_{\text{vap}}}{RT_2}$$

which can be combined into:

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{\text{vap}}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Example 8.4.3: Estimating Enthalpy of Vaporization

Isooctane (2,2,4-trimethylpentane) has an octane rating of 100. It is used as one of the standards for the octane-rating system for gasoline. At 34.0 °C, the vapor pressure of isooctane is 10.0 kPa, and at 98.8 °C, its vapor pressure is 100.0 kPa. Use this information to estimate the enthalpy of vaporization for isooctane.

Solution

The enthalpy of vaporization, ΔH_{vap}, can be determined by using the Clausius-Clapeyron equation:

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{\text{vap}}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Since we have two vapor pressure-temperature values (T_1 = 34.0 °C = 307.2 K, P_1 = 10.0 kPa and T_2 = 98.8 °C = 372.0 K, P_2 = 100 kPa), we can substitute them into this equation and solve for ΔH_{vap}. Rearranging the Clausius-Clapeyron equation and solving for ΔH_{vap} yields:

$$\Delta H_{\text{vap}} = \frac{R \cdot \ln\left(\frac{P_2}{P_1}\right)}{\left(\frac{1}{T_1} - \frac{1}{T_2}\right)} = \frac{(8.3145 J/mol \cdot K) \cdot \ln\left(\frac{100 kPa}{10.0 kPa}\right)}{\left(\frac{1}{307.2 K} - \frac{1}{372.0 K}\right)} = 33,800 J/mol = 33.8 kJ/mol$$

Note that the pressure can be in any units, so long as they agree for both *P* values, but the temperature must be in kelvin for the Clausius-Clapeyron equation to be valid.

Exercise 8.4.3

At 20.0 °C, the vapor pressure of ethanol is 5.95 kPa, and at 63.5 °C, its vapor pressure is 53.3 kPa. Use this information to estimate the enthalpy of vaporization for ethanol.

Answer

41,360 J/mol or 41.4 kJ/mol

Example 8.4.4: Estimating Temperature (or Vapor Pressure)

For benzene (C_6H_6), the normal boiling point is 80.1 °C and the enthalpy of vaporization is 30.8 kJ/mol. What is the boiling point of benzene in Denver, where atmospheric pressure = 83.4 kPa?

Solution

If the temperature and vapor pressure are known at one point, along with the enthalpy of vaporization, ΔH_{vap}, then the temperature that corresponds to a different vapor pressure (or the vapor pressure that corresponds to a different temperature) can be determined by using the Clausius-Clapeyron equation:

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{\text{vap}}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Since the normal boiling point is the temperature at which the vapor pressure equals atmospheric pressure at sea level, we know one vapor pressure-temperature value (T_1 = 80.1 °C = 353.3 K, P_1 = 101.3 kPa, ΔH_{vap} = 30.8 kJ/mol) and want to find the temperature (T_2) that corresponds to vapor pressure P_2 = 83.4 kPa. We can substitute these values into the Clausius-Clapeyron equation and then solve for T_2. Rearranging the Clausius-Clapeyron equation and solving for T_2 yields:

$$T_2 = \left(\frac{-R \cdot \ln\left(\frac{P_2}{P_1}\right)}{\Delta H_{\text{vap}}} + \frac{1}{T_1}\right)^{-1} = \left(\frac{-(8.3145 J/mol \cdot K) \cdot \ln\left(\frac{83.4 kPa}{101.3 kPa}\right)}{30,800 J/mol} + \frac{1}{353.3 K}\right)^{-1} = 346.9$$

Exercise 8.4.4

For acetone $(CH_3)_2CO$, the normal boiling point is 56.5 °C and the enthalpy of vaporization is 31.3 kJ/mol. What is the vapor pressure of acetone at 25.0 °C?

Answer

30.1 kPa

Enthalpy of Vaporization

Vaporization is an endothermic process. The cooling effect can be evident when you leave a swimming pool or a shower. When the water on your skin evaporates, it removes heat from your skin and causes you to feel cold. The energy change associated with the vaporization process is the enthalpy of vaporization, ΔH_{vap}. For example, the vaporization of water at standard temperature is represented by:

$$H_2O(l) \longrightarrow H_2O(g) \qquad \Delta H_{vap} = 44.01 kJ/mol$$

As described in the chapter on thermochemistry, the reverse of an endothermic process is exothermic. And so, the condensation of a gas releases heat:

$$\mathrm{H_2O(g) \longrightarrow H_2O(l)} \qquad \Delta H_{\text{con}} = -\Delta H_{\text{vap}} = -44.01 kJ/mol$$

Example 8.4.5: Using Enthalpy of Vaporization

One way our body is cooled is by evaporation of the water in sweat (Figure 8.4.4). In very hot climates, we can lose as much as 1.5 L of sweat per day. Although sweat is not pure water, we can get an approximate value of the amount of heat removed by evaporation by assuming that it is. How much heat is required to evaporate 1.5 L of water (1.5 kg) at T = 37 °C (normal body temperature); ΔH_{vap} = 43.46 kJ/mol at 37 °C.



Figure 8.4.4: Evaporation of sweat helps cool the body. (credit: “Kullez”/Flickr)

Solution

We start with the known volume of sweat (approximated as just water) and use the given information to convert to the amount of heat needed:

$$1.5 \cancel{\text{ L}} \times \frac{1000 \cancel{\text{ g}}}{1 \cancel{\text{ L}}} \times \frac{1 \cancel{\text{ mol}}}{18 \cancel{\text{ g}}} \times \frac{43.46 \text{ kJ}}{1 \cancel{\text{ mol}}} = 3.6 \times 10^3 \text{ kJ}$$

Thus, 3600 kJ of heat are removed by the evaporation of 1.5 L of water.

Exercise 8.4.1

How much heat is required to evaporate 100.0 g of liquid ammonia, NH_3, at its boiling point if its enthalpy of vaporization is 4.8 kJ/mol?

Answer

28 kJ

Melting and Freezing

When we heat a crystalline solid, we increase the average energy of its atoms, molecules, or ions and the solid gets hotter. At some point, the added energy becomes large enough to partially overcome the forces holding the molecules or ions of the solid in their fixed positions, and the solid begins the process of transitioning to the liquid state, or melting. At this point, the temperature of the solid stops rising, despite the continual input of heat, and it remains constant until all of the solid is melted. Only after all of the solid has melted will continued heating increase the temperature of the liquid (Figure 8.4.5).

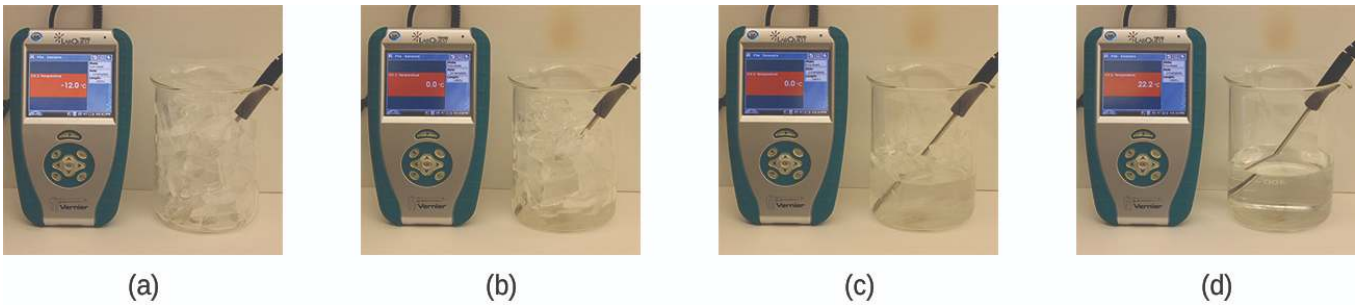


Figure 8.4.5: (a) This beaker of ice has a temperature of −12.0 °C. (b) After 10 minutes the ice has absorbed enough heat from the air to warm to 0 °C. A small amount has melted. (c) Thirty minutes later, the ice has absorbed more heat, but its temperature is still 0 °C. The ice melts without changing its temperature. (d) Only after all the ice has melted does the heat absorbed cause the temperature to increase to 22.2 °C. (credit: modification of work by Mark Ott)

If we stop heating during melting and place the mixture of solid and liquid in a perfectly insulated container so no heat can enter or escape, the solid and liquid phases remain in equilibrium. This is almost the situation with a mixture of ice and water in a very good thermos bottle; almost no heat gets in or out, and the mixture of solid ice and liquid water remains for hours. In a mixture of solid and liquid at equilibrium, the reciprocal processes of melting and **freezing** occur at equal rates, and the quantities of solid and liquid therefore remain constant. The temperature at which the solid and liquid phases of a given substance are in equilibrium is called the **melting point** of the solid or the freezing point of the liquid. Use of one term or the other is normally dictated by the direction of the phase transition being considered, for example, solid to liquid (melting) or liquid to solid (freezing).

The enthalpy of fusion and the melting point of a crystalline solid depend on the strength of the attractive forces between the units present in the crystal. Molecules with weak attractive forces form crystals with low melting points. Crystals consisting of particles with stronger attractive forces melt at higher temperatures.

The amount of heat required to change one mole of a substance from the solid state to the liquid state is the enthalpy of fusion, ΔH_{fus} of the substance. The enthalpy of fusion of ice is 6.0 kJ/mol at 0 °C. Fusion (melting) is an endothermic process:

$$\mathrm{H_2O(s)} \longrightarrow \mathrm{H_2O(l)} \quad \Delta H_{\text{fus}} = 6.01 kJ/mol$$

The reciprocal process, freezing, is an exothermic process whose enthalpy change is −6.0 kJ/mol at 0 °C:

$$\mathrm{H_2O(l)} \longrightarrow \mathrm{H_2O(s)} \quad \Delta H_{\text{frz}} = -\Delta H_{\text{fus}} = -6.01 kJ/mol$$

Sublimation and Deposition

Some solids can transition directly into the gaseous state, bypassing the liquid state, via a process known as **sublimation**. At room temperature and standard pressure, a piece of dry ice (solid $\mathrm{CO_2}$) sublimes, appearing to gradually disappear without ever forming any liquid. Snow and ice sublime at temperatures below the melting point of water, a slow process that may be accelerated by winds and the reduced atmospheric pressures at high altitudes. When solid iodine is warmed, the solid sublimes and a vivid purple vapor forms (Figure 8.4.6). The reverse of sublimation is called **deposition**, a process in which gaseous substances condense directly into the solid state, bypassing the liquid state. The formation of frost is an example of deposition.



Figure 8.4.6: Sublimation of solid iodine in the bottom of the tube produces a purple gas that subsequently deposits as solid iodine on the colder part of the tube above. (credit: modification of work by Mark Ott)

Like vaporization, the process of sublimation requires an input of energy to overcome intermolecular attractions. The enthalpy of sublimation, ΔH_{sub}, is the energy required to convert one mole of a substance from the solid to the gaseous state. For example, the sublimation of carbon dioxide is represented by:

$$\mathrm{CO_2(s) \longrightarrow CO_2(g)} \qquad \Delta H_{\text{sub}} = 26.1 kJ/mol$$

Likewise, the enthalpy change for the reverse process of deposition is equal in magnitude but opposite in sign to that for sublimation:

$$\mathrm{CO_2(g) \longrightarrow CO_2(s)} \qquad \Delta H_{\text{dep}} = -\Delta H_{\text{sub}} = -26.1 kJ/mol$$

Consider the extent to which intermolecular attractions must be overcome to achieve a given phase transition. Converting a solid into a liquid requires that these attractions be only partially overcome; transition to the gaseous state requires that they be completely overcome. As a result, the enthalpy of fusion for a substance is less than its enthalpy of vaporization. This same logic can be used to derive an approximate relation between the enthalpies of all phase changes for a given substance. Though not an entirely accurate description, sublimation may be conveniently modeled as a sequential two-step process of melting followed by vaporization in order to apply Hess's Law. Viewed in this manner, the enthalpy of sublimation for a substance may be estimated as the sum of its enthalpies of fusion and vaporization, as illustrated in Figure 8.4.7. For example:

$$\begin{array}{ll} \text{solid} \longrightarrow \text{liquid} & \Delta H_{\text{fus}} \\ \text{liquid} \longrightarrow \text{gas} & \Delta H_{\text{vap}} \\ \hline \text{solid} \longrightarrow \text{gas} & \Delta H_{\text{sub}} = \Delta H_{\text{fus}} + \Delta H_{\text{vap}} \end{array}$$

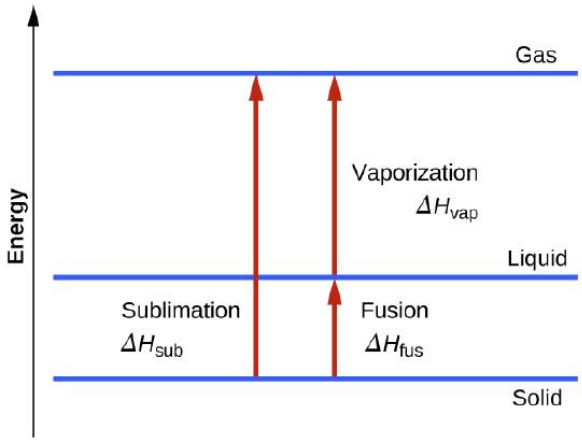


Figure 8.4.7: For a given substance, the sum of its enthalpy of fusion and enthalpy of vaporization is approximately equal to its enthalpy of sublimation.

Heating and Cooling Curves

In the chapter on thermochemistry, the relation between the amount of heat absorbed or released by a substance, q, and its accompanying temperature change, ΔT, was introduced:

$$q = mc\Delta T$$

where m is the mass of the substance and c is its specific heat. The relation applies to matter being heated or cooled, but not undergoing a change in state. When a substance being heated or cooled reaches a temperature corresponding to one of its phase transitions, further gain or loss of heat is a result of diminishing or enhancing intermolecular attractions, instead of increasing or decreasing molecular kinetic energies. While a substance is undergoing a change in state, its temperature remains constant. Figure 8.4.8 shows a typical heating curve.

Consider the example of heating a pot of water to boiling. A stove burner will supply heat at a roughly constant rate; initially, this heat serves to increase the water's temperature. When the water reaches its boiling point, the temperature remains constant despite the continued input of heat from the stove burner. This same temperature is maintained by the water as long as it is boiling. If the burner setting is increased to provide heat at a greater rate, the water temperature does not rise, but instead the boiling becomes more vigorous (rapid). This behavior is observed for other phase transitions as well: For example, temperature remains constant while the change of state is in progress.

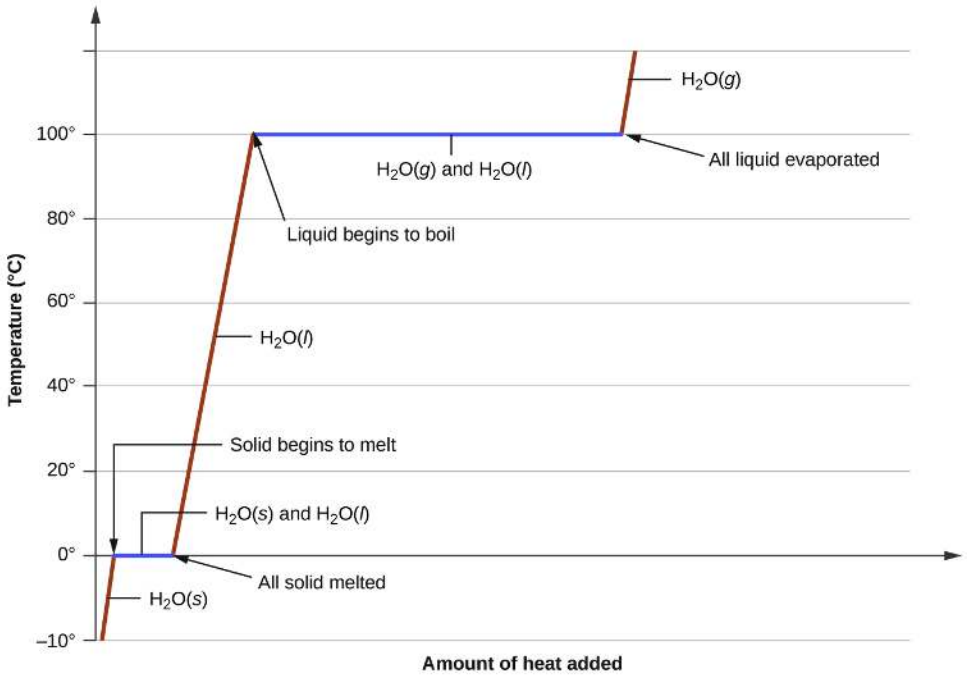


Figure 8.4.8: A typical heating curve for a substance depicts changes in temperature that result as the substance absorbs increasing amounts of heat. Plateaus in the curve (regions of constant temperature) are exhibited when the substance undergoes phase transitions.

Example 8.4.6: Total Heat Needed to Change Temperature and Phase for a Substance

How much heat is required to convert 135 g of ice at −15 °C into water vapor at 120 °C?

Solution

The transition described involves the following steps:

1. Heat ice from −15 °C to 0 °C
2. Melt ice
3. Heat water from 0 °C to 100 °C
4. Boil water
5. Heat steam from 100 °C to 120 °C

The heat needed to change the temperature of a given substance (with no change in phase) is: $q = mc\Delta T$. The heat needed to induce a given change in phase is given by $q = n\Delta H$.

Using these equations with the appropriate values for specific heat of ice, water, and steam, and enthalpies of fusion and vaporization, we have:

$$\begin{aligned} q_{\text{total}} = \; & (m \cdot c \cdot \Delta T)_{\text{ice}} + n \cdot \Delta H_{\text{fus}} + (m \cdot c \cdot \Delta T)_{\text{water}} + n \cdot \Delta H_{\text{vap}} + (m \cdot c \cdot \Delta T)_{\text{steam}} \\ = \; & (135g \cdot 2.09J/g \cdot {}^\circ C \cdot 15^\circ C) + \left(135 \cdot \frac{1mol}{18.02g} \cdot 6.01kJ/mol\right) \\ & + (135g \cdot 4.18J/g \cdot {}^\circ C \cdot 100^\circ C) + \left(135g \cdot \frac{1mol}{18.02g} \cdot 40.67kJ/mol\right) \\ & + (135g \cdot 1.86J/g \cdot {}^\circ C \cdot 20^\circ C) \\ = \; & 4230J + 45.0kJ + 56,500J + 305kJ + 5022J \end{aligned}$$

Converting the quantities in J to kJ permits them to be summed, yielding the total heat required:

$$q_{\text{total}} = 4.23kJ + 45.0kJ + 56.5kJ + 305kJ + 5.02kJ = 416kJ$$

NOTE: The value of ΔH_{vap} at the boiling point of water (40.67 kJ/mol) is used here instead of the value at standard temperature (44.01 kJ/mol).

Exercise 8.4.6

What is the total amount of heat released when 94.0 g water at 80.0 °C cools to form ice at −30.0 °C?

Answer

68.7 kJ

This page titled 8.4: Phase Transitions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **10.3: Phase Transitions** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.5: Phase Diagrams

Learning Objectives

By the end of this section, you will be able to:

- Explain the construction and use of a typical phase diagram
- Use phase diagrams to identify stable phases at given temperatures and pressures, and to describe phase transitions resulting from changes in these properties
- Describe the supercritical fluid phase of matter

In the previous module, the variation of a liquid's equilibrium vapor pressure with temperature was described. Considering the definition of boiling point, plots of vapor pressure versus temperature represent how the boiling point of the liquid varies with pressure. Also described was the use of heating and cooling curves to determine a substance's melting (or freezing) point. Making such measurements over a wide range of pressures yields data that may be presented graphically as a phase diagram. A **phase diagram** combines plots of pressure versus temperature for the liquid-gas, solid-liquid, and solid-gas phase-transition equilibria of a substance. These diagrams indicate the physical states that exist under specific conditions of pressure and temperature, and also provide the pressure dependence of the phase-transition temperatures (melting points, sublimation points, boiling points). A typical phase diagram for a pure substance is shown in Figure 8.5.1.

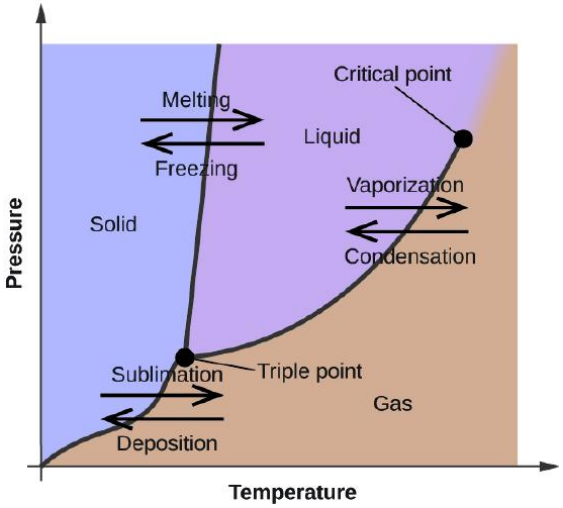


Figure 8.5.1: The physical state of a substance and its phase-transition temperatures are represented graphically in a phase diagram.

To illustrate the utility of these plots, consider the phase diagram for water shown in Figure 8.5.2.

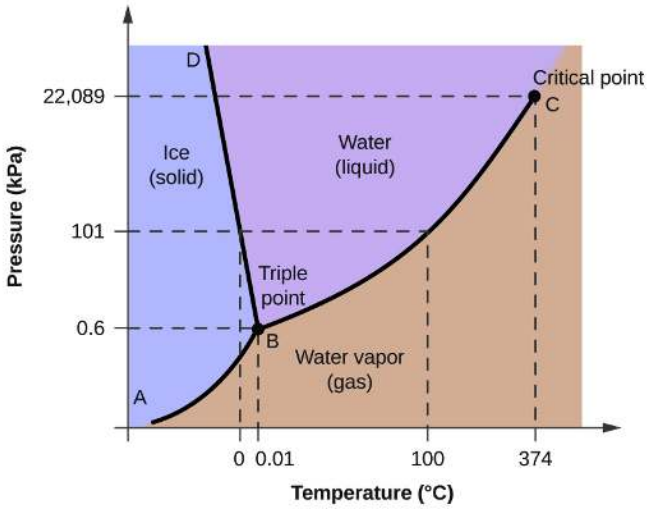


Figure 8.5.2: The pressure and temperature axes on this phase diagram of water are not drawn to constant scale in order to illustrate several important properties.

We can use the phase diagram to identify the physical state of a sample of water under specified conditions of pressure and temperature. For example, a pressure of 50 kPa and a temperature of −10 °C correspond to the region of the diagram labeled “ice.” Under these conditions, water exists only as a solid (ice). A pressure of 50 kPa and a temperature of 50 °C correspond to the “water” region—here, water exists only as a liquid. At 25 kPa and 200 °C, water exists only in the gaseous state. Note that on the H_2O phase diagram, the pressure and temperature axes are not drawn to a constant scale in order to permit the illustration of several important features as described here.

The curve BC in Figure 8.5.2 is the plot of vapor pressure versus temperature as described in the previous module of this chapter. This “liquid-vapor” curve separates the liquid and gaseous regions of the phase diagram and provides the boiling point for water at any pressure. For example, at 1 atm, the boiling point is 100 °C. Notice that the liquid-vapor curve terminates at a temperature of 374 °C and a pressure of 218 atm, indicating that water cannot exist as a liquid above this temperature, regardless of the pressure. The physical properties of water under these conditions are intermediate between those of its liquid and gaseous phases. This unique state of matter is called a supercritical fluid, a topic that will be described in the next section of this module.

The solid-vapor curve, labeled AB in Figure 8.5.2, indicates the temperatures and pressures at which ice and water vapor are in equilibrium. These temperature-pressure data pairs correspond to the sublimation, or deposition, points for water. If we could zoom in on the solid-gas line in Figure 8.5.2, we would see that ice has a vapor pressure of about 0.20 kPa at −10 °C. Thus, if we place a frozen sample in a vacuum with a pressure less than 0.20 kPa, ice will sublime. This is the basis for the “freeze-drying” process often used to preserve foods, such as the ice cream shown in Figure 8.5.3.



Figure 8.5.3: Freeze-dried foods, like this ice cream, are dehydrated by sublimation at pressures below the triple point for water. (credit: "lwao"/Flickr)

The solid-liquid curve labeled BD shows the temperatures and pressures at which ice and liquid water are in equilibrium, representing the melting/freezing points for water. Note that this curve exhibits a slight negative slope (greatly exaggerated for

clarity), indicating that the melting point for water decreases slightly as pressure increases. Water is an unusual substance in this regard, as most substances exhibit an increase in melting point with increasing pressure. This behavior is partly responsible for the movement of glaciers, like the one shown in Figure 8.5.4. The bottom of a glacier experiences an immense pressure due to its weight that can melt some of the ice, forming a layer of liquid water on which the glacier may more easily slide.

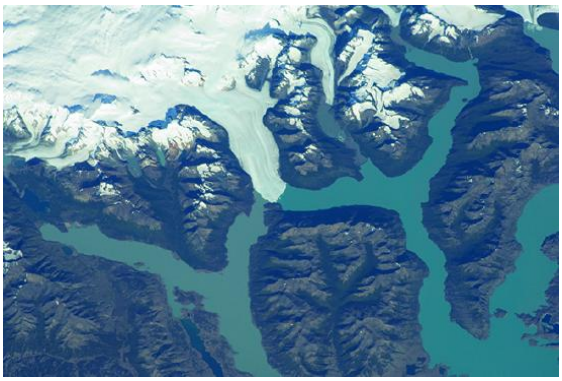


Figure 8.5.4: The immense pressures beneath glaciers result in partial melting to produce a layer of water that provides lubrication to assist glacial movement. This satellite photograph shows the advancing edge of the Perito Moreno glacier in Argentina. (credit: NASA)

The point of intersection of all three curves is labeled B in Figure 8.5.2. At the pressure and temperature represented by this point, three phases of water coexist in equilibrium. This temperature-pressure data pair is called the triple point. At pressures lower than the triple point, water cannot exist as a liquid, regardless of the temperature.

✔ Example 8.5.1: Determining the State of Water

Using the phase diagram for water given in Figure 8.5.2, determine the state of water at the following temperatures and pressures:

a. −10 °C and 50 kPa
b. 25 °C and 90 kPa
c. 50 °C and 40 kPa
d. 80 °C and 5 kPa
e. −10 °C and 0.3 kPa
f. 50 °C and 0.3 kPa

Solution

Using the phase diagram for water, we can determine that the state of water at each temperature and pressure given are as follows: (a) solid; (b) liquid; (c) liquid; (d) gas; (e) solid; (f) gas.

? Exercise 8.5.1

What phase changes can water undergo as the temperature changes if the pressure is held at 0.3 kPa? If the pressure is held at 50 kPa?

Answer

At 0.3 kPa: at −58 °C. At 50 kPa: at 0 °C, l ⟶ g at 78 °C

Consider the phase diagram for carbon dioxide shown in Figure 8.5.5 as another example. The solid-liquid curve exhibits a positive slope, indicating that the melting point for CO_2 increases with pressure as it does for most substances (water being a notable exception as described previously). Notice that the triple point is well above 1 atm, indicating that carbon dioxide cannot exist as a liquid under ambient pressure conditions. Instead, cooling gaseous carbon dioxide at 1 atm results in its deposition into the solid state. Likewise, solid carbon dioxide does not melt at 1 atm pressure but instead sublimes to yield gaseous CO_2. Finally, notice that the critical point for carbon dioxide is observed at a relatively modest temperature and pressure in comparison to water.

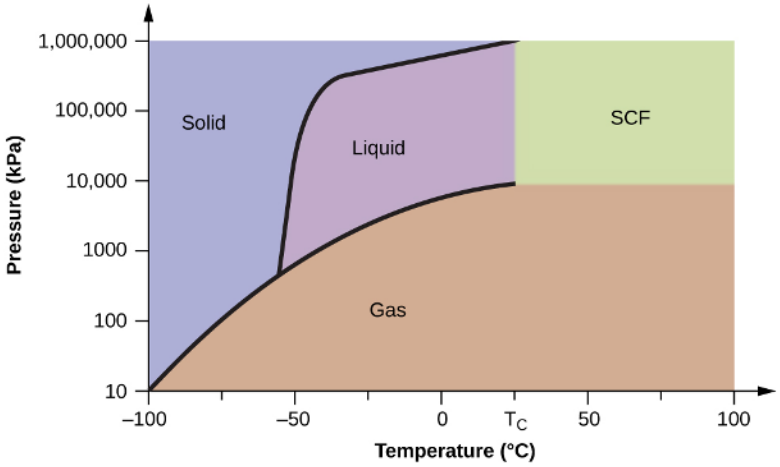


Figure 8.5.5: A phase diagram for carbon dioxide is shown. The pressure axis is plotted on a logarithmic scale to accommodate the large range of values.

Example 8.5.2: Determining the State of Carbon Dioxide

Using the phase diagram for carbon dioxide shown in Figure 8.5.5, determine the state of CO_2 at the following temperatures and pressures:

a. −30 °C and 2000 kPa
b. −90 °C and 1000 kPa
c. −60 °C and 100 kPa
d. −40 °C and 1500 kPa
e. 0 °C and 100 kPa
f. 20 °C and 100 kPa

Solution

Using the phase diagram for carbon dioxide provided, we can determine that the state of CO_2 at each temperature and pressure given are as follows: (a) liquid; (b) solid; (c) gas; (d) liquid; (e) gas; (f) gas.

Exercise 8.5.2

Identify the phase changes that carbon dioxide will undergo as its temperature is increased from −100 °C while holding its pressure constant at 1500 kPa. At 50 kPa. At what approximate temperatures do these phase changes occur?

Answer

at 1500 kPa: at −55 °C, at −10 °C;

at 50 kPa: at −60 °C

Supercritical Fluids

If we place a sample of water in a sealed container at 25 °C, remove the air, and let the vaporization-condensation equilibrium establish itself, we are left with a mixture of liquid water and water vapor at a pressure of 0.03 atm. A distinct boundary between the more dense liquid and the less dense gas is clearly observed. As we increase the temperature, the pressure of the water vapor increases, as described by the liquid-gas curve in the phase diagram for water (Figure 8.5.2), and a two-phase equilibrium of liquid and gaseous phases remains. At a temperature of 374 °C, the vapor pressure has risen to 218 atm, and any further increase in temperature results in the disappearance of the boundary between liquid and vapor phases. All of the water in the container is now present in a single phase whose physical properties are intermediate between those of the gaseous and liquid states. This phase of matter is called a **supercritical fluid**, and the temperature and pressure above which this phase exists is the **critical point** (Figure 8.5.6). Above its critical temperature, a gas cannot be liquefied no matter how much pressure is applied. The pressure required to

liquefy a gas at its critical temperature is called the critical pressure. The critical temperatures and critical pressures of some common substances are given in the following table.

Substance	Critical Temperature (°C)	Critical Pressure (kPa)
hydrogen	−240.0	1300
nitrogen	−147.2	3400
oxygen	−118.9	5000
carbon dioxide	31.1	7400
ammonia	132.4	11,300
sulfur dioxide	157.2	7800
water	374.0	22,000

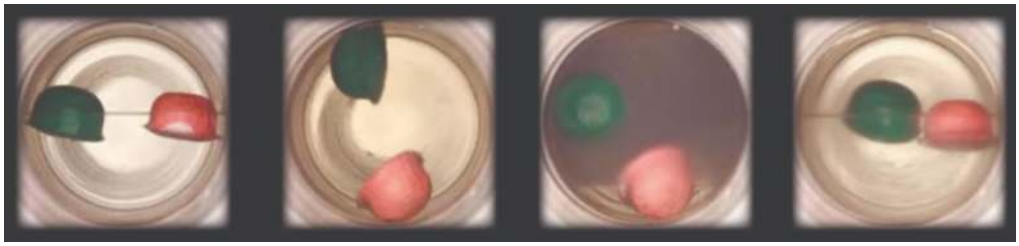


Figure 8.5.6: (a) A sealed container of liquid carbon dioxide slightly below its critical point is heated, resulting in (b) the formation of the supercritical fluid phase. Cooling the supercritical fluid lowers its temperature and pressure below the critical point, resulting in the reestablishment of separate liquid and gaseous phases (c and d). Colored floats illustrate differences in density between the liquid, gaseous, and supercritical fluid states. (credit: modification of work by “mrmrobin”/YouTube)

Link to Learning

Observe the liquid-to-supercritical fluid transition for carbon dioxide.

Like a gas, a supercritical fluid will expand and fill a container, but its density is much greater than typical gas densities, typically being close to those for liquids. Similar to liquids, these fluids are capable of dissolving nonvolatile solutes. They exhibit essentially no surface tension and very low viscosities, however, so they can more effectively penetrate very small openings in a solid mixture and remove soluble components. These properties make supercritical fluids extremely useful solvents for a wide range of applications. For example, supercritical carbon dioxide has become a very popular solvent in the food industry, being used to decaffeinate coffee, remove fats from potato chips, and extract flavor and fragrance compounds from citrus oils. It is nontoxic, relatively inexpensive, and not considered to be a pollutant. After use, the CO_2 can be easily recovered by reducing the pressure and collecting the resulting gas.

Example 8.5.3: The Critical Temperature of Carbon Dioxide

If we shake a carbon dioxide fire extinguisher on a cool day (18 °C), we can hear liquid CO_2 sloshing around inside the cylinder. However, the same cylinder appears to contain no liquid on a hot summer day (35 °C). Explain these observations.

Solution

On the cool day, the temperature of the CO_2 is below the critical temperature of CO_2, 304 K or 31 °C, so liquid CO_2 is present in the cylinder. On the hot day, the temperature of the CO_2 is greater than its critical temperature of 31 °C. Above this temperature no amount of pressure can liquefy CO_2 so no liquid CO_2 exists in the fire extinguisher.

Exercise 8.5.3

Ammonia can be liquefied by compression at room temperature; oxygen cannot be liquefied under these conditions. Why do the two gases exhibit different behavior?

Answer

The critical temperature of ammonia is 405.5 K, which is higher than room temperature. The critical temperature of oxygen is below room temperature; thus oxygen cannot be liquefied at room temperature.

Chemistry in Everyday Life: Decaffeinating Coffee Using Supercritical CO_2

Coffee is the world's second most widely traded commodity, following only petroleum. Across the globe, people love coffee's aroma and taste. Many of us also depend on one component of coffee—caffeine—to help us get going in the morning or stay alert in the afternoon. But late in the day, coffee's stimulant effect can keep you from sleeping, so you may choose to drink decaffeinated coffee in the evening.

Since the early 1900s, many methods have been used to decaffeinate coffee. All have advantages and disadvantages, and all depend on the physical and chemical properties of caffeine. Because caffeine is a somewhat polar molecule, it dissolves well in water, a polar liquid. However, since many of the other 400-plus compounds that contribute to coffee's taste and aroma also dissolve in H_2O, hot water decaffeination processes can also remove some of these compounds, adversely affecting the smell and taste of the decaffeinated coffee. Dichloromethane (CH_2Cl_2) and ethyl acetate ($CH_3CO_2C_2H_5$) have similar polarity to caffeine, and are therefore very effective solvents for caffeine extraction, but both also remove some flavor and aroma components, and their use requires long extraction and cleanup times. Because both of these solvents are toxic, health concerns have been raised regarding the effect of residual solvent remaining in the decaffeinated coffee.

Supercritical fluid extraction using carbon dioxide is now being widely used as a more effective and environmentally friendly decaffeination method (Figure 8.5.7). At temperatures above 304.2 K and pressures above 7376 kPa, CO_2 is a supercritical fluid, with properties of both gas and liquid. Like a gas, it penetrates deep into the coffee beans; like a liquid, it effectively dissolves certain substances. Supercritical carbon dioxide extraction of steamed coffee beans removes 97−99% of the caffeine, leaving coffee's flavor and aroma compounds intact. Because CO_2 is a gas under standard conditions, its removal from the extracted coffee beans is easily accomplished, as is the recovery of the caffeine from the extract. The caffeine recovered from coffee beans via this process is a valuable product that can be used subsequently as an additive to other foods or drugs.

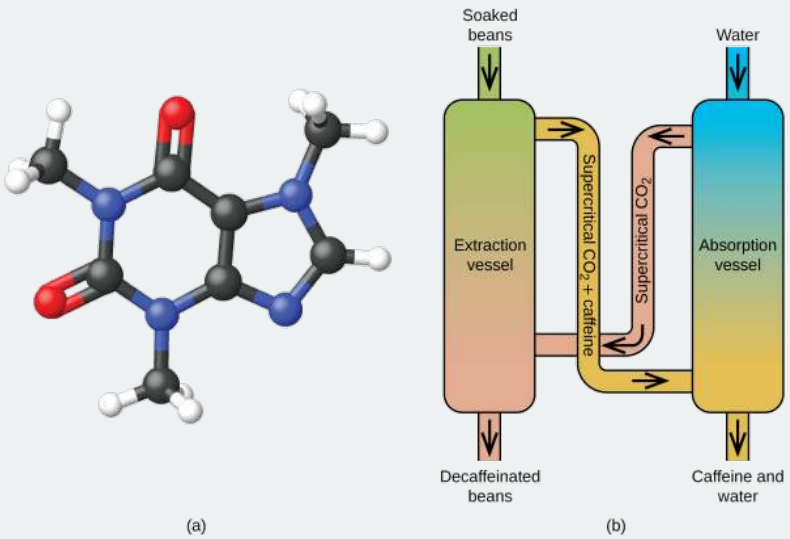


Figure 8.5.7: (a) Caffeine molecules have both polar and nonpolar regions, making it soluble in solvents of varying polarities. (b) The schematic shows a typical decaffeination process involving supercritical carbon dioxide.

This page titled 8.5: Phase Diagrams is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **10.4: Phase Diagrams** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.6: The Solid State of Matter

Learning Objectives

By the end of this section, you will be able to:

- Define and describe the bonding and properties of ionic, molecular, metallic, and covalent network crystalline solids
- Describe the main types of crystalline solids: ionic solids, metallic solids, covalent network solids, and molecular solids
- Explain the ways in which crystal defects can occur in a solid

When most liquids are cooled, they eventually freeze and form **crystalline solids**, solids in which the atoms, ions, or molecules are arranged in a definite repeating pattern. It is also possible for a liquid to freeze before its molecules become arranged in an orderly pattern. The resulting materials are called **amorphous solids** or noncrystalline solids (or, sometimes, glasses). The particles of such solids lack an ordered internal structure and are randomly arranged (Figure 8.6.1).

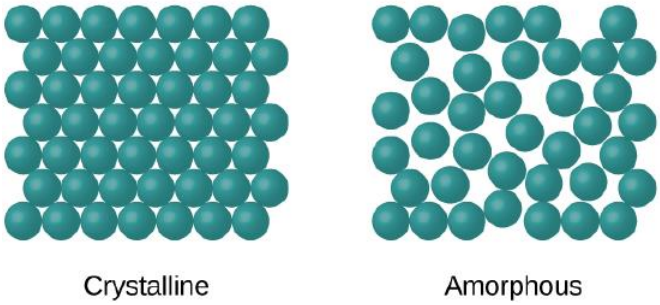


Figure 8.6.1: The entities of a solid phase may be arranged in a regular, repeating pattern (crystalline solids) or randomly (amorphous).

Metals and ionic compounds typically form ordered, crystalline solids. Substances that consist of large molecules, or a mixture of molecules whose movements are more restricted, often form amorphous solids. For examples, candle waxes are amorphous solids composed of large hydrocarbon molecules. Some substances, such as silicon dioxide (shown in Figure 8.6.2), can form either crystalline or amorphous solids, depending on the conditions under which it is produced. Also, amorphous solids may undergo a transition to the crystalline state under appropriate conditions.

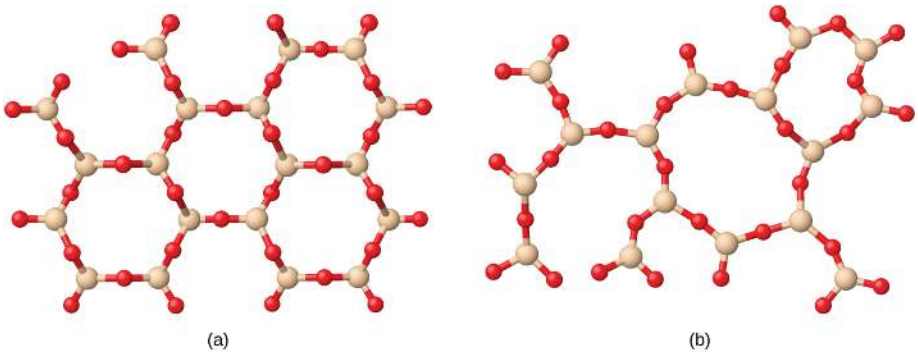


Figure 8.6.2: (a) Silicon dioxide, SiO_2, is abundant in nature as one of several crystalline forms of the mineral quartz. (b) Rapid cooling of molten SiO_2 yields an amorphous solid known as “fused silica”.

Crystalline solids are generally classified according to the nature of the forces that hold its particles together. These forces are primarily responsible for the physical properties exhibited by the bulk solids. The following sections provide descriptions of the major types of crystalline solids: ionic, metallic, covalent network, and molecular.

Ionic Solids

Ionic solids, such as sodium chloride and nickel oxide, are composed of positive and negative ions that are held together by electrostatic attractions, which can be quite strong (Figure 8.6.3). Many ionic crystals also have high melting points. This is due to the very strong attractions between the ions—in ionic compounds, the attractions between full charges are (much) larger than those between the partial charges in polar molecular compounds. This will be looked at in more detail in a later discussion of lattice

energies. Although they are hard, they also tend to be brittle, and they shatter rather than bend. Ionic solids do not conduct electricity; however, they do conduct when molten or dissolved because their ions are free to move. Many simple compounds formed by the reaction of a metallic element with a nonmetallic element are ionic.

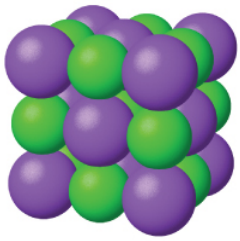


Figure 8.6.3: Sodium chloride is an ionic solid.

Metallic Solids

Metallic solids such as crystals of copper, aluminum, and iron are formed by metal atoms Figure 8.6.4. The structure of metallic crystals is often described as a uniform distribution of atomic nuclei within a "sea" of delocalized electrons. The atoms within such a metallic solid are held together by a unique force known as *metallic bonding* that gives rise to many useful and varied bulk properties. All exhibit high thermal and electrical conductivity, metallic luster, and malleability. Many are very hard and quite strong. Because of their malleability (the ability to deform under pressure or hammering), they do not shatter and, therefore, make useful construction materials. The melting points of the metals vary widely. Mercury is a liquid at room temperature, and the alkali metals melt below 200 °C. Several post-transition metals also have low melting points, whereas the transition metals melt at temperatures above 1000 °C. These differences reflect differences in strengths of metallic bonding among the metals.

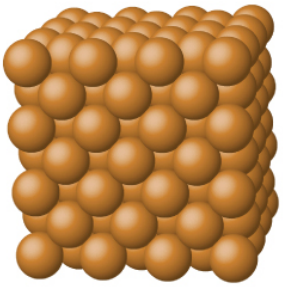


Figure 8.6.4: Copper is a metallic solid.

Covalent Network Solids

Covalent network solids include crystals of diamond, silicon, some other nonmetals, and some covalent compounds such as silicon dioxide (sand) and silicon carbide (carborundum, the abrasive on sandpaper). Many minerals have networks of covalent bonds. The atoms in these solids are held together by a network of covalent bonds, as shown in Figure 8.6.5. To break or to melt a covalent network solid, covalent bonds must be broken. Because covalent bonds are relatively strong, covalent network solids are typically characterized by hardness, strength, and high melting points. For example, diamond is one of the hardest substances known and melts above 3500 °C.

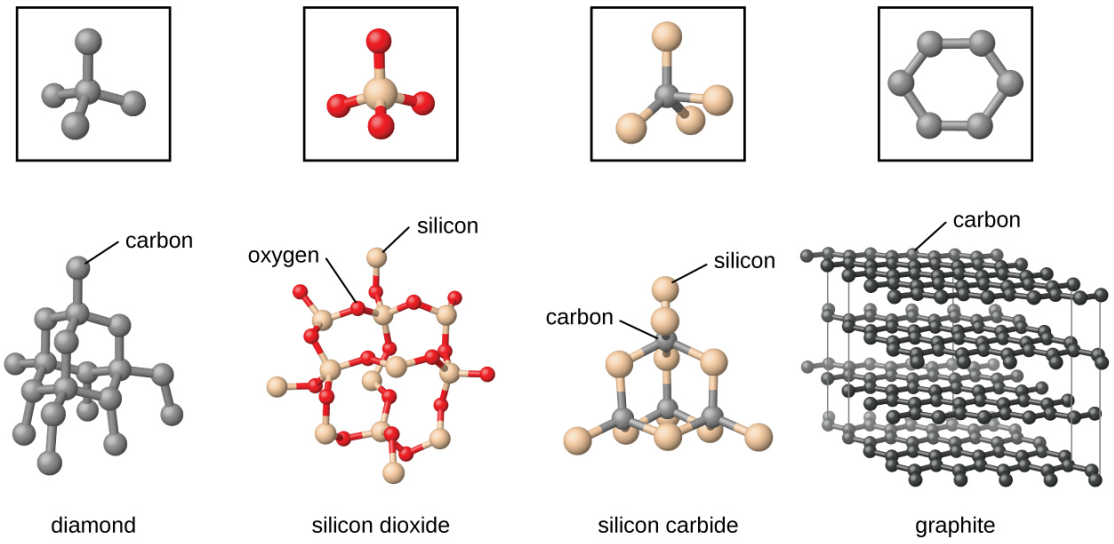


Figure 8.6.5: A covalent crystal contains a three-dimensional network of covalent bonds, as illustrated by the structures of diamond, silicon dioxide, silicon carbide, and graphite. Graphite is an exceptional example, composed of planar sheets of covalent crystals that are held together in layers by noncovalent forces. Unlike typical covalent solids, graphite is very soft and electrically conductive.

Molecular Solids

Molecular solids, such as ice, sucrose (table sugar), and iodine, as shown in Figure 8.6.6: , are composed of neutral molecules. The strengths of the attractive forces between the units present in different crystals vary widely, as indicated by the melting points of the crystals. Small symmetrical molecules (nonpolar molecules), such as H_2, N_2, O_2, and F_2, have weak attractive forces and form molecular solids with very low melting points (below −200 °C). Substances consisting of larger, nonpolar molecules have larger attractive forces and melt at higher temperatures. Molecular solids composed of molecules with permanent dipole moments (polar molecules) melt at still higher temperatures. Examples include ice (melting point, 0 °C) and table sugar (melting point, 185 °C).

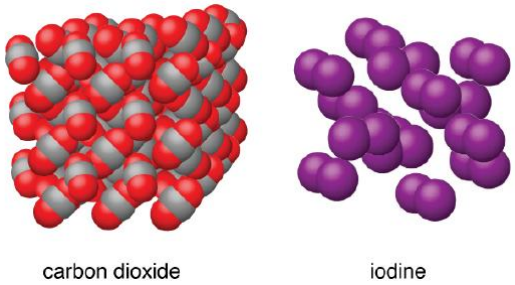


Figure 8.6.6: Carbon dioxide (CO_2) consists of small, nonpolar molecules and forms a molecular solid with a melting point of −78 °C. Iodine (I_2) consists of larger, nonpolar molecules and forms a molecular solid that melts at 114 °C.

Properties of Solids

A crystalline solid, like those listed in Table 8.6.1, has a precise melting temperature because each atom or molecule of the same type is held in place with the same forces or energy. Thus, the attractions between the units that make up the crystal all have the same strength and all require the same amount of energy to be broken. The gradual softening of an amorphous material differs dramatically from the distinct melting of a crystalline solid. This results from the structural nonequivalence of the molecules in the amorphous solid. Some forces are weaker than others, and when an amorphous material is heated, the weakest intermolecular attractions break first. As the temperature is increased further, the stronger attractions are broken. Thus amorphous materials soften over a range of temperatures.

Table 8.6.1: Types of Crystalline Solids and Their Properties

Type of Solid	Type of Particles	Type of Attractions	Properties	Examples

Type of Solid	Type of Particles	Type of Attractions	Properties	Examples
ionic	ions	ionic bonds	hard, brittle, conducts electricity as a liquid but not as a solid, high to very high melting points	NaCl, Al_2O_3
metallic	atoms of electropositive elements	metallic bonds	shiny, malleable, ductile, conducts heat and electricity well, variable hardness and melting temperature	Cu, Fe, Ti, Pb, U
covalent network	atoms of electronegative elements	covalent bonds	very hard, not conductive, very high melting points	C (diamond), SiO_2, SiC
molecular	molecules (or atoms)	IMFs	variable hardness, variable brittleness, not conductive, low melting points	H_2O, CO_2, I_2, $C_{12}H_{22}O_{11}$

How Sciences Interconnect: Graphene: Material of the Future

Carbon is an essential element in our world. The unique properties of carbon atoms allow the existence of carbon-based life forms such as ourselves. Carbon forms a huge variety of substances that we use on a daily basis, including those shown in Figure 8.6.7. You may be familiar with diamond and graphite, the two most common *allotropes* of carbon. (Allotropes are different structural forms of the same element.) Diamond is one of the hardest-known substances, whereas graphite is soft enough to be used as pencil lead. These very different properties stem from the different arrangements of the carbon atoms in the different allotropes.

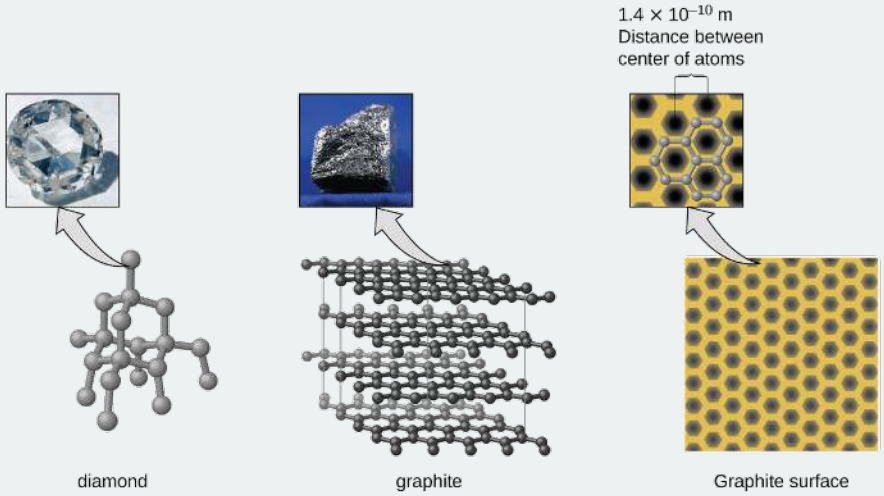


Figure 8.6.7: Diamond is extremely hard because of the strong bonding between carbon atoms in all directions. Graphite (in pencil lead) rubs off onto paper due to the weak attractions between the carbon layers. An image of a graphite surface shows the distance between the centers of adjacent carbon atoms. (credit left photo: modification of work by Steve Jurvetson; credit middle photo: modification of work by United States Geological Survey)

You may be less familiar with a recently discovered form of carbon: graphene. Graphene was first isolated in 2004 by using tape to peel off thinner and thinner layers from graphite. It is essentially a single sheet (one atom thick) of graphite. Graphene, illustrated in Figure 8.6.8: , is not only strong and lightweight, but it is also an excellent conductor of electricity and heat. These properties may prove very useful in a wide range of applications, such as vastly improved computer chips and circuits, better batteries and solar cells, and stronger and lighter structural materials. The 2010 Nobel Prize in Physics was awarded to Andre Geim and Konstantin Novoselov for their pioneering work with graphene.

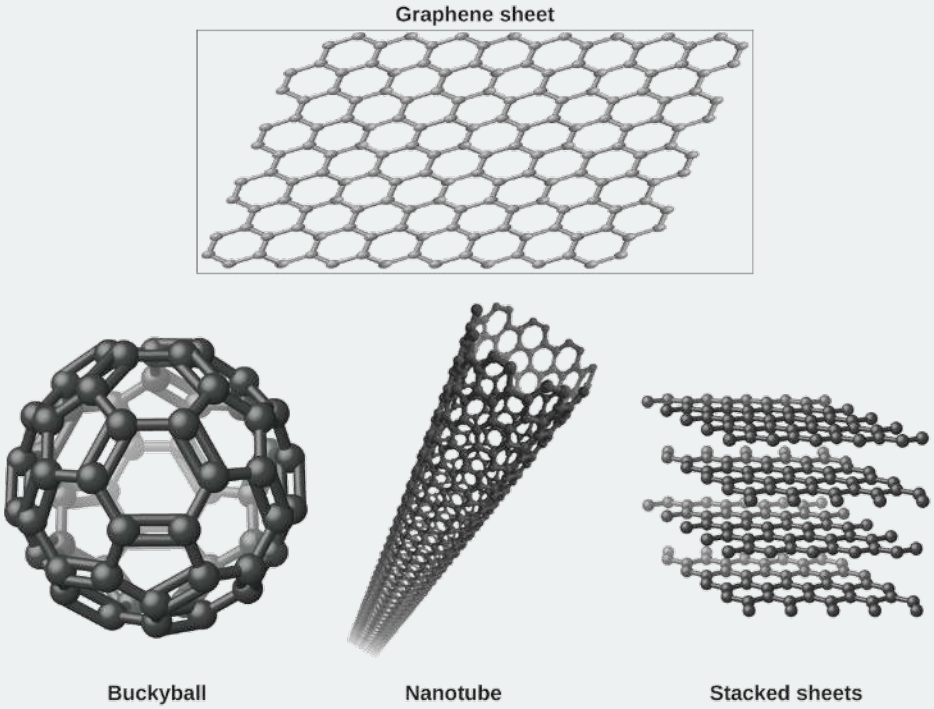


Figure 8.6.8: Graphene sheets can be formed into buckyballs, nanotubes, and stacked layers.

This page titled 8.6: The Solid State of Matter is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **10.5: The Solid State of Matter** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

8.7: Key Terms

adhesive force | force of attraction between molecules of different chemical identities

amorphous solid | (also, noncrystalline solid) solid in which the particles lack an ordered internal structure

body-centered cubic (BCC) solid | crystalline structure that has a cubic unit cell with lattice points at the corners and in the center of the cell

body-centered cubic unit cell | simplest repeating unit of a body-centered cubic crystal; it is a cube containing lattice points at each corner and in the center of the cube

boiling point | temperature at which the vapor pressure of a liquid equals the pressure of the gas above it

Bragg equation | equation that relates the angles at which X-rays are diffracted by the atoms within a crystal

capillary action | flow of liquid within a porous material due to the attraction of the liquid molecules to the surface of the material and to other liquid molecules

Clausius-Clapeyron equation | mathematical relationship between the temperature, vapor pressure, and enthalpy of vaporization for a substance

cohesive force | force of attraction between identical molecules

condensation | change from a gaseous to a liquid state

coordination number | number of atoms closest to any given atom in a crystal or to the central metal atom in a complex

covalent network solid | solid whose particles are held together by covalent bonds

critical point | temperature and pressure above which a gas cannot be condensed into a liquid

crystalline solid | solid in which the particles are arranged in a definite repeating pattern

cubic closest packing (CCP) | crystalline structure in which planes of closely packed atoms or ions are stacked as a series of three alternating layers of different relative orientations (ABC)

deposition | change from a gaseous state directly to a solid state

diffraction | redirection of electromagnetic radiation that occurs when it encounters a physical barrier of appropriate dimensions

dipole-dipole attraction | intermolecular attraction between two permanent dipoles

dispersion force | (also, London dispersion force) attraction between two rapidly fluctuating, temporary dipoles; significant only when particles are very close together

dynamic equilibrium | state of a system in which reciprocal processes are occurring at equal rates

face-centered cubic (FCC) solid | crystalline structure consisting of a cubic unit cell with lattice points on the corners and in the center of each face

face-centered cubic unit cell | simplest repeating unit of a face-centered cubic crystal; it is a cube containing lattice points at each corner and in the center of each face

freezing | change from a liquid state to a solid state

freezing point | temperature at which the solid and liquid phases of a substance are in equilibrium; see also *melting point*

hexagonal closest packing (HCP) | crystalline structure in which close packed layers of atoms or ions are stacked as a series of two alternating layers of different relative orientations (AB)

hole | (also, interstice) space between atoms within a crystal

hydrogen bonding | occurs when exceptionally strong dipoles attract; bonding that exists when hydrogen is bonded to one of the three most electronegative elements: F, O, or N

induced dipole | temporary dipole formed when the electrons of an atom or molecule are distorted by the instantaneous dipole of a neighboring atom or molecule

instantaneous dipole | temporary dipole that occurs for a brief moment in time when the electrons of an atom or molecule are distributed asymmetrically

intermolecular force | noncovalent attractive force between atoms, molecules, and/or ions

interstitial sites | spaces between the regular particle positions in any array of atoms or ions

ionic solid | solid composed of positive and negative ions held together by strong electrostatic attractions

isomorphous | possessing the same crystalline structure

melting | change from a solid state to a liquid state

melting point | temperature at which the solid and liquid phases of a substance are in equilibrium; see also *freezing point*

metallic solid | solid composed of metal atoms

molecular solid | solid composed of neutral molecules held together by intermolecular forces of attraction

normal boiling point | temperature at which a liquid's vapor pressure equals 1 atm (760 torr)

octahedral hole | open space in a crystal at the center of six particles located at the corners of an octahedron

phase diagram | pressure-temperature graph summarizing conditions under which the phases of a substance can exist

polarizability | measure of the ability of a charge to distort a molecule's charge distribution (electron cloud)

simple cubic structure | crystalline structure with a cubic unit cell with lattice points only at the corners

simple cubic unit cell | (also, primitive cubic unit cell) unit cell in the simple cubic structure

space lattice | all points within a crystal that have identical environments

sublimation | change from solid state directly to gaseous state

supercritical fluid | substance at a temperature and pressure higher than its critical point; exhibits properties intermediate between those of gaseous and liquid states

surface tension | energy required to increase the area, or length, of a liquid surface by a given amount

tetrahedral hole | tetrahedral space formed by four atoms or ions in a crystal

triple point | temperature and pressure at which three phases of a substance are in equilibrium

unit cell | smallest portion of a space lattice that is repeated in three dimensions to form the entire lattice

vacancy | defect that occurs when a position that should contain an atom or ion is vacant

van der Waals force | attractive or repulsive force between molecules, including dipole-dipole, dipole-induced dipole, and London dispersion forces; does not include forces due to covalent or ionic bonding, or the attraction between ions and molecules

vapor pressure | (also, equilibrium vapor pressure) pressure exerted by a vapor in equilibrium with a solid or a liquid at a given temperature

vaporization | change from liquid state to gaseous state

viscosity | measure of a liquid's resistance to flow

X-ray crystallography | experimental technique for determining distances between atoms in a crystal by measuring the angles at which X-rays are diffracted when passing through the crystal

---

This page titled 8.7: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **10.7: Key Terms** is licensed CC BY 4.0.

8.8: Key Equations

This page titled 8.8: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **10.8: Key Equations** is licensed CC BY 4.0.

8.9: Summary

The physical properties of condensed matter (liquids and solids) can be explained in terms of the kinetic molecular theory. In a liquid, intermolecular attractive forces hold the molecules in contact, although they still have sufficient KE to move past each other.

Intermolecular attractive forces, collectively referred to as van der Waals forces, are responsible for the behavior of liquids and solids and are electrostatic in nature. Dipole-dipole attractions result from the electrostatic attraction of the partial negative end of one polar molecule for the partial positive end of another. The temporary dipole that results from the motion of the electrons in an atom can induce a dipole in an adjacent atom and give rise to the London dispersion force. London forces increase with increasing molecular size. Hydrogen bonds are a special type of dipole-dipole attraction that results when hydrogen is bonded to one of the three most electronegative elements: F, O, or N.

The intermolecular forces between molecules in the liquid state vary depending upon their chemical identities and result in corresponding variations in various physical properties. Cohesive forces between like molecules are responsible for a liquid's viscosity (resistance to flow) and surface tension (elasticity of a liquid surface). Adhesive forces between the molecules of a liquid and different molecules composing a surface in contact with the liquid are responsible for phenomena such as surface wetting and capillary rise.

Phase transitions are processes that convert matter from one physical state into another. There are six phase transitions between the three phases of matter. Melting, vaporization, and sublimation are all endothermic processes, requiring an input of heat to overcome intermolecular attractions. The reciprocal transitions of freezing, condensation, and deposition are all exothermic processes, involving heat as intermolecular attractive forces are established or strengthened. The temperatures at which phase transitions occur are determined by the relative strengths of intermolecular attractions and are, therefore, dependent on the chemical identity of the substance.

The temperature and pressure conditions at which a substance exists in solid, liquid, and gaseous states are summarized in a phase diagram for that substance. Phase diagrams are combined plots of pressure-temperature equilibrium curves representing the relationships between phase transition temperatures and pressures. The point of intersection of any three curves in a phase diagram represents a substance's triple point—the temperature and pressure at which three different phases are in equilibrium. At pressures below a solid-liquid-gas triple point, a substance cannot exist in the liquid state, regardless of its temperature. The terminus of the liquid-gas curve represents the substance's critical point, the pressure and temperature above which a liquid phase cannot exist.

Some substances form crystalline solids consisting of particles in a very organized structure; others form amorphous (noncrystalline) solids with an internal structure that is not ordered. The main types of crystalline solids are ionic solids, metallic solids, covalent network solids, and molecular solids. The properties of the different kinds of crystalline solids are due to the types of particles of which they consist, the arrangements of the particles, and the strengths of the attractions between them. Because their particles experience identical attractions, crystalline solids have distinct melting temperatures; the particles in amorphous solids experience a range of interactions, so they soften gradually and melt over a range of temperatures. Some crystalline solids have defects in the definite repeating pattern of their particles. These defects (which include vacancies, atoms or ions not in the regular positions, and impurities) change physical properties such as electrical conductivity, which is exploited in the silicon crystals used to manufacture computer chips.

The structures of crystalline metals and simple ionic compounds can be described in terms of packing of spheres. Metal atoms can pack in hexagonal closest-packed structures, cubic closest-packed structures, body-centered structures, and simple cubic structures. The anions in simple ionic structures commonly adopt one of these structures, and the cations occupy the spaces remaining between the anions. Small cations usually occupy tetrahedral holes in a closest-packed array of anions. Larger cations usually occupy octahedral holes. Still larger cations can occupy cubic holes in a simple cubic array of anions. The structure of a solid can be described by indicating the size and shape of a unit cell and the contents of the cell. The type of structure and dimensions of the unit cell can be determined by X-ray diffraction measurements.

This page titled 8.9: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **10.9: Summary** is licensed CC BY 4.0.

8.10: Exercises

1.

In terms of their bulk properties, how do liquids and solids differ? How are they similar?

2.

In terms of the kinetic molecular theory, in what ways are liquids similar to solids? In what ways are liquids different from solids?

3.

In terms of the kinetic molecular theory, in what ways are liquids similar to gases? In what ways are liquids different from gases?

4.

Explain why liquids assume the shape of any container into which they are poured, whereas solids are rigid and retain their shape.

5.

What is the evidence that all neutral atoms and molecules exert attractive forces on each other?

6.

Open the PhET States of Matter Simulation to answer the following questions:

(a) Select the Solid, Liquid, Gas tab. Explore by selecting different substances, heating and cooling the systems, and changing the state. What similarities do you notice between the four substances for each phase (solid, liquid, gas)? What differences do you notice?

(b) For each substance, select each of the states and record the given temperatures. How do the given temperatures for each state correlate with the strengths of their intermolecular attractions? Explain.

(c) Select the Interaction Potential tab, and use the default neon atoms. Move the Ne atom on the right and observe how the potential energy changes. Select the Total Force button, and move the Ne atom as before. When is the total force on each atom attractive and large enough to matter? Then select the Component Forces button, and move the Ne atom. When do the attractive (van der Waals) and repulsive (electron overlap) forces balance? How does this relate to the potential energy versus the distance between atoms graph? Explain.

7.

Define the following and give an example of each:

(a) dispersion force

(b) dipole-dipole attraction

(c) hydrogen bond

8.

The types of intermolecular forces in a substance are identical whether it is a solid, a liquid, or a gas. Why then does a substance change phase from a gas to a liquid or to a solid?

9.

Why do the boiling points of the noble gases increase in the order He < Ne < Ar < Kr < Xe?

10.

Neon and HF have approximately the same molecular masses.

(a) Explain why the boiling points of Neon and HF differ.

(b) Compare the change in the boiling points of Ne, Ar, Kr, and Xe with the change of the boiling points of HF, HCl, HBr, and HI, and explain the difference between the changes with increasing atomic or molecular mass.

11.

Arrange each of the following sets of compounds in order of increasing boiling point temperature:

(a) HCl, H_2O, SiH_4

(b) F_2, Cl_2, Br_2

(c) CH_4, C_2H_6, C_3H_8

(d) O_2, NO, N_2

12.

The molecular mass of butanol, C_4H_9OH, is 74.14; that of ethylene glycol, $CH_2(OH)CH_2OH$, is 62.08, yet their boiling points are 117.2 °C and 174 °C, respectively. Explain the reason for the difference.

13.

On the basis of intermolecular attractions, explain the differences in the boiling points of *n*–butane (−1 °C) and chloroethane (12 °C), which have similar molar masses.

14.

On the basis of dipole moments and/or hydrogen bonding, explain in a qualitative way the differences in the boiling points of acetone (56.2 °C) and 1-propanol (97.4 °C), which have similar molar masses.

15.

The melting point of $H_2O(s)$ is 0 °C. Would you expect the melting point of $H_2S(s)$ to be −85 °C, 0 °C, or 185 °C? Explain your answer.

16.

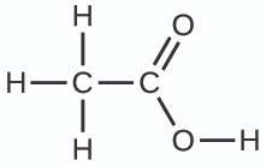
Silane (SiH_4), phosphine (PH_3), and hydrogen sulfide (H_2S) melt at −185 °C, −133 °C, and −85 °C, respectively. What does this suggest about the polar character and intermolecular attractions of the three compounds?

17.

Explain why a hydrogen bond between two water molecules is weaker than a hydrogen bond between two hydrogen fluoride molecules.

18.

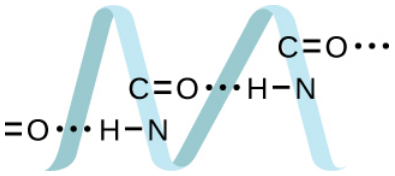
Under certain conditions, molecules of acetic acid, CH_3COOH, form “dimers,” pairs of acetic acid molecules held together by strong intermolecular attractions:



Draw a dimer of acetic acid, showing how two CH_3COOH molecules are held together, and stating the type of IMF that is responsible.

19.

Proteins are chains of amino acids that can form in a variety of arrangements, one of which is a helix. What kind of IMF is responsible for holding the protein strand in this shape? On the protein image, show the locations of the IMFs that hold the protein together:



20.

The density of liquid NH_3 is 0.64 g/mL; the density of gaseous NH_3 at STP is 0.0007 g/mL. Explain the difference between the densities of these two phases.

21.

Identify the intermolecular forces present in the following solids:

(a) CH_3CH_2OH

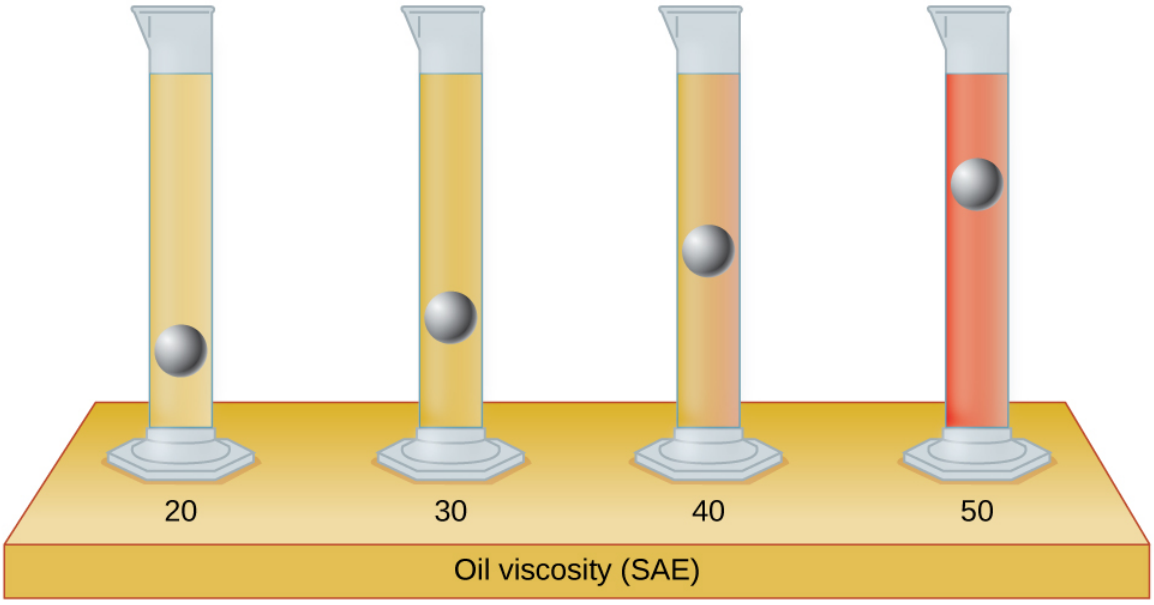
(b) $CH_3CH_2CH_3$

(c) CH_3CH_2Cl

22.

The test tubes shown here contain equal amounts of the specified motor oils. Identical metal spheres were dropped at the same time into each of the tubes, and a brief moment later, the spheres had fallen to the heights indicated in the illustration.

Rank the motor oils in order of increasing viscosity, and explain your reasoning:



23.

Although steel is denser than water, a steel needle or paper clip placed carefully lengthwise on the surface of still water can be made to float. Explain at a molecular level how this is possible.

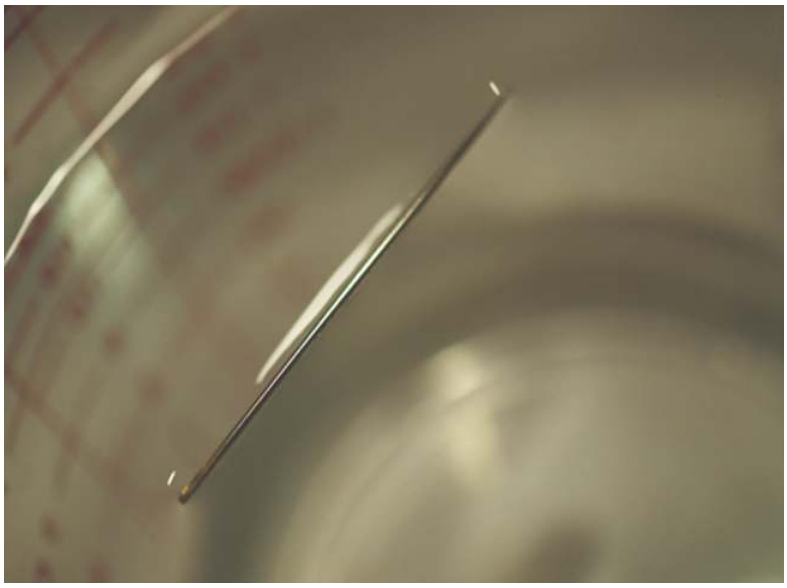
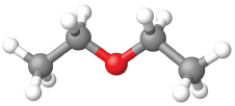
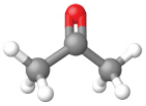
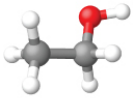
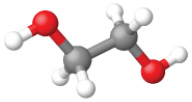


Figure 10.67 (credit: Cory Zanker)

24.

The surface tension and viscosity values for diethyl ether, acetone, ethanol, and ethylene glycol are shown here.

Compound	Molecule	Surface Tension (mN/m)	Viscosity (mPa s)
diethyl ether $C_2H_5OC_2H_5$		17	0.22
acetone CH_3COCH_3		23	0.31
ethanol C_2H_5OH		22	1.07
ethylene glycol $CH_2(OH)CH_2(OH)$		48	16.1

(a) Explain their differences in viscosity in terms of the size and shape of their molecules and their IMFs.

(b) Explain their differences in surface tension in terms of the size and shape of their molecules and their IMFs:

25.

You may have heard someone use the figure of speech “slower than molasses in winter” to describe a process that occurs slowly. Explain why this is an apt idiom, using concepts of molecular size and shape, molecular interactions, and the effect of changing temperature.

26.

It is often recommended that you let your car engine run idle to warm up before driving, especially on cold winter days. While the benefit of prolonged idling is dubious, it is certainly true that a warm engine is more fuel efficient than a cold one. Explain the reason for this.

27.

The surface tension and viscosity of water at several different temperatures are given in this table.

Water	Surface Tension (mN/m)	Viscosity (mPa s)
0 °C	75.6	1.79
20 °C	72.8	1.00
60 °C	66.2	0.47
100 °C	58.9	0.28

(a) As temperature increases, what happens to the surface tension of water? Explain why this occurs, in terms of molecular interactions and the effect of changing temperature.

(b) As temperature increases, what happens to the viscosity of water? Explain why this occurs, in terms of molecular interactions and the effect of changing temperature.

28.

At 25 °C, how high will water rise in a glass capillary tube with an inner diameter of 0.63 mm? Refer to Example 10.4 for the required information.

29.

Water rises in a glass capillary tube to a height of 17 cm. What is the diameter of the capillary tube?

30.

Heat is added to boiling water. Explain why the temperature of the boiling water does not change. What does change?

31.

Heat is added to ice at 0 °C. Explain why the temperature of the ice does not change. What does change?

32.

What feature characterizes the dynamic equilibrium between a liquid and its vapor in a closed container?

33.

Identify two common observations indicating some liquids have sufficient vapor pressures to noticeably evaporate?

34.

Identify two common observations indicating some solids, such as dry ice and mothballs, have vapor pressures sufficient to sublime?

35.

What is the relationship between the intermolecular forces in a liquid and its vapor pressure?

36.

What is the relationship between the intermolecular forces in a solid and its melting temperature?

37.

Why does spilled gasoline evaporate more rapidly on a hot day than on a cold day?

38.

Carbon tetrachloride, CCl_4, was once used as a dry cleaning solvent, but is no longer used because it is carcinogenic. At 57.8 °C, the vapor pressure of CCl_4 is 54.0 kPa, and its enthalpy of vaporization is 33.05 kJ/mol. Use this information to estimate the normal boiling point for CCl_4.

39.

When is the boiling point of a liquid equal to its normal boiling point?

40.

How does the boiling of a liquid differ from its evaporation?

41.

Use the information in Figure 10.24 to estimate the boiling point of water in Denver when the atmospheric pressure is 83.3 kPa.

42.

A syringe at a temperature of 20 °C is filled with liquid ether in such a way that there is no space for any vapor. If the temperature is kept constant and the plunger is withdrawn to create a volume that can be occupied by vapor, what would be the approximate pressure of the vapor produced?

43.

Explain the following observations:

(a) It takes longer to cook an egg in Ft. Davis, Texas (altitude, 5000 feet above sea level) than it does in Boston (at sea level).

(b) Perspiring is a mechanism for cooling the body.

44.

The enthalpy of vaporization of water is larger than its enthalpy of fusion. Explain why.

45.

Explain why the molar enthalpies of vaporization of the following substances increase in the order $CH_4 < C_2H_6 < C_3H_8$, even though the type of IMF (dispersion) is the same.

46.

Explain why the enthalpies of vaporization of the following substances increase in the order $CH_4 < NH_3 < H_2O$, even though all three substances have approximately the same molar mass.

47.

The enthalpy of vaporization of $CO_2(l)$ is 9.8 kJ/mol. Would you expect the enthalpy of vaporization of $CS_2(l)$ to be 28 kJ/mol, 9.8 kJ/mol, or −8.4 kJ/mol? Discuss the plausibility of each of these answers.

48.

The hydrogen fluoride molecule, HF, is more polar than a water molecule, H_2O (for example, has a greater dipole moment), yet the molar enthalpy of vaporization for liquid hydrogen fluoride is lesser than that for water. Explain.

49.

Ethyl chloride (boiling point, 13 °C) is used as a local anesthetic. When the liquid is sprayed on the skin, it cools the skin enough to freeze and numb it. Explain the cooling effect of liquid ethyl chloride.

50.

Which contains the compounds listed correctly in order of increasing boiling points?

(a) $N_2 < CS_2 < H_2O < KCl$

(b) $H_2O < N_2 < CS_2 < KCl$

(c) $N_2 < KCl < CS_2 < H_2O$

(d) $CS_2 < N_2 < KCl < H_2O$

(e) $KCl < H_2O < CS_2 < N_2$

51.

How much heat is required to convert 422 g of liquid H_2O at 23.5 °C into steam at 150 °C?

52.

Evaporation of sweat requires energy and thus take excess heat away from the body. Some of the water that you drink may eventually be converted into sweat and evaporate. If you drink a 20-ounce bottle of water that had been in the refrigerator at 3.8 °C, how much heat is needed to convert all of that water into sweat and then to vapor? (Note: Your body temperature is 36.6 °C. For the purpose of solving this problem, assume that the thermal properties of sweat are the same as for water.)

53.

Titanium tetrachloride, $TiCl_4$, has a melting point of −23.2 °C and has a $\Delta H_{\text{fusion}} = 9.37$ kJ/mol.

(a) How much energy is required to melt 263.1 g $TiCl_4$?

(b) For $TiCl_4$, which will likely have the larger magnitude: ΔH_{fusion} or $\Delta H_{\text{vaporization}}$? Explain your reasoning.

54.

From the phase diagram for water (Figure 10.31), determine the state of water at:

(a) 35 °C and 85 kPa

(b) −15 °C and 40 kPa

(c) −15 °C and 0.1 kPa

(d) 75 °C and 3 kPa

(e) 40 °C and 0.1 kPa

(f) 60 °C and 50 kPa

55.

What phase changes will take place when water is subjected to varying pressure at a constant temperature of 0.005 °C? At 40 °C? At −40 °C?

56.

Pressure cookers allow food to cook faster because the higher pressure inside the pressure cooker increases the boiling temperature of water. A particular pressure cooker has a safety valve that is set to vent steam if the pressure exceeds 3.4 atm. What is the approximate maximum temperature that can be reached inside this pressure cooker? Explain your reasoning.

57.

From the phase diagram for carbon dioxide in Figure 10.34, determine the state of CO_2 at:

(a) 20 °C and 1000 kPa

(b) 10 °C and 2000 kPa

(c) 10 °C and 100 kPa

(d) −40 °C and 500 kPa

(e) −80 °C and 1500 kPa

(f) −80 °C and 10 kPa

58.

Determine the phase changes that carbon dioxide undergoes as pressure is increased at a constant temperature of (a) −50 °C and (b) 50 °C. If the temperature is held at −40 °C? At 20 °C? (See the phase diagram in Figure 10.34.)

59.

Consider a cylinder containing a mixture of liquid carbon dioxide in equilibrium with gaseous carbon dioxide at an initial pressure of 65 atm and a temperature of 20 °C. Sketch a plot depicting the change in the cylinder pressure with time as gaseous carbon dioxide is released at constant temperature.

60.

Dry ice, $CO_2(s)$, does not melt at atmospheric pressure. It sublimes at a temperature of −78 °C. What is the lowest pressure at which $CO_2(s)$ will melt to give $CO_2(l)$? At approximately what temperature will this occur? (See Figure 10.34 for the phase diagram.)

61.

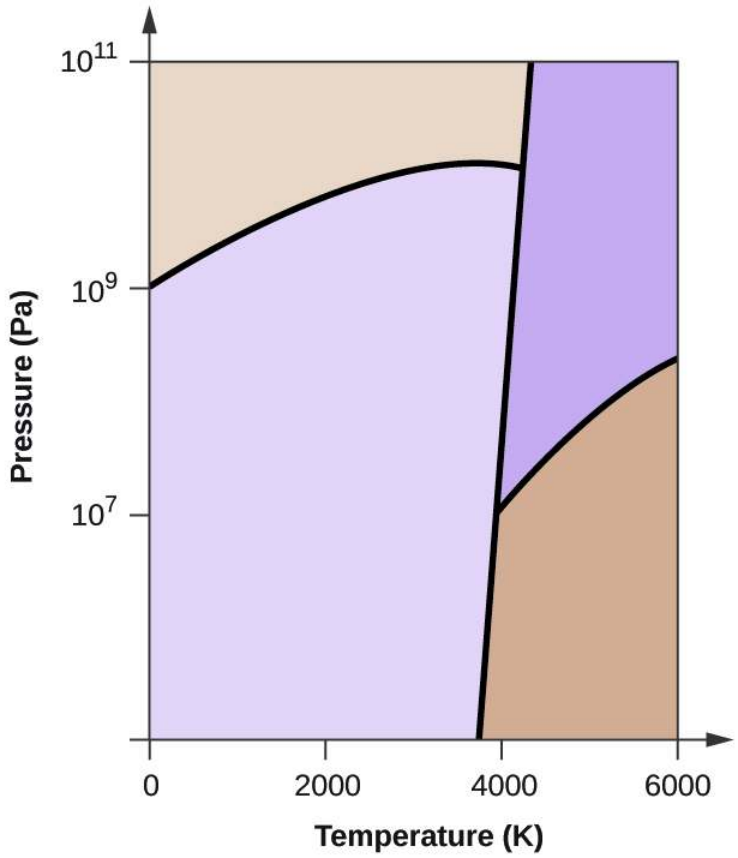
If a severe storm results in the loss of electricity, it may be necessary to use a clothesline to dry laundry. In many parts of the country in the dead of winter, the clothes will quickly freeze when they are hung on the line. If it does not snow, will they dry anyway? Explain your answer.

62.

Is it possible to liquefy nitrogen at room temperature (about 25 °C)? Is it possible to liquefy sulfur dioxide at room temperature? Explain your answers.

63.

Elemental carbon has one gas phase, one liquid phase, and two different solid phases, as shown in the phase diagram:



(a) On the phase diagram, label the gas and liquid regions.

(b) Graphite is the most stable phase of carbon at normal conditions. On the phase diagram, label the graphite phase.

(c) If graphite at normal conditions is heated to 2500 K while the pressure is increased to 10^{10} Pa, it is converted into diamond. Label the diamond phase.

(d) Circle each triple point on the phase diagram.

(e) In what phase does carbon exist at 5000 K and 10^{8} Pa?

(f) If the temperature of a sample of carbon increases from 3000 K to 5000 K at a constant pressure of 10^{6} Pa, which phase transition occurs, if any?

64.

What types of liquids typically form amorphous solids?

65.

At very low temperatures oxygen, O_2, freezes and forms a crystalline solid. Which best describes these crystals?

(a) ionic

(b) covalent network

(c) metallic

(d) amorphous

(e) molecular crystals

66.

As it cools, olive oil slowly solidifies and forms a solid over a range of temperatures. Which best describes the solid?

(a) ionic

(b) covalent network

(c) metallic

(d) amorphous

(e) molecular crystals

67.

Explain why ice, which is a crystalline solid, has a melting temperature of 0 °C, whereas butter, which is an amorphous solid, softens over a range of temperatures.

68.

Identify the type of crystalline solid (metallic, network covalent, ionic, or molecular) formed by each of the following substances:

(a) SiO_2

(b) KCl

(c) Cu

(d) CO_2

(e) C (diamond)

(f) $BaSO_4$

(g) NH_3

(h) NH_4F

(i) C_2H_5OH

69.

Identify the type of crystalline solid (metallic, network covalent, ionic, or molecular) formed by each of the following substances:

(a) $CaCl_2$

(b) SiC

(c) N_2

(d) Fe

(e) C (graphite)

(f) $CH_3CH_2CH_2CH_3$

(g) HCl

(h) NH_4NO_3

(i) K_3PO_4

70.

Classify each substance in the table as either a metallic, ionic, molecular, or covalent network solid:

Substance	Appearance	Melting Point	Electrical Conductivity	Solubility in Water
X	lustrous, malleable	1500 °C	high	insoluble
Y	soft, yellow	113 °C	none	insoluble

Substance	Appearance	Melting Point	Electrical Conductivity	Solubility in Water
Z	hard, white	800 °C	only if melted/dissolved	soluble

71.

Classify each substance in the table as either a metallic, ionic, molecular, or covalent network solid:

Substance	Appearance	Melting Point	Electrical Conductivity	Solubility in Water
X	brittle, white	800 °C	only if melted/dissolved	soluble
Y	shiny, malleable	1100 °C	high	insoluble
Z	hard, colorless	3550 °C	none	insoluble

72.

Identify the following substances as ionic, metallic, covalent network, or molecular solids:

Substance A is malleable, ductile, conducts electricity well, and has a melting point of 1135 °C. Substance B is brittle, does not conduct electricity as a solid but does when molten, and has a melting point of 2072 °C. Substance C is very hard, does not conduct electricity, and has a melting point of 3440 °C. Substance D is soft, does not conduct electricity, and has a melting point of 185 °C.

73.

Substance A is shiny, conducts electricity well, and melts at 975 °C. Substance A is likely a(n):

(a) ionic solid

(b) metallic solid

(c) molecular solid

(d) covalent network solid

74.

Substance B is hard, does not conduct electricity, and melts at 1200 °C. Substance B is likely a(n):

(a) ionic solid

(b) metallic solid

(c) molecular solid

(d) covalent network solid

75.

Describe the crystal structure of iron, which crystallizes with two equivalent metal atoms in a cubic unit cell.

76.

Describe the crystal structure of Pt, which crystallizes with four equivalent metal atoms in a cubic unit cell.

77.

What is the coordination number of a chromium atom in the body-centered cubic structure of chromium?

78.

What is the coordination number of an aluminum atom in the face-centered cubic structure of aluminum?

79.

Cobalt metal crystallizes in a hexagonal closest packed structure. What is the coordination number of a cobalt atom?

80.

Nickel metal crystallizes in a cubic closest packed structure. What is the coordination number of a nickel atom?

81.

Tungsten crystallizes in a body-centered cubic unit cell with an edge length of 3.165 Å.

(a) What is the atomic radius of tungsten in this structure?

(b) Calculate the density of tungsten.

82.

Platinum (atomic radius = 1.38 Å) crystallizes in a cubic closely packed structure. Calculate the edge length of the face-centered cubic unit cell and the density of platinum.

83.

Barium crystallizes in a body-centered cubic unit cell with an edge length of 5.025 Å

(a) What is the atomic radius of barium in this structure?

(b) Calculate the density of barium.

84.

Aluminum (atomic radius = 1.43 Å) crystallizes in a cubic closely packed structure. Calculate the edge length of the face-centered cubic unit cell and the density of aluminum.

85.

The density of aluminum is 2.7 g/cm^3; that of silicon is 2.3 g/cm^3. Explain why Si has the lower density even though it has heavier atoms.

86.

The free space in a metal may be found by subtracting the volume of the atoms in a unit cell from the volume of the cell. Calculate the percentage of free space in each of the three cubic lattices if all atoms in each are of equal size and touch their nearest neighbors. Which of these structures represents the most efficient packing? That is, which packs with the least amount of unused space?

87.

Cadmium sulfide, sometimes used as a yellow pigment by artists, crystallizes with cadmium, occupying one-half of the tetrahedral holes in a closest packed array of sulfide ions. What is the formula of cadmium sulfide? Explain your answer.

88.

A compound of cadmium, tin, and phosphorus is used in the fabrication of some semiconductors. It crystallizes with cadmium occupying one-fourth of the tetrahedral holes and tin occupying one-fourth of the tetrahedral holes in a closest packed array of phosphide ions. What is the formula of the compound? Explain your answer.

89.

What is the formula of the magnetic oxide of cobalt, used in recording tapes, that crystallizes with cobalt atoms occupying one-eighth of the tetrahedral holes and one-half of the octahedral holes in a closely packed array of oxide ions?

90.

A compound containing zinc, aluminum, and sulfur crystallizes with a closest-packed array of sulfide ions. Zinc ions are found in one-eighth of the tetrahedral holes and aluminum ions in one-half of the octahedral holes. What is the empirical formula of the compound?

91.

A compound of thallium and iodine crystallizes in a simple cubic array of iodide ions with thallium ions in all of the cubic holes. What is the formula of this iodide? Explain your answer.

92.

Which of the following elements reacts with sulfur to form a solid in which the sulfur atoms form a closest-packed array with all of the octahedral holes occupied: Li, Na, Be, Ca, or Al?

93.

What is the percent by mass of titanium in rutile, a mineral that contains titanium and oxygen, if structure can be described as a closest packed array of oxide ions with titanium ions in one-half of the octahedral holes? What is the oxidation number of titanium?

94.

Explain why the chemically similar alkali metal chlorides NaCl and CsCl have different structures, whereas the chemically different NaCl and MnS have the same structure.

95.

As minerals were formed from the molten magma, different ions occupied the same cites in the crystals. Lithium often occurs along with magnesium in minerals despite the difference in the charge on their ions. Suggest an explanation.

96.

Rubidium iodide crystallizes with a cubic unit cell that contains iodide ions at the corners and a rubidium ion in the center. What is the formula of the compound?

97.

One of the various manganese oxides crystallizes with a cubic unit cell that contains manganese ions at the corners and in the center. Oxide ions are located at the center of each edge of the unit cell. What is the formula of the compound?

98.

NaH crystallizes with the same crystal structure as NaCl. The edge length of the cubic unit cell of NaH is 4.880 Å.

(a) Calculate the ionic radius of H^-. (The ionic radius of Li^+ is 0.0.95 Å.)

(b) Calculate the density of NaH.

99.

Thallium(I) iodide crystallizes with the same structure as CsCl. The edge length of the unit cell of TlI is 4.20 Å. Calculate the ionic radius of TI^+. (The ionic radius of I^- is 2.16 Å.)

100.

A cubic unit cell contains manganese ions at the corners and fluoride ions at the center of each edge.

(a) What is the empirical formula of this compound? Explain your answer.

(b) What is the coordination number of the Mn^{3+} ion?

(c) Calculate the edge length of the unit cell if the radius of a Mn^{3+} ion is 0.65 A.

(d) Calculate the density of the compound.

101.

What is the spacing between crystal planes that diffract X-rays with a wavelength of 1.541 nm at an angle θ of 15.55° (first order reflection)?

102.

A diffractometer using X-rays with a wavelength of 0.2287 nm produced first order diffraction peak for a crystal angle θ = 16.21°. Determine the spacing between the diffracting planes in this crystal.

103.

A metal with spacing between planes equal to 0.4164 nm diffracts X-rays with a wavelength of 0.2879 nm. What is the diffraction angle for the first order diffraction peak?

104.

Gold crystallizes in a face-centered cubic unit cell. The second-order reflection (n = 2) of X-rays for the planes that make up the tops and bottoms of the unit cells is at θ = 22.20°. The wavelength of the X-rays is 1.54 Å. What is the density of metallic gold?

105.

When an electron in an excited molybdenum atom falls from the L to the K shell, an X-ray is emitted. These X-rays are diffracted at an angle of 7.75° by planes with a separation of 2.64 Å. What is the difference in energy between the K shell and the L shell in molybdenum assuming a first order diffraction?

This page titled 8.10: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

9: Gases

In this chapter, we examine the relationships between gas temperature, pressure, amount, and volume. We will study a simple theoretical model and use it to analyze the experimental behavior of gases. The results of these analyses will show us the limitations of the theory and how to improve on it.

9.1: Introduction
9.2: Gas Pressure
9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law
9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions
9.5: Effusion and Diffusion of Gases
9.6: The Kinetic-Molecular Theory
9.7: Non-Ideal Gas Behavior
9.8: Key Terms
9.9: Key Equations
9.10: Summary
9.11: Exercises

This page titled 9: Gases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

9.1: Introduction

We are surrounded by an ocean of gas—the atmosphere—and many of the properties of gases are familiar to us from our daily activities. Heated gases expand, which can make a hot air balloon rise (Figure 9.1.1) or cause a blowout in a bicycle tire left in the sun on a hot day. Gases have played an important part in the development of chemistry. In the seventeenth and eighteenth centuries, many scientists investigated gas behavior, providing the first mathematical descriptions of the behavior of matter.



Figure 9.1.1: The hot air inside these balloons is less dense than the surrounding cool air. This results in a buoyant force that causes the balloons to rise when their guy lines are untied. (credit: modification of work by Anthony Quintano)

In this chapter, we will examine the relationships between gas temperature, pressure, amount, and volume. We will study a simple theoretical model and use it to analyze the experimental behavior of gases. The results of these analyses will show us the limitations of the theory and how to improve on it.

This page titled 9.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.0: Introduction** is licensed CC BY 4.0.

9.2: Gas Pressure

Learning Objectives

By the end of this section, you will be able to:

- Define the property of pressure
- Define and convert among the units of pressure measurements
- Describe the operation of common tools for measuring gas pressure

The earth's atmosphere exerts a pressure, as does any other gas. Although we do not normally notice atmospheric pressure, we are sensitive to pressure changes—for example, when your ears "pop" during take-off and landing while flying, or when you dive underwater. Gas pressure is caused by the force exerted by gas molecules colliding with the surfaces of objects (Figure 9.2.1). Although the force of each collision is very small, any surface of appreciable area experiences a large number of collisions in a short time, which can result in a high pressure. In fact, normal air pressure is strong enough to crush a metal container when not balanced by equal pressure from inside the container.

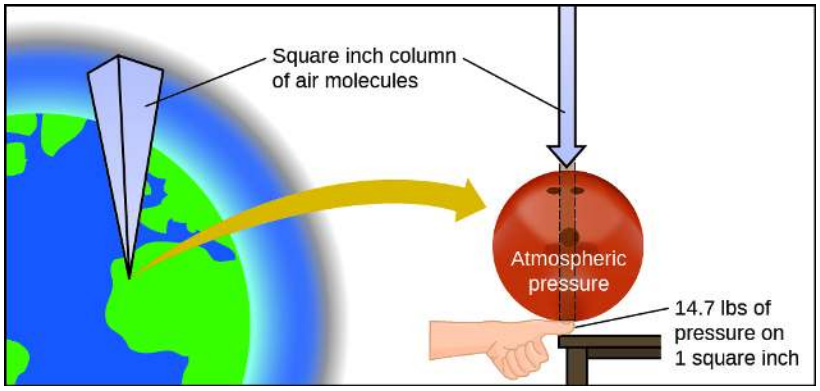


Figure 9.2.1: The atmosphere above us exerts a large pressure on objects at the surface of the earth, roughly equal to the weight of a bowling ball pressing on an area the size of a human thumbnail.

Link to Learning

A dramatic illustration of atmospheric pressure is provided in this brief video, which shows a railway tanker car imploding when its internal pressure is decreased.

A smaller scale demonstration of this phenomenon is briefly explained.

Atmospheric pressure is caused by the weight of the column of air molecules in the atmosphere above an object, such as the tanker car. At sea level, this pressure is roughly the same as that exerted by a full-grown African elephant standing on a doormat, or a typical bowling ball resting on your thumbnail. These may seem like huge amounts, and they are, but life on earth has evolved under such atmospheric pressure. If you actually perch a bowling ball on your thumbnail, the pressure experienced is *twice* the usual pressure, and the sensation is unpleasant.

In general, **pressure** is defined as the force exerted on a given area: \(P=\frac{F}{A}). Note that pressure is directly proportional to force and inversely proportional to area. Thus, pressure can be increased either by increasing the amount of force or by decreasing the area over which it is applied; pressure can be decreased by decreasing the force or increasing the area.

Let's apply this concept to determine which exerts a greater pressure in Figure 9.2.2—the elephant or the figure skater? A large African elephant can weigh 7 tons, supported on four feet, each with a diameter of about 1.5 ft (footprint area of 250 in^2), so the pressure exerted by each foot is about 14 lb/in^2:

$$\text{pressure per elephant foot } = 14,000\,\frac{lb}{\text{elephant}} \times \frac{1 \text{ elephant}}{4 \text{ feet}} \times \frac{1 \text{ foot}}{250 in^2} = 14 lb/in^2$$

The figure skater weighs about 120 lbs, supported on two skate blades, each with an area of about 2 in^2, so the pressure exerted by each blade is about 30 lb/in^2:

$$\text{pressure per skate blade } = 120\frac{lb}{\text{skater}} \times \frac{1 \text{ skater}}{2 \text{ blades}} \times \frac{1 \text{ blade}}{2 \text{ in }^2} = 30 lb/in^2$$

Even though the elephant is more than one hundred-times heavier than the skater, it exerts less than one-half of the pressure. On the other hand, if the skater removes their skates and stands with bare feet (or regular footwear) on the ice, the larger area over which their weight is applied greatly reduces the pressure exerted:

$$\text{pressure per human foot } = 120\frac{lb}{\text{skater}} \times \frac{1 \text{ skater}}{2 \text{ feet}} \times \frac{1 \text{ foot}}{30 in^2} = 2 lb/in^2$$



Figure 9.2.2: Although (a) an elephant's weight is large, creating a very large force on the ground, (b) the figure skater exerts a much higher pressure on the ice due to the small surface area of the skates. (credit a: modification of work by Guido da Rozze; credit b: modification of work by Ryosuke Yagi)

The SI unit of pressure is the **pascal (Pa)**, with 1 Pa = 1 N/m^2, where N is the newton, a unit of force defined as 1 kg m/s^2. One pascal is a small pressure; in many cases, it is more convenient to use units of kilopascal (1 kPa = 1000 Pa) or **bar** (1 bar = 100,000 Pa). In the United States, pressure is often measured in pounds of force on an area of one square inch—**pounds per square inch (psi)**—for example, in car tires. Pressure can also be measured using the unit **atmosphere** (atm), which originally represented the average sea level air pressure at the approximate latitude of Paris (45°). Table 9.2.1 provides some information on these and a few other common units for pressure measurements

Table 9.2.1: Pressure Units

Unit Name and Abbreviation	Definition or Relation to Other Unit
pascal (Pa)	1 Pa = 1 N/m^2 recommended IUPAC unit
kilopascal (kPa)	1 kPa = 1000 Pa
pounds per square inch (psi)	air pressure at sea level is ~14.7 psi
atmosphere (atm)	1 atm = 101,325 Pa = 760 torr air pressure at sea level is ~1 atm
bar (bar, or b)	1 bar = 100,000 Pa (exactly) commonly used in meteorology
millibar (mbar, or mb)	1000 mbar = 1 bar
inches of mercury (in. Hg)	1 in. Hg = 3386 Pa used by aviation industry, also some weather reports
torr	named after Evangelista Torricelli, inventor of the barometer
millimeters of mercury (mm Hg)	1 mm Hg ~1 torr

Example 9.2.1: Conversion of Pressure Units

The United States National Weather Service reports pressure in both inches of Hg and millibars. Convert a pressure of 29.2 in. Hg into:

a. torr
b. atm
c. kPa
d. mbar

Solution

This is a unit conversion problem. The relationships between the various pressure units are given in Table 9.2.1.

a. $$29.2 \text{ in } Hg_g \times \frac{25.4mm}{1im} \times \frac{1 \text{ torr}}{1mmHg_g} = 742 \text{ torr}$$

b. $$742 \text{ torr } \times \frac{1atm}{760 \text{ torr}} = 0.976atm$$

c. $$742 \text{ torr } \times \frac{101.325kPa}{760 \text{ torr}} = 98.9kPa$$

d. $$98.9kPa \times \frac{1000Pa}{1kPa} \times \frac{1bar}{100,000Pa} \times \frac{1000mbar}{1bar} = 989mbar$$

Exercise 9.2.1

A typical barometric pressure in Kansas City is 740 torr. What is this pressure in atmospheres, in millimeters of mercury, in kilopascals, and in bar?

Answer

0.974 atm; 740 mm Hg; 98.7 kPa; 0.987 bar

We can measure atmospheric pressure, the force exerted by the atmosphere on the earth's surface, with a barometer (Figure 9.2.3). A barometer is a glass tube that is closed at one end, filled with a nonvolatile liquid such as mercury, and then inverted and immersed in a container of that liquid. The atmosphere exerts pressure on the liquid outside the tube, the column of liquid exerts pressure inside the tube, and the pressure at the liquid surface is the same inside and outside the tube. The height of the liquid in the tube is therefore proportional to the pressure exerted by the atmosphere.

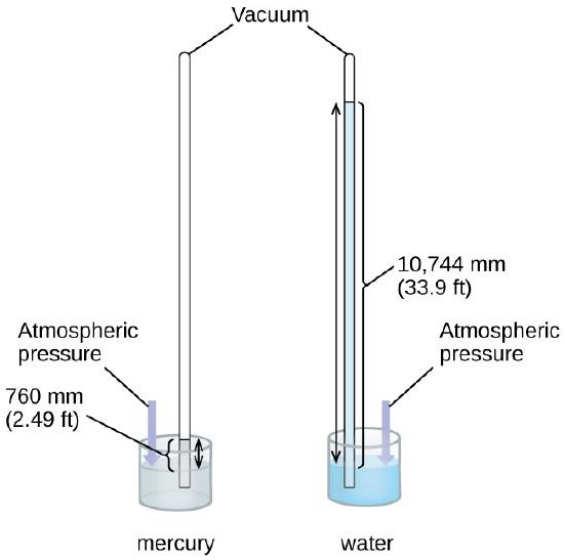


Figure 9.2.3: In a barometer, the height, *h*, of the column of liquid is used as a measurement of the air pressure. Using very dense liquid mercury (left) permits the construction of reasonably sized barometers, whereas using water (right) would require a barometer more than 30 feet tall.

If the liquid is water, normal atmospheric pressure will support a column of water over 10 meters high, which is rather inconvenient for making (and reading) a barometer. Because mercury (Hg) is about 13.6-times denser than water, a mercury barometer only needs to be \frac{1}{13.6} as tall as a water barometer—a more suitable size. Standard atmospheric pressure of 1 atm at sea level (101,325 Pa) corresponds to a column of mercury that is about 760 mm (29.92 in.) high. The **torr** was originally intended to be a unit equal to one millimeter of mercury, but it no longer corresponds exactly. The pressure exerted by a fluid due to gravity is known as **hydrostatic pressure**, p:

$$p = h\rho g$$

where *h* is the height of the fluid, *ρ* (lowercase Greek letter rho) is the density of the fluid, and *g* is acceleration due to gravity.

Example 9.2.2: Calculation of Barometric Pressure

Show the calculation supporting the claim that atmospheric pressure near sea level corresponds to the pressure exerted by a column of mercury that is about 760 mm high. The density of mercury = 13.6 g/cm^3.

Solution

The hydrostatic pressure is given by $p = h\rho g$, with h = 760 mm, ρ = 13.6 g/cm^3, and g = 9.81 m/s^2. Plugging these values into the equation and doing the necessary unit conversions will give us the value we seek. (Note: We are expecting to find a pressure of ~101,325 Pa.)

$$101,325 N/m^2 = 101,325 \frac{kg \cdot m/s^2}{m^2} = 101,325 \frac{kg}{m \cdot s^2}$$

$$p = \left(760 mm \times \frac{1m}{1000mm}\right) \times \left(\frac{13.6g}{1cm^3} \times \frac{1kg}{1000g} \times \frac{(100cm)^3}{(1m)^3}\right) \times \left(\frac{9.81m}{1s^2}\right)$$

$$= (0.760m)\left(13,600 kg/m^3\right)\left(9.81 m/s^2\right) = 1.01 \times 10^5 kg/ms^2 = 1.01 \times 10^5 N/m^2$$

$$= 1.01 \times 10^5 Pa$$

Exercise 9.2.2

Calculate the height of a column of water at 25 °C that corresponds to normal atmospheric pressure. The density of water at this temperature is 1.0 g/cm^3.

Answer

10.3 m

A **manometer** is a device similar to a barometer that can be used to measure the pressure of a gas trapped in a container. A closed-end manometer is a U-shaped tube with one closed arm, one arm that connects to the gas to be measured, and a nonvolatile liquid (usually mercury) in between. As with a barometer, the distance between the liquid levels in the two arms of the tube (h in the diagram) is proportional to the pressure of the gas in the container. An open-end manometer (Figure 9.2.4) is the same as a closed-end manometer, but one of its arms is open to the atmosphere. In this case, the distance between the liquid levels corresponds to the difference in pressure between the gas in the container and the atmosphere.

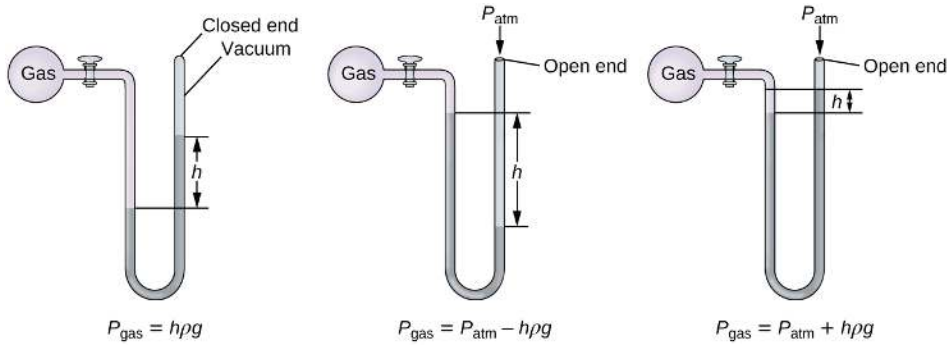
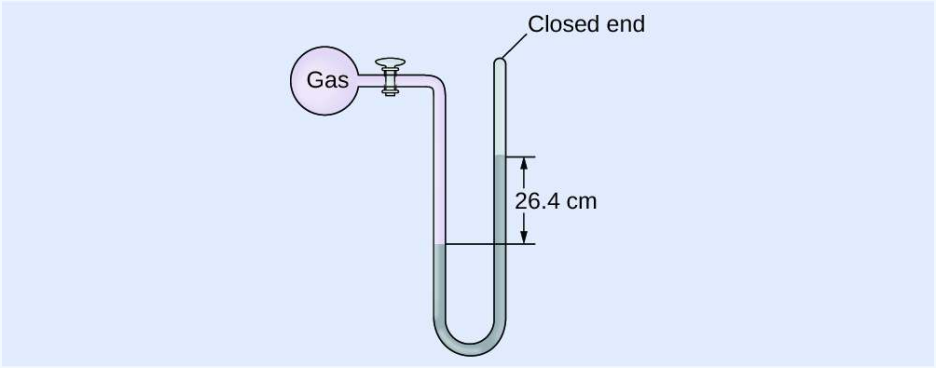


Figure 9.2.4: A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.

Example 9.2.3: Calculation of Pressure Using a Closed-End Manometer

The pressure of a sample of gas is measured with a closed-end manometer, as shown to the right. The liquid in the manometer is mercury. Determine the pressure of the gas in:

a. torr
b. Pa
c. bar



Solution

The pressure of the gas is equal to a column of mercury of height 26.4 cm. (The pressure at the bottom horizontal line is equal on both sides of the tube. The pressure on the left is due to the gas and the pressure on the right is due to 26.4 cm Hg, or mercury.) We could use the equation $p = h\rho g$ as in Example 9.2.2, but it is simpler to just convert between units using Table 9.2.1.

a. $$26.4 emHg_g \times \frac{10 mmHg_g}{1 emHg} \times \frac{1 \text{ torr}}{1 mm\Pi Hg} = 264 \text{ torr}$$

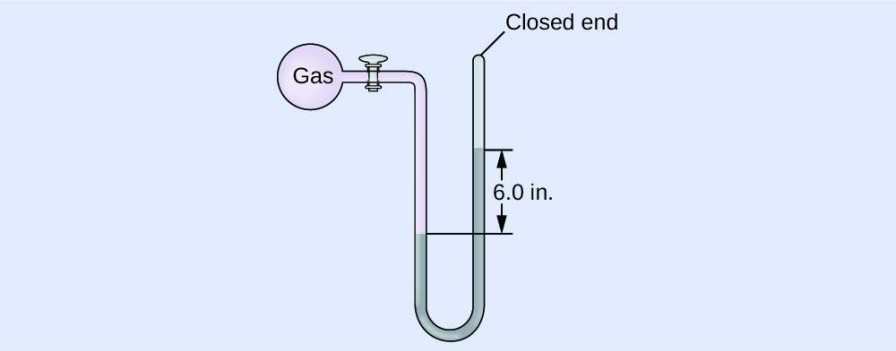
b. $$264 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{101,325 Pa}{1 \text{ atm}} = 35,200 Pa$$

c. $$35,200 Pa \times \frac{1 bar}{100,000 Pa} = 0.352 bar$$

Exercise 9.2.3

The pressure of a sample of gas is measured with a closed-end manometer. The liquid in the manometer is mercury. Determine the pressure of the gas in:

a. torr
b. Pa
c. bar



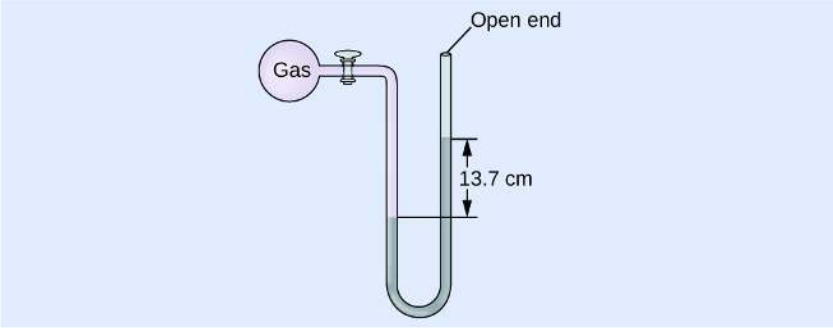
Answer

(a) ~150 torr; (b) ~20,000 Pa; (c) ~0.20 bar

Example 9.2.4: Calculation of Pressure Using an Open-End Manometer

The pressure of a sample of gas is measured at sea level with an open-end Hg (mercury) manometer, as shown to the right. Determine the pressure of the gas in:

a. mm Hg
b. atm
c. kPa



Solution

The pressure of the gas equals the hydrostatic pressure due to a column of mercury of height 13.7 cm plus the pressure of the atmosphere at sea level. (The pressure at the bottom horizontal line is equal on both sides of the tube. The pressure on the left is due to the gas and the pressure on the right is due to 13.7 cm of Hg plus atmospheric pressure.)

a. In mm Hg, this is: 137 mm Hg + 760 mm Hg = 897 mm Hg

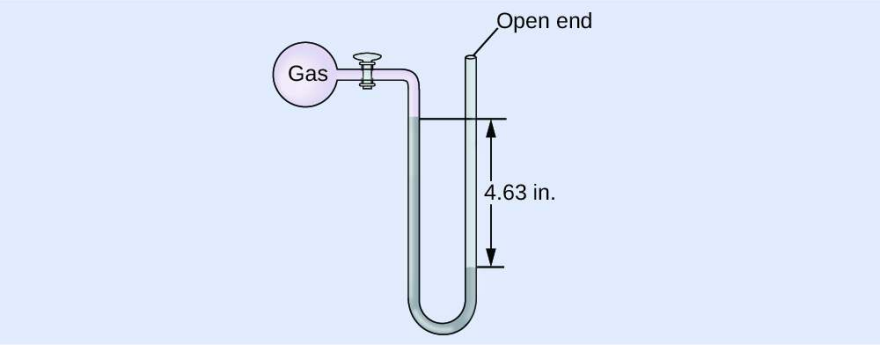
b. $$897 mm Hg_g \times \frac{1 atm}{760 mm Hg_g} = 1.18 atm$$

c. $$1.18 \text{ atm} \times \frac{101.325 kPa}{1 - 20 \times 10^2 kPa} = 1.20$$

Exercise 9.2.4

The pressure of a sample of gas is measured at sea level with an open-end Hg manometer, as shown to the right. Determine the pressure of the gas in:

a. mm Hg
b. atm
c. kPa



Answer

(a) 642 mm Hg; (b) 0.845 atm; (c) 85.6 kPa

Chemistry in Everyday Life: Measuring Blood Pressure

Blood pressure is measured using a device called a sphygmomanometer (Greek *sphygmos* = "pulse"). It consists of an inflatable cuff to restrict blood flow, a manometer to measure the pressure, and a method of determining when blood flow begins and when it becomes impeded (Figure 9.2.5). Since its invention in 1881, it has been an essential medical device. There are many types of sphygmomanometers: manual ones that require a stethoscope and are used by medical professionals; mercury ones, used when the most accuracy is required; less accurate mechanical ones; and digital ones that can be used with little training but that have limitations. When using a sphygmomanometer, the cuff is placed around the upper arm and inflated until blood flow is completely blocked, then slowly released. As the heart beats, blood forced through the arteries causes a rise in pressure. This rise in pressure at which blood flow begins is the *systolic pressure*—the peak pressure in the cardiac cycle. When the cuff's pressure equals the arterial systolic pressure, blood flows past the cuff, creating audible sounds that can be heard using a stethoscope. This is followed by a decrease in pressure as the heart's ventricles prepare for another beat. As cuff pressure continues to decrease, eventually sound is no longer heard; this is the *diastolic pressure*—the lowest pressure (resting phase) in the cardiac cycle. Blood pressure units from a sphygmomanometer are in terms of millimeters of mercury (mm Hg).



Figure 9.2.5: (a) A medical technician prepares to measure a patient's blood pressure with a sphygmomanometer. (b) A typical sphygmomanometer uses a valved rubber bulb to inflate the cuff and a diaphragm gauge to measure pressure. (credit a: modification of work by Master Sgt. Jeffrey Allen)

How Sciences Interconnect: Meteorology, Climatology, and Atmospheric Science

Throughout the ages, people have observed clouds, winds, and precipitation, trying to discern patterns and make predictions: when it is best to plant and harvest; whether it is safe to set out on a sea voyage; and much more. We now face complex weather and atmosphere-related challenges that will have a major impact on our civilization and the ecosystem. Several different scientific disciplines use chemical principles to help us better understand weather, the atmosphere, and climate. These are meteorology, climatology, and atmospheric science. Meteorology is the study of the atmosphere, atmospheric phenomena, and atmospheric effects on earth's weather. Meteorologists seek to understand and predict the weather in the short term, which can save lives and benefit the economy. Weather forecasts (Figure 9.2.6) are the result of thousands of measurements of air pressure, temperature, and the like, which are compiled, modeled, and analyzed in weather centers worldwide.

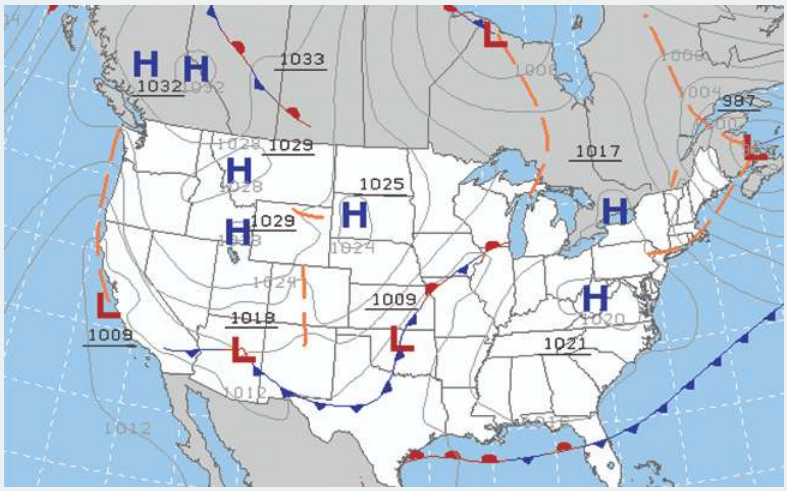


Figure 9.2.6: Meteorologists use weather maps to describe and predict weather. Regions of high (H) and low (L) pressure have large effects on weather conditions. The gray lines represent locations of constant pressure known as isobars. (credit: modification of work by National Oceanic and Atmospheric Administration)

In terms of weather, low-pressure systems occur when the earth's surface atmospheric pressure is lower than the surrounding environment: Moist air rises and condenses, producing clouds. Movement of moisture and air within various weather fronts instigates most weather events.

The atmosphere is the gaseous layer that surrounds a planet. Earth's atmosphere, which is roughly 100–125 km thick, consists of roughly 78.1% nitrogen and 21.0% oxygen, and can be subdivided further into the regions shown in Figure 9.2.7 the exosphere (furthest from earth, > 700 km above sea level), the thermosphere (80–700 km), the mesosphere (50–80 km), the stratosphere (second lowest level of our atmosphere, 12–50 km above sea level), and the troposphere (up to 12 km above sea level, roughly 80% of the earth's atmosphere by mass and the layer where most weather events originate). As you go higher in the troposphere, air density and temperature both decrease.

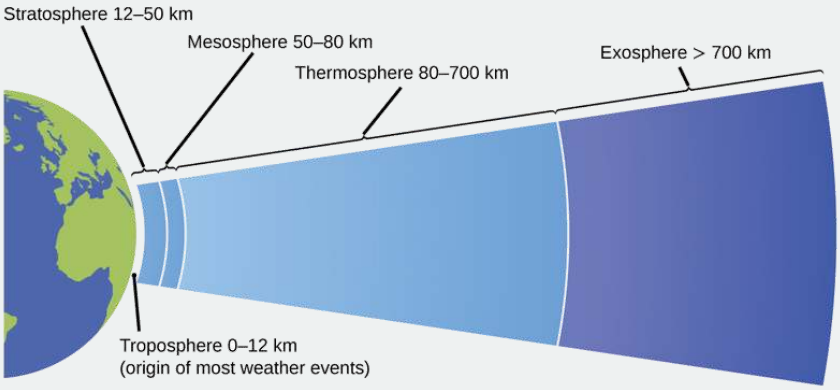


Figure 9.2.7: Earth's atmosphere has five layers: the troposphere, the stratosphere, the mesosphere, the thermosphere, and the exosphere.

Climatology is the study of the climate, averaged weather conditions over long time periods, using atmospheric data. However, climatologists study patterns and effects that occur over decades, centuries, and millennia, rather than shorter time frames of hours, days, and weeks like meteorologists. Atmospheric science is an even broader field, combining meteorology, climatology, and other scientific disciplines that study the atmosphere.

This page titled 9.2: Gas Pressure is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.1: Gas Pressure** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law

Learning Objectives

By the end of this section, you will be able to:

- Identify the mathematical relationships between the various properties of gases
- Use the ideal gas law, and related gas laws, to compute the values of various gas properties under specified conditions

During the seventeenth and especially eighteenth centuries, driven both by a desire to understand nature and a quest to make balloons in which they could fly (Figure 9.3.1), a number of scientists established the relationships between the macroscopic physical properties of gases, that is, pressure, volume, temperature, and amount of gas. Although their measurements were not precise by today's standards, they were able to determine the mathematical relationships between pairs of these variables (e.g., pressure and temperature, pressure and volume) that hold for an *ideal* gas—a hypothetical construct that real gases approximate under certain conditions. Eventually, these individual laws were combined into a single equation—the *ideal gas law*—that relates gas quantities for gases and is quite accurate for low pressures and moderate temperatures. We will consider the key developments in individual relationships (for pedagogical reasons not quite in historical order), then put them together in the ideal gas law.

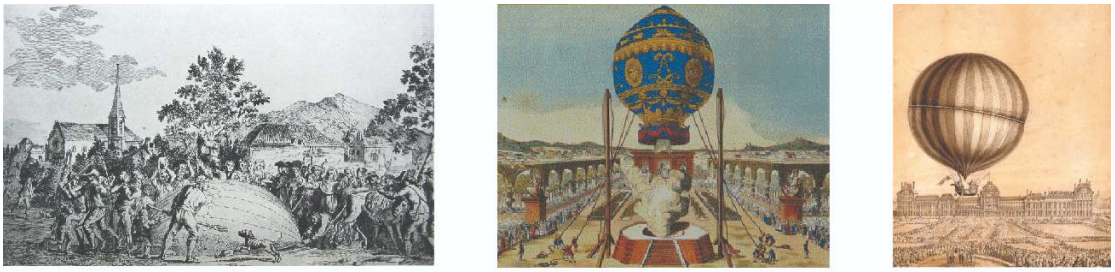


Figure 9.3.1: In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. The launch of the latter was reportedly viewed by 400,000 people in Paris.

Pressure and Temperature: Amontons's Law

Imagine filling a rigid container attached to a pressure gauge with gas and then sealing the container so that no gas may escape. If the container is cooled, the gas inside likewise gets colder and its pressure is observed to decrease. Since the container is rigid and tightly sealed, both the volume and number of moles of gas remain constant. If we heat the sphere, the gas inside gets hotter (Figure 9.3.2) and the pressure increases.

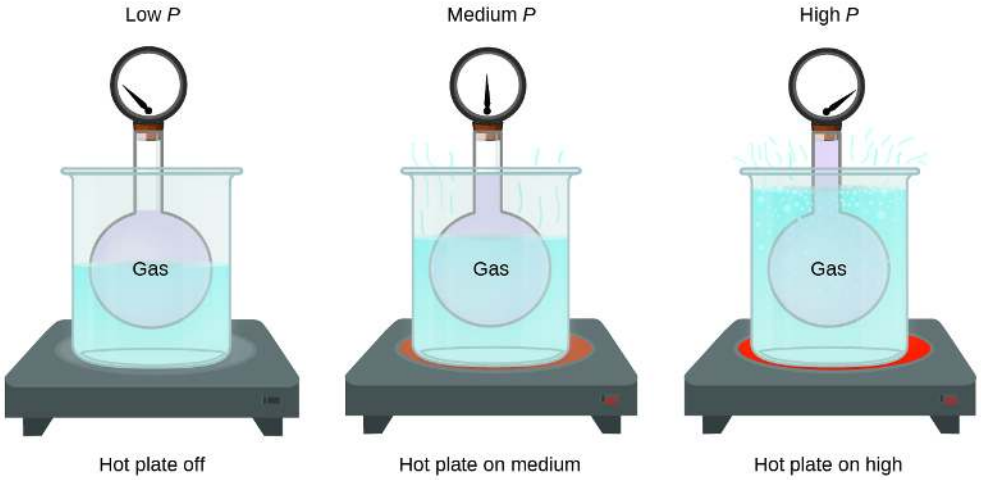


Figure 9.3.2: The effect of temperature on gas pressure: When the hot plate is off, the pressure of the gas in the sphere is relatively low. As the gas is heated, the pressure of the gas in the sphere increases.

This relationship between temperature and pressure is observed for any sample of gas confined to a constant volume. An example of experimental pressure-temperature data is shown for a sample of air under these conditions in Figure 9.3.3. We find that temperature and pressure are linearly related, and if the temperature is on the kelvin scale, then *P* and *T* are directly proportional (again, when *volume and moles of gas are held constant*); if the temperature on the kelvin scale increases by a certain factor, the gas pressure increases by the same factor.

Temperature (°C)	Temperature (K)	Pressure (kPa)
−100	173	36.0
−50	223	46.4
0	273	56.7
50	323	67.1
100	373	77.5
150	423	88.0

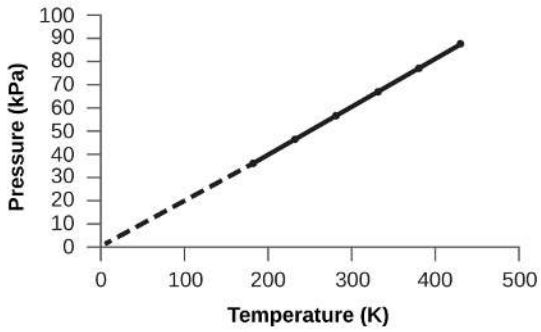


Figure 9.3.3: For a constant volume and amount of air, the pressure and temperature are directly proportional, provided the temperature is in kelvin. (Measurements cannot be made at lower temperatures because of the condensation of the gas.) When this line is extrapolated to lower pressures, it reaches a pressure of 0 at –273 °C, which is 0 on the kelvin scale and the lowest possible temperature, called absolute zero.

Guillaume Amontons was the first to empirically establish the relationship between the pressure and the temperature of a gas (~1700), and Joseph Louis Gay-Lussac determined the relationship more precisely (~1800). Because of this, the *P*-*T* relationship for gases is known as either **Amontons's law** or **Gay-Lussac's law**. Under either name, it states that *the pressure of a given amount of gas is directly proportional to its temperature on the kelvin scale when the volume is held constant*. Mathematically, this can be written:

$$P \propto T \text{ or } P = \text{ constant } \times T \text{ or } P = k \times T$$

where $\propto$ means "is proportional to," and k is a proportionality constant that depends on the identity, amount, and volume of the gas.

For a confined, constant volume of gas, the $\frac{P}{T}$ ratio is therefore constant (i.e., $\frac{P}{T} = k$). If the gas is initially in "Condition 1" (with $P = P_1$ and $T = T_1$), and then changes to "Condition 2" (with $P = P_2$ and $T = T_2$), we have that $\frac{P_1}{T_1} = k$ and $\frac{P_2}{T_2} = k$ which reduces to

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}.$$

This equation is useful for pressure-temperature calculations for a confined gas at constant volume. Note that temperatures must be on the kelvin scale for any gas law calculations (0 on the kelvin scale and the lowest possible temperature is called **absolute zero**). (Also note that there are at least three ways we can describe how the pressure of a gas changes as its temperature changes: We can use a table of values, a graph, or a mathematical equation.)

Example 9.3.1: Predicting Change in Pressure with Temperature

A can of hair spray is used until it is empty except for the propellant, isobutane gas.

a. On the can is the warning "Store only at temperatures below 120 °F (48.8 °C). Do not incinerate." Why?
b. The gas in the can is initially at 24 °C and 360 kPa, and the can has a volume of 350 mL. If the can is left in a car that reaches 50 °C on a hot day, what is the new pressure in the can?

Solution

(a) The can contains an amount of isobutane gas at a constant volume, so if the temperature is increased by heating, the pressure will increase proportionately. High temperature could lead to high pressure, causing the can to burst. (Also, isobutane is combustible, so incineration could cause the can to explode.)

(b) We are looking for a pressure change due to a temperature change at constant volume, so we will use Amontons's/Gay-Lussac's law. Taking P_1 and T_1 as the initial values, T_2 as the temperature where the pressure is unknown and P_2 as the unknown pressure, and converting °C to K, we have:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ which means that } \frac{360 \text{ kPa}}{297 \text{ K}} = \frac{P_2}{323 \text{ K}}$$

Rearranging and solving gives:

$$P_2 = \frac{360 \text{ kPa} \times 323 \cancel{\text{ K}}}{297 \cancel{\text{ K}}} = 390 \text{ kPa}$$

? Exercise 9.3.1

A sample of nitrogen, N_2, occupies 45.0 mL at 27 °C and 600 torr. What pressure will it have if cooled to –73 °C while the volume remains constant?

Answer

400 torr

Volume and Temperature: Charles's Law

If we fill a balloon with air and seal it, the balloon contains a specific amount of air at atmospheric pressure, let's say 1 atm. If we put the balloon in a refrigerator, the gas inside gets cold and the balloon shrinks (although both the amount of gas and its pressure remain constant). If we make the balloon very cold, it will shrink a great deal, and it expands again when it warms up.

Link to Learning

This video shows how cooling and heating a gas causes its volume to decrease or increase, respectively.

These examples of the effect of temperature on the volume of a given amount of a confined gas at constant pressure are true in general: The volume increases as the temperature increases, and decreases as the temperature decreases. Volume-temperature data for a 1-mole sample of methane gas at 1 atm are listed and graphed in Figure 9.12.

Temperature (°C)	Temperature (K)	Volume (L)
–3	270	22
–23	250	21
–53	220	18
–162	111	9

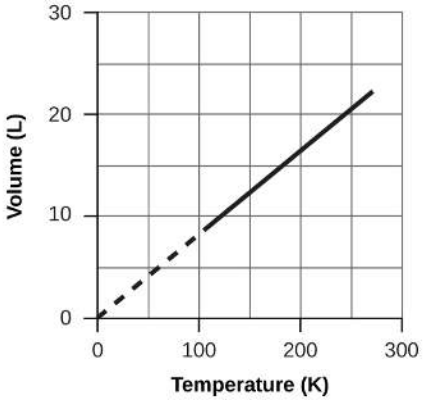


Figure 9.12 The volume and temperature are linearly related for 1 mole of methane gas at a constant pressure of 1 atm. If the temperature is in kelvin, volume and temperature are directly proportional. The line stops at 111 K because methane liquefies at this temperature; when extrapolated, it intersects the graph's origin, representing a temperature of absolute zero.

The relationship between the volume and temperature of a given amount of gas at constant pressure is known as Charles's law in recognition of the French scientist and balloon flight pioneer Jacques Alexandre César Charles. **Charles's law** states that *the volume of a given amount of gas is directly proportional to its temperature on the kelvin scale when the pressure is held constant.*

Mathematically, this can be written as:

$$V \alpha T \text{ or } V = \text{ constant } \cdot T \text{ or } V = k \cdot T \text{ or } V_1/T_1 = V_2/T_2$$

with k being a proportionality constant that depends on the amount and pressure of the gas.

For a confined, constant pressure gas sample, $\frac{V}{T}$ is constant (i.e., the ratio = *k*), and as seen with the *P*-*T* relationship, this leads to another form of Charles's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}. \tag{9.3.1}$$

Example 9.3.2: Predicting Change in Volume with Temperature

A sample of carbon dioxide, CO_2, occupies 0.300 L at 10 °C and 750 torr. What volume will the gas have at 30 °C and 750 torr?

Solution

Because we are looking for the volume change caused by a temperature change at constant pressure, this is a job for Charles's law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume, and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{0.300L}{283K} = \frac{V_2}{303K}$$

Rearranging and solving gives:

$$V_2 = \frac{0.300L \times 303 \cancel{\text{ K}}}{283 \cancel{\text{ K}}} = 0.321 \text{ L}$$

This answer supports our expectation from Charles's law, namely, that raising the gas temperature (from 283 K to 303 K) at a constant pressure will yield an increase in its volume (from 0.300 L to 0.321 L).

Exercise 9.3.2

A sample of oxygen, O_2, occupies 32.2 mL at 30 °C and 452 torr. What volume will it occupy at –70 °C and the same pressure?

Answer

21.6 mL

Example 9.3.3: Measuring Temperature with a Volume Change

Temperature is sometimes measured with a gas thermometer by observing the change in the volume of the gas as the temperature changes at constant pressure. The hydrogen in a particular hydrogen gas thermometer has a volume of 150.0 cm^3 when immersed in a mixture of ice and water (0.00 °C). When immersed in boiling liquid ammonia, the volume of the hydrogen, at the same pressure, is 131.7 cm^3. Find the temperature of boiling ammonia on the kelvin and Celsius scales.

Solution

A volume change caused by a temperature change at constant pressure means we should use Charles's law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume, and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{150.0cm^3}{273.15K} = \frac{131.7cm^3}{T_2}$$

Rearrangement gives

$$T_2 = \frac{131.7 \cancel{\text{ cm}^3} \times 273.15 \text{ K}}{150.0 \cancel{\text{ cm}^3}} = 239.8 \text{ K}$$

Subtracting 273.15 from 239.8 K, we find that the temperature of the boiling ammonia on the Celsius scale is –33.4 °C.

Exercise 9.3.3

What is the volume of a sample of ethane at 467 K and 1.1 atm if it occupies 405 mL at 298 K and 1.1 atm?

Answer

635 mL

Volume and Pressure: Boyle's Law

If we partially fill an airtight syringe with air, the syringe contains a specific amount of air at constant temperature, say 25 °C. If we slowly push in the plunger while keeping temperature constant, the gas in the syringe is compressed into a smaller volume and its pressure increases; if we pull out the plunger, the volume increases and the pressure decreases. This example of the effect of volume on the pressure of a given amount of a confined gas is true in general. Decreasing the volume of a contained gas will increase its pressure, and increasing its volume will decrease its pressure. In fact, if the volume increases by a certain factor, the pressure decreases by the same factor, and vice versa. Volume-pressure data for an air sample at room temperature are graphed in Figure 9.3.4.

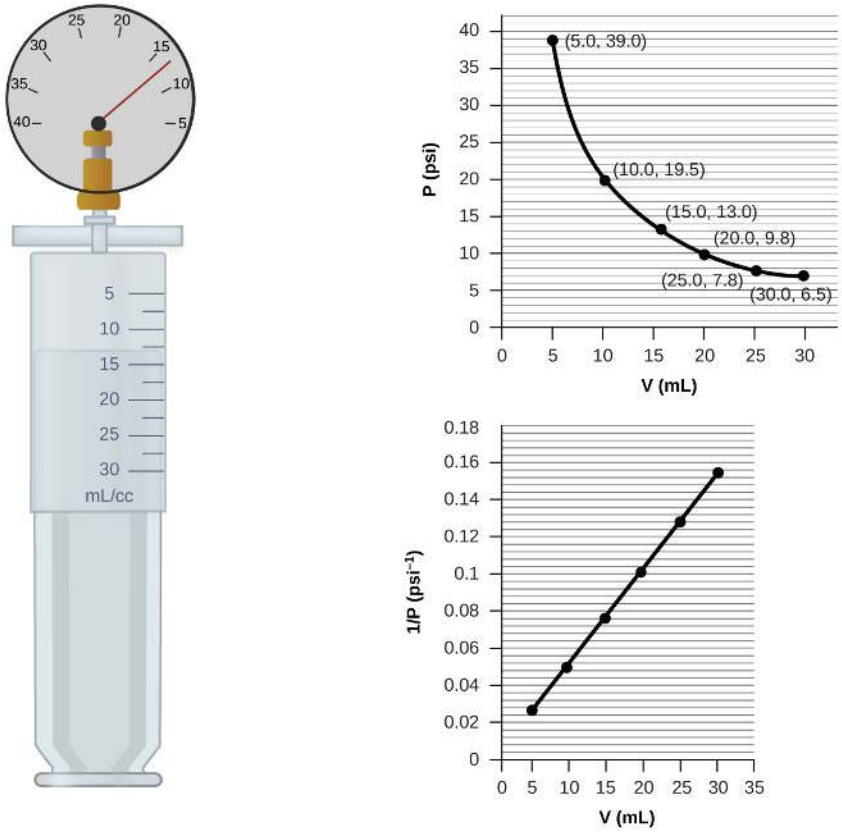


Figure 9.3.4: When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since P and V are inversely proportional, a graph of $\frac{1}{P}$ vs. V is linear.

Unlike the P-T and V-T relationships, pressure and volume are not directly proportional to each other. Instead, P and V exhibit inverse proportionality: Increasing the pressure results in a decrease of the volume of the gas. Mathematically this can be written:

$$P \propto \frac{1}{V} \tag{9.3.2}$$

or

$$P = k \cdot \frac{1}{V} \tag{9.3.3}$$

or

$$P \cdot V = k \tag{9.3.4}$$

or

$$P_1V_1 = P_2V_2 \tag{9.3.5}$$

with *k* being a constant. Graphically, this relationship is shown by the straight line that results when plotting the inverse of the pressure

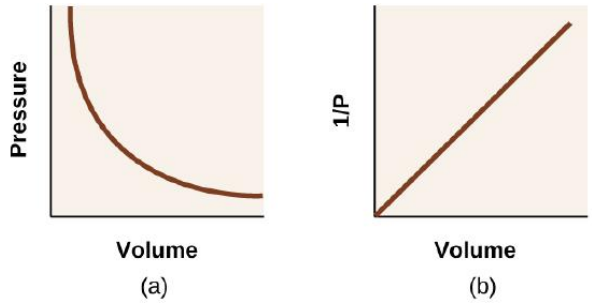


Figure 9.3.5: The relationship between pressure and volume is inversely proportional. (a) The graph of *P* vs. *V* is a hyperbola, whereas (b) the graph of $\frac{1}{P}$ vs. *V* is linear.

The relationship between the volume and pressure of a given amount of gas at constant temperature was first published by the English natural philosopher Robert Boyle over 300 years ago. It is summarized in the statement now known as **Boyle's law**: *The volume of a given amount of gas held at constant temperature is inversely proportional to the pressure under which it is measured.*

Example 9.3.4: Volume of a Gas Sample

The sample of gas in Figure 9.3.4: has a volume of 15.0 mL at a pressure of 13.0 psi. Determine the pressure of the gas at a volume of 7.5 mL, using:

a. the *P*-*V* graph in Figure 9.3.5
b. the $\frac{1}{P}$ vs. V graph in Figure 9.3.5
c. the Boyle's law equations (Equations 9.3.2 - 9.3.5)

Comment on the likely accuracy of each method.

Solution

(a) Estimating from the *P*-*V* graph gives a value for *P* somewhere around 27 psi.

(b) Estimating from the \frac{1}{P}\) versus V graph give a value of about 26 psi.

(c) From Boyle's law, we know that the product of pressure and volume (*PV*) for a given sample of gas at a constant temperature is always equal to the same value. Therefore we have $P_1V_1 = k$ and $P_2V_2 = k$ which means that $P_1V_1 = P_2V_2$.

Using P_1 and V_1 as the known values 13.0 psi and 15.0 mL, P_2 as the pressure at which the volume is unknown, and V_2 as the unknown volume, we have:

$$P_1\ V_1 = P_2\ V_2$$
$$(13.0\ \text{psi}) \times (15.0\ \text{mL} = P_2 \times (7.5\ \text{mL})$$

Solving:

$$P_2 = \frac{13.0\ \text{psi} \times 15.0\ \cancel{\text{mL}}}{7.5\ \cancel{\text{mL}}} = 26\ \text{psi}$$

It was more difficult to estimate well from the *P*-*V* graph, so (a) is likely more inaccurate than (b) or (c). The calculation will be as accurate as the equation and measurements allow.

Exercise 9.3.4

The sample of gas in Figure 9.3.4 has a volume of 30.0 mL at a pressure of 6.5 psi. Determine the volume of the gas at a pressure of 11.0 psi, using:

a. the *P*-*V* graph in Figure 9.3.5
b. the $\frac{1}{P}$ vs. \(*V*\) graph in Figure 9.3.5
c. the Boyle's law equations (9.3.2 - 9.3.5)

Comment on the likely accuracy of each method.

Answer

(a) about 17–18 mL; (b) ~18 mL; (c) 17.7 mL; it was more difficult to estimate well from the *P*-*V* graph, so (a) is likely more inaccurate than (b); the calculation will be as accurate as the equation and measurements allow

Chemistry in Everyday Life: Breathing and Boyle's Law

What do you do about 20 times per minute for your whole life, without break, and often without even being aware of it? The answer, of course, is respiration, or breathing. How does it work? It turns out that the gas laws apply here. Your lungs take in gas that your body needs (oxygen) and get rid of waste gas (carbon dioxide). Lungs are made of spongy, stretchy tissue that expands and contracts while you breathe. When you inhale, your diaphragm and intercostal muscles (the muscles between your ribs) contract, expanding your chest cavity and making your lung volume larger. The increase in volume leads to a decrease in pressure (Boyle's law). This causes air to flow into the lungs (from high pressure to low pressure). When you exhale, the process reverses: Your diaphragm and rib muscles relax, your chest cavity contracts, and your lung volume decreases, causing the pressure to increase (Boyle's law again), and air flows out of the lungs (from high pressure to low pressure). You then breathe in and out again, and again, repeating this Boyle's law cycle for the rest of your life (Figure 9.3.6).

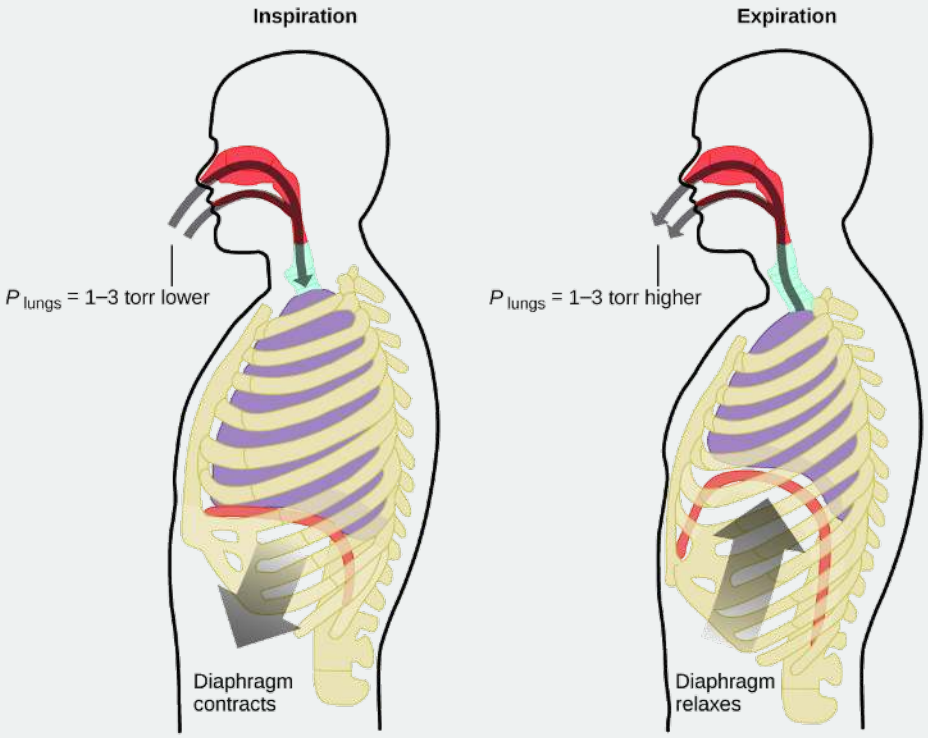


Figure 9.3.6: Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.

Moles of Gas and Volume: Avogadro's Law

The Italian scientist Amedeo Avogadro advanced a hypothesis in 1811 to account for the behavior of gases, stating that equal volumes of all gases, measured under the same conditions of temperature and pressure, contain the same number of molecules. Over time, this relationship was supported by many experimental observations as expressed by **Avogadro's law**: *For a confined gas, the volume (V) and number of moles (n) are directly proportional if the pressure and temperature both remain constant.*

In equation form, this is written as:

$$V \propto n \tag{9.3.6}$$

or

$$V = k \times n \tag{9.3.7}$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \tag{9.3.8}$$

Mathematical relationships can also be determined for the other variable pairs, such as *P* versus *n*, and *n* versus *T*.

Link to Learning

Visit this interactive PhET simulation to investigate the relationships between pressure, volume, temperature, and amount of gas. Use the simulation to examine the effect of changing one parameter on another while holding the other parameters constant (as described in the preceding sections on the various gas laws).

The Ideal Gas Law

To this point, four separate laws have been discussed that relate pressure, volume, temperature, and the number of moles of the gas:

- Boyle's law: PV = constant at constant *T* and *n*
- Amontons's law: $\frac{P}{T}$ = constant at constant *V* and *n*
- Charles's law: $\frac{V}{T}$ = constant at constant *P* and *n*
- Avogadro's law: $\frac{V}{n}$ = constant at constant *P* and *T*

Combining these four laws yields the **ideal gas law**, a relation between the pressure, volume, temperature, and number of moles of a gas:

$$PV = nRT$$

where *P* is the pressure of a gas, *V* is its volume, *n* is the number of moles of the gas, *T* is its temperature on the kelvin scale, and *R* is a constant called the **ideal gas constant** or the universal gas constant. The units used to express pressure, volume, and temperature will determine the proper form of the gas constant as required by dimensional analysis, the most commonly encountered values being 0.08206 L atm mol^{-1} K^{-1} and 8.314 kPa L mol^{-1} K^{-1}.

Gases whose properties of *P*, *V*, and *T* are accurately described by the ideal gas law (or the other gas laws) are said to exhibit *ideal behavior* or to approximate the traits of an **ideal gas**. An ideal gas is a hypothetical construct that may be used along with *kinetic molecular theory* to effectively explain the gas laws as will be described in a later module of this chapter. Although all the calculations presented in this module assume ideal behavior, this assumption is only reasonable for gases under conditions of relatively low pressure and high temperature. In the final module of this chapter, a modified gas law will be introduced that accounts for the *non-ideal* behavior observed for many gases at relatively high pressures and low temperatures.

The ideal gas equation contains five terms, the gas constant *R* and the variable properties *P*, *V*, *n*, and *T*. Specifying any four of these terms will permit use of the ideal gas law to calculate the fifth term as demonstrated in the following example exercises.

Example 9.3.5: Using the Ideal Gas Law

Methane, CH_4, is being considered for use as an alternative automotive fuel to replace gasoline. One gallon of gasoline could be replaced by 655 g of CH_4. What is the volume of this much methane at 25 °C and 745 torr?

Solution

We must rearrange $PV = nRT$ to solve for V:

$$V = \frac{nRT}{P}$$

If we choose to use R = 0.08206 L atm mol^{-1} K^{-1}, then the amount must be in moles, temperature must be in kelvin, and pressure must be in atm.

Converting into the "right" units:

$$n = 655 \text{ g } \cancel{CH_4} \times \frac{1 \text{ mol}}{16.043 \text{ g } \cancel{CH_4}} = 40.8 \text{ mol}$$

$$T = 25^{\circ}C + 273 = 298 \text{ K}$$

$$P = 745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.980 \text{ atm}$$

$$V = \frac{nRT}{P} = \frac{(40.8 \cancel{\text{ mol}}) \left(0.08206 \text{ L atm } \cancel{\text{mol}^{-1}\text{K}^{-1}}\right) (298 \cancel{\text{ K}})}{0.980 \cancel{\text{ atm}}} = 1.02 \times 10^3 \text{ L}$$

It would require 1020 L (269 gal) of gaseous methane at about 1 atm of pressure to replace 1 gal of gasoline. It requires a large container to hold enough methane at 1 atm to replace several gallons of gasoline.

? Exercise 9.3.5

Calculate the pressure in bar of 2520 moles of hydrogen gas stored at 27 °C in the 180-L storage tank of a modern hydrogen-powered car.

Answer

350 bar

If the number of moles of an ideal gas are kept constant under two different sets of conditions, a useful mathematical relationship called the **combined gas law** is obtained:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

using units of atm, L, and K. Both sets of conditions are equal to the product of n R (where n = the number of moles of the gas and R is the ideal gas law constant).

✓ Example 9.3.6: Using the Combined Gas Law

When filled with air, a typical scuba tank with a volume of 13.2 L has a pressure of 153 atm (Figure 9.3.7). If the water temperature is 27 °C, how many liters of air will such a tank provide to a diver's lungs at a depth of approximately 70 feet in the ocean where the pressure is 3.13 atm?



Figure 9.3.7: Scuba divers use compressed air to breathe while underwater. (credit: modification of work by Mark Goodchild)

Letting *1* represent the air in the scuba tank and *2* represent the air in the lungs, and noting that body temperature (the temperature the air will be in the lungs) is 37 °C, we have:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \rightarrow \frac{(153atm)(13.2L)}{(300K)} = \frac{(3.13atm)\,(V_2)}{(310K)}$$

Solving for V_2:

$$V_2 = \frac{(153\text{ atm})(13.2L)(310K)}{(300K)(3.13\text{ atm})} = 667L$$

(Note: Be advised that this particular example is one in which the assumption of ideal gas behavior is not very reasonable, since it involves gases at relatively high pressures and low temperatures. Despite this limitation, the calculated volume can be viewed as a good “ballpark” estimate.)

Exercise 9.3.6

A sample of ammonia is found to occupy 0.250 L under laboratory conditions of 27 °C and 0.850 atm. Find the volume of this sample at 0 °C and 1.00 atm.

Answer

0.193 L

Chemistry in Everyday Life: The Interdependence between Ocean Depth and Pressure in Scuba Diving

Whether scuba diving at the Great Barrier Reef in Australia (shown in Figure 9.3.8) or in the Caribbean, divers must understand how pressure affects a number of issues related to their comfort and safety.



Figure 9.3.8: Scuba divers, whether at the Great Barrier Reef or in the Caribbean, must be aware of buoyancy, pressure equalization, and the amount of time they spend underwater, to avoid the risks associated with pressurized gases in the body. (credit: Kyle Taylor)

Pressure increases with ocean depth, and the pressure changes most rapidly as divers reach the surface. The pressure a diver experiences is the sum of all pressures above the diver (from the water and the air). Most pressure measurements are given in units of atmospheres, expressed as “atmospheres absolute” or ATA in the diving community: Every 33 feet of salt water represents 1 ATA of pressure in addition to 1 ATA of pressure from the atmosphere at sea level. As a diver descends, the increase in pressure causes the body’s air pockets in the ears and lungs to compress; on the ascent, the decrease in pressure causes these air pockets to expand, potentially rupturing eardrums or bursting the lungs. Divers must therefore undergo equalization by adding air to body airspaces on the descent by breathing normally and adding air to the mask by breathing out of the nose or adding air to the ears and sinuses by equalization techniques; the corollary is also true on ascent, divers must release air from the body to maintain equalization. Buoyancy, or the ability to control whether a diver sinks or floats, is controlled by the buoyancy compensator (BCD). If a diver is ascending, the air in their BCD expands because of lower pressure according to Boyle’s law (decreasing the pressure of gases increases the volume). The expanding air increases the buoyancy of the diver, and they begin to ascend. The diver must vent air from the BCD or risk an uncontrolled ascent that could rupture the lungs. In descending, the increased pressure causes the air in the BCD to compress and the diver sinks much more quickly; the diver must add air to the BCD or risk an uncontrolled descent, facing much higher pressures near the ocean floor. The pressure also impacts how long a diver can stay underwater before ascending. The deeper a diver dives, the more compressed the air that is breathed because of increased pressure: If a diver dives 33 feet, the pressure is 2 ATA and the air would be compressed to one-half of its original volume. The diver uses up available air twice as fast as at the surface.

Standard Conditions of Temperature and Pressure

We have seen that the volume of a given quantity of gas and the number of molecules (moles) in a given volume of gas vary with changes in pressure and temperature. Chemists sometimes make comparisons against a **standard temperature and pressure (STP)** for reporting properties of gases: 273.15 K and 1 atm (101.325 kPa).[1] At STP, one mole of an ideal gas has a volume of about 22.4 L—this is referred to as the **standard molar volume** (Figure 9.3.9).

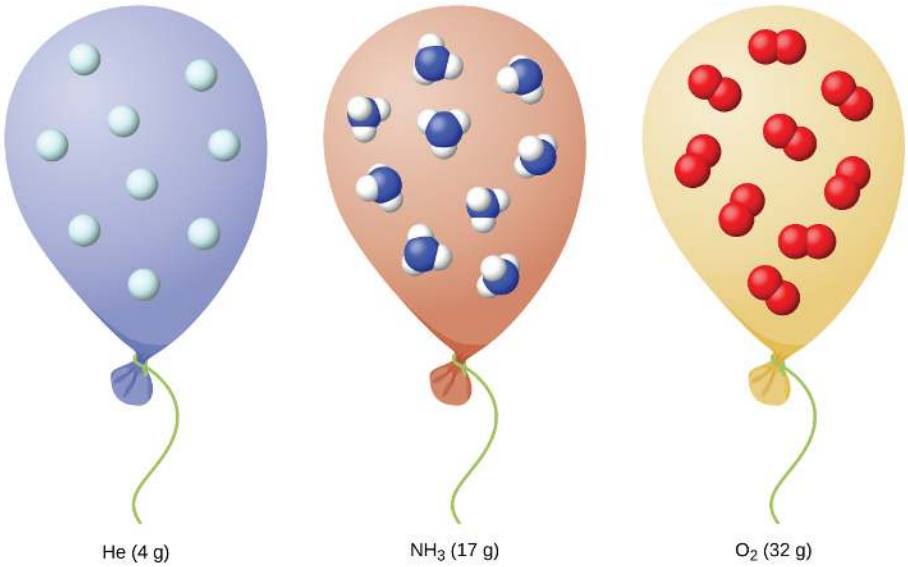


Figure 9.3.9: Regardless of its chemical identity, one mole of gas behaving ideally occupies a volume of ~22.4 L at STP.

Footnotes

- 1The IUPAC definition of standard pressure was changed from 1 atm to 1 bar (100 kPa) in 1982, but the prior definition remains in use by many literature resources and will be used in this text.

This page titled 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **9.2: Relating Pressure, Volume, Amount, and Temperature - The Ideal Gas Law** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions

Learning Objectives

By the end of this section, you will be able to:

- Use the ideal gas law to compute gas densities and molar masses
- Perform stoichiometric calculations involving gaseous substances
- State Dalton's law of partial pressures and use it in calculations involving gaseous mixtures

The study of the chemical behavior of gases was part of the basis of perhaps the most fundamental chemical revolution in history. French nobleman Antoine Lavoisier, widely regarded as the "father of modern chemistry," changed chemistry from a qualitative to a quantitative science through his work with gases. He discovered the law of conservation of matter, discovered the role of oxygen in combustion reactions, determined the composition of air, explained respiration in terms of chemical reactions, and more. He was a casualty of the French Revolution, guillotined in 1794. Of his death, mathematician and astronomer Joseph-Louis Lagrange said, "It took the mob only a moment to remove his head; a century will not suffice to reproduce it."[2] Much of the knowledge we do have about Lavoisier's contributions is due to his wife, Marie-Anne Paulze Lavoisier, who worked with him in his lab. A trained artist fluent in several languages, she created detailed illustrations of the equipment in his lab, and translated texts from foreign scientists to complement his knowledge. After his execution, she was instrumental in publishing Lavoisier's major treatise, which unified many concepts of chemistry and laid the groundwork for significant further study.

As described in an earlier chapter of this text, we can turn to chemical stoichiometry for answers to many of the questions that ask "How much?" The essential property involved in such use of stoichiometry is the amount of substance, typically measured in moles (*n*). For gases, molar amount can be derived from convenient experimental measurements of pressure, temperature, and volume. Therefore, these measurements are useful in assessing the stoichiometry of pure gases, gas mixtures, and chemical reactions involving gases. This section will not introduce any new material or ideas, but will provide examples of applications and ways to integrate concepts already discussed.

Gas Density and Molar Mass

The ideal gas law described previously in this chapter relates the properties of pressure *P*, volume *V*, temperature *T*, and molar amount *n*. This law is universal, relating these properties in identical fashion regardless of the chemical identity of the gas:

$$PV = nRT$$

The density *d* of a gas, on the other hand, is determined by its identity. As described in another chapter of this text, the density of a substance is a characteristic property that may be used to identify the substance.

$$d = \frac{m}{V}$$

Rearranging the ideal gas equation to isolate *V* and substituting into the density equation yields

$$d = \frac{mP}{nRT} = \left(\frac{m}{n}\right)\frac{P}{RT}$$

The ratio *m*/*n* is the definition of molar mass, $\mathcal{M}$:

$$\mathcal{M} = \frac{m}{n}$$

The density equation can then be written

$$d = \frac{\mathcal{M}P}{RT}$$

This relation may be used for calculating the densities of gases of known identities at specified values of pressure and temperature as demonstrated in Example 9.4.1.

Example 9.4.1: Measuring Gas Density

What is the density of molecular nitrogen gas at STP?

Solution

The molar mass of molecular nitrogen, N_2, is 28.01 g/mol. Substituting this value along with standard temperature and pressure into the gas density equation yields

$$d = \frac{\mathscr{M}P}{RT} = \frac{(28.01\ \text{g/mol})(1.00\ \text{atm})}{\left(0.0821\ \text{L}\cdot\text{atm}\cdot\text{mol}^{-1}\text{K}^{-1}\right)(273\ \text{K})} = 1.25\ \text{g / L}$$

Exercise 9.4.1

What is the density of molecular hydrogen gas at 17.0 °C and a pressure of 760 torr?

Answer

d = 0.0847 g/L

When the identity of a gas is unknown, measurements of the mass, pressure, volume, and temperature of a sample can be used to calculate the molar mass of the gas (a useful property for identification purposes). Combining the ideal gas equation

$$PV = nRT$$

and the definition of molar mass

$$\mathscr{M} = \frac{m}{n}$$

yields the following equation:

$$\mathscr{M} = \frac{mRT}{PV}$$

Determining the molar mass of a gas via this approach is demonstrated in Example 9.4.2.

Example 9.4.2: Determining the Molecular Formula of a Gas from its Molar Mass and Empirical Formula

Cyclopropane, a gas once used with oxygen as a general anesthetic, is composed of 85.7% carbon and 14.3% hydrogen by mass. Find the empirical formula. If 1.56 g of cyclopropane occupies a volume of 1.00 L at 0.984 atm and 50 °C, what is the molecular formula for cyclopropane?

Solution

First determine the empirical formula of the gas. Assume 100 g and convert the percentage of each element into grams. Determine the number of moles of carbon and hydrogen in the 100-g sample of cyclopropane. Divide by the smallest number of moles to relate the number of moles of carbon to the number of moles of hydrogen.

$$85.7\ \text{g}\ C \times \frac{1\ \text{mol}\ C}{12.01\ \text{g}\ C} = 7.136\ \text{mol}\ C$$

$$14.3\ \text{g}\ H \times \frac{1\ \text{mol}\ H}{1.01\ \text{g}\ H} = 14.158\ \text{mol}\ H$$

In the last step, realize that the smallest whole number ratio is the empirical formula:

$$\frac{7.136}{7.136} = 1.00\ \text{mol}\ C$$

$$\frac{14.158}{7.136} = 1.98\ \text{mol}\ H$$

Empirical formula is CH_2 [empirical mass (EM) of 14.03 g/empirical unit].

Next, use the provided values for mass, pressure, temperature and volume to compute the molar mass of the gas:

$$\mathcal{M} = \frac{mRT}{PV} = \frac{(1.56\text{ g})\left(0.0821\text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\text{K}^{-1}\right)(323\text{ K})}{(0.984\text{ atm})(1.00\text{ L})} = 42.0\text{ g/mol}$$

Comparing the molar mass to the empirical formula mass shows how many empirical formula units make up a molecule:

$$\frac{\mathcal{M}}{EM} = \frac{42.0\text{ g/mol}}{14.0\text{ g/mol}} = 3$$

The molecular formula is thus derived from the empirical formula by multiplying each of its subscripts by three:

$$(CH_2)_3 = C_3H_6$$

Exercise 9.4.2

Acetylene, a fuel used welding torches, is composed of 92.3% C and 7.7% H by mass. Find the empirical formula. If 1.10 g of acetylene occupies of volume of 1.00 L at 1.15 atm and 59.5 °C, what is the molecular formula for acetylene?

Answer

Empirical formula, CH; Molecular formula, C_2H_2

Example 9.4.3: Determining the Molar Mass of a Volatile Liquid

The approximate molar mass of a volatile liquid can be determined by:

1. Heating a sample of the liquid in a flask with a tiny hole at the top, which converts the liquid into gas that may escape through the hole
2. Removing the flask from heat at the instant when the last bit of liquid becomes gas, at which time the flask will be filled with only gaseous sample at ambient pressure
3. Sealing the flask and permitting the gaseous sample to condense to liquid, and then weighing the flask to determine the sample's mass (see Figure 9.4.1)

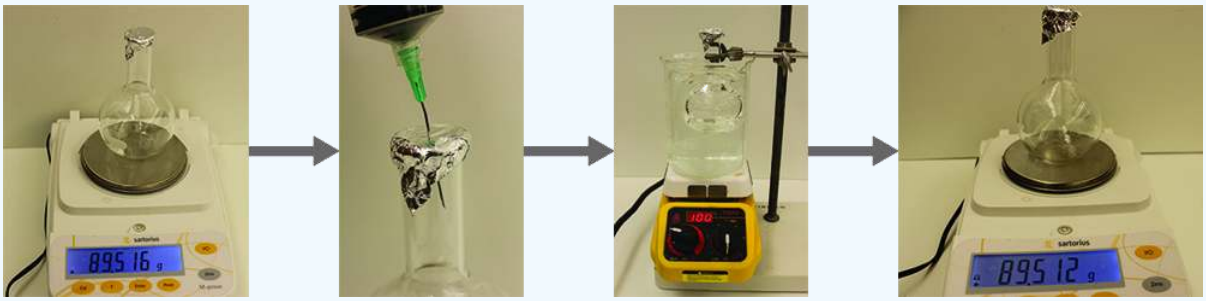


Figure 9.4.1: When the volatile liquid in the flask is heated past its boiling point, it becomes gas and drives air out of the flask. At the flask is filled with volatile liquid gas at the same pressure as the atmosphere. If the flask is then cooled to room temperature, the gas condenses and the mass of the gas that filled the flask, and is now liquid, can be measured. (credit: modification of work by Mark Ott)

Using this procedure, a sample of chloroform gas weighing 0.494 g is collected in a flask with a volume of 129 cm^3 at 99.6 °C when the atmospheric pressure is 742.1 mm Hg. What is the approximate molar mass of chloroform?

Solution

Since $\mathcal{M} = \frac{m}{n}$ and $n = \frac{PV}{RT}$ substituting and rearranging gives $\mathcal{M} = \frac{mRT}{PV}$,

then

$$\mathcal{M} = \frac{mRT}{PV} = \frac{(0.494\text{ g}) \times 0.08206\text{ L}\cdot\text{atm/mol K} \times 372.8\text{ K}}{0.976\text{ atm} \times 0.129\text{ L}} = 120\text{ g/mol}$$

? Exercise 9.4.3

A sample of phosphorus that weighs 3.243×10^{-2} g exerts a pressure of 31.89 kPa in a 56.0-mL bulb at 550 °C. What are the molar mass and molecular formula of phosphorus vapor?

Answer

124 g/mol P_4

The Pressure of a Mixture of Gases: Dalton's Law

Unless they chemically react with each other, the individual gases in a mixture of gases do not affect each other's pressure. Each individual gas in a mixture exerts the same pressure that it would exert if it were present alone in the container (Figure 9.4.2). The pressure exerted by each individual gas in a mixture is called its **partial pressure**. This observation is summarized by **Dalton's law of partial pressures**: *The total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases*:

$$P_{\text{Total}} = P_A + P_B + P_C + \ldots = \sum_i P_i$$

In the equation P_{Total} is the total pressure of a mixture of gases, P_A is the partial pressure of gas A; P_B is the partial pressure of gas B; P_C is the partial pressure of gas C; and so on.

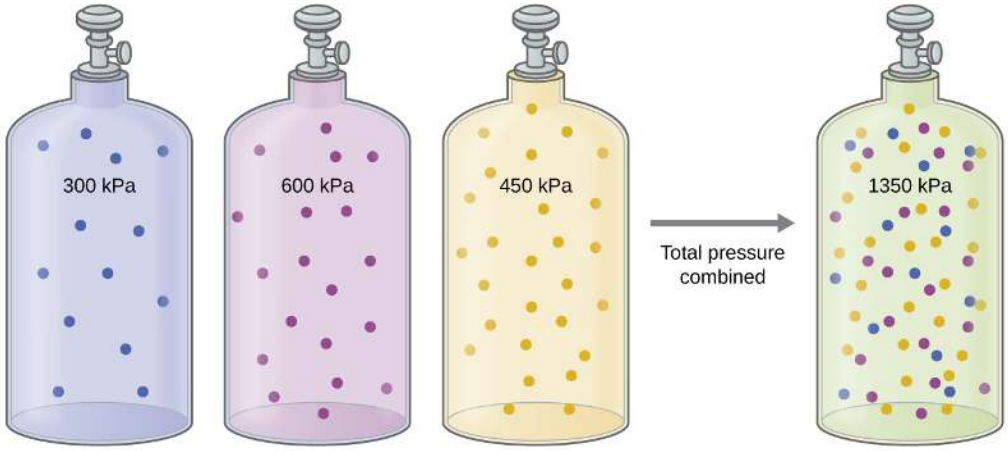


Figure 9.4.2: If equal-volume cylinders containing gasses at pressures of 300 kPa, 450 kPa, and 600 kPa are all combined in the same-size cylinder, the total pressure of the gas mixture is 1350 kPa.

The partial pressure of gas A is related to the total pressure of the gas mixture via its **mole fraction** (X), a unit of concentration defined as the number of moles of a component of a solution divided by the total number of moles of all components:

$$P_A = X_A \times P_{\text{Total}}$$

where

$$X_A = \frac{n_A}{n_{\text{Total}}}$$

where P_A, X_A, and n_A are the partial pressure, mole fraction, and number of moles of gas A, respectively, and n_{Total} is the number of moles of all components in the mixture.

✓ Example 9.4.4: The Pressure of a Mixture of Gases

A 10.0-L vessel contains 2.50 10^{-3} mol of H_2, 1.00 10^{-3} mol of He, and 3.00 10^{-4} mol of Ne at 35 °C.

a. What are the partial pressures of each of the gases?
b. What is the total pressure in atmospheres?

Solution

The gases behave independently, so the partial pressure of each gas can be determined from the ideal gas equation, using $P = \frac{nRT}{V}$:

$$P_{H_2} = \frac{\left(2.50 \times 10^{-3} \ \cancel{\text{mol}}\right)\left(0.08206 \ \cancel{\text{L}}\,\text{atm} \ \text{mol}^{-1}\,\cancel{\text{K}^{-1}}\right)(308 \ \cancel{\text{K}})}{10.0 \ \cancel{\text{L}}} = 6.32 \times 10^{-3} \text{ atm}$$

$$= \frac{\left(1.00 \times 10^{-3} \ \cancel{\text{mol}}\right)\left(0.08206 \ \cancel{\text{L}}\,\text{atm} \ \text{mol}^{-1}\,\cancel{\text{K}^{-1}}\right)(308 \ \cancel{\text{K}})}{10.0 \ \cancel{\text{L}}} = 2.53 \times 10^{-3} \text{ atm}$$

$$= \frac{\left(3.00 \times 10^{-4} \ \cancel{\text{mol}}\right)\left(0.08206 \ \cancel{\text{L}}\,\text{atm} \ \text{mol}^{-1}\,\cancel{\text{K}^{-1}}\right)(308 \ \cancel{\text{K}})}{10.0 \ \cancel{\text{L}}} = 7.58 \times 10^{-4} \text{ atm}$$

The total pressure is given by the sum of the partial pressures:

$$P_T = P_{H_2} + P_{He} + P_{Ne} = (0.00632 + 0.00253 + 0.00076) \text{ atm} = 9.61 \times 10^{-3} \text{ atm}$$

Exercise 9.4.4

A 5.73-L flask at 25 °C contains 0.0388 mol of N_2, 0.147 mol of CO, and 0.0803 mol of H_2. What is the total pressure in the flask in atmospheres?

Answer

1.137 atm

Here is another example of this concept, but dealing with mole fraction calculations.

Example 9.4.5: The Pressure of a Mixture of Gases

A gas mixture used for anesthesia contains 2.83 mol oxygen, O_2, and 8.41 mol nitrous oxide, N_2O. The total pressure of the mixture is 192 kPa.

a. What are the mole fractions of O_2 and N_2O?
b. What are the partial pressures of O_2 and N_2O?

Solution

The mole fraction is given by $X_A = \frac{n_A}{n_{\text{Total}}}$ and the partial pressure is $P_A = X_A P_{Total}$.

For O_2,

$$X_{O_2} = \frac{n_{O_2}}{n_{\text{Total}}} = \frac{2.83mol}{(2.83+8.41)mol} = 0.252$$

and $P_{O_2} = X_{O_2} \times P_{\text{Total}} = 0.252 \times 192kPa = 48.4kPa$

For N_2O,

$$X_{N_2O} = \frac{n_{N_2O}}{n_{\text{Total}}} = \frac{8.41mol}{(2.83+8.41)mol} = 0.748$$

and

$$P_{N_2O} = X_{N_2O} \times P_{\text{Total}} = 0.748 \times 192kPa = 144kPa$$

Exercise 9.4.5

What is the pressure of a mixture of 0.200 g of H_2, 1.00 g of N_2, and 0.820 g of Ar in a container with a volume of 2.00 L at 20 °C?

Answer

1.87 atm

Collection of Gases over Water

A simple way to collect gases that do not react with water is to capture them in a bottle that has been filled with water and inverted into a dish filled with water. The pressure of the gas inside the bottle can be made equal to the air pressure outside by raising or lowering the bottle. When the water level is the same both inside and outside the bottle (Figure 9.4.3), the pressure of the gas is equal to the atmospheric pressure, which can be measured with a barometer.

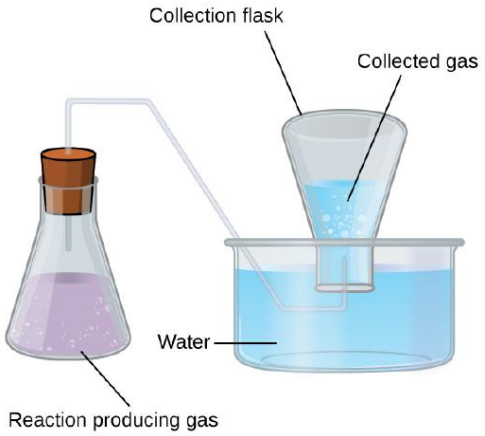


Figure 9.4.3: When a reaction produces a gas that is collected above water, the trapped gas is a mixture of the gas produced by the reaction and water vapor. If the collection flask is appropriately positioned to equalize the water levels both within and outside the flask, the pressure of the trapped gas mixture will equal the atmospheric pressure outside the flask (see the earlier discussion of manometers).

However, there is another factor we must consider when we measure the pressure of the gas by this method. Water evaporates and there is always gaseous water (water vapor) above a sample of liquid water. As a gas is collected over water, it becomes saturated with water vapor and the total pressure of the mixture equals the partial pressure of the gas plus the partial pressure of the water vapor. The pressure of the pure gas is therefore equal to the total pressure minus the pressure of the water vapor—this is referred to as the "dry" gas pressure, that is, the pressure of the gas only, without water vapor. The **vapor pressure of water**, which is the pressure exerted by water vapor in equilibrium with liquid water in a closed container, depends on the temperature (Figure 9.4.4); more detailed information on the temperature dependence of water vapor can be found in Table 9.4.1, and vapor pressure will be discussed in more detail in the next chapter on liquids.

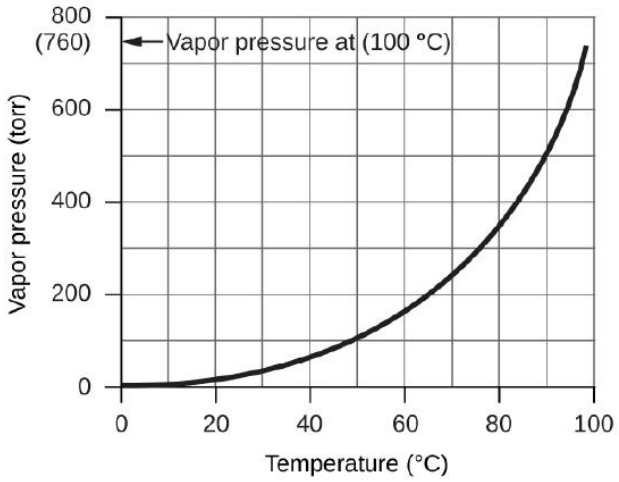


Figure 9.4.4: This graph shows the vapor pressure of water at sea level as a function of temperature.

Table 9.4.1: Vapor Pressure of Ice and Water in Various Temperatures at Sea Level

Temperature (°C)	Pressure (torr)	Temperature (°C)	Pressure (torr)	Temperature (°C)	Pressure (torr)
–10	1.95	18	15.5	30	31.8
–5	3.0	19	16.5	35	42.2
–2	3.9	20	17.5	40	55.3
0	4.6	21	18.7	50	92.5
2	5.3	22	19.8	60	149.4
4	6.1	23	21.1	70	233.7
6	7.0	24	22.4	80	355.1
8	8.0	25	23.8	90	525.8
10	9.2	26	25.2	95	633.9
12	10.5	27	26.7	99	733.2
14	12.0	28	28.3	100.0	760.0
16	13.6	29	30.0	101.0	787.6

Example 9.4.6: Pressure of a Gas Collected Over Water

If 0.200 L of argon is collected over water at a temperature of 26 °C and a pressure of 750 torr in a system like that shown in Figure 9.4.3, what is the partial pressure of argon?

Solution

According to Dalton's law, the total pressure in the bottle (750 torr) is the sum of the partial pressure of argon and the partial pressure of gaseous water:

$$P_T = P_{Ar} + P_{H_2O}$$

Rearranging this equation to solve for the pressure of argon gives:

$$P_{Ar} = P_T - P_{H_2O}$$

The pressure of water vapor above a sample of liquid water at 26 °C is 25.2 torr (Appendix E), so:

$$P_{Ar} = 750 \text{ torr } - 25.2 \text{ torr } = 725 \text{ torr}$$

Exercise 9.4.6

A sample of oxygen collected over water at a temperature of 29.0 °C and a pressure of 764 torr has a volume of 0.560 L. What volume would the dry oxygen from this sample have under the same conditions of temperature and pressure?

Answer

0.537 L

Chemical Stoichiometry and Gases

Chemical stoichiometry describes the quantitative relationships between reactants and products in chemical reactions.

We have previously measured quantities of reactants and products using masses for solids and volumes in conjunction with the molarity for solutions; now we can also use gas volumes to indicate quantities. If we know the volume, pressure, and temperature of a gas, we can use the ideal gas equation to calculate how many moles of the gas are present. If we know how many moles of a gas are involved, we can calculate the volume of a gas at any temperature and pressure.

Avogadro's Law Revisited

Sometimes we can take advantage of a simplifying feature of the stoichiometry of gases that solids and solutions do not exhibit: All gases that show ideal behavior contain the same number of molecules in the same volume (at the same temperature and pressure). Thus, the ratios of volumes of gases involved in a chemical reaction are given by the coefficients in the equation for the reaction, provided that the gas volumes are measured at the same temperature and pressure.

We can extend Avogadro's law (that the volume of a gas is directly proportional to the number of moles of the gas) to chemical reactions with gases: Gases combine, or react, in definite and simple proportions by volume, provided that all gas volumes are measured at the same temperature and pressure. For example, since nitrogen and hydrogen gases react to produce ammonia gas according to

$$\mathrm{N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)}$$

a given volume of nitrogen gas reacts with three times that volume of hydrogen gas to produce two times that volume of ammonia gas, if pressure and temperature remain constant.

The explanation for this is illustrated in Figure 9.4.5. According to Avogadro's law, equal volumes of gaseous N_2, H_2, and NH_3, at the same temperature and pressure, contain the same number of molecules. Because one molecule of N_2 reacts with three molecules of H_2 to produce two molecules of NH_3, the volume of H_2 required is three times the volume of N_2, and the volume of NH_3 produced is two times the volume of N_2.

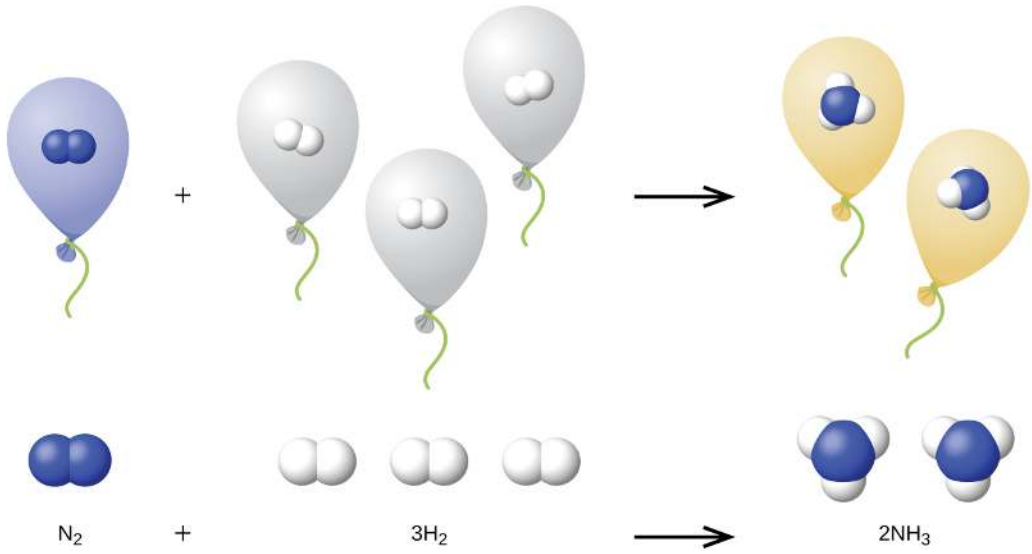


Figure 9.4.5: One volume of N_2 combines with three volumes of H_2 to form two volumes of NH_3.

https://chem.libretexts.org/@go/page/482978

Example 9.4.7: Reaction of Gases

Propane, $C_3H_8(g)$, is used in gas grills to provide the heat for cooking. What volume of $O_2(g)$ measured at 25 °C and 760 torr is required to react with 2.7 L of propane measured under the same conditions of temperature and pressure? Assume that the propane undergoes complete combustion.

Solution

The ratio of the volumes of C_3H_8 and O_2 will be equal to the ratio of their coefficients in the balanced equation for the reaction:

$$\underset{C_3H_8(g)+5\ O_2(g)}{1\text{ volume} + 5\text{ volumes}} \longrightarrow \underset{3\ CO_2(g)+4\ H_2O(l)}{3\text{ volumes} + 4\text{ volumes}}$$

From the equation, we see that one volume of C_3H_8 will react with five volumes of O_2:

$$2.7\ \cancel{L}\ C_3H_8 \times \frac{5\ L\ O_2}{1\ \cancel{L C_3H_8}} = 13.5\ L\ O_2$$

A volume of 13.5 L of O_2 will be required to react with 2.7 L of C_3H_8.

Exercise 9.4.7

An acetylene tank for an oxyacetylene welding torch provides 9340 L of acetylene gas, C_2H_2, at 0 °C and 1 atm. How many tanks of oxygen, each providing 7.00 10^3 L of O_2 at 0 °C and 1 atm, will be required to burn the acetylene?

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

Answer

3.34 tanks (2.34 10^4 L)

Example 9.4.8: Volumes of Reacting Gases

Ammonia is an important fertilizer and industrial chemical. Suppose that a volume of 683 billion cubic feet of gaseous ammonia, measured at 25 °C and 1 atm, was manufactured. What volume of $H_2(g)$, measured under the same conditions, was required to prepare this amount of ammonia by reaction with N_2?

$$N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

Solution

Because equal volumes of H_2 and NH_3 contain equal numbers of molecules and each three molecules of H_2 that react produce two molecules of NH_3, the ratio of the volumes of H_2 and NH_3 will be equal to 3:2. Two volumes of NH_3, in this case in units of billion ft^3, will be formed from three volumes of H_2:

$$683\text{ billion } ft^3 NH_3 \times \frac{3\text{ billion } ft^3 H_2}{2\text{ billion } ft^3 NH_3} = 1.02 \times 10^3\text{ billion } ft^3 H_2$$

The manufacture of 683 billion ft^3 of NH_3 required 1020 billion ft^3 of H_2. (At 25 °C and 1 atm, this is the volume of a cube with an edge length of approximately 1.9 miles.)

Exercise 9.4.8

What volume of $O_2(g)$ measured at 25 °C and 760 torr is required to react with 17.0 L of ethylene, $C_2H_4(g)$, measured under the same conditions of temperature and pressure? The products are CO_2 and water vapor.

Answer

51.0 L

Example 9.4.9: Volume of Gaseous Product

What volume of hydrogen at 27 °C and 723 torr may be prepared by the reaction of 8.88 g of gallium with an excess of hydrochloric acid?

$$2\,\mathrm{Ga(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{GaCl_3(aq)} + 3\,\mathrm{H_2(g)}$$

Solution

Convert the provided mass of the limiting reactant, Ga, to moles of hydrogen produced:

$$8.88\ \cancel{\text{g Ga}} \times \frac{1\ \cancel{\text{mol Ga}}}{69.723\ \cancel{\text{g Ga}}} \times \frac{3\ \text{mol H}_2}{2\ \cancel{\text{mol Ga}}} = 0.191\ \text{mol H}_2$$

Convert the provided temperature and pressure values to appropriate units (K and atm, respectively), and then use the molar amount of hydrogen gas and the ideal gas equation to calculate the volume of gas:

$$V = \left(\frac{nRT}{P}\right) = \frac{0.191\ \cancel{\text{mol}} \times 0.08206\ \text{L atm}\ \cancel{\text{mol}^{-1}}\text{K}^{-1} \times 300\ \text{K}}{0.951\ \cancel{\text{atm}}} = 4.94\ \text{L}$$

Exercise 9.4.9

Sulfur dioxide is an intermediate in the preparation of sulfuric acid. What volume of SO_2 at 343 °C and 1.21 atm is produced by burning l.00 kg of sulfur in excess oxygen?

Answer

1.30 10^3 L

How Sciences Interconnect: Greenhouse Gases and Climate Change

The thin skin of our atmosphere keeps the earth from being an ice planet and makes it habitable. In fact, this is due to less than 0.5% of the air molecules. Of the energy from the sun that reaches the earth, almost $\frac{1}{3}$ is reflected back into space, with the rest absorbed by the atmosphere and the surface of the earth. Some of the energy that the earth absorbs is re-emitted as infrared (IR) radiation, a portion of which passes back out through the atmosphere into space. Most if this IR radiation, however, is absorbed by certain atmospheric gases, effectively trapping heat within the atmosphere in a phenomenon known as the greenhouse effect. This effect maintains global temperatures within the range needed to sustain life on earth. Without our atmosphere, the earth's average temperature would be lower by more than 30 °C (nearly 60 °F). The major greenhouse gases (GHGs) are water vapor, carbon dioxide, methane, and ozone. Since the Industrial Revolution, human activity has been increasing the concentrations of GHGs, which have changed the energy balance and are significantly altering the earth’s climate (Figure 9.4.6).

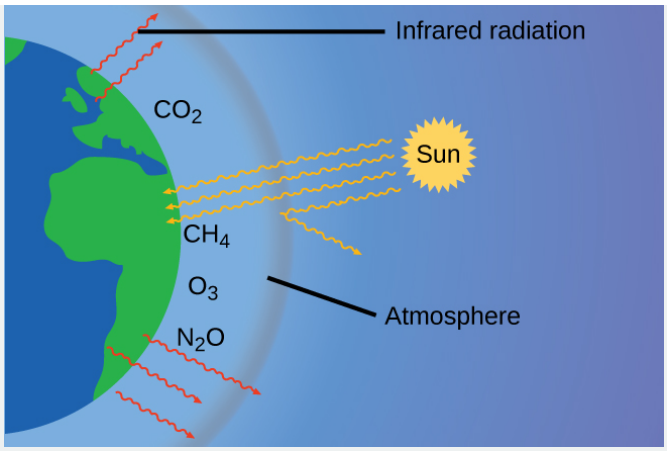


Figure 9.4.6: Greenhouse gases trap enough of the sun's energy to make the planet habitable—this is known as the greenhouse effect. Human activities are increasing greenhouse gas levels, warming the planet and causing more extreme weather events.

There is strong evidence from multiple sources that higher atmospheric levels of CO2 are caused by human activity, with fossil fuel burning accounting for about $\frac{3}{4}$ of the recent increase in CO2. Reliable data from ice cores reveals that CO2 concentration in the atmosphere is at the highest level in the past 800,000 years; other evidence indicates that it may be at its highest level in 20 million years. In recent years, the CO2 concentration has increased preindustrial levels of ~280 ppm to more than 400 ppm today (Figure 9.4.7).

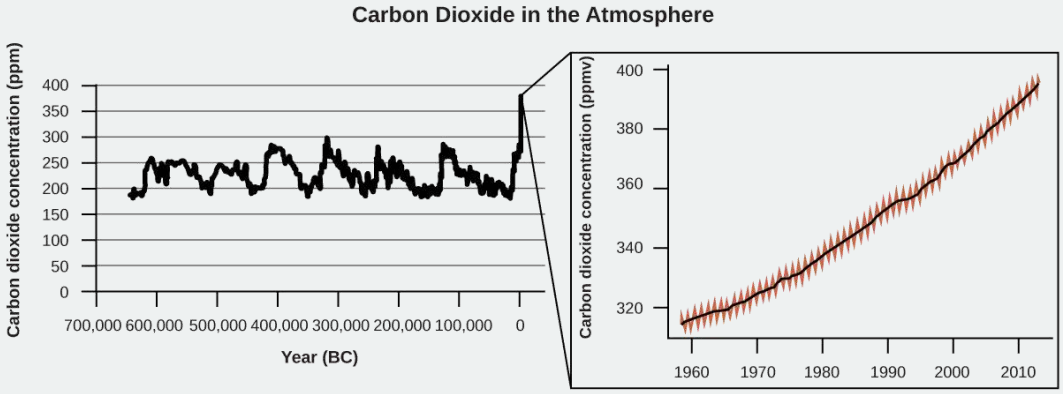


Figure 9.4.7: CO_2 levels over the past 700,000 years were typically from 200–300 ppm, with a steep, unprecedented increase over the past 50 years.

Link to Learning

Click here to see a 2-minute video explaining greenhouse gases and global warming.

Portrait of a Chemist: Susan Solomon

Atmospheric and climate scientist Susan Solomon (Figure 9.4.8) is the author of one of *The New York Times* books of the year (*The Coldest March*, 2001), one of Time magazine's 100 most influential people in the world (2008), and a working group leader of the Intergovernmental Panel on Climate Change (IPCC), which was the recipient of the 2007 Nobel Peace Prize. She helped determine and explain the cause of the formation of the ozone hole over Antarctica, and has authored many important papers on climate change. She has been awarded the top scientific honors in the US and France (the National Medal of Science and the Grande Medaille, respectively), and is a member of the National Academy of Sciences, the Royal Society, the French Academy of Sciences, and the European Academy of Sciences. Formerly a professor at the University of Colorado, she is now at MIT, and continues to work at NOAA.



Figure 9.4.8: Susan Solomon's research focuses on climate change and has been instrumental in determining the cause of the ozone hole over Antarctica. (credit: National Oceanic and Atmospheric Administration)

For more information, watch this video about Susan Solomon.

Footnotes

- 2"Quotations by Joseph-Louis Lagrange," last modified February 2006, accessed February 10, 2015, www-history.mcs.st-andrews.ac.../Lagrange.html

---

This page titled 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.3: Stoichiometry of Gaseous Substances, Mixtures, and Reactions** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.5: Effusion and Diffusion of Gases

Learning Objectives

By the end of this section, you will be able to:

- Define and explain effusion and diffusion
- State Graham's law and use it to compute relevant gas properties

If you have ever been in a room when a piping hot pizza was delivered, you have been made aware of the fact that gaseous molecules can quickly spread throughout a room, as evidenced by the pleasant aroma that soon reaches your nose. Although gaseous molecules travel at tremendous speeds (hundreds of meters per second), they collide with other gaseous molecules and travel in many different directions before reaching the desired target. At room temperature, a gaseous molecule will experience billions of collisions per second. The **mean free path** is the average distance a molecule travels between collisions. The mean free path increases with decreasing pressure; in general, the mean free path for a gaseous molecule will be hundreds of times the diameter of the molecule

In general, we know that when a sample of gas is introduced to one part of a closed container, its molecules very quickly disperse throughout the container; this process by which molecules disperse in space in response to differences in concentration is called **diffusion** (shown in Figure 9.5.1). The gaseous atoms or molecules are, of course, unaware of any concentration gradient, they simply move randomly—regions of higher concentration have more particles than regions of lower concentrations, and so a net movement of species from high to low concentration areas takes place. In a closed environment, diffusion will ultimately result in equal concentrations of gas throughout, as depicted in Figure 9.5.1. The gaseous atoms and molecules continue to move, but since their concentrations are the same in both bulbs, the rates of transfer between the bulbs are equal (no *net* transfer of molecules occurs).

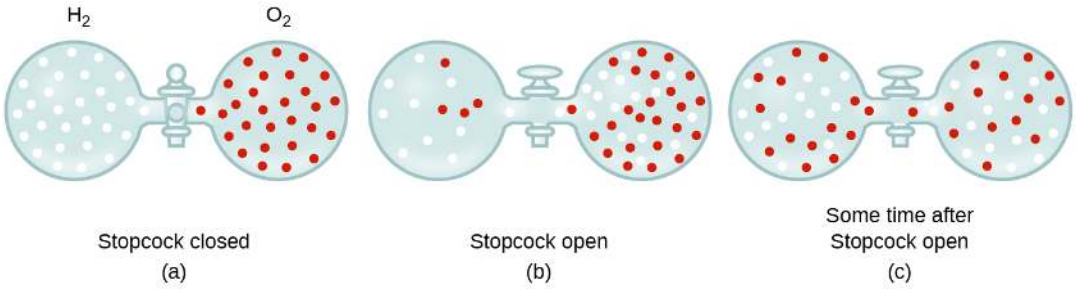


Figure 9.5.1: (a) Two gases, H_2 and O_2, are initially separated. (b) When the stopcock is opened, they mix together. The lighter gas, H_2, passes through the opening faster than O_2, so just after the stopcock is opened, more H_2 molecules move to the O_2 side than O_2 molecules move to the H_2 side. (c) After a short time, both the slower-moving O_2 molecules and the faster-moving H_2 molecules have distributed themselves evenly on both sides of the vessel.

We are often interested in the **rate of diffusion**, the amount of gas passing through some area per unit time:

$$\text{rate of diffusion} = \frac{\text{amount of gas passing through an area}}{\text{unit of time}}$$

The diffusion rate depends on several factors: temperature; the mass of the atoms or molecules; the concentration gradient (the increase or decrease in concentration from one point to another); the amount of surface area available for diffusion; and the distance the gas particles must travel. Note also that the time required for diffusion to occur is inversely proportional to the rate of diffusion, as shown in the rate of diffusion equation.

A process involving movement of gaseous species similar to diffusion is **effusion**, the escape of gas molecules through a tiny hole such as a pinhole in a balloon into a vacuum (Figure 9.5.2). Although diffusion and effusion rates both depend on the molar mass of the gas involved, their rates are not equal; however, the ratios of their rates are the same.

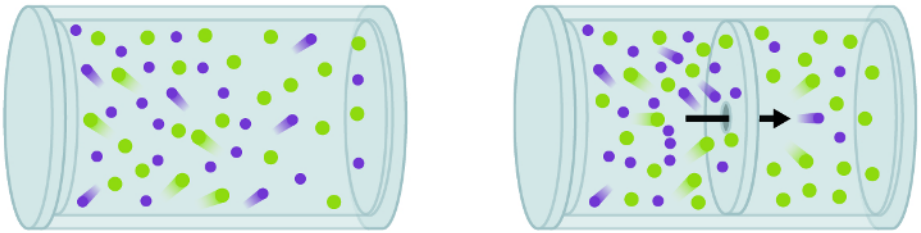


Figure 9.5.2: Diffusion involves the unrestricted dispersal of molecules throughout space due to their random motion. When this process is restricted to passage of molecules through very small openings in a physical barrier, the process is called effusion.

If a mixture of gases is placed in a container with porous walls, the gases effuse through the small openings in the walls. The lighter gases pass through the small openings more rapidly (at a higher rate) than the heavier ones (Figure 9.5.3). In 1832, Thomas Graham studied the rates of effusion of different gases and formulated **Graham's law of effusion**: *The rate of effusion of a gas is inversely proportional to the square root of the mass of its particles*:

$$\text{rate of effusion} \propto \frac{1}{\sqrt{M}}$$

This means that if two gases A and B are at the same temperature and pressure, the ratio of their effusion rates is inversely proportional to the ratio of the square roots of the masses of their particles:

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

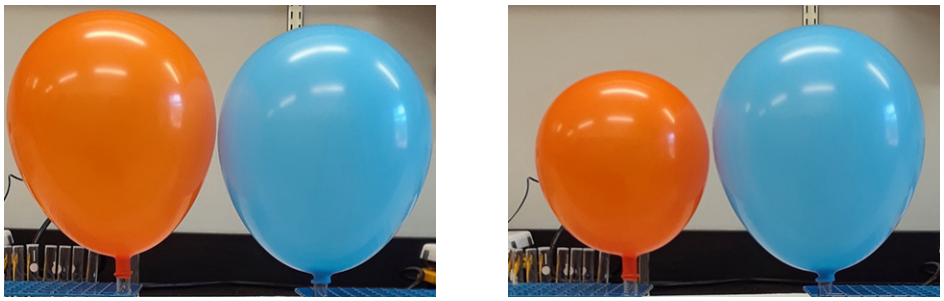


Figure 9.5.3: The left photograph shows two balloons inflated with different gases, helium (orange) and argon (blue).The right-side photograph shows the balloons approximately 12 hours after being filled, at which time the helium balloon has become noticeably more deflated than the argon balloon, due to the greater effusion rate of the lighter helium gas. (credit: modification of work by Paul Flowers)

Example 9.5.1: Applying Graham's Law to Rates of Effusion

Calculate the ratio of the rate of effusion of hydrogen to the rate of effusion of oxygen.

Solution

From Graham's law, we have:

$$\frac{\text{rate of effusion of hydrogen}}{\text{rate of effusion of oxygen}} = \frac{\sqrt{32 gmol^{-1}}}{\sqrt{2 gmol^{-1}}} = \frac{\sqrt{16}}{\sqrt{1}} = \frac{4}{1}$$

Hydrogen effuses four times as rapidly as oxygen.

Exercise 9.5.1

At a particular pressure and temperature, nitrogen gas effuses at the rate of 79 mL/s. Under the same conditions, at what rate will sulfur dioxide effuse?

Answer

52 mL/s

Example 9.5.2: Effusion Time Calculations

It takes 243 s for 4.46 10^{-5} mol Xe to effuse through a tiny hole. Under the same conditions, how long will it take 4.46 10^{-5} mol Ne to effuse?

Solution

It is important to resist the temptation to use the times directly, and to remember how rate relates to time as well as how it relates to mass. Recall the definition of rate of effusion:

$$\text{rate of effusion} = \frac{\text{amount of gas transferred}}{\text{time}}$$

and combine it with Graham's law:

$$\frac{\text{rate of effusion of gas Xe}}{\text{rate of effusion of gas Ne}} = \frac{\sqrt{M_{Ne}}}{\sqrt{M_{Xe}}}$$

To get:

$$\frac{\frac{\text{amount of Xe transferred}}{\text{time for Xe}}}{\frac{\text{amount of Ne transferred}}{\text{time for Ne}}} = \frac{\sqrt{M_{Ne}}}{\sqrt{M_{Xe}}}$$

Noting that *amount of A* = *amount of B*, and solving for *time for Ne*:

$$\frac{\frac{\text{amount of } Xe}{\text{time for Xe}}}{\frac{\text{amount of } Ne}{\text{time for } Ne}} = \frac{\text{time for } Ne}{\text{time for Xe}} = \frac{\sqrt{M_{Ne}}}{\sqrt{M_{Xe}}} = \frac{\sqrt{M_{Ne}}}{\sqrt{M_{Xe}}}$$

and substitute values:

$$\frac{\text{time for Ne}}{243s} = \sqrt{\frac{20.2 \ \cancel{\text{g mol}}}{131.3 \ \cancel{\text{g mol}}}} = 0.392$$

Finally, solve for the desired quantity:

$$\text{time for } Ne = 0.392 \times 243 \text{ s} = 95.3 \text{ s}$$

Note that this answer is reasonable: Since Ne is lighter than Xe, the effusion rate for Ne will be larger than that for Xe, which means the time of effusion for Ne will be smaller than that for Xe.

Exercise 9.5.2

A party balloon filled with helium deflates to $\frac{2}{3}$ of its original volume in 8.0 hours. How long will it take an identical balloon filled with the same number of moles of air ($\mathcal{M}$ = 28.2 g/mol) to deflate to $\frac{1}{2}$ of its original volume?

Answer

32 h

Example 9.5.3: Determining Molar Mass Using Graham's Law

An unknown gas effuses 1.66 times more rapidly than CO_2. What is the molar mass of the unknown gas? Can you make a reasonable guess as to its identity?

Solution

From Graham's law, we have:

$$\frac{\text{rate of effusion of Unknown}}{\text{rate of effusion of } CO_2} = \frac{\sqrt{M_{CO_2}}}{\sqrt{M_{\text{Unknown}}}}$$

Plug in known data:

$$\frac{1.66}{1} = \frac{\sqrt{44.0 g/mol}}{\sqrt{M_{\text{Unknown}}}}$$

Solve:

$$M_{\text{Unknown}} = \frac{44.0 g/mol}{(1.66)^2} = 16.0 g/mol$$

The gas could well be CH_4, the only gas with this molar mass.

Exercise 9.5.3

Hydrogen gas effuses through a porous container 8.97-times faster than an unknown gas. Estimate the molar mass of the unknown gas.

Answer

163 g/mol

How Sciences Interconnect: Use of Diffusion for Nuclear Energy Applications: Uranium Enrichment

Gaseous diffusion has been used to produce enriched uranium for use in nuclear power plants and weapons. Naturally occurring uranium contains only 0.72% of ^{235}U, the kind of uranium that is "fissile," that is, capable of sustaining a nuclear fission chain reaction. Nuclear reactors require fuel that is 2–5% ^{235}U, and nuclear bombs need even higher concentrations. One way to enrich uranium to the desired levels is to take advantage of Graham's law. In a gaseous diffusion enrichment plant, uranium hexafluoride (UF_6, the only uranium compound that is volatile enough to work) is slowly pumped through large cylindrical vessels called diffusers, which contain porous barriers with microscopic openings. The process is one of diffusion because the other side of the barrier is not evacuated. The $^{235}UF_6$ molecules have a higher average speed and diffuse through the barrier a little faster than the heavier $^{238}UF_6$ molecules. The gas that has passed through the barrier is slightly enriched in $^{235}UF_6$ and the residual gas is slightly depleted. The small difference in molecular weights between $^{235}UF_6$ and $^{238}UF_6$ only about 0.4% enrichment, is achieved in one diffuser (Figure 9.5.4). But by connecting many diffusers in a sequence of stages (called a cascade), the desired level of enrichment can be attained.

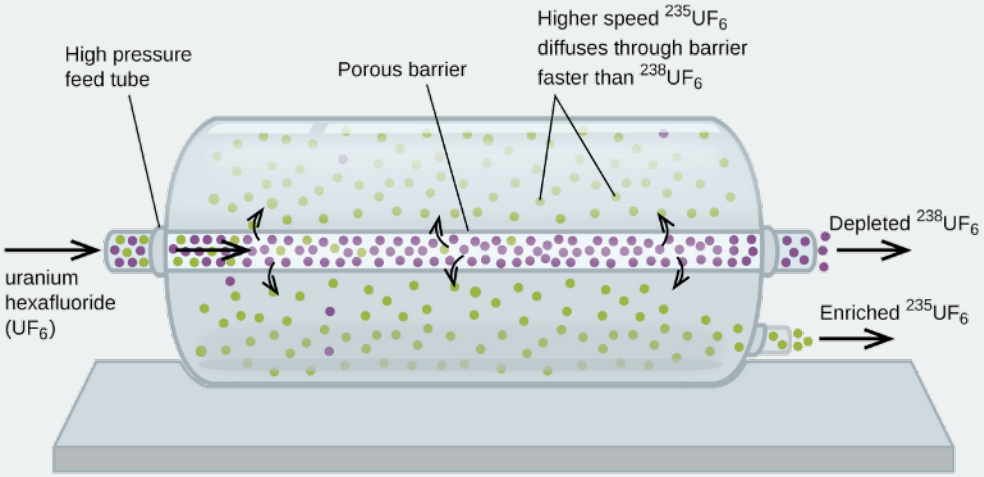


Figure 9.5.4: In a diffuser, gaseous UF_6 is pumped through a porous barrier, which partially separates $^{235}UF_6$ from $^{238}UF_6$ The UF_6 must pass through many large diffuser units to achieve sufficient enrichment in ^{235}U.

The large scale separation of gaseous $^{235}UF_6$ from $^{238}UF_6$ was first done during the World War II, at the atomic energy installation in Oak Ridge, Tennessee, as part of the Manhattan Project (the development of the first atomic bomb). Although the theory is simple, this required surmounting many daunting technical challenges to make it work in practice. The barrier must have tiny, uniform holes (about 10^{-6} cm in diameter) and be porous enough to produce high flow rates. All materials (the barrier, tubing, surface coatings, lubricants, and gaskets) need to be able to contain, but not react with, the highly reactive and corrosive UF_6.

Because gaseous diffusion plants require very large amounts of energy (to compress the gas to the high pressures required and drive it through the diffuser cascade, to remove the heat produced during compression, and so on), it is now being replaced by gas centrifuge technology, which requires far less energy. A current hot political issue is how to deny this technology to Iran, to prevent it from producing enough enriched uranium for them to use to make nuclear weapons.

This page titled 9.5: Effusion and Diffusion of Gases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.4: Effusion and Diffusion of Gases** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.6: The Kinetic-Molecular Theory

Learning Objectives

- State the postulates of the kinetic-molecular theory
- Use this theory's postulates to explain the gas laws

The gas laws that we have seen to this point, as well as the ideal gas equation, are empirical, that is, they have been derived from experimental observations. The mathematical forms of these laws closely describe the macroscopic behavior of most gases at pressures less than about 1 or 2 atm. Although the gas laws describe relationships that have been verified by many experiments, they do not tell us why gases follow these relationships.

The **kinetic molecular theory** (KMT) is a simple microscopic model that effectively explains the gas laws described in previous modules of this chapter. This theory is based on the following five postulates described here. (Note: The term "molecule" will be used to refer to the individual chemical species that compose the gas, although some gases are composed of atomic species, for example, the noble gases.)

1. Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they collide with other molecules or with the walls of a container.
2. The molecules composing the gas are negligibly small compared to the distances between them.
3. The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls.
4. Gas molecules exert no attractive or repulsive forces on each other or the container walls; therefore, their collisions are *elastic* (do not involve a loss of energy).
5. The average kinetic energy of the gas molecules is proportional to the kelvin temperature of the gas.

The test of the KMT and its postulates is its ability to explain and describe the behavior of a gas. The various gas laws can be derived from the assumptions of the KMT, which have led chemists to believe that the assumptions of the theory accurately represent the properties of gas molecules. We will first look at the individual gas laws (Boyle's, Charles's, Amontons's, Avogadro's, and Dalton's laws) conceptually to see how the KMT explains them. Then, we will more carefully consider the relationships between molecular masses, speeds, and kinetic energies with temperature, and explain Graham's law.

The Kinetic-Molecular Theory Explains the Behavior of Gases, Part I

Recalling that gas pressure is exerted by rapidly moving gas molecules and depends directly on the number of molecules hitting a unit area of the wall per unit of time, we see that the KMT conceptually explains the behavior of a gas as follows:

- *Amontons's law.* If the temperature is increased, the average speed and kinetic energy of the gas molecules increase. If the volume is held constant, the increased speed of the gas molecules results in more frequent and more forceful collisions with the walls of the container, therefore increasing the pressure (Figure 9.6.1).
- *Charles's law.* If the temperature of a gas is increased, a constant pressure may be maintained only if the volume occupied by the gas increases. This will result in greater average distances traveled by the molecules to reach the container walls, as well as increased wall surface area. These conditions will decrease the both the frequency of molecule-wall collisions and the number of collisions per unit area, the combined effects of which balance the effect of increased collision forces due to the greater kinetic energy at the higher temperature.
- *Boyle's law.* If the gas volume volume of a given amount of gas at a given temperature is decreased (that is, if the gas is *compressed*), the molecules will be exposed to a decreased container wall area. Collisions with the container wall will therefore occur more frequently and the pressure exerted by the gas will increase (Figure 9.6.1).
- *Avogadro's law.* At constant pressure and temperature, the frequency and force of molecule-wall collisions are constant. Under such conditions, increasing the number of gaseous molecules will require a proportional increase in the container volume in order to yield a decrease in the number of collisions per unit area to compensate for the increased frequency of collisions (Figure 9.6.1).
- *Dalton's Law.* Because of the large distances between them, the molecules of one gas in a mixture bombard the container walls with the same frequency whether other gases are present or not, and the total pressure of a gas mixture equals the sum of the (partial) pressures of the individual gases.

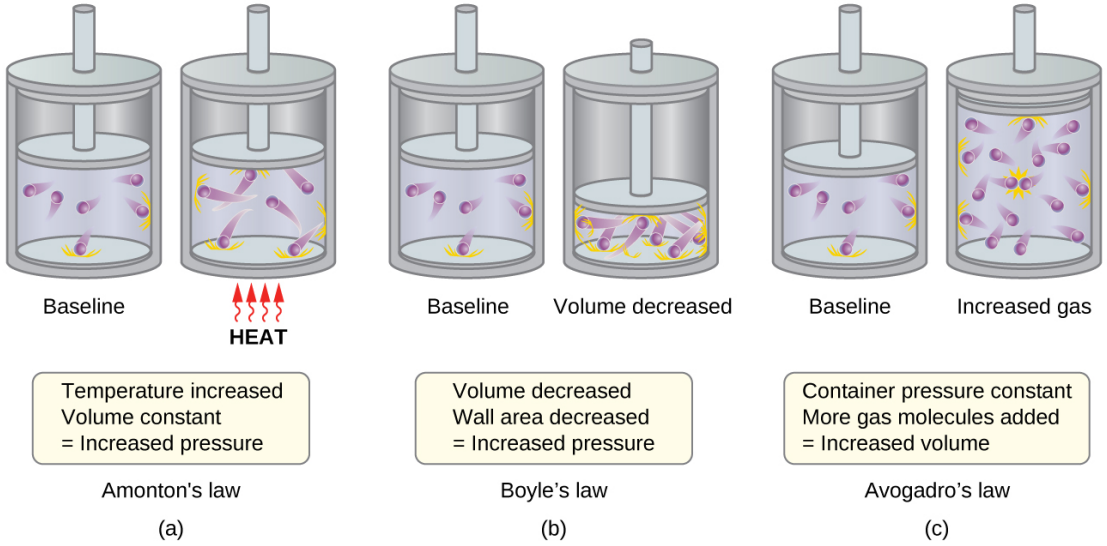


Figure 9.6.1: (a) When gas temperature increases, gas pressure increases due to increased force and frequency of molecular collisions. (b) When volume decreases, gas pressure increases due to increased frequency of molecular collisions. (c) When the amount of gas increases at a constant pressure, volume increases to yield a constant number of collisions per unit wall area per unit time.

molecular speeds and Kinetic Energy

The previous discussion showed that the KMT qualitatively explains the behaviors described by the various gas laws. The postulates of this theory may be applied in a more quantitative fashion to derive these individual laws. To do this, we must first look at speeds and kinetic energies of gas molecules, and the temperature of a gas sample.

In a gas sample, individual molecules have widely varying speeds; however, because of the *vast* number of molecules and collisions involved, the molecular speed distribution and average speed are constant. This molecular speed distribution is known as a Maxwell-Boltzmann distribution, and it depicts the relative numbers of molecules in a bulk sample of gas that possesses a given speed (Figure 9.6.2).

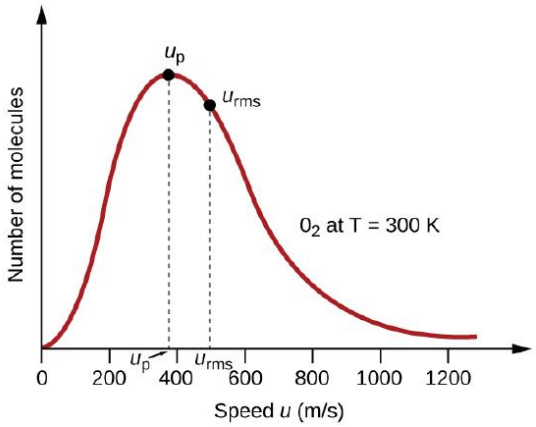


Figure 9.6.2: The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, v_p, is a little less than 400 m/s, while the root mean square speed, u_{rms}, is closer to 500 m/s.

The kinetic energy (KE) of a particle of mass (m) and speed (u) is given by:

$$KE = \frac{1}{2}mu^2$$

Expressing mass in kilograms and speed in meters per second will yield energy values in units of joules (J = kg m^2 s^{-2}). To deal with a large number of gas molecules, we use averages for both speed and kinetic energy. In the KMT, the root mean square speed of a particle, u_{rms}, is defined as the square root of the average of the squares of the speeds with n = the number of particles:

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \ldots}{n}}$$

The average kinetic energy for a mole of particles, KE_{avg}, is then equal to:

$$KE_{avg} = \frac{1}{2} M u_{rms}^2$$

where M is the molar mass expressed in units of kg/mol. The KE_{avg} of a mole of gas molecules is also directly proportional to the temperature of the gas and may be described by the equation:

$$KE_{avg} = \frac{3}{2} RT$$

where R is the gas constant and T is the kelvin temperature. When used in this equation, the appropriate form of the gas constant is 8.314 J/mol·K (8.314 kg $m^2 s^{-2} mol^{-1} K^{-1}$). These two separate equations for KE_{avg} may be combined and rearranged to yield a relation between molecular speed and temperature:

$$\frac{1}{2} M u_{rms}^2 = \frac{3}{2} RT$$

with

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

Example 9.6.1: Calculation of u_{rms}

Calculate the root-mean-square speed for a nitrogen molecule at 30 °C.

Solution

Convert the temperature into Kelvin:

$$30^{\circ} C + 273 = 303 K$$

Determine the molar mass of nitrogen in kilograms:

\frac{28.0 ~ \cancel{\text{g}}}{1 ~\text{mol} } \times \frac{1~\text{kg} }{1000 ~ \cancel{\text{g}}}=0.028~ \text{kg/mol} \nonumber

Replace the variables and constants in the root-mean-square speed equation, replacing Joules with the equivalent kg $m^2 s^{-2}$:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

$$u_{rms} = \sqrt{\frac{3(8.314 J/molK)(303 K)}{(0.028 kg/mol)}} = \sqrt{2.70 \times 10^5 m^2 s^{-2}} = 519 m/s$$

Exercise 9.6.1

Calculate the root-mean-square speed for a mole of oxygen molecules at –23 °C.

Answer

441 m/s

If the temperature of a gas increases, its KE_{avg} increases, more molecules have higher speeds and fewer molecules have lower speeds, and the distribution shifts toward higher speeds overall, that is, to the right. If temperature decreases, KE_{avg} decreases, more molecules have lower speeds and fewer molecules have higher speeds, and the distribution shifts toward lower speeds overall, that is, to the left. This behavior is illustrated for nitrogen gas in Figure 9.6.3.

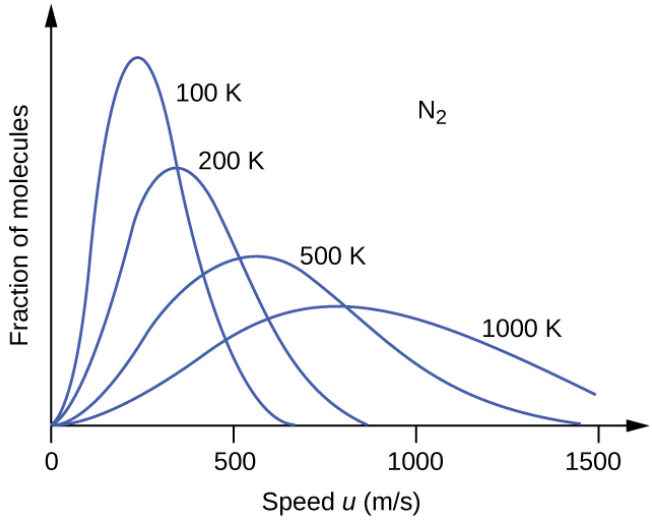


Figure 9.6.3: The molecular speed distribution for nitrogen gas (N_2) shifts to the right and flattens as the temperature increases; it shifts to the left and heightens as the temperature decreases.

At a given temperature, all gases have the same KE_{avg} for their molecules. Gases composed of lighter molecules have more high-speed particles and a higher u_{rms}, with a speed distribution that peaks at relatively higher speeds. Gases consisting of heavier molecules have more low-speed particles, a lower u_{rms}, and a speed distribution that peaks at relatively lower speeds. This trend is demonstrated by the data for a series of noble gases shown in Figure 9.6.4.

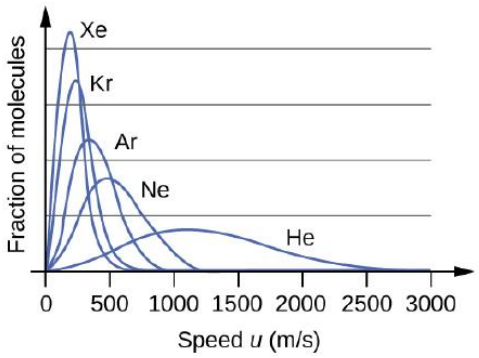


Figure 9.6.4: molecular speed is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules.

Link to Learning

The gas simulator may be used to examine the effect of temperature on molecular speeds. Examine the simulator's "energy histograms" (molecular speed distributions) and "species information" (which gives average speed values) for molecules of different masses at various temperatures.

The Kinetic-Molecular Theory Explains the Behavior of Gases, Part II

According to Graham's law, the molecules of a gas are in rapid motion and the molecules themselves are small. The average distance between the molecules of a gas is large compared to the size of the molecules. As a consequence, gas molecules can move past each other easily and diffuse at relatively fast rates.

The rate of effusion of a gas depends directly on the (average) speed of its molecules:

$$\text{effusion rate} \propto u_{\text{rms}}$$

Using this relation, and the equation relating molecular speed to mass, Graham's law may be easily derived as shown here:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

$$M = \frac{3RT}{u_{rms}^2} = \frac{3RT}{\bar{u}^2}$$

$$\frac{\text{effusion rate } A}{\text{effusion rate } B} = \frac{u_{rmsA}}{u_{rmsB}} = \frac{\sqrt{\frac{3RT}{M_A}}}{\sqrt{\frac{3RT}{M_B}}} = \sqrt{\frac{M_B}{M_A}}$$

The ratio of the rates of effusion is thus derived to be inversely proportional to the ratio of the square roots of their masses. This is the same relation observed experimentally and expressed as Graham's law.

This page titled 9.6: The Kinetic-Molecular Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.5: The Kinetic-Molecular Theory** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.7: Non-Ideal Gas Behavior

Learning Objectives

By the end of this section, you will be able to:

- Describe the physical factors that lead to deviations from ideal gas behavior
- Explain how these factors are represented in the van der Waals equation
- Define compressibility (Z) and describe how its variation with pressure reflects non-ideal behavior
- Quantify non-ideal behavior by comparing computations of gas properties using the ideal gas law and the van der Waals equation

Thus far, the ideal gas law, $PV = nRT$, has been applied to a variety of different types of problems, ranging from reaction stoichiometry and empirical and molecular formula problems to determining the density and molar mass of a gas. As mentioned in the previous modules of this chapter, however, the behavior of a gas is often non-ideal, meaning that the observed relationships between its pressure, volume, and temperature are not accurately described by the gas laws. In this section, the reasons for these deviations from ideal gas behavior are considered.

One way in which the accuracy of *PV* = *nRT* can be judged is by comparing the actual volume of 1 mole of gas (its molar volume, V_m) to the molar volume of an ideal gas at the same temperature and pressure. This ratio is called the **compressibility factor** (Z) with:

$$Z = \frac{\text{molar volume of gas at same } T \text{ and } P}{\text{molar volume of ideal gas at same } T \text{ and } P} = \left(\frac{PV_m}{RT}\right)_{\text{measured}}$$

Ideal gas behavior is therefore indicated when this ratio is equal to 1, and any deviation from 1 is an indication of non-ideal behavior. Figure 9.7.1 shows plots of Z over a large pressure range for several common gases.

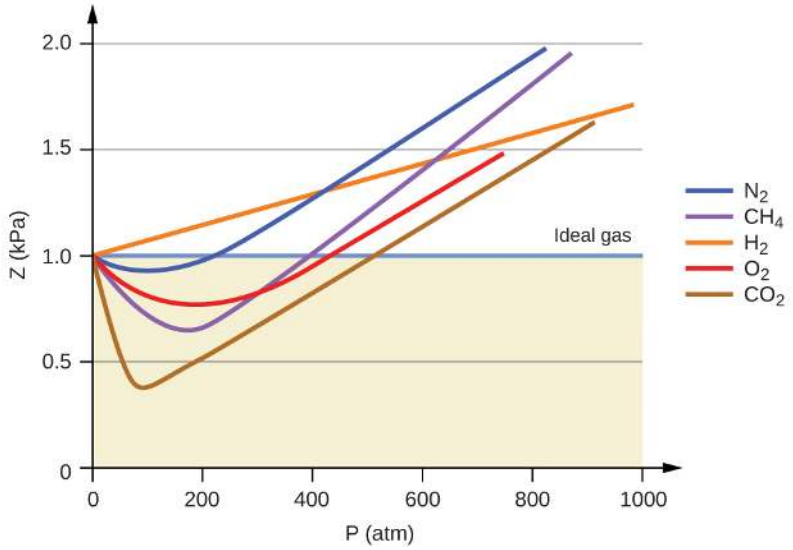


Figure 9.7.1: A graph of the compressibility factor (Z) vs. pressure shows that gases can exhibit significant deviations from the behavior predicted by the ideal gas law.

As is apparent from Figure 9.7.1, the ideal gas law does not describe gas behavior well at relatively high pressures. To determine why this is, consider the differences between real gas properties and what is expected of a hypothetical ideal gas.

Particles of a hypothetical ideal gas have no significant volume and do not attract or repel each other. In general, real gases approximate this behavior at relatively low pressures and high temperatures. However, at high pressures, the molecules of a gas are crowded closer together, and the amount of empty space between the molecules is reduced. At these higher pressures, the volume of the gas molecules themselves becomes appreciable relative to the total volume occupied by the gas. The gas therefore becomes less compressible at these high pressures, and although its volume continues to decrease with increasing pressure, this decrease is not *proportional* as predicted by Boyle's law.

At relatively low pressures, gas molecules have practically no attraction for one another because they are (on average) so far apart, and they behave almost like particles of an ideal gas. At higher pressures, however, the force of attraction is also no longer insignificant. This force pulls the molecules a little closer together, slightly decreasing the pressure (if the volume is constant) or decreasing the volume (at constant pressure) (Figure 9.7.2). This change is more pronounced at low temperatures because the molecules have lower KE relative to the attractive forces, and so they are less effective in overcoming these attractions after colliding with one another.

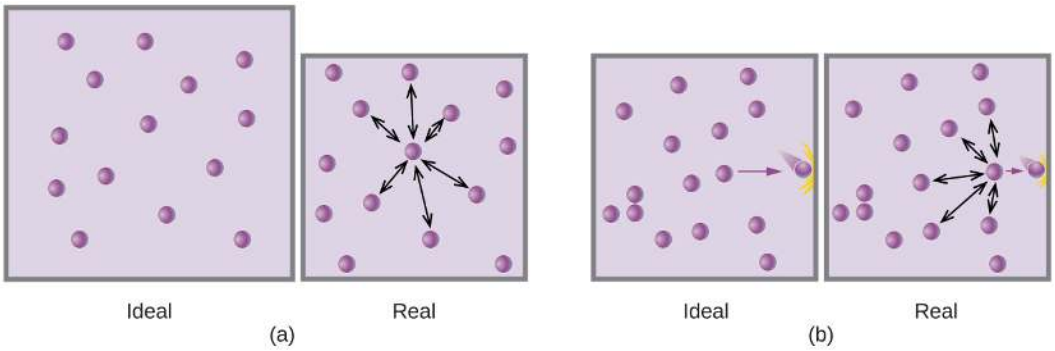


Figure 9.7.2: (a) Attractions between gas molecules serve to decrease the gas volume at constant pressure compared to an ideal gas whose molecules experience no attractive forces. (b) These attractive forces will decrease the force of collisions between the molecules and container walls, therefore reducing the pressure exerted at constant volume compared to an ideal gas.

There are several different equations that better approximate gas behavior than does the ideal gas law. The first, and simplest, of these was developed by the Dutch scientist Johannes van der Waals in 1879. The **van der Waals equation** improves upon the ideal gas law by adding two terms: one to account for the volume of the gas molecules and another for the attractive forces between them.

$$PV = nRT \longrightarrow \left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

Correction for molecular attraction

Correction for volume of molecules

The constant a corresponds to the strength of the attraction between molecules of a particular gas, and the constant b corresponds to the size of the molecules of a particular gas. The "correction" to the pressure term in the ideal gas law is

Table 9.7.1: Values of van der Waals Constants for Some Common Gases

Gas	a (L^2 atm/mol^2)	b (L/mol)
N_2	1.39	0.0391
O_2	1.36	0.0318
CO_2	3.59	0.0427
H_2O	5.46	0.0305
He	0.0342	0.0237
CCl_4	20.4	0.1383

At low pressures, the correction for intermolecular attraction, a, is more important than the one for molecular volume, b. At high pressures and small volumes, the correction for the volume of the molecules becomes important because the molecules themselves are incompressible and constitute an appreciable fraction of the total volume. At some intermediate pressure, the two corrections have opposing influences and the gas appears to follow the relationship given by $PV = nRT$ over a small range of pressures. This behavior is reflected by the "dips" in several of the compressibility curves shown in Figure 9.7.1. The attractive force between molecules initially makes the gas more compressible than an ideal gas, as pressure is raised (Z decreases with increasing P). At very high pressures, the gas becomes less compressible (Z increases with P), as the gas molecules begin to occupy an increasingly significant fraction of the total gas volume.

Strictly speaking, the ideal gas equation functions well when intermolecular attractions between gas molecules are negligible and the gas molecules themselves do not occupy an appreciable part of the whole volume. These criteria are satisfied under conditions of *low pressure and high temperature*. Under such conditions, the gas is said to behave ideally, and deviations from the gas laws are small enough that they may be disregarded—this is, however, very often not the case.

Example 9.7.1: Comparison of Ideal Gas Law and van der Waals Equation

A 4.25-L flask contains 3.46 mol CO_2 at 229 °C. Calculate the pressure of this sample of CO_2:

a. from the ideal gas law
b. from the van der Waals equation
c. Explain the reason(s) for the difference.

Solution

(a) From the ideal gas law:

$$P = \frac{nRT}{V} = \frac{3.46\text{ mol} \times 0.08206 Eatmmol^{-1}K^{-1} \times 502K}{4.25I} = 33.5atm$$

(b) From the van der Waals equation:

$$\left(P + \frac{n^2a}{V^2}\right) \times (V - nb) = nRT \longrightarrow P = \frac{nRT}{(V - nb)} - \frac{n^2a}{V^2}$$

$$P = \frac{3.46\text{ mol} \times 0.08206\text{ L atm mol}^{-1}\text{ K}^{-1} \times 502\text{ K}}{\left(4.25\text{ L} - 3.46\text{ mol} \times 0.0427\text{ L mol}^{-1}\right)} - \frac{(3.46\text{ mol})^2 \times 3.59\text{ L}^2\text{ atm mol}^2}{(4.24\text{ L})^2}$$

This finally yields P = 32.4 atm.

(c) This is not very different from the value from the ideal gas law because the pressure is not very high and the temperature is not very low. The value is somewhat different because CO_2 molecules do have some volume and attractions between molecules, and the ideal gas law assumes they do not have volume or attractions.

Exercise 9.7.1

A 560-mL flask contains 21.3 g N_2 at 145 °C. Calculate the pressure of N_2:

a. from the ideal gas law
b. from the van der Waals equation
c. Explain the reason(s) for the difference.

Answer

(a) 46.562 atm; (b) 46.594 atm; (c) The van der Waals equation takes into account the volume of the gas molecules themselves as well as intermolecular attractions.

This page titled 9.7: Non-Ideal Gas Behavior is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **9.6: Non-Ideal Gas Behavior** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

9.8: Key Terms

absolute zero | temperature at which the volume of a gas would be zero according to Charles's law.

Amontons's law | (also, Gay-Lussac's law) pressure of a given number of moles of gas is directly proportional to its kelvin temperature when the volume is held constant

atmosphere (atm) | unit of pressure; 1 atm = 101,325 Pa

Avogadro's law | volume of a gas at constant temperature and pressure is proportional to the number of gas molecules

bar | (bar or b) unit of pressure; 1 bar = 100,000 Pa

barometer | device used to measure atmospheric pressure

Boyle's law | volume of a given number of moles of gas held at constant temperature is inversely proportional to the pressure under which it is measured

Charles's law | volume of a given number of moles of gas is directly proportional to its kelvin temperature when the pressure is held constant

compressibility factor (Z) | ratio of the experimentally measured molar volume for a gas to its molar volume as computed from the ideal gas equation

Dalton's law of partial pressures | total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases

diffusion | movement of an atom or molecule from a region of relatively high concentration to one of relatively low concentration (discussed in this chapter with regard to gaseous species, but applicable to species in any phase)

effusion | transfer of gaseous atoms or molecules from a container to a vacuum through very small openings

Graham's law of effusion | rates of diffusion and effusion of gases are inversely proportional to the square roots of their molecular masses

hydrostatic pressure | pressure exerted by a fluid due to gravity

ideal gas | hypothetical gas whose physical properties are perfectly described by the gas laws

ideal gas constant (*R*) | constant derived from the ideal gas equation $R = 0.08206$ L atm mol^{-1} K^{-1} or 8.314 L kPa mol^{-1} K^{-1}

ideal gas law | relation between the pressure, volume, amount, and temperature of a gas under conditions derived by combination of the simple gas laws

kinetic molecular theory | theory based on simple principles and assumptions that effectively explains ideal gas behavior

manometer | device used to measure the pressure of a gas trapped in a container

mean free path | average distance a molecule travels between collisions

mole fraction (*X*) | concentration unit defined as the ratio of the molar amount of a mixture component to the total number of moles of all mixture components

partial pressure | pressure exerted by an individual gas in a mixture

pascal (Pa) | SI unit of pressure; 1 Pa = 1 N/m^2

pounds per square inch (psi) | unit of pressure common in the US

pressure | force exerted per unit area

rate of diffusion | amount of gas diffusing through a given area over a given time

root mean square speed (u_{rms}) | measure of average speed for a group of particles calculated as the square root of the average squared speed

standard conditions of temperature and pressure (STP) | 273.15 K (0 °C) and 1 atm (101.325 kPa)

standard molar volume | volume of 1 mole of gas at STP, approximately 22.4 L for gases behaving ideally

torr | unit of pressure;

van der Waals equation | modified version of the ideal gas equation containing additional terms to account for non-ideal gas behavior

vapor pressure of water | pressure exerted by water vapor in equilibrium with liquid water in a closed container at a specific temperature

This page titled 9.8: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **9.7: Key Terms** is licensed CC BY 4.0.

9.9: Key Equations

$$p = h\rho g$$

$$PV = nRT$$

$$P_{Total} = P_A + P_B + P_C + \ldots = \Sigma_i P_i$$

$$P_A = X_A P_{Total}$$

This page titled 9.9: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **9.8: Key Equations** is licensed CC BY 4.0.

9.10: Summary

Gases exert pressure, which is force per unit area. The pressure of a gas may be expressed in the SI unit of pascal or kilopascal, as well as in many other units including torr, atmosphere, and bar. Atmospheric pressure is measured using a barometer; other gas pressures can be measured using one of several types of manometers.

The behavior of gases can be described by several laws based on experimental observations of their properties. The pressure of a given amount of gas is directly proportional to its absolute temperature, provided that the volume does not change (Amontons's law). The volume of a given gas sample is directly proportional to its absolute temperature at constant pressure (Charles's law). The volume of a given amount of gas is inversely proportional to its pressure when temperature is held constant (Boyle's law). Under the same conditions of temperature and pressure, equal volumes of all gases contain the same number of molecules (Avogadro's law).

The equations describing these laws are special cases of the ideal gas law, $PV = nRT$, where P is the pressure of the gas, V is its volume, n is the number of moles of the gas, T is its kelvin temperature, and R is the ideal (universal) gas constant.

The ideal gas law can be used to derive a number of convenient equations relating directly measured quantities to properties of interest for gaseous substances and mixtures. Appropriate rearrangement of the ideal gas equation may be made to permit the calculation of gas densities and molar masses. Dalton's law of partial pressures may be used to relate measured gas pressures for gaseous mixtures to their compositions. Avogadro's law may be used in stoichiometric computations for chemical reactions involving gaseous reactants or products.

Gaseous atoms and molecules move freely and randomly through space. Diffusion is the process whereby gaseous atoms and molecules are transferred from regions of relatively high concentration to regions of relatively low concentration. Effusion is a similar process in which gaseous species pass from a container to a vacuum through very small orifices. The rates of effusion of gases are inversely proportional to the square roots of their densities or to the square roots of their atoms/molecules' masses (Graham's law).

The kinetic molecular theory is a simple but very effective model that effectively explains ideal gas behavior. The theory assumes that gases consist of widely separated molecules of negligible volume that are in constant motion, colliding elastically with one another and the walls of their container with average speeds determined by their absolute temperatures. The individual molecules of a gas exhibit a range of speeds, the distribution of these speeds being dependent on the temperature of the gas and the mass of its molecules.

Gas molecules possess a finite volume and experience forces of attraction for one another. Consequently, gas behavior is not necessarily described well by the ideal gas law. Under conditions of low pressure and high temperature, these factors are negligible, the ideal gas equation is an accurate description of gas behavior, and the gas is said to exhibit ideal behavior. However, at lower temperatures and higher pressures, corrections for molecular volume and molecular attractions are required to account for finite molecular size and attractive forces. The van der Waals equation is a modified version of the ideal gas law that can be used to account for the non-ideal behavior of gases under these conditions.

This page titled 9.10: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **9.9: Summary** is licensed CC BY 4.0.

9.11: Exercises

1.

Why are sharp knives more effective than dull knives? (Hint: Think about the definition of pressure.)

2.

Why do some small bridges have weight limits that depend on how many wheels or axles the crossing vehicle has?

3.

Why should you roll or belly-crawl rather than walk across a thinly-frozen pond?

4.

A typical barometric pressure in Redding, California, is about 750 mm Hg. Calculate this pressure in atm and kPa.

5.

A typical barometric pressure in Denver, Colorado, is 615 mm Hg. What is this pressure in atmospheres and kilopascals?

6.

A typical barometric pressure in Kansas City is 740 torr. What is this pressure in atmospheres, in millimeters of mercury, and in kilopascals?

7.

Canadian tire pressure gauges are marked in units of kilopascals. What reading on such a gauge corresponds to 32 psi?

8.

During the Viking landings on Mars, the atmospheric pressure was determined to be on the average about 6.50 millibars (1 bar = 0.987 atm). What is that pressure in torr and kPa?

9.

The pressure of the atmosphere on the surface of the planet Venus is about 88.8 atm. Compare that pressure in psi to the normal pressure on earth at sea level in psi.

10.

A medical laboratory catalog describes the pressure in a cylinder of a gas as 14.82 MPa. What is the pressure of this gas in atmospheres and torr?

11.

Consider this scenario and answer the following questions: On a mid-August day in the northeastern United States, the following information appeared in the local newspaper: atmospheric pressure at sea level 29.97 in. Hg, 1013.9 mbar.

(a) What was the pressure in kPa?

(b) The pressure near the seacoast in the northeastern United States is usually reported near 30.0 in. Hg. During a hurricane, the pressure may fall to near 28.0 in. Hg. Calculate the drop in pressure in torr.

12.

Why is it necessary to use a nonvolatile liquid in a barometer or manometer?

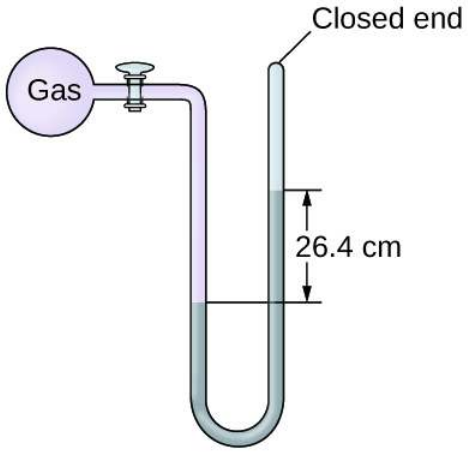
13.

The pressure of a sample of gas is measured at sea level with a closed-end manometer. The liquid in the manometer is mercury. Determine the pressure of the gas in:

(a) torr

(b) Pa

(c) bar



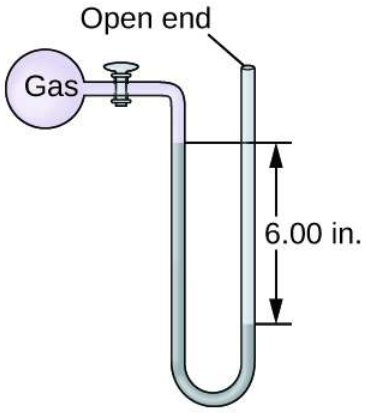
14.

The pressure of a sample of gas is measured with an open-end manometer, partially shown to the right. The liquid in the manometer is mercury. Assuming atmospheric pressure is 29.92 in. Hg, determine the pressure of the gas in:

(a) torr

(b) Pa

(c) bar



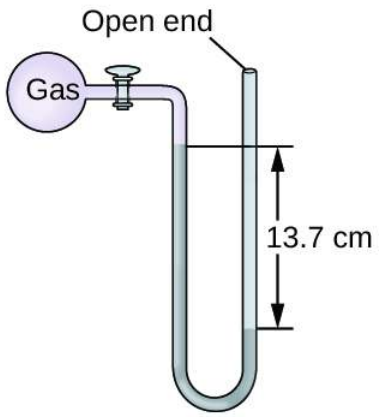
15.

The pressure of a sample of gas is measured at sea level with an open-end mercury manometer. Assuming atmospheric pressure is 760.0 mm Hg, determine the pressure of the gas in:

(a) mm Hg

(b) atm

(c) kPa



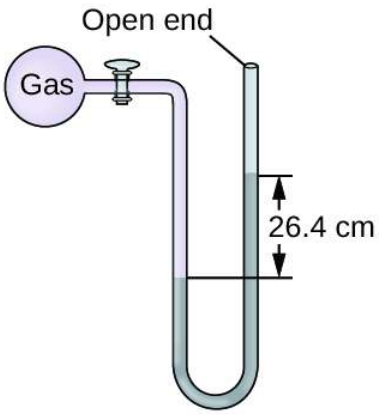
16.

The pressure of a sample of gas is measured at sea level with an open-end mercury manometer. Assuming atmospheric pressure is 760 mm Hg, determine the pressure of the gas in:

(a) mm Hg

(b) atm

(c) kPa



17.

How would the use of a volatile liquid affect the measurement of a gas using open-ended manometers vs. closed-end manometers?

18.

Sometimes leaving a bicycle in the sun on a hot day will cause a blowout. Why?

19.

Explain how the volume of the bubbles exhausted by a scuba diver (Figure 8.16) change as they rise to the surface, assuming that they remain intact.

20.

One way to state Boyle's law is "All other things being equal, the pressure of a gas is inversely proportional to its volume." (a) What is the meaning of the term "inversely proportional?" (b) What are the "other things" that must be equal?

21.

An alternate way to state Avogadro's law is "All other things being equal, the number of molecules in a gas is directly proportional to the volume of the gas." (a) What is the meaning of the term "directly proportional?" (b) What are the "other things" that must be equal?

22.

How would the graph in Figure 8.12 change if the number of moles of gas in the sample used to determine the curve were doubled?

23.

How would the graph in Figure 8.13 change if the number of moles of gas in the sample used to determine the curve were doubled?

24.

In addition to the data found in Figure 8.13, what other information do we need to find the mass of the sample of air used to determine the graph?

25.

Determine the volume of 1 mol of CH_4 gas at 150 K and 1 atm, using Figure 8.12.

26.

Determine the pressure of the gas in the syringe shown in Figure 8.13 when its volume is 12.5 mL, using:

(a) the appropriate graph

(b) Boyle's law

27.

A spray can is used until it is empty except for the propellant gas, which has a pressure of 1344 torr at 23 °C. If the can is thrown into a fire (T = 475 °C), what will be the pressure in the hot can?

28.

What is the temperature of an 11.2-L sample of carbon monoxide, CO, at 744 torr if it occupies 13.3 L at 55 °C and 744 torr?

29.

A 2.50-L volume of hydrogen measured at –196 °C is warmed to 100 °C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

30.

A balloon inflated with three breaths of air has a volume of 1.7 L. At the same temperature and pressure, what is the volume of the balloon if five more same-sized breaths are added to the balloon?

31.

A weather balloon contains 8.80 moles of helium at a pressure of 0.992 atm and a temperature of 25 °C at ground level. What is the volume of the balloon under these conditions?



32.

The volume of an automobile air bag was 66.8 L when inflated at 25 °C with 77.8 g of nitrogen gas. What was the pressure in the bag in kPa?

33.

How many moles of gaseous boron trifluoride, BF_3, are contained in a 4.3410-L bulb at 788.0 K if the pressure is 1.220 atm? How many grams of BF_3?

34.

Iodine, I_2, is a solid at room temperature but sublimes (converts from a solid into a gas) when warmed. What is the temperature in a 73.3-mL bulb that contains 0.292 g of I_2 vapor at a pressure of 0.462 atm?

35.

How many grams of gas are present in each of the following cases?

(a) 0.100 L of CO_2 at 307 torr and 26 °C

(b) 8.75 L of C_2H_4, at 378.3 kPa and 483 K

(c) 221 mL of Ar at 0.23 torr and –54 °C

36.

A high altitude balloon is filled with 1.41×10^4 L of hydrogen at a temperature of 21 °C and a pressure of 745 torr. What is the volume of the balloon at a height of 20 km, where the temperature is –48 °C and the pressure is 63.1 torr?

37.

A cylinder of medical oxygen has a volume of 35.4 L, and contains O_2 at a pressure of 151 atm and a temperature of 25 °C. What volume of O_2 does this correspond to at normal body conditions, that is, 1 atm and 37 °C?

38.

A large scuba tank (Figure 8.16) with a volume of 18 L is rated for a pressure of 220 bar. The tank is filled at 20 °C and contains enough air to supply 1860 L of air to a diver at a pressure of 2.37 atm (a depth of 45 feet). Was the tank filled to capacity at 20 °C?

39.

A 20.0-L cylinder containing 11.34 kg of butane, C_4H_{10}, was opened to the atmosphere. Calculate the mass of the gas remaining in the cylinder if it were opened and the gas escaped until the pressure in the cylinder was equal to the atmospheric pressure, 0.983 atm, and a temperature of 27 °C.

40.

While resting, the average 70-kg human male consumes 14 L of pure O_2 per hour at 25 °C and 100 kPa. How many moles of O_2 are consumed by a 70 kg man while resting for 1.0 h?

41.

For a given amount of gas showing ideal behavior, draw labeled graphs of:

(a) the variation of P with V

(b) the variation of V with T

(c) the variation of P with T

(d) the variation of $\frac{1}{P}$ with V

42.

A liter of methane gas, CH_4, at STP contains more atoms of hydrogen than does a liter of pure hydrogen gas, H_2, at STP. Using Avogadro's law as a starting point, explain why.

43.

The effect of chlorofluorocarbons (such as CCl_2F_2) on the depletion of the ozone layer is well known. The use of substitutes, such as $CH_3CH_2F(g)$, for the chlorofluorocarbons, has largely corrected the problem. Calculate the volume occupied by 10.0 g of each of these compounds at STP:

(a) $CCl_2F_2(g)$

(b) $CH_3CH_2F(g)$

44.

As 1 g of the radioactive element radium decays over 1 year, it produces 1.16×10^{18} alpha particles (helium nuclei). Each alpha particle becomes an atom of helium gas. What is the pressure in pascal of the helium gas produced if it occupies a volume of 125 mL at a temperature of 25 °C?

45.

A balloon with a volume of 100.21 L at 21 °C and 0.981 atm is released and just barely clears the top of Mount Crumpit in British Columbia. If the final volume of the balloon is 144.53 L at a temperature of 5.24 °C, what is the pressure experienced by the balloon as it clears Mount Crumpit?

46.

If the temperature of a fixed amount of a gas is doubled at constant volume, what happens to the pressure?

47.

If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

48.

What is the density of laughing gas, dinitrogen monoxide, N_2O, at a temperature of 325 K and a pressure of 113.0 kPa?

49.

Calculate the density of Freon 12, CF_2Cl_2, at 30.0 °C and 0.954 atm.

50.

Which is denser at the same temperature and pressure, dry air or air saturated with water vapor? Explain.

51.

A cylinder of $O_2(g)$ used in breathing by patients with emphysema has a volume of 3.00 L at a pressure of 10.0 atm. If the temperature of the cylinder is 28.0 °C, what mass of oxygen is in the cylinder?

52.

What is the molar mass of a gas if 0.0494 g of the gas occupies a volume of 0.100 L at a temperature 26 °C and a pressure of 307 torr?

53.

What is the molar mass of a gas if 0.281 g of the gas occupies a volume of 125 mL at a temperature 126 °C and a pressure of 777 torr?

54.

How could you show experimentally that the molecular formula of propene is C_3H_6, not CH_2?

55.

The density of a certain gaseous fluoride of phosphorus is 3.93 g/L at STP. Calculate the molar mass of this fluoride and determine its molecular formula.

56.

Consider this question: What is the molecular formula of a compound that contains 39% C, 45% N, and 16% H if 0.157 g of the compound occupies 125 mL with a pressure of 99.5 kPa at 22 °C?

(a) Outline the steps necessary to answer the question.

(b) Answer the question.

57.

A 36.0–L cylinder of a gas used for calibration of blood gas analyzers in medical laboratories contains 350 g CO_2, 805 g O_2, and 4,880 g N_2. At 25 degrees C, what is the pressure in the cylinder in atmospheres?

58.

A cylinder of a gas mixture used for calibration of blood gas analyzers in medical laboratories contains 5.0% CO_2, 12.0% O_2, and the remainder N_2 at a total pressure of 146 atm. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

59.

A sample of gas isolated from unrefined petroleum contains 90.0% CH_4, 8.9% C_2H_6, and 1.1% C_3H_8 at a total pressure of 307.2 kPa. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

60.

A mixture of 0.200 g of H_2, 1.00 g of N_2, and 0.820 g of Ar is stored in a closed container at STP. Find the volume of the container, assuming that the gases exhibit ideal behavior.

61.

Most mixtures of hydrogen gas with oxygen gas are explosive. However, a mixture that contains less than 3.0 % O_2 is not. If enough O_2 is added to a cylinder of H_2 at 33.2 atm to bring the total pressure to 34.5 atm, is the mixture explosive?

62.

A commercial mercury vapor analyzer can detect, in air, concentrations of gaseous Hg atoms (which are poisonous) as low as 2×10^{-6} mg/L of air. At this concentration, what is the partial pressure of gaseous mercury if the atmospheric pressure is 733 torr at 26 °C?

63.

A sample of carbon monoxide was collected over water at a total pressure of 756 torr and a temperature of 18 °C. What is the pressure of the carbon monoxide? (See Table 8.2 for the vapor pressure of water.)

64.

In an experiment in a general chemistry laboratory, a student collected a sample of a gas over water. The volume of the gas was 265 mL at a pressure of 753 torr and a temperature of 27 °C. The mass of the gas was 0.472 g. What was the molar mass of the gas?

65.

Joseph Priestley first prepared pure oxygen by heating mercuric oxide, HgO:
$2\mathrm{HgO}\,(s) \longrightarrow 2\mathrm{Hg}\,(l) + \mathrm{O_2}(g)$

(a) Outline the steps necessary to answer the following question: What volume of O_2 at 23 °C and 0.975 atm is produced by the decomposition of 5.36 g of HgO?

(b) Answer the question.

66.

Cavendish prepared hydrogen in 1766 by the novel method of passing steam through a red-hot gun barrel:
$4\mathrm{H_2O}\,(g) + 3\mathrm{Fe}\,(s) \longrightarrow \mathrm{Fe_3O_4}(s) + 4\mathrm{H_2}(g)$

(a) Outline the steps necessary to answer the following question: What volume of H_2 at a pressure of 745 torr and a temperature of 20 °C can be prepared from the reaction of 15.O g of H_2O?

(b) Answer the question.

67.

The chlorofluorocarbon CCl_2F_2 can be recycled into a different compound by reaction with hydrogen to produce $CH_2F_2(g)$, a compound useful in chemical manufacturing:
$\mathrm{CCl_2F_2}(g) + 4\mathrm{H_2}(g) \longrightarrow \mathrm{CH_2F_2}(g) + 2\mathrm{HCl}\,(g)$

(a) Outline the steps necessary to answer the following question: What volume of hydrogen at 225 atm and 35.5 °C would be required to react with 1 ton (1.000×10^3 kg) of CCl_2F_2?

(b) Answer the question.

68.

Automobile air bags are inflated with nitrogen gas, which is formed by the decomposition of solid sodium azide (NaN_3). The other product is sodium metal. Calculate the volume of nitrogen gas at 27 °C and 756 torr formed by the decomposition of 125 g of sodium azide.

69.

Lime, CaO, is produced by heating calcium carbonate, $CaCO_3$; carbon dioxide is the other product.

(a) Outline the steps necessary to answer the following question: What volume of carbon dioxide at 875 K and 0.966 atm is produced by the decomposition of 1 ton (1.000×10^3 kg) of calcium carbonate?

(b) Answer the question.

70.

Before small batteries were available, carbide lamps were used for bicycle lights. Acetylene gas, C_2H_2, and solid calcium hydroxide were formed by the reaction of calcium carbide, CaC_2, with water. The ignition of the acetylene gas provided the light. Currently, the same lamps are used by some cavers, and calcium carbide is used to produce acetylene for carbide cannons.

(a) Outline the steps necessary to answer the following question: What volume of C_2H_2 at 1.005 atm and 12.2 °C is formed by the reaction of 15.48 g of CaC_2 with water?

(b) Answer the question.

71.

Calculate the volume of oxygen required to burn 12.00 L of ethane gas, C_2H_6, to produce carbon dioxide and water, if the volumes of C_2H_6 and O_2 are measured under the same conditions of temperature and pressure.

72.

What volume of O_2 at STP is required to oxidize 8.0 L of NO at STP to NO_2? What volume of NO_2 is produced at STP?

73.

Consider the following questions:

(a) What is the total volume of the $CO_2(g)$ and $H_2O(g)$ at 600 °C and 0.888 atm produced by the combustion of 1.00 L of $C_2H_6(g)$ measured at STP?

(b) What is the partial pressure of H_2O in the product gases?

74.

Methanol, CH_3OH, is produced industrially by the following reaction:

$$\mathrm{CO}\,(g) + 2\mathrm{H_2}\,(g) \xrightarrow{\text{copper catalyst } 300\ ^\circ\mathrm{C},\ 300\ \text{atm}} \mathrm{CH_3OH}\,(g)$$

Assuming that the gases behave as ideal gases, find the ratio of the total volume of the reactants to the final volume.

75.

What volume of oxygen at 423.0 K and a pressure of 127.4 kPa is produced by the decomposition of 129.7 g of BaO_2 to BaO and O_2?

76.

A 2.50-L sample of a colorless gas at STP decomposed to give 2.50 L of N_2 and 1.25 L of O_2 at STP. What is the colorless gas?

77.

Ethanol, C_2H_5OH, is produced industrially from ethylene, C_2H_4, by the following sequence of reactions:

$$3\mathrm{C_2H_4} + 2\mathrm{H_2SO_4} \longrightarrow \mathrm{C_2H_5HSO_4} + (\mathrm{C_2H_5})_2\mathrm{SO_4}$$
$$\mathrm{C_2H_5HSO_4} + (\mathrm{C_2H_5})_2\mathrm{SO_4} + 3\mathrm{H_2O} \longrightarrow 3\mathrm{C_2H_5OH} + 2\mathrm{H_2SO_4}$$

What volume of ethylene at STP is required to produce 1.000 metric ton (1000 kg) of ethanol if the overall yield of ethanol is 90.1%?

78.

One molecule of hemoglobin will combine with four molecules of oxygen. If 1.0 g of hemoglobin combines with 1.53 mL of oxygen at body temperature (37 °C) and a pressure of 743 torr, what is the molar mass of hemoglobin?

79.

A sample of a compound of xenon and fluorine was confined in a bulb with a pressure of 18 torr. Hydrogen was added to the bulb until the pressure was 72 torr. Passage of an electric spark through the mixture produced Xe and HF. After the HF was removed by reaction with solid KOH, the final pressure of xenon and unreacted hydrogen in the bulb was 36 torr. What is the empirical formula of the xenon fluoride in the original sample? (Note: Xenon fluorides contain only one xenon atom per molecule.)

80.

One method of analyzing amino acids is the van Slyke method. The characteristic amino groups ($-NH_2$) in protein material are allowed to react with nitrous acid, HNO_2, to form N_2 gas. From the volume of the gas, the amount of amino acid can be determined. A 0.0604-g sample of a biological sample containing glycine, $CH_2(NH_2)COOH$, was analyzed by the van Slyke method and yielded 3.70 mL of N_2 collected over water at a pressure of 735 torr and 29 °C. What was the percentage of glycine in the sample?

$$\mathrm{CH_2\,(NH_2)\,CO_2H + HNO_2 \longrightarrow CH_2\,(OH)\,CO_2H + H_2O + N_2}$$

81.

A balloon filled with helium gas takes 6 hours to deflate to 50% of its original volume. How long will it take for an identical balloon filled with the same volume of hydrogen gas (instead of helium) to decrease its volume by 50%?

82.

Explain why the numbers of molecules are not identical in the left- and right-hand bulbs shown in the center illustration of Figure 8.27.

83.

Starting with the definition of rate of effusion and Graham's finding relating rate and molar mass, show how to derive the Graham's law equation, relating the relative rates of effusion for two gases to their molecular masses.

84.

Heavy water, D_2O (molar mass = 20.03 g mol^{-1}), can be separated from ordinary water, H_2O (molar mass = 18.01), as a result of the difference in the relative rates of diffusion of the molecules in the gas phase. Calculate the relative rates of diffusion of H_2O and D_2O.

85.

Which of the following gases diffuse more slowly than oxygen? F_2, Ne, N_2O, C_2H_2, NO, Cl_2, H_2S

86.

During the discussion of gaseous diffusion for enriching uranium, it was claimed that $^{235}UF_6$ diffuses 0.4% faster than $^{238}UF_6$. Show the calculation that supports this value. The molar mass of $^{235}UF_6$ = 235.043930 + 6 × 18.998403 = 349.034348 g/mol, and the molar mass of $^{238}UF_6$ = 238.050788 + 6 × 18.998403 = 352.041206 g/mol.

87.

Calculate the relative rate of diffusion of $^1H_2$ (molar mass 2.0 g/mol) compared with $^2H_2$ (molar mass 4.0 g/mol) and the relative rate of diffusion of O_2 (molar mass 32 g/mol) compared with O_3 (molar mass 48 g/mol).

88.

A gas of unknown identity diffuses at a rate of 83.3 mL/s in a diffusion apparatus in which carbon dioxide diffuses at the rate of 102 mL/s. Calculate the molecular mass of the unknown gas.

89.

When two cotton plugs, one moistened with ammonia and the other with hydrochloric acid, are simultaneously inserted into opposite ends of a glass tube that is 87.0 cm long, a white ring of NH_4Cl forms where gaseous NH_3 and gaseous HCl first come into contact. $NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$ At approximately what distance from the ammonia moistened plug does this occur? (Hint: Calculate the rates of diffusion for both NH_3 and HCl, and find out how much faster NH_3 diffuses than HCl.)

90.

Using the postulates of the kinetic molecular theory, explain why a gas uniformly fills a container of any shape.

91.

Can the speed of a given molecule in a gas double at constant temperature? Explain your answer.

92.

Describe what happens to the average kinetic energy of ideal gas molecules when the conditions are changed as follows:

(a) The pressure of the gas is increased by reducing the volume at constant temperature.

(b) The pressure of the gas is increased by increasing the temperature at constant volume.

(c) The average speed of the molecules is increased by a factor of 2.

93.

The distribution of molecular speeds in a sample of helium is shown in Figure 8.34. If the sample is cooled, will the distribution of speeds look more like that of H_2 or of H_2O? Explain your answer.

94.

What is the ratio of the average kinetic energy of a SO_2 molecule to that of an O_2 molecule in a mixture of two gases? What is the ratio of the root mean square speeds, u_{rms}, of the two gases?

95.

A 1-L sample of CO initially at STP is heated to 546 K, and its volume is increased to 2 L.

(a) What effect do these changes have on the pressure exerted by the gas?

(b) What is the effect on the average kinetic energy of the molecules?

(c) What is the effect on the root mean square speed of the molecules?

96.

The root mean square speed of H_2 molecules at 25 °C is about 1.6 km/s. What is the root mean square speed of a N_2 molecule at 25 °C?

97.

Answer the following questions:

(a) Is the pressure of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?

(b) Is the density of the gas in the hot-air balloon shown at the opening of this chapter greater than, less than, or equal to that of the atmosphere outside the balloon?

(c) At a pressure of 1 atm and a temperature of 20 °C, dry air has a density of 1.2256 g/L. What is the (average) molar mass of dry air?

(d) The average temperature of the gas in a hot-air balloon is 1.30×10^2 °F. Calculate its density, assuming the molar mass equals that of dry air.

(e) The lifting capacity of a hot-air balloon is equal to the difference in the mass of the cool air displaced by the balloon and the mass of the gas in the balloon. What is the difference in the mass of 1.00 L of the cool air in part (c) and the hot air in part (d)?

(f) An average balloon has a diameter of 60 feet and a volume of 1.1×10^5 ft^3. What is the lifting power of such a balloon? If the weight of the balloon and its rigging is 500 pounds, what is its capacity for carrying passengers and cargo?

(g) A balloon carries 40.0 gallons of liquid propane (density 0.5005 g/L). What volume of CO_2 and H_2O gas is produced by the combustion of this propane at STP?

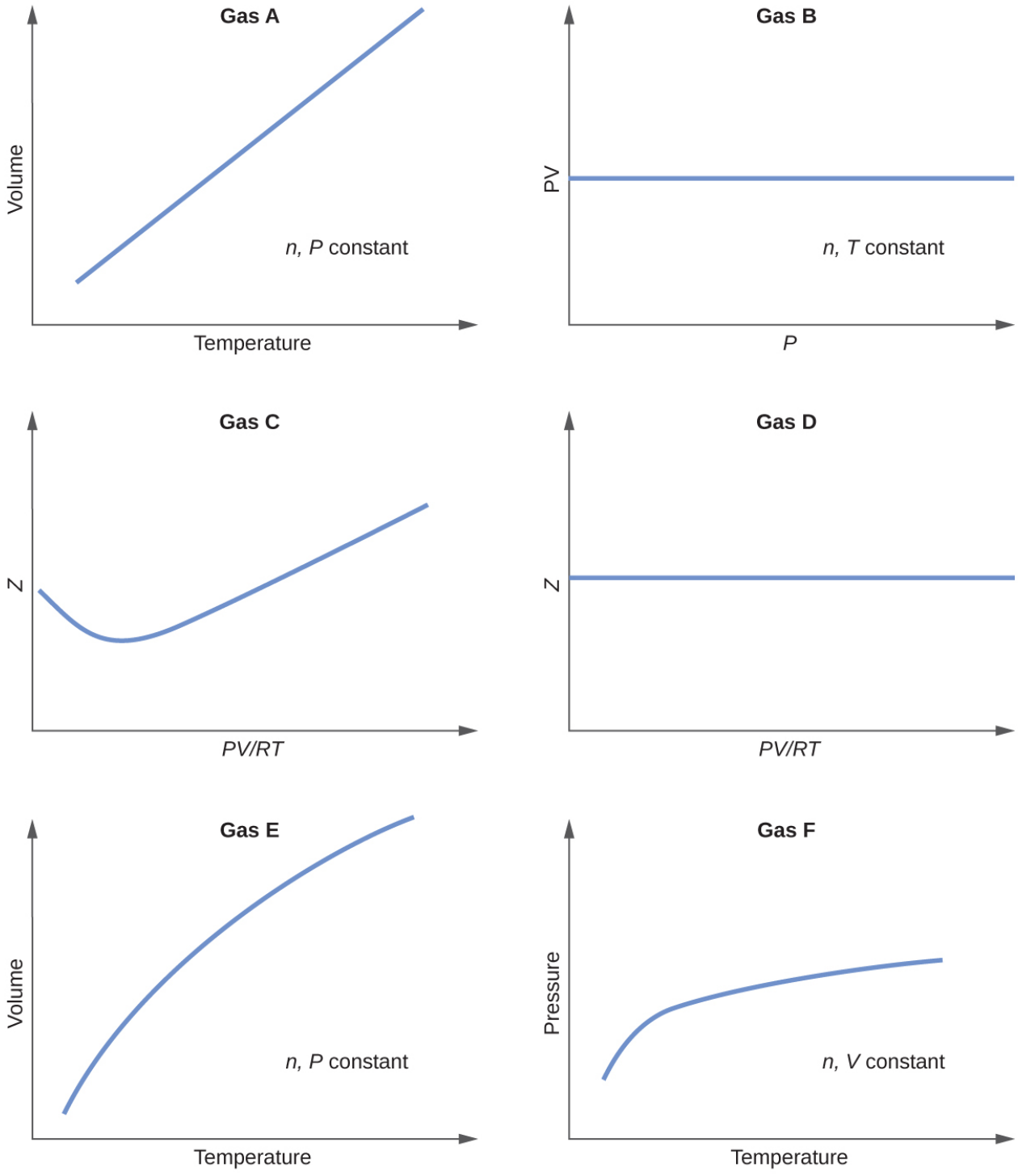
(h) A balloon flight can last about 90 minutes. If all of the fuel is burned during this time, what is the approximate rate of heat loss (in kJ/min) from the hot air in the bag during the flight?

98.

Show that the ratio of the rate of diffusion of Gas 1 to the rate of diffusion of Gas 2, $\frac{R_1}{R_2}$, is the same at 0 °C and 100 °C.

99.

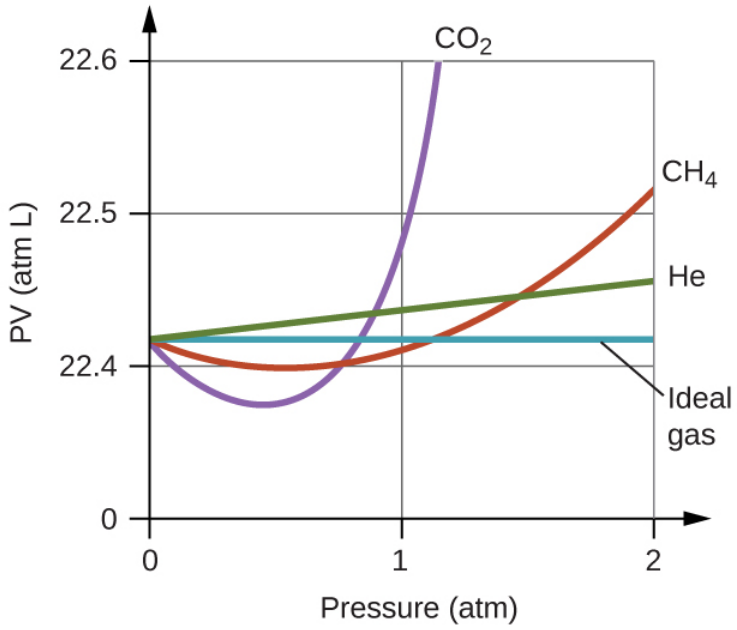
Graphs showing the behavior of several different gases follow. Which of these gases exhibit behavior significantly different from that expected for ideal gases?



100.

Explain why the plot of PV for CO_2 differs from that of an ideal gas.

https://chem.libretexts.org/@go/page/482985



101.

Under which of the following sets of conditions does a real gas behave most like an ideal gas, and for which conditions is a real gas expected to deviate from ideal behavior? Explain.

(a) high pressure, small volume

(b) high temperature, low pressure

(c) low temperature, high pressure

102.

Describe the factors responsible for the deviation of the behavior of real gases from that of an

ideal gas.

103.

For which of the following gases should the correction for the molecular volume be largest:

CO, CO_2, H_2, He, NH_3, SF_6?

104.

A 0.245-L flask contains 0.467 mol CO_2 at 159 °C. Calculate the pressure:

(a) using the ideal gas law

(b) using the van der Waals equation

(c) Explain the reason for the difference.

(d) Identify which correction (that for P or V) is dominant and why.

105.

Answer the following questions:

(a) If XX behaved as an ideal gas, what would its graph of Z vs. P look like?

(b) For most of this chapter, we performed calculations treating gases as ideal. Was this justified?

(c) What is the effect of the volume of gas molecules on Z? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.

(d) What is the effect of intermolecular attractions on the value of Z? Under what conditions is this effect small? When is it large? Explain using an appropriate diagram.

(e) In general, under what temperature conditions would you expect Z to have the largest deviations from the Z for an ideal gas?

This page titled 9.11: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

10: Three Major Classes of Chemical Reactions

10.1: Formula Mass
10.2: Writing and Balancing Chemical Equations
10.3: Reaction Stoichiometry
10.4: Reaction Yields
10.5: Molarity
10.6: Precipitation Reactions
10.7: Acid-Base Reactions
10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations
10.9: Oxidation-Reduction- Some General Principles
10.10: Balancing Oxidation-Reduction Equations
10.11: Oxidizing and Reducing Agents
10.12: Key Terms
10.13: Key Equations
10.E: Reactions in Aqueous Solution (Exercises)
10.S: Summary

This page titled 10: Three Major Classes of Chemical Reactions is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

10.1: Formula Mass

Learning Objectives

By the end of this section, you will be able to:

- Calculate formula masses for covalent and ionic compounds
- Define the amount unit mole and the related quantity Avogadro's number Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another

Many argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Dalton's atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances.

Formula Mass

An earlier chapter of this text described the development of the atomic mass unit, the concept of average atomic masses, and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the formula mass of a substance by summing the average atomic masses of all the atoms represented in the substance's formula.

Formula Mass for Covalent Substances

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. Figure 10.1.1 outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37

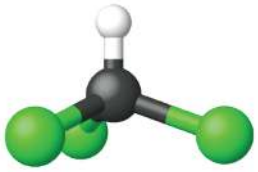


Figure 10.1.1: The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

Likewise, the molecular mass of an aspirin molecule, $C_9H_8O_4$, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms, and four oxygen atoms, which amounts to 180.15 amu (Figure 10.1.2).

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	9	×	12.01	=	108.09
H	8	×	1.008	=	8.064
O	4	×	16.00	=	64.00
				Molecular mass	180.15

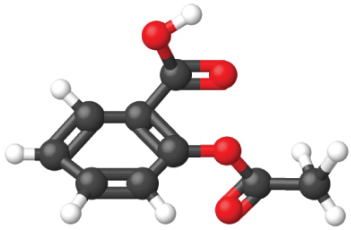


Figure 10.1.2: The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $C_9H_8O_4$.

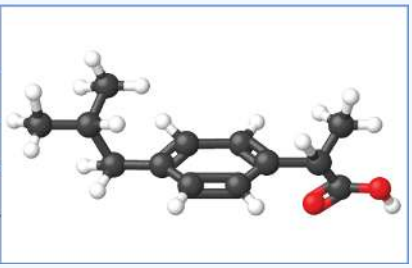
Example 10.1.1: Computing Molecular Mass for a Covalent Compound

Ibuprofen, $C_{13}H_{18}O_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound?

Solution

Molecules of this compound are composed of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.144
O	2	×	16.00	=	32.00
				Molecular mass	206.27



Exercise 10.1.1

Acetaminophen, $C_8H_9NO_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound?

Answer

151.16 amu

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (see Figure 10.1.3).

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

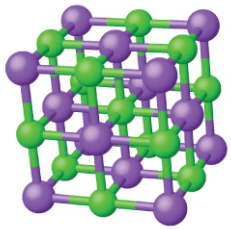


Figure 10.1.3: Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses.

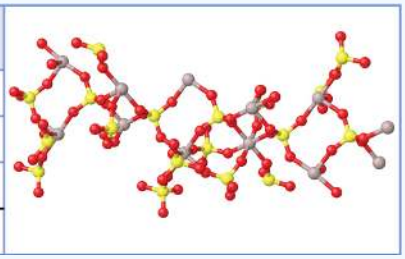
Example 10.1.2: Computing Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Molecular mass	342.14



Exercise 10.1.2

Calcium phosphate, $Ca_3(PO_4)_2$, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate?

Answer

310.18 amu

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O, and hydrogen peroxide, H_2O_2, are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, sophisticated instruments allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the *mole*, which remains indispensable in modern chemical science.

The *mole* is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a sample of matter. One Latin connotation for the word “mole” is “large mass” or “bulk,” which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth. A mole of substance is that amount in which there are $6.02214076 \times 10^{23}$.

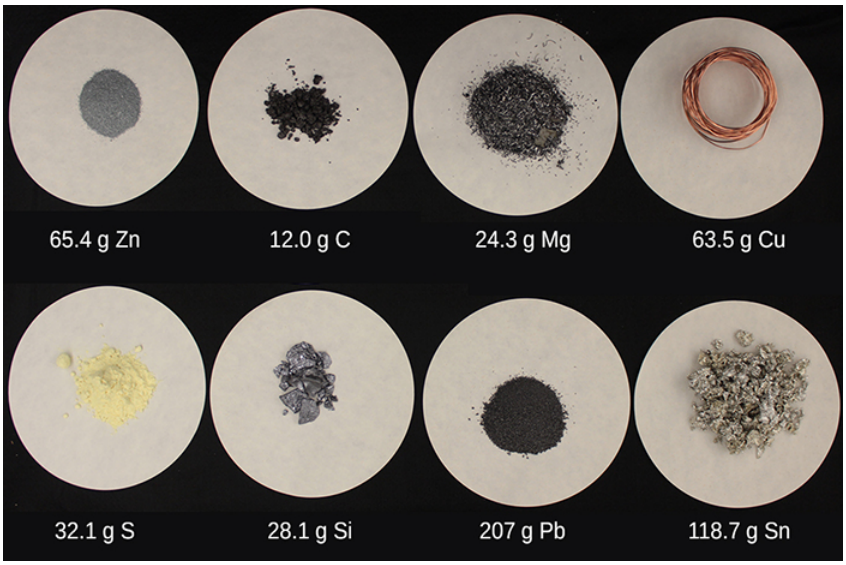


Figure 10.1.4: Each sample contains 6.022 10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott)

The molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). A mole of ^{12}C weighs 12 g (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, ^{12}C. Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu (Figure 10.1.5).



Figure 10.1.5: Each sample contains 6.02 10^{23} molecules or formula units—1.00 mol of the compound or element. Clock-wise from the upper left: 130.2 g of $C_8H_{17}OH$ (1-octanol, formula mass 130.2 amu), 454.4 g of HgI_2 (mercury(II) iodide, formula mass 454.4 amu), 32.0 g of CH_3OH (methanol, formula mass 32.0 amu) and 256.5 g of S_8 (sulfur, formula mass 256.5 amu). (credit: Sahar Atwa)

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022 10^{23}
H	1.008	1.008	6.022 10^{23}
O	16.00	16.00	6.022 10^{23}
Na	22.99	22.99	6.022 10^{23}
Cl	35.45	35.45	6.022 10^{23}

While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the

mole, consider a small drop of water weighing about 0.03 g (see Figure 10.1.6). Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules.



Figure 10.1.6: The number of molecules in a single droplet of water is roughly 100 billion times greater than the number of people on earth. (credit: “tanakawho”/Wikimedia commons)

Link to Learning

The mole is used in chemistry to represent 6.022 10^{23} of something, but it can be difficult to conceptualize such a large number. Watch this video and then complete the “Think” questions that follow. Explore more about the mole by reviewing the information under “Dig Deeper.”

The relationships between formula mass, the mole, and Avogadro’s number can be applied to compute various quantities that describe the composition of substances and compounds, as demonstrated in the next several example problems.

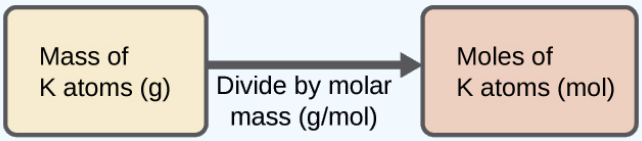
Example 10.1.3: Deriving Moles from Grams for an Element

According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?

Solution

The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable “ballpark” estimate of the number of moles would be slightly greater than 0.1 mol.

The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol):



The factor-label method supports this mathematical approach since the unit “g” cancels and the answer has units of “mol:”

$$4.7\ \cancel{\text{g K}}\left(\frac{\text{mol K}}{39.10\ \cancel{\text{g K}}}\right) = 0.12\ \text{mol K}$$

The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol,

Exercise 10.1.3

Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g?

Answer

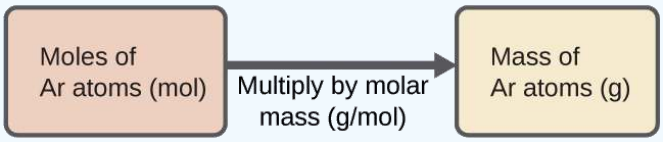
0.360 mol

Example 10.1.4: Deriving Grams from Moles for an Element

A liter of air contains 9.2 10^{-4} mol argon. What is the mass of Ar in a liter of air?

Solution

The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth (~10^{-3}) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~0.04 g):



In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol):

$$9.2 \times 10^{-4}\ \text{mol Ar} \left(\frac{39.95\ \text{g Ar}}{\text{mol Ar}} \right) = 0.037\ \text{g Ar}$$

The result is in agreement with our expectations, around 0.04 g Ar.

Exercise 10.1.1

What is the mass of 2.561 mol of gold?

Answer

504.4 g

Example 10.1.5: Deriving Number of Atoms from Mass for an Element

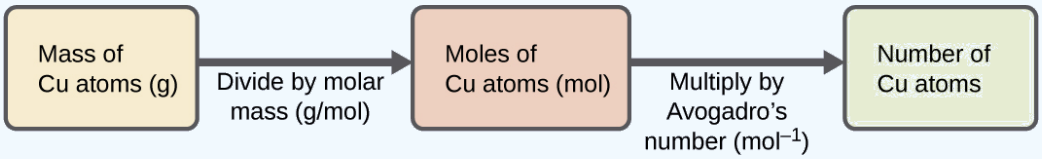
Copper is commonly used to fabricate electrical wire (Figure 10.1.7). How many copper atoms are in 5.00 g of copper wire?



Figure 10.1.7: Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol)

Solution

The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (N_A) to convert this molar amount to number of Cu atoms:



Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth N_A, or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$5.00\ \text{g}\ \cancel{\text{Cu}}\left(\frac{\text{mol}\ \cancel{\text{Cu}}}{63.55\ \text{g}\ \cancel{\text{Cu}}}\right)\left(\frac{6.022\times10^{23}\,\text{Cu atoms}}{\text{mol}\ \cancel{\text{Cu}}}\right)=4.74\times10^{22}\,\text{Cu atoms}$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 10^{22} as expected.

? Exercise 10.1.5

A prospector panning for gold in a river collects 15.00 g of pure gold. How many Au atoms are in this quantity of gold?

Answer

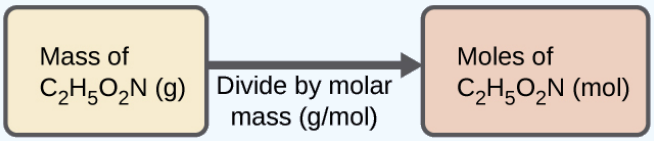
4.586×10^{22} Au atoms

✓ Example 10.1.6: Deriving Moles from Grams for a Compound

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $C_2H_5O_2N$. How many moles of glycine molecules are contained in 28.35 g of glycine?

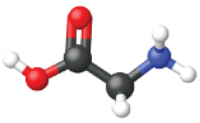
Solution

Derive the number of moles of a compound from its mass following the same procedure used for an element in Example 10.1.3:



The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $C_2H_5O_2N$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07



The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so the computed result is expected to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$28.35\ \cancel{\text{g glycine}}\left(\frac{\text{mol glycine}}{75.07\ \cancel{\text{g glycine}}}\right) = 0.378\ \text{mol glycine}$$

This result is consistent with the rough estimate,

? Exercise 10.1.6

How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a 25-g sample of sucrose?

Answer

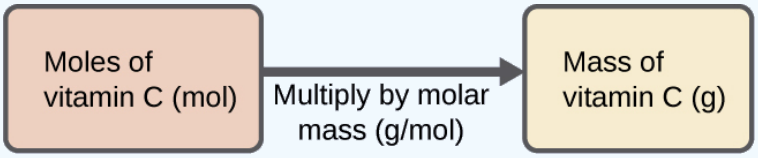
0.073 mol

✓ Example 10.1.7: Deriving Grams from Moles for a Compound

Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is $1.42\ 10^{-4}$ mol. What is the mass of this allowance in grams?

Solution

As for elements, the mass of a compound can be derived from its molar amount as shown:



The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole (~10^{-4} or one-ten thousandth); therefore, the corresponding mass is expected to be about one-ten thousandth of the molar mass (~0.02 g). Performing the calculation yields:

$$1.42 \times 10^{-4}\ \cancel{\text{mol vitamin C}}\left(\frac{176.124\ \text{g vitamin C}}{\cancel{\text{mol vitamin C}}}\right) = 0.0250\ \text{g vitamin C}$$

This is consistent with the anticipated result.

? Exercise 10.1.7

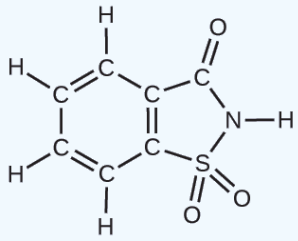
What is the mass of 0.443 mol of hydrazine, N_2H_4?

Answer

14.2 g

✓ Example 10.1.8: Deriving the Number of Atoms and Molecules from the Mass of a Compound

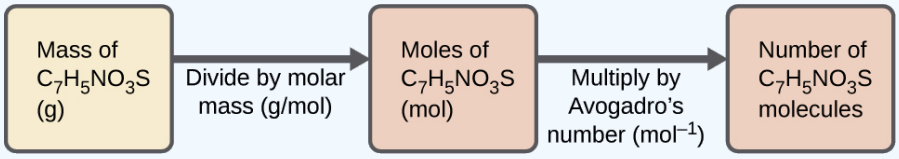
A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution

The number of molecules in a given mass of compound is computed by first deriving the number of moles, as demonstrated in Example 10.1.6, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400\ \text{g}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}}\left(\frac{\text{mol}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}}}{183.18\ \text{g}\ \cancel{\text{CH}_5\text{NO}_3\text{S}}}\right)\left(\frac{6.022\times10^{23}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}}{1\ \text{mol}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}}}\right)$$
$$= 1.31\times10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is:

$$1.31\times10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}\left(\frac{7\ \text{C atoms}}{1\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecule}}\right) = 9.17\times10^{20}\ \text{C atoms}$$

Exercise 10.1.8

How many C_4H_{10} molecules are contained in 9.213 g of this compound? How many hydrogen atoms?

Answer

9.545×10^{22} molecules C_4H_{10}

9.545×10^{23} atoms H

How Sciences Interconnect: Counting Neurotransmitter Molecules in the Brain

The brain is the control center of the central nervous system (Figure 10.1.8). It sends and receives signals to and from muscles and other internal organs to monitor and control their functions; it processes stimuli detected by sensory organs to guide interactions with the external world; and it houses the complex physiological processes that give rise to our intellect and emotions. The broad field of neuroscience spans all aspects of the structure and function of the central nervous system, including research on the anatomy and physiology of the brain. Great progress has been made in brain research over the past few decades, and the BRAIN Initiative, a federal initiative announced in 2013, aims to accelerate and capitalize on these advances through the concerted efforts of various industrial, academic, and government agencies (more details available at www.whitehouse.gov/share/brain-initiative).

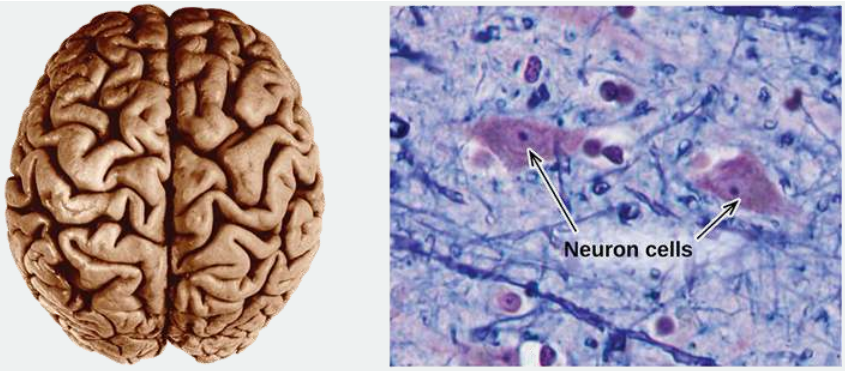


Figure 10.1.8: (a) A typical human brain weighs about 1.5 kg and occupies a volume of roughly 1.1 L. (b) Information is transmitted in brain tissue and throughout the central nervous system by specialized cells called neurons (micrograph shows cells at 1600× magnification).

Specialized cells called neurons transmit information between different parts of the central nervous system by way of electrical and chemical signals. Chemical signaling occurs at the interface between different neurons when one of the cells releases molecules (called neurotransmitters) that diffuse across the small gap between the cells (called the synapse) and bind to the surface of the other cell. These neurotransmitter molecules are stored in small intracellular structures called vesicles that fuse to the cell membrane and then break open to release their contents when the neuron is appropriately stimulated. This process is called exocytosis (see Figure 10.1.9). One neurotransmitter that has been very extensively studied is dopamine, $C_8H_{11}NO_2$. Dopamine is involved in various neurological processes that impact a wide variety of human behaviors. Dysfunctions in the dopamine systems of the brain underlie serious neurological diseases such as Parkinson's and schizophrenia.

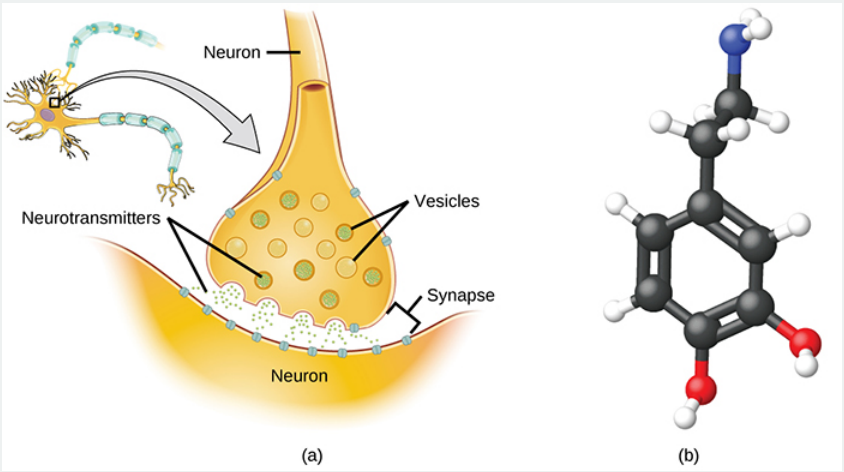


Figure 10.1.9: (a) Chemical signals are transmitted from neurons to other cells by the release of neurotransmitter molecules into the small gaps (synapses) between the cells. (b) Dopamine, $C_8H_{11}NO_2$, is a neurotransmitter involved in a number of neurological processes.

One important aspect of the complex processes related to dopamine signaling is the number of neurotransmitter molecules released during exocytosis. Since this number is a central factor in determining neurological response (and subsequent human thought and action), it is important to know how this number changes with certain controlled stimulations, such as the administration of drugs. It is also important to understand the mechanism responsible for any changes in the number of neurotransmitter molecules released—for example, some dysfunction in exocytosis, a change in the number of vesicles in the neuron, or a change in the number of neurotransmitter molecules in each vesicle.

Significant progress has been made recently in directly measuring the number of dopamine molecules stored in individual vesicles and the amount actually released when the vesicle undergoes exocytosis. Using miniaturized probes that can selectively detect dopamine molecules in very small amounts, scientists have determined that the vesicles of a certain type of mouse brain neuron contain an average of 30,000 dopamine molecules per vesicle (about 5×10^{-20} mol or 50 zmol). Analysis of these neurons from mice subjected to various drug therapies shows significant changes in the average number of dopamine molecules contained in individual vesicles, increasing or decreasing by up to three-fold, depending on the specific drug used.

> These studies also indicate that not all of the dopamine in a given vesicle is released during exocytosis, suggesting that it may be possible to regulate the fraction released using pharmaceutical therapies.

Footnotes

- 1Omiatek, Donna M., Amanda J. Bressler, Ann-Sofie Cans, Anne M. Andrews, Michael L. Heien, and Andrew G. Ewing. “The Real Catecholamine Content of Secretory Vesicles in the CNS Revealed by Electrochemical Cytometry.” *Scientific Report* 3 (2013): 1447, accessed January 14, 2015, doi:10.1038/srep01447.

This page titled 10.1: Formula Mass is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **3.1: Formula Mass and the Mole Concept** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

10.2: Writing and Balancing Chemical Equations

Learning Objectives

By the end of this section, you will be able to:

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations in molecular, total ionic, and net ionic formats.

An earlier chapter of this text introduced the use of element symbols to represent individual atoms. When atoms gain or lose electrons to yield ions, or combine with other atoms to form molecules, their symbols are modified or combined to generate chemical formulas that appropriately represent these species. Extending this symbolism to represent both the identities and the relative quantities of substances undergoing a chemical (or physical) change involves writing and balancing a **chemical equation**. Consider as an example the reaction between one methane molecule (CH_4) and two diatomic oxygen molecules (O_2) to produce one carbon dioxide molecule (CO_2) and two water molecules (H_2O). The chemical equation representing this process is provided in the upper half of Figure 10.2.1: , with space-filling molecular models shown in the lower half of the figure.

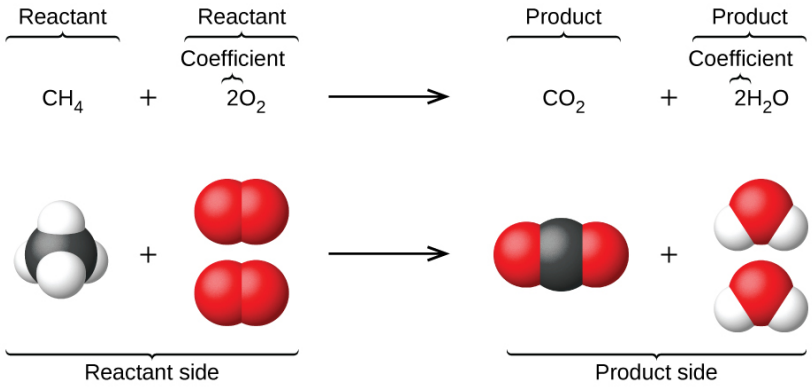


Figure 10.2.1: The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be represented by a chemical equation using formulas (top).

This example illustrates the fundamental aspects of any chemical equation:

1. The substances undergoing reaction are called reactants, and their formulas are placed on the left side of the equation.
2. The substances generated by the reaction are called products, and their formulas are placed on the right side of the equation.
3. Plus signs (+) separate individual reactant and product formulas, and an arrow separates the reactant and product (left and right) sides of the equation.
4. The relative numbers of reactant and product species are represented by coefficients (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

It is common practice to use the smallest possible whole-number coefficients in a chemical equation, as is done in this example. Realize, however, that these coefficients represent the *relative* numbers of reactants and products, and, therefore, they may be correctly interpreted as ratios. Methane and oxygen react to yield carbon dioxide and water in a 1:2:1:2 ratio. This ratio is satisfied if the numbers of these molecules are, respectively, 1-2-1-2, or 2-4-2-4, or 3-6-3-6, and so on (Figure 10.2.2). Likewise, these coefficients may be interpreted with regard to any amount (number) unit, and so this equation may be correctly read in many ways, including:

- *One* methane molecule and *two* oxygen molecules react to yield *one* carbon dioxide molecule and *two* water molecules.
- *One dozen* methane molecules and *two dozen* oxygen molecules react to yield *one dozen* carbon dioxide molecules and *two dozen* water molecules.
- *One mole* of methane molecules and *2 moles* of oxygen molecules react to yield *1 mole* of carbon dioxide molecules and *2 moles* of water molecules.

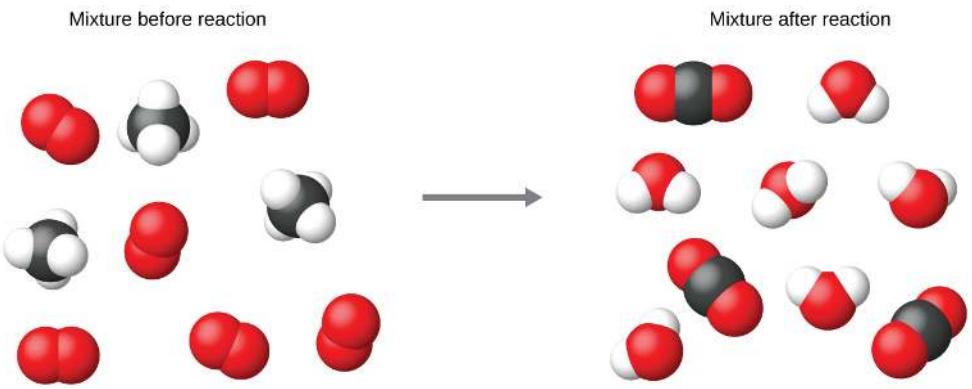


Figure 10.2.2: Regardless of the absolute numbers of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given by the chemical reaction equation.

Balancing Equations

The chemical equation described above in Figure 10.2.1 is **balanced**, meaning that equal numbers of atoms for each element involved in the reaction are represented on the reactant and product sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example reaction, CO_2 and H_2O, contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

$$\left(1CO_2 \text{ molecule} \times \frac{2O \text{ atoms}}{CO_2 \text{ molecule}}\right) + \left(2H_2O \text{ molecules} \times \frac{1O \text{ atom}}{H_2O \text{ molecule}}\right) = 4O \text{ atoms}$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:

$$\mathrm{CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O}$$

Element	Reactants	Products	Balanced?
C	1 × 1 = 1	1 × 1 = 1	1 = 1, yes
H	4 × 1 = 4	2 × 2 = 4	4 = 4, yes
O	2 × 2 = 4	(1 × 2) + (2 × 1) = 4	4 = 4, yes

A balanced chemical equation often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an *unbalanced* chemical equation:

$$\mathrm{H_2O \longrightarrow H_2 + O_2} \qquad \text{(unbalanced)}$$

Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

Element	Reactants	Products	Balanced?
H	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 1 = 1	1 × 2 = 2	1 ≠ 2, no

The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the *coefficients* of the equation may be changed as needed. Keep in mind, of course, that the *formula subscripts* define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For

example, changing the reactant formula from H_2O to H_2O_2 would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for H_2O to 2.

$$2\,H_2O \longrightarrow H_2 + O_2 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
H	**2** × 2 = 4	1 × 2 = 2	4 ≠ 2, no
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the H_2 product to 2.

$$2\,H_2O \longrightarrow 2\,H_2 + O_2 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
H	2 × 2 = 4	**2** × 2 = 4	4 = 4, yes
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides, and the balanced equation is, therefore:

$$H_2O \longrightarrow H_2 + O_2 \qquad \text{(balanced)}$$

Example 10.2.1: Balancing Chemical Equations

Write a balanced equation for the reaction of molecular nitrogen (N_2) and oxygen (O_2) to form dinitrogen pentoxide.

Solution

First, write the unbalanced equation.

$$N_2 + O_2 \longrightarrow N_2O_5 \qquad \text{(unbalanced)}$$

Next, count the number of each type of atom present in the unbalanced equation.

Element	Reactants	Products	Balanced?
N	1 2 = 2	1 2 = 2	2 = 2, yes
O	1 2 = 2	1 5 = 5	2 ≠ 5, no

Though nitrogen is balanced, changes in coefficients are needed to balance the number of oxygen atoms. To balance the number of oxygen atoms, a reasonable first attempt would be to change the coefficients for the O_2 and N_2O_5 to integers that will yield 10 O atoms (the least common multiple for the O atom subscripts in these two formulas).

$$N_2 + 5\,O_2 \longrightarrow 2\,N_2O_5 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	**2** × 2 = 4	2 ≠ 4, no
O	**5** × 2 = 10	**2** × 5 = 10	10 = 10, yes

The N atom balance has been upset by this change; it is restored by changing the coefficient for the reactant N_2 to 2.

$$2\,N_2 + 5\,O_2 \longrightarrow 2\,N_2O_5 \qquad \text{(balanced)}$$

Element	Reactants	Products	Balanced?

Element	Reactants	Products	Balanced?
N	**2** × 2 = 4	2 × 2 = 4	4 = 4, yes
O	5 × 2 = 10	2 × 5 = 10	10 = 10, yes

The numbers of N and O atoms on either side of the equation are now equal, and so the equation is balanced.

? Exercise 10.2.1

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen, and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)

Answer

$$2\,NH_4NO_3 \longrightarrow 2\,N_2{=}O_2 + 4\,H_2O$$

It is sometimes convenient to use fractions instead of integers as intermediate coefficients in the process of balancing a chemical equation. When balance is achieved, all the equation's coefficients may then be multiplied by a whole number to convert the fractional coefficients to integers without upsetting the atom balance. For example, consider the reaction of ethane (C_2H_6) with oxygen to yield H_2O and CO_2, represented by the unbalanced equation:

$$C_2H_6 + O_2 \longrightarrow H_2O + CO_2 \qquad \text{(unbalanced)}$$

Following the usual inspection approach, one might first balance C and H atoms by changing the coefficients for the two product species, as shown:

$$C_2H_6 + O_2 \longrightarrow 3\,H_2O + 2\,CO_2 \qquad \text{(unbalanced)}$$

This results in seven O atoms on the product side of the equation, an odd number—no integer coefficient can be used with the O_2 reactant to yield an odd number, so a fractional coefficient, is used instead to yield a provisional balanced equation:

$$C_2H_6 + \tfrac{7}{2}O_2 \longrightarrow 3\,H_2O + 2\,CO_2 \qquad \text{(unbalanced)}$$

A conventional balanced equation with integer-only coefficients is derived by multiplying each coefficient by 2:

$$2\,C_2H_6 + 7\,O_2 \longrightarrow 6\,H_2O + 4\,CO_2 \qquad \text{(balanced)}$$

Finally with regard to balanced equations, recall that convention dictates use of the *smallest whole-number coefficients*. Although the equation for the reaction between molecular nitrogen and molecular hydrogen to produce ammonia is, indeed, balanced,

$$3\,N_2 + 9\,H_2 \longrightarrow 6\,NH_3$$

the coefficients are not the smallest possible integers representing the relative numbers of reactant and product molecules. Dividing each coefficient by the greatest common factor, 3, gives the preferred equation:

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

Link to Learning

Use this interactive tutorial for additional practice balancing equations.

Additional Information in Chemical Equations

The physical states of reactants and products in chemical equations very often are indicated with a parenthetical abbreviation following the formulas. Common abbreviations include *s* for solids, *l* for liquids, *g* for gases, and *aq* for substances dissolved in water (*aqueous solutions*, as introduced in the preceding chapter). These notations are illustrated in the example equation here:

$$2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

This equation represents the reaction that takes place when sodium metal is placed in water. The solid sodium reacts with liquid water to produce molecular hydrogen gas and the ionic compound sodium hydroxide (a solid in pure form, but readily dissolved in water).

Special conditions necessary for a reaction are sometimes designated by writing a word or symbol above or below the equation's arrow. For example, a reaction carried out by heating may be indicated by the uppercase Greek letter delta (Δ) over the arrow.

$$\mathrm{CaCO_3(s)} \xrightarrow{\Delta} \mathrm{CaO(s) + CO_2(g)}$$

Other examples of these special conditions will be encountered in more depth in later chapters.

Equations for Ionic Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an aqueous solution. When aqueous solutions of $CaCl_2$ and $AgNO_3$ are mixed, a reaction takes place producing aqueous $Ca(NO_3)_2$ and solid AgCl:

$$\mathrm{CaCl_2(aq) + 2\,AgNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2\,AgCl(s)}$$

This balanced equation, derived in the usual fashion, is called a **molecular equation** because it doesn't explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may *dissociate* into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution (a thorough discussion of this important process is provided in the chapter on solutions). Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$\begin{aligned}
\mathrm{CaCl_2(aq)} &\longrightarrow \mathrm{Ca^{2+}(aq) + 2\,Cl^-(aq)} \\
\mathrm{2\,AgNO_3(aq)} &\longrightarrow \mathrm{2\,Ag^+(aq) + 2\,NO_3^-(aq)} \\
\mathrm{Ca(NO_3)_2(aq)} &\longrightarrow \mathrm{Ca^{2+}(aq) + 2\,NO_3^-(aq)}
\end{aligned}$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, *s*.

Explicitly representing all dissolved ions results in a **complete ionic equation**. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$\mathrm{Ca^{2+}(aq) + 2\,Cl^-(aq) + 2\,Ag^+(aq) + 2\,NO_3^-(aq) \longrightarrow Ca^{2+}(aq) + 2\,NO_3^-(aq) + 2\,AgCl(s)}$$

Examining this equation shows that two chemical species are present in identical form on both sides of the arrow, $Ca^{2+}(aq)$ and These spectator ions—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield a more succinct representation called a **net ionic equation**:

$$\begin{aligned}
&\cancel{\mathrm{Ca^{2+}(aq)}} + \mathrm{2\,Cl^-(aq) + 2\,Ag^+(aq) + \cancel{2\,NO_3^-(aq)} \longrightarrow \cancel{Ca^{2+}(aq)} + \cancel{2\,NO_3^-(aq)} + 2\,AgCl(s)} \\
&\mathrm{2\,Cl^-(aq) + 2\,Ag^+(aq) \longrightarrow 2\,AgCl(s)}
\end{aligned}$$

Following the convention of using the smallest possible integers as coefficients, this equation is then written:

$$\mathrm{Cl^-(aq) + Ag^+(aq) \longrightarrow AgCl(s)}$$

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver(I) ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of Cl^- and Ag^+.

✔ Example 10.2.2: Ionic and Molecular Equations

When carbon dioxide is dissolved in an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

Solution

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$$\mathrm{CO_2(aq) + NaOH(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)} \qquad \text{(unbalanced)}$$

Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction:

$$CO_2(aq) + 2\,NaOH(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

The two dissolved ionic compounds, NaOH and Na_2CO_3, can be represented as dissociated ions to yield the complete ionic equation:

$$CO_2(aq) + 2\,Na^+(aq) + OH^-(aq) \longrightarrow 2\,Na^+(aq) + CO_3^{2-}(aq) + H_2O(l)$$

Finally, identify the spectator ion(s), in this case $Na^+(aq)$, and remove it from each side of the equation to generate the net ionic equation:

$$CO_2(aq) + \cancel{2\,Na^+(aq)} + OH^-(aq) \longrightarrow \cancel{2\,Na^+(aq)} + CO_3^{2-}(aq) + H_2O(l)$$

$$CO_2(aq) + OH^-(aq) \longrightarrow CO_3^{2-}(aq) + H_2O(l)$$

Exercise 10.2.2

Diatomic chlorine and sodium hydroxide (lye) are commodity chemicals produced in large quantities, along with diatomic hydrogen, via the electrolysis of brine, according to the following unbalanced equation:

$$NaCl(aq) + H_2O(l) \xrightarrow{\text{electricity}} NaOH(aq) + H_2(g) + Cl_2(g)$$

Write balanced molecular, complete ionic, and net ionic equations for this process.

Answer

$$2\,NaCl(aq) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g) + Cl_2(g) \quad \text{(molecular)}$$

$$2\,Na^+(aq) + 2\,Cl^-(aq) + 2\,H_2O(l) \longrightarrow 2\,Na^+(aq) + 2\,OH^-(aq) + H_2(g) + Cl_2(g) \quad \text{(complete ionic)}$$

$$2\,Cl^-(aq) + 2\,H_2O(l) \longrightarrow 2\,OH^-(aq) + H_2(g) + Cl_2(g) \quad \text{(net ionic)}$$

This page titled 10.2: Writing and Balancing Chemical Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **4.1: Writing and Balancing Chemical Equations** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

10.3: Reaction Stoichiometry

Learning Objectives

By the end of this section, you will be able to:

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass, moles, and solution molarity

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's **stoichiometry**, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Food preparation, for example, offers an appropriate comparison. A recipe for making eight pancakes calls for 1 cup pancake mix, cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is

$$1 \text{ cup mix} + \frac{3}{4} \text{ cup milk} + 1 \text{ egg} \longrightarrow 8 \text{ pancakes}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24 \ \cancel{\text{paneakes}} \times \frac{1 \text{ egg}}{8 \ \cancel{\text{paneakes}}} = 3 \text{ eggs}$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive **stoichiometric factors** that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

$$\mathrm{N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g)}$$

This equation shows ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2 \text{ NH}_3 \text{ molecules}}{3 \text{ H}_2 \text{ molecules}} \text{ or } \frac{2 \text{ doz NH}_3 \text{ molecules}}{3 \text{ doz H}_2 \text{ molecules}} \text{ or } \frac{2 \text{ mol NH}_3 \text{ molecules}}{3 \text{ mol H}_2 \text{ molecules}}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Example 10.3.1: Moles of Reactant Required in a Reaction

How many moles of $\mathrm{I_2}$ are required to react with 0.429 mol of Al according to the following equation (see Figure 10.3.1)?

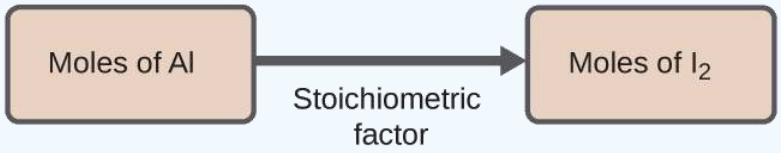
$$\mathrm{2\,Al + 3\,I_2 \longrightarrow 2\,AlI_3}$$



Figure 10.3.1: Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapor. (credit: modification of work by Mark Ott)

Solution

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest is The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$mol\,\mathrm{I_2} = 0.429\,\mathrm{mol\,Al} \times \frac{3\,\mathrm{mol\,I_2}}{2\,\mathrm{mol\,Al}}$$
$$= 0.644\,\mathrm{mol\,I_2}$$

? Exercise 10.3.1

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation

Answer

2.04 mol

✓ Example 10.3.2: Number of Product Molecules Generated by a Reaction

How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$$

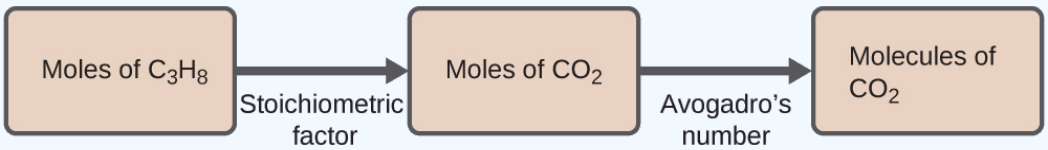
Solution

The approach here is the same as for Example 10.3.1, though the absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

$$\frac{3\,\mathrm{mol\,CO_2}}{1\,\mathrm{mol\,C_3H_8}}$$

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number,



$$0.75 mol\,\mathrm{C_3H_8} \times \frac{3 mol \mathrm{CO_2}}{1 mol\,\mathrm{C_3H_8}} \times \frac{6.022 \times 10^{23}\,\mathrm{CO_2\ molecules}}{mol\,\mathrm{CO_2}} = 1.4 \times 10^{24}\,\mathrm{CO_2\ molecules}$$

Exercise 10.3.2

How many NH_3 molecules are produced by the reaction of 4.0 mol of $Ca(OH)_2$ according to the following equation:

$$(NH_4)_2SO_4 + Ca(OH)_2 \longrightarrow 2\,NH_3 + CaSO_4 + 2\,H_2O$$

Answer

$4.8\ 10^{24}$ NH_3 molecules

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related. Directly measuring numbers of atoms and molecules is, however, not an easy task, and the practical application of stoichiometry requires that we use the more readily measured property of mass.

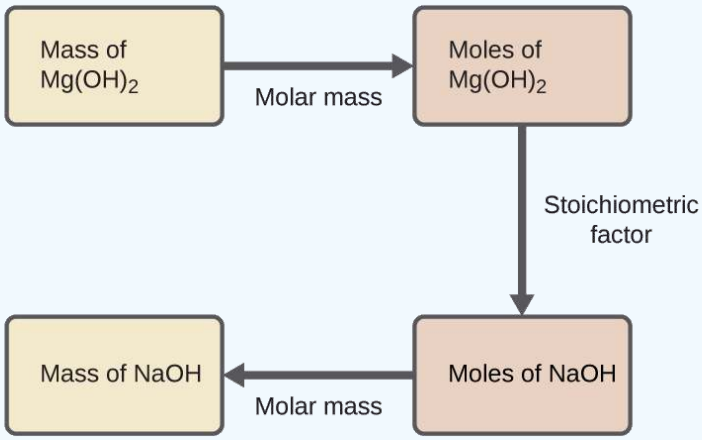
Example 10.3.3: Relating Masses of Reactants and Products

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$MgCl_2(aq) + 2\,NaOH(aq) \longrightarrow Mg(OH)_2(s) + 2\,NaCl(aq)$$

Solution

The approach used previously in Example 10.3.1 and Example 10.3.2 is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16g\,Mg(OH)_2 \times \frac{1mol\,Mg(OH)_2}{58.3g\,Mg(OH)_2} \times \frac{2\,mol\,NaOH}{1\,mol\,Mg(OH)_2} \times \frac{40.0\,g\,NaOH}{mol\,NaOH} = 22\,g\,NaOH$$

Exercise 10.3.3

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is:

$$4\,Ga + 3\,O_2 \longrightarrow 2\,Ga_2O_3$$

Answer

39.0 g

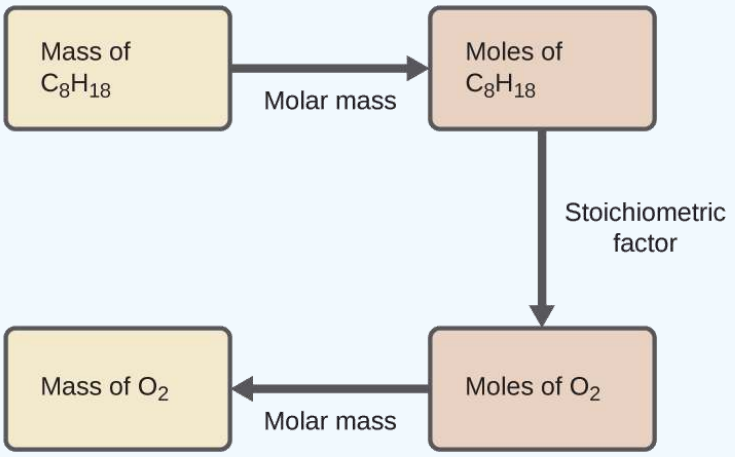
Example 10.3.4: Relating Masses of Reactants

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$2\,C_8H_{18} + 25\,O_2 \longrightarrow 16\,CO_2 + 18\,H_2O$$

Solution

The approach required here is the same as for the Example 10.3.3, differing only in that the provided and requested masses are both for reactant species.



$$702 g C_8H_{18} \times \frac{1 mol C_8H_{18}}{114.23 g G_8H_{18}} \times \frac{25 mol O_2}{2 mol C_8H_{18}} \times \frac{32.00 g O_2}{mol O_2} = 2.46 \times 10^3 g O_2$$

Exercise 10.3.4

What mass of CO is required to react with 25.13 g of Fe_2O_3 according to the equation $Fe_2O_3 + 3\,CO \longrightarrow 2\,Fe + 3\,CO_2$?

Answer

13.22 g

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure 10.3.2 provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

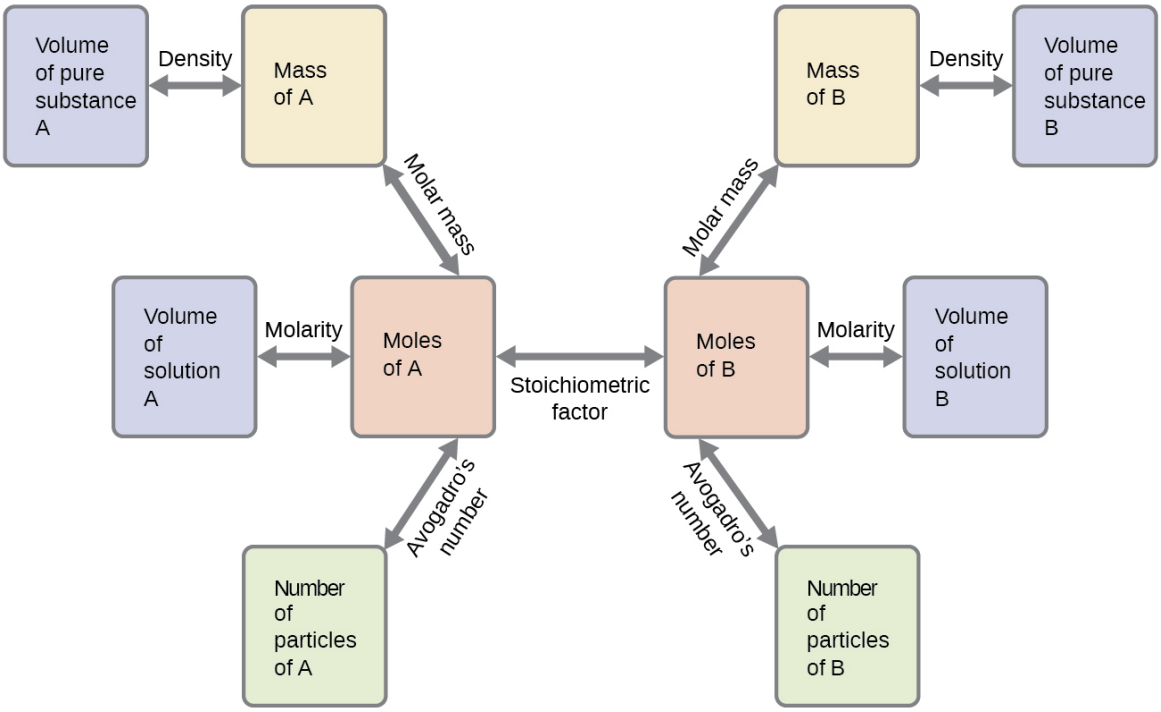


Figure 10.3.2: The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations.

Chemistry in Everyday Life: Airbags

Airbags (Figure 10.3.3) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2\,NaN_3(s) \longrightarrow 3\,N_2(g) + 2\,Na(s)$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.



Figure 10.3.3: Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

This page titled 10.3: Reaction Stoichiometry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **4.3: Reaction Stoichiometry** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

10.4: Reaction Yields

Learning Objectives

By the end of this section, you will be able to:

- Explain the concepts of theoretical yield and limiting reactants/reagents.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as *stoichiometric amounts*. All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Limiting Reactant

Consider another food analogy, making grilled cheese sandwiches (Figure 10.4.1):

$$1 \text{ slice of cheese} + 2 \text{ slices of bread} \longrightarrow 1 \text{ sandwich}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

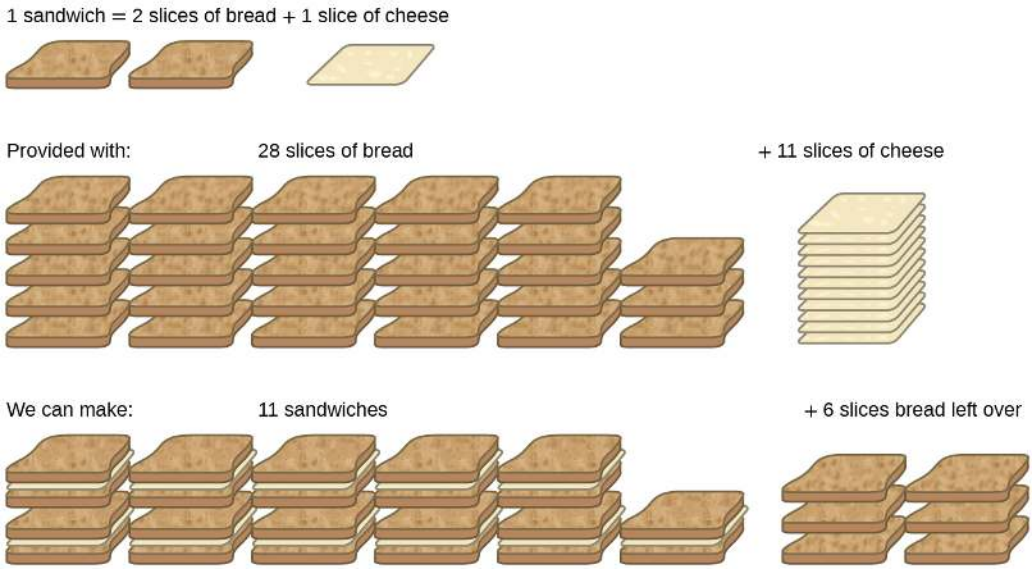


Figure 10.4.1: Sandwich making can illustrate the concepts of limiting and excess reactants.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)}$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the **limiting reactant**, and the other substance(s) is the **excess reactant**. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric

ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen unreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 3\ \text{mol H}_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol H}_2} = 6\ \text{mol HCl}$$

Complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 2\ \text{mol Cl}_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol Cl}_2} = 4\ \text{mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be unreacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure 10.4.2).

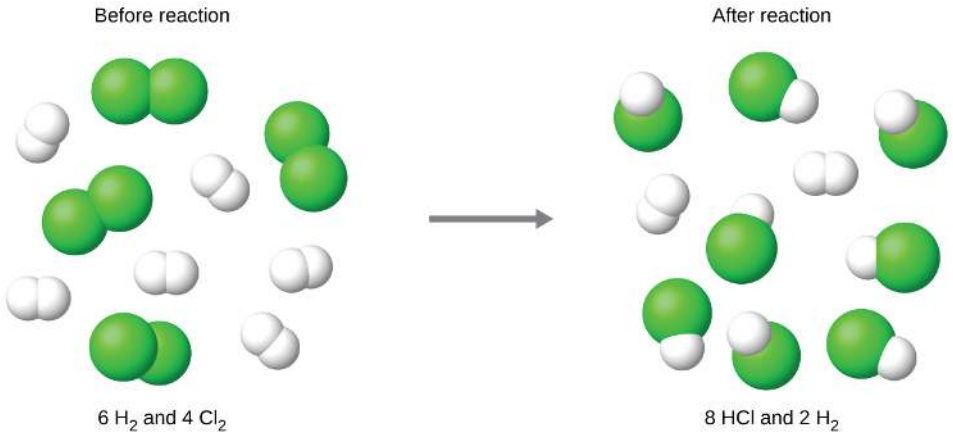


Figure 10.4.2: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

Link to Learning

View this interactive simulation illustrating the concepts of limiting and excess reactants.

Example 10.4.1: Identifying the Limiting Reactant

Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$$3Si(s) + 2N_2(g) \longrightarrow Si_3N_4(s)$$

Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

Solution

Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$molSi = 2.00gSi \times \frac{1molSi}{28.09gSi} = 0.0712molSi$$

$$molSi = 2.00gSi \times \frac{1molSi}{28.09gSi} = 0.0712molSi$$

The provided Si:N_2 molar ratio is:

$$\frac{0.0712 mol Si}{0.0535 mol N_2} = \frac{1.33 mol Si}{1 mol N_2}$$

The stoichiometric Si:N_2 ratio is:

$$\frac{3 mol Si}{2 mol N_2} = \frac{1.5 mol Si}{1 mol N_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$mol Si_3N_4 \text{ produced } = 0.0712 mol Si \times \frac{1 mol Si_3N_4}{3 mol Si^2} = 0.0237 mol Si_3N_4$$

while the 0.0535 moles of nitrogen would produce

$$\text{mol Si } N_4 \text{ produced } = 0.0535 mol N \times \frac{1 mol Si_3N_4}{2 mol N_2} = 0.0268 mol Si_3N_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant.

? Exercise 10.4.1

Which is the limiting reactant when 5.00 g of H_2 and 10.0 g of O_2 react and form water?

Answer

O_2

Percent Yield

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the **theoretical yield** of the reaction. In practice, the amount of product obtained is called the **actual yield**, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this chapter). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its **percent yield**:

$$\text{percent yield } = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{10.4.1}$$

Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when percent yield is calculated.

✓ Example 10.4.2: Calculation of Percent Yield

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$CuSO_4(aq) + Zn(s) \longrightarrow Cu(s) + ZnSO_4(aq)$$

What is the percent yield?

Solution

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

$$1.274\ \text{g}\ \text{CuSO}_4 \times \frac{1\ \text{mol}\ \text{CuSO}_4}{159.62\ \text{g}\ \text{CuSO}_4} \times \frac{1\ \text{mol}\ \text{Cu}}{1\ \text{mol}\ \text{CuSO}_4} \times \frac{63.55\ \text{g}\ \text{Cu}}{1\ \text{mol}\ \text{Cu}} = 0.5072\ \text{g}\ \text{Cu}$$

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated (via Equation 10.4.1 to be

$$\begin{aligned}\text{percent yield} &= \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100\% \\ &= \left(\frac{0.392\ \text{g}\ \text{Cu}}{0.5072\ \text{g}\ \text{Cu}_2}\right) \times 100\% \\ &= 77.3\%\end{aligned}$$

Exercise 10.4.2

What is the percent yield of a reaction that produces 12.5 g of the gas Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$\text{CCl}_4 + 2\,\text{HF} \longrightarrow \text{CF}_2\text{Cl}_2 + 2\,\text{HCl}$$

Answer

48.3%

How Sciences Interconnect: Green Chemistry and Atom Economy

The purposeful design of chemical products and processes that minimize the use of environmentally hazardous substances and the generation of waste is known as *green chemistry*. Green chemistry is a philosophical approach that is being applied to many areas of science and technology, and its practice is summarized by guidelines known as the "Twelve Principles of Green Chemistry" (see details at this website). One of the 12 principles is aimed specifically at maximizing the efficiency of processes for synthesizing chemical products. The *atom economy* of a process is a measure of this efficiency, defined as the percentage by mass of the final product of a synthesis relative to the masses of *all* the reactants used:

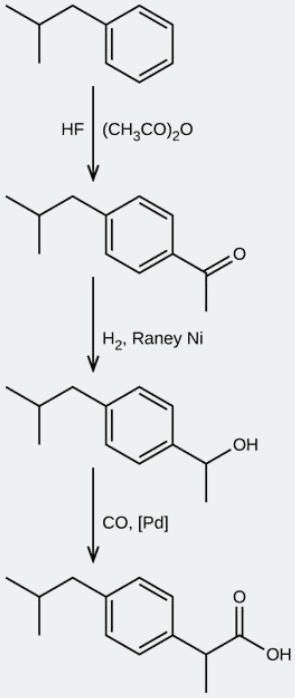
$$\text{atom economy} = \frac{\text{mass of product}}{\text{mass of reactants}} \times 100\%$$

Though the definition of atom economy at first glance appears very similar to that for percent yield, be aware that this property represents a difference in the *theoretical* efficiencies of *different* chemical processes. The percent yield of a given chemical process, on the other hand, evaluates the efficiency of a process by comparing the yield of product actually obtained to the maximum yield predicted by stoichiometry.

The synthesis of the common nonprescription pain medication, ibuprofen, nicely illustrates the success of a green chemistry approach (Figure 10.4.3). First marketed in the early 1960s, ibuprofen was produced using a six-step synthesis that required 514 g of reactants to generate each mole (206 g) of ibuprofen, an atom economy of 40%. In the 1990s, an alternative process was developed by the BHC Company (now BASF Corporation) that requires only three steps and has an atom economy of ~80%, nearly twice that of the original process. The BHC process generates significantly less chemical waste; uses less-hazardous and recyclable materials; and provides significant cost-savings to the manufacturer (and, subsequently, the consumer). In recognition of the positive environmental impact of the BHC process, the company received the Environmental Protection Agency's Greener Synthetic Pathways Award in 1997.



(a)



(b)

Figure 10.4.3: (a) Ibuprofen is a popular nonprescription pain medication commonly sold as 200 mg tablets. (b) The BHC process for synthesizing ibuprofen requires only three steps and exhibits an impressive atom economy. (credit a: modification of work by Derrick Coetzee)

This page titled 10.4: Reaction Yields is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **4.4: Reaction Yields** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

10.5: Molarity

Learning Objectives

By the end of this section, you will be able to:

- Describe the fundamental properties of solutions
- Calculate solution concentrations using molarity
- Perform dilution calculations using the dilution equation

Preceding sections of this chapter focused on the composition of substances: samples of matter that contain only one type of element or compound. However, mixtures—samples of matter containing two or more substances physically combined—are more commonly encountered in nature than are pure substances. Similar to a pure substance, the relative composition of a mixture plays an important role in determining its properties. The relative amount of oxygen in a planet's atmosphere determines its ability to sustain aerobic life. The relative amounts of iron, carbon, nickel, and other elements in steel (a mixture known as an "alloy") determine its physical strength and resistance to corrosion. The relative amount of the active ingredient in a medicine determines its effectiveness in achieving the desired pharmacological effect. The relative amount of sugar in a beverage determines its sweetness (see Figure 10.5.1). This section will describe one of the most common ways in which the relative compositions of mixtures may be quantified.



Figure 10.5.1: Sugar is one of many components in the complex mixture known as coffee. The amount of sugar in a given amount of coffee is an important determinant of the beverage's sweetness. (credit: Jane Whitney)

Solutions

Solutions have previously been defined as homogeneous mixtures, meaning that the composition of the mixture (and therefore its properties) is uniform throughout its entire volume. Solutions occur frequently in nature and have also been implemented in many forms of manmade technology. A more thorough treatment of solution properties is provided in the chapter on solutions and colloids, but provided here is an introduction to some of the basic properties of solutions.

The relative amount of a given solution component is known as its **concentration**. Often, though not always, a solution contains one component with a concentration that is significantly greater than that of all other components. This component is called the solvent and may be viewed as the medium in which the other components are dispersed, or **dissolved**. Solutions in which water is the solvent are, of course, very common on our planet. A solution in which water is the solvent is called an **aqueous solution**.

A **solute** is a component of a solution that is typically present at a much lower concentration than the solvent. Solute concentrations are often described with qualitative terms such as **dilute** (of relatively low concentration) and **concentrated** (of relatively high concentration).

Concentrations may be quantitatively assessed using a wide variety of measurement units, each convenient for particular applications. **Molarity** (*M*) is a useful concentration unit for many applications in chemistry. Molarity is defined as the number of moles of solute in exactly 1 liter (1 L) of the solution:

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

Example 10.5.1: Calculating Molar Concentrations

A 355-mL soft drink sample contains 0.133 mol of sucrose (table sugar). What is the molar concentration of sucrose in the beverage?

Solution

Since the molar amount of solute and the volume of solution are both given, the molarity can be calculated using the definition of molarity. Per this definition, the solution volume must be converted from mL to L:

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{0.133\ \text{mol}}{355\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}}} = 0.375 M$$

Exercise 10.5.1

A teaspoon of table sugar contains about 0.01 mol sucrose. What is the molarity of sucrose if a teaspoon of sugar has been dissolved in a cup of tea with a volume of 200 mL?

Answer

0.05 M

Example 10.5.2: Deriving Moles and Volumes from Molar Concentrations

How much sugar (mol) is contained in a modest sip (~10 mL) of the soft drink from Example 10.5.1?

Solution

Rearrange the definition of molarity to isolate the quantity sought, moles of sugar, then substitute the value for molarity derived in Example 10.5.1, 0.375 M:

$$\begin{aligned} M &= \frac{\text{mol solute}}{L\ \text{solution}} \\ \text{mol solute} &= M \times L\ \text{solution} \\ &= 0.375 \frac{\text{mol sugar}}{L} \times \left(10\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}}\right) \\ &= 0.004\ \text{mol sugar} \end{aligned}$$

Exercise 10.5.2

What volume (mL) of the sweetened tea described in Example 10.5.1 contains the same amount of sugar (mol) as 10 mL of the soft drink in this example?

Answer

80 mL

Example 10.5.3: Calculating Molar Concentrations from the Mass of Solute

Distilled white vinegar (Figure 10.5.2) is a solution of acetic acid, CH_3CO_2H, in water. A 0.500-L vinegar solution contains 25.2 g of acetic acid. What is the concentration of the acetic acid solution in units of molarity?



Figure 10.5.2: Distilled white vinegar is a solution of acetic acid in water.

Solution

As in previous examples, the definition of molarity is the primary equation used to calculate the quantity sought. Since the mass of solute is provided instead of its molar amount, use the solute's molar mass to obtain the amount of solute in moles:

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{25.2\,\text{g}\,\text{CH}_3\text{CO}_2\text{H} \times \frac{1\,\text{mol}]\,\text{CH}_3\text{CO}_2\text{H}}{60.052\,\text{g}\,\text{CH}_3\text{CO}_2\text{H}}}{0.500\,\text{L solution}} = 0.839\,\text{M}$$

? Exercise 10.5.3

Calculate the molarity of 6.52 g of $CoCl_2$ (128.9 g/mol) dissolved in an aqueous solution with a total volume of 75.0 mL.

Answer

0.674 *M*

✔ Example 10.5.4: Determining the Mass of Solute in a Given Volume of Solution

How many grams of NaCl are contained in 0.250 L of a 5.30-*M* solution?

Solution

The volume and molarity of the solution are specified, so the amount (mol) of solute is easily computed as demonstrated in Example 10.5.2

$$\begin{aligned} M &= \frac{\text{mol solute}}{L\,\text{solution}} \\ \text{mol solute} &= M \times L\,\text{solution} \\ &= 5.30\frac{\text{mol NaCl}}{\text{L}} \times 0.250L \\ &= 1.325\,\text{mol NaCl} \end{aligned}$$

Finally, this molar amount is used to derive the mass of NaCl:

$$1.325\,\text{mol NaCl} \times \frac{58.44\,\text{g NaCl}}{1\,\text{mol NaCl}} = 77.4\,\text{g NaCl}$$

Exercise 10.5.4

How many grams of $CaCl_2$ (110.98 g/mol) are contained in 250.0 mL of a 0.200-*M* solution of calcium chloride?

Answer

5.55 g $CaCl_2$

When performing calculations stepwise, as in Example 10.5.4, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In Example 10.5.4, the molar amount of NaCl computed in the first step, 1.325 mol, would be properly rounded to 1.32 mol if it were to be reported; however, although the last digit (5) is not significant, it must be retained as a guard digit in the intermediate calculation. If the guard digit had not been retained, the final calculation for the mass of NaCl would have been 77.1 g, a difference of 0.3 g.

In addition to retaining a guard digit for intermediate calculations, rounding errors may also be avoided by performing computations in a single step (see Example 10.5.5). This eliminates intermediate steps so that only the final result is rounded.

Example 10.5.5: Determining the Volume of Solution Containing a Given Mass of Solute

In Example 10.5.3, the concentration of acetic acid in white vinegar was determined to be 0.839 *M*. What volume of vinegar contains 75.6 g of acetic acid?

Solution

First, use the molar mass to calculate moles of acetic acid from the given mass:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} = \text{mol solute}$$

Then, use the molarity of the solution to calculate the volume of solution containing this molar amount of solute:

$$\text{mol solute} \times \frac{\text{L solution}}{\text{mol solute}} = \text{L solution}$$

Combining these two steps into one yields:

$$\text{g solute} \times \left(\frac{\text{mol solute}}{\text{g solute}}\right) \times \left(\frac{\text{L solution}}{\text{mol solute}}\right) = \text{L solution}$$

$$75.6\ \text{g}\ CH_3CO_2H \left(\frac{\text{mol}\ CH_3CO_2H}{60.05\ \text{g}}\right) \left(\frac{\text{L solution}}{0.839\ \text{mol CH3}\ CO_2H}\right) = 1.50\ \text{L solution}$$

Exercise 10.5.5

What volume of a 1.50-*M* KBr solution contains 66.0 g KBr?

Answer

0.370 L

Dilution of Solutions

Dilution is the process whereby the concentration of a solution is lessened by the addition of solvent. For example, a glass of iced tea becomes increasingly diluted as the ice melts. The water from the melting ice increases the volume of the solvent (water) and the overall volume of the solution (iced tea), thereby reducing the relative concentrations of the solutes that give the beverage its taste (Figure 10.5.3).

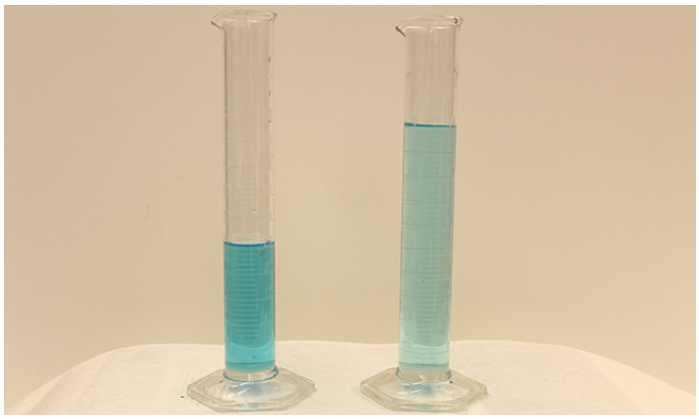


Figure 10.5.3: Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott)

Dilution is also a common means of preparing solutions of a desired concentration. By adding solvent to a measured portion of a more concentrated *stock solution*, a solution of lesser concentration may be prepared. For example, commercial pesticides are typically sold as solutions in which the active ingredients are far more concentrated than is appropriate for their application. Before they can be used on crops, the pesticides must be diluted. This is also a very common practice for the preparation of a number of common laboratory reagents.

A simple mathematical relationship can be used to relate the volumes and concentrations of a solution before and after the dilution process. According to the definition of molarity, the number of moles of solute in a solution (n) is equal to the product of the solution's molarity (M) and its volume in liters (L):

$$n = ML$$

Expressions like these may be written for a solution before and after it is diluted:

$$n_1 = M_1 L_1$$

$$n_2 = M_2 L_2$$

where the subscripts "1" and "2" refer to the solution before and after the dilution, respectively. Since the dilution process *does not change the amount of solute in the solution*, $n_1 = n_2$. Thus, these two equations may be set equal to one another:

$$M_1 L_1 = M_2 L_2$$

This relation is commonly referred to as the dilution equation. Although this equation uses molarity as the unit of concentration and liters as the unit of volume, other units of concentration and volume may be used as long as the units properly cancel per the factor-label method. Reflecting this versatility, the dilution equation is often written in the more general form:

$$C_1 V_1 = C_2 V_2$$

where \(C\) and \(V\) are concentration and volume, respectively.

Link to Learning

Use the simulation to explore the relations between solute amount, solution volume, and concentration and to confirm the dilution equation.

Example 10.5.6: Determining the Concentration of a Diluted Solution

If 0.850 L of a 5.00-*M* solution of copper nitrate, $Cu(NO_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

Solution

The stock concentration, C_1, and volume, V_1, are provided as well as the volume of the diluted solution, V_2. Rearrange the dilution equation to isolate the unknown property, the concentration of the diluted solution, C_2:

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{C_1V_1}{V_2}$$

Since the stock solution is being diluted by more than two-fold (volume is increased from 0.85 L to 1.80 L), the diluted solution's concentration is expected to be less than one-half 5 *M*. This ballpark estimate will be compared to the calculated result to check for any gross errors in computation (for example, such as an improper substitution of the given quantities). Substituting the given values for the terms on the right side of this equation yields:

$$C_2 = \frac{0.850L \times 5.00\,\frac{mol}{L}}{1.80L} = 2.36M$$

This result compares well to our ballpark estimate (it's a bit less than one-half the stock concentration, 5 *M*).

? Exercise 10.5.6

What is the concentration of the solution that results from diluting 25.0 mL of a 2.04-*M* solution of CH_3OH to 500.0 mL?

Answer

0.102 *M* CH_3OH

✓ Example 10.5.7: Volume of a Diluted Solution

What volume of 0.12 *M* HBr can be prepared from 11 mL (0.011 L) of 0.45 *M* HBr?

Solution

Provided are the volume and concentration of a stock solution, V_1 and C_1, and the concentration of the resultant diluted solution, C_2. Find the volume of the diluted solution, V_2 by rearranging the dilution equation to isolate V_2:

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1V_1}{C_2}$$

Since the diluted concentration (0.12 *M*) is slightly more than one-fourth the original concentration (0.45 *M*), the volume of the diluted solution is expected to be roughly four times the original volume, or around 44 mL. Substituting the given values and solving for the unknown volume yields:

$$V_2 = \frac{(0.45\,\text{M})(0.011\,\text{L})}{(0.12\,\text{M})}$$

$$= 0.041L$$

The volume of the 0.12-*M* solution is 0.041 L (41 mL). The result is reasonable and compares well with the rough estimate.

? Exercise 10.5.7

A laboratory experiment calls for 0.125 *M* HNO_3. What volume of 0.125 *M* HNO_3 can be prepared from 0.250 L of 1.88 *M* HNO_3?

Answer

3.76 L

Example 10.5.8: Volume of a Concentrated Solution Needed for Dilution

What volume of 1.59 *M* KOH is required to prepare 5.00 L of 0.100 *M* KOH?

Solution

Given are the concentration of a stock solution, C_1, and the volume and concentration of the resultant diluted solution, V_2 and C_2. Find the volume of the stock solution, V_1 by rearranging the dilution equation to isolate V_1:

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1}$$

Since the concentration of the diluted solution 0.100 *M* is roughly one-sixteenth that of the stock solution (1.59 *M*), the volume of the stock solution is expected to be about one-sixteenth that of the diluted solution, or around 0.3 liters. Substituting the given values and solving for the unknown volume yields:

$$V_1 = \frac{(0.100\,\text{M})(5.00\,\text{L})}{1.59\,\text{M}}$$

$$= 0.314\,\text{L}$$

Thus, 0.314 L of the 1.59-*M* solution is needed to prepare the desired solution. This result is consistent with the rough estimate.

Exercise 10.5.8

What volume of a 0.575-*M* solution of glucose ($C_6H_{12}O_6$) can be prepared from 50.00 mL of a 3.00-*M* glucose solution?

Answer

0.261 L

This page titled 10.5: Molarity is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **3.3: Molarity** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

10.6: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubilities.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. We described a precipitation reaction in which a colorless solution of silver nitrate was mixed with a yellow-orange solution of potassium dichromate to give a reddish precipitate of silver dichromate:

$$AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + KNO_3(aq) \tag{10.6.1}$$

This equation has the general form of an exchange reaction:

$$AC + BD \rightarrow \underset{insoluble}{AD} + BC \tag{10.6.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called *double-displacement reactions*. Two important uses of precipitation reactions are to isolate metals that have been extracted from their ores and to recover precious metals for recycling.



***Video**: Mixing Potassium Chromate and Silver Nitrate together to initiate a precipitation reaction (Equation 10.6.3).*

While chemical equations show the identities of the reactants and the products and gave the stoichiometries of the reactions, but they told us very little about what was occurring in solution. In contrast, equations that show only the hydrated species focus our attention on the chemistry that is taking place and allow us to see similarities between reactions that might not otherwise be apparent.

Let's consider the reaction of silver nitrate with potassium dichromate. When aqueous solutions of silver nitrate and potassium dichromate are mixed, silver dichromate forms as a red solid. The overall balanced chemical equation for the reaction shows each reactant and product as undissociated, electrically neutral compounds:

$$2AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + 2KNO_3(aq) \tag{10.6.3}$$

Although Equation 10.6.3 gives the identity of the reactants and the products, it does not show the identities of the actual species in solution. Because ionic substances such as $AgNO_3$ and $K_2Cr_2O_7$ are strong electrolytes, they dissociate completely in aqueous solution to form ions. In contrast, because $Ag_2Cr_2O_7$ is not very soluble, it separates from the solution as a solid. To find out what is actually occurring in solution, it is more informative to write the reaction as a complete ionic equation showing which ions and molecules are hydrated and which are present in other forms and phases:

$$2Ag^+(aq) + 2NO_3^-(aq) + 2K^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s) + 2K^+(aq) + 2NO_3^-(aq) \tag{10.6.4}$$

Note that K^+(aq) and NO_3^-(aq) ions are present on both sides of the equation, and their coefficients are the same on both sides. These ions are called spectator ions because they do not participate in the actual reaction. Canceling the spectator ions gives the net ionic equation, which shows only those species that participate in the chemical reaction:

$$2Ag^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s) \tag{10.6.5}$$

Both mass and charge must be conserved in chemical reactions because the numbers of electrons and protons do not change. For charge to be conserved, the sum of the charges of the ions multiplied by their coefficients must be the same on both sides of the equation. In Equation 10.6.5, the charge on the left side is 2(+1) + 1(−2) = 0, which is the same as the charge of a neutral $Ag_2Cr_2O_7$ formula unit.

By eliminating the spectator ions, we can focus on the chemistry that takes place in a solution. For example, the overall chemical equation for the reaction between silver fluoride and ammonium dichromate is as follows:

$$2AgF(aq) + (NH_4)_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + 2NH_4F(aq) \tag{10.6.6}$$

The complete ionic equation for this reaction is as follows:

$$2Ag^+(aq) + 2F^-(aq) + 2NH_4^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s) + 2NH_4^+(aq) + 2F^-(aq) \tag{10.6.7}$$

Because two NH_4^+(aq) and two F^-(aq) ions appear on both sides of Equation 10.6.7, they are spectator ions. They can therefore be canceled to give the net ionic equation (Equation 10.6.8), which is identical to Equation 10.6.5:

$$2Ag^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow Ag_2Cr_2O_7(s) \tag{10.6.8}$$

If we look at net ionic equations, it becomes apparent that many different combinations of reactants can result in the same net chemical reaction. For example, we can predict that silver fluoride could be replaced by silver nitrate in the preceding reaction without affecting the outcome of the reaction.

✔ Example 10.6.1

Write the overall chemical equation, the complete ionic equation, and the net ionic equation for the reaction of aqueous barium nitrate with aqueous sodium phosphate to give solid barium phosphate and a solution of sodium nitrate.

Given: reactants and products

Asked for: overall, complete ionic, and net ionic equations

Strategy:

Write and balance the overall chemical equation. Write all the soluble reactants and products in their dissociated form to give the complete ionic equation; then cancel species that appear on both sides of the complete ionic equation to give the net ionic equation.

Solution:

From the information given, we can write the unbalanced chemical equation for the reaction:

$$Ba(NO_3)_2(aq) + Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + NaNO_3(aq)$$

Because the product is $Ba_3(PO_4)_2$, which contains three Ba^{2+} ions and two PO_4^{3-} ions per formula unit, we can balance the equation by inspection:

$$3Ba(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ba_3(PO_4)_2(s) + 6NaNO_3(aq)$$

This is the overall balanced chemical equation for the reaction, showing the reactants and products in their undissociated form. To obtain the complete ionic equation, we write each soluble reactant and product in dissociated form:

$$3Ba^{2+}(aq) + 6NO_3^-(aq) + 6Na^+(aq) + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s) + 6Na^+(aq) + 6NO_3^-(aq)$$

The six NO_3^-(aq) ions and the six Na^+(aq) ions that appear on both sides of the equation are spectator ions that can be canceled to give the net ionic equation:

$$3Ba^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ba_3(PO_4)_2(s)$$

? Exercise 10.6.1: Mixing Silver Fluoride with Sodium Phosphate

Write the overall chemical equation, the complete ionic equation, and the net ionic equation for the reaction of aqueous silver fluoride with aqueous sodium phosphate to give solid silver phosphate and a solution of sodium fluoride.

Answer

overall chemical equation:

$$3AgF(aq) + Na_3PO_4(aq) \rightarrow Ag_3PO_4(s) + 3NaF(aq)$$

complete ionic equation:

$$3Ag^+(aq) + 3F^-(aq) + 3Na^+(aq) + PO_4^{3-}(aq) \rightarrow Ag_3PO_4(s) + 3Na^+(aq) + 3F^-(aq)$$

net ionic equation:

$$3Ag^+(aq) + PO_4^{3-}(aq) \rightarrow Ag_3PO_4(s)$$

So far, we have always indicated whether a reaction will occur when solutions are mixed and, if so, what products will form. As you advance in chemistry, however, you will need to predict the results of mixing solutions of compounds, anticipate what kind of reaction (if any) will occur, and predict the identities of the products. Students tend to think that this means they are supposed to "just know" what will happen when two substances are mixed. Nothing could be further from the truth: an infinite number of chemical reactions is possible, and neither you nor anyone else could possibly memorize them all. Instead, you must begin by identifying the various reactions that *could* occur and then assessing which is the most probable (or least improbable) outcome.

The most important step in analyzing an unknown reaction is to *write down all the species—whether molecules or dissociated ions—that are actually present in the solution* (not forgetting the solvent itself) so that you can assess which species are most likely to react with one another. The easiest way to make that kind of prediction is to attempt to place the reaction into one of several familiar classifications, refinements of the five general kinds of reactions (acid–base, exchange, condensation, cleavage, and oxidation–reduction reactions). In the sections that follow, we discuss three of the most important kinds of reactions that occur in aqueous solutions: precipitation reactions (also known as exchange reactions), acid–base reactions, and oxidation–reduction reactions.



Determining the Products for Precipitation Reactions: https://youtu.be/r0kYeZVuTAM

Predicting Solubilities

Table 10.6.2 gives guidelines for predicting the solubility of a wide variety of ionic compounds. To determine whether a precipitation reaction will occur, we identify each species in the solution and then refer to Table 10.6.2 to see which, if any, combination(s) of cation and anion are likely to produce an insoluble salt. In doing so, it is important to recognize that *soluble* and *insoluble* are relative terms that span a wide range of actual solubilities. We will discuss solubilities in more detail later, where you will learn that very small amounts of the constituent ions remain in solution even after precipitation of an "insoluble" salt. For our purposes, however, we will assume that precipitation of an insoluble salt is complete.

Table 10.6.2: Guidelines for Predicting the Solubility of Ionic Compounds in Water

Soluble			**Exceptions**
Rule 1	most salts that contain an alkali metal (Li^+, Na^+, K^+, Rb^+, and Cs^+) and ammonium (NH_4^+)		
Rule 2	most salts that contain the nitrate (NO_3^-) anion		
Rule 3	most salts of anions derived from monocarboxylic acids (e.g., $CH_3CO_2^-$)	but not	silver acetate and salts of long-chain carboxylates
Rule 4	most chloride, bromide, and iodide salts	but not	salts of metal ions located on the lower right side of the periodic table (e.g., Cu^+, Ag^+, Pb^{2+}, and Hg_2^{2+}).
	Insoluble		**Exceptions**

Soluble			Exceptions
Rule 5	most salts that contain the hydroxide (OH^-) and sulfide (S^{2-}) anions	but not	salts of the alkali metals (group 1), the heavier alkaline earths (Ca^{2+}, Sr^{2+}, and Ba^{2+} in group 2), and the NH_4^+ ion.
Rule 6	most carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}) salts	but not	salts of the alkali metals or the NH_4^+ ion.
Rule 7	most sulfate (SO_4^{2-}) salts that contain main group cations with a charge ≥ +2	but not	salts of +1 cations, Mg^{2+}, and dipositive transition metal cations (e.g., Ni^{2+})

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of a 1.0 M aqueous NaCl solution is mixed with 500 mL of a 1.0 M aqueous KBr solution, the final solution has a volume of 1.00 L and contains 0.50 M Na^+(aq), 0.50 M Cl^-(aq), 0.50 M K^+(aq), and 0.50 M Br^-(aq). As you will see in the following sections, none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other (Figure 10.6.1).

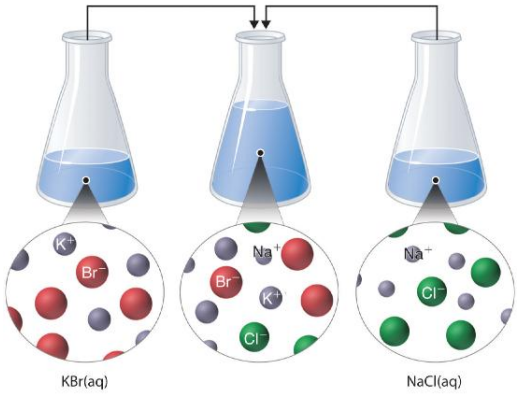


Figure 10.6.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Example 10.6.2

Using the information in Table 10.6.2 predict what will happen in each case involving strong electrolytes. Write the net ionic equation for any reaction that occurs.

a. Aqueous solutions of barium chloride and lithium sulfate are mixed.
b. Aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.
c. Aqueous solutions of strontium bromide and aluminum nitrate are mixed.
d. Solid lead(II) acetate is added to an aqueous solution of ammonium iodide.

Given: reactants

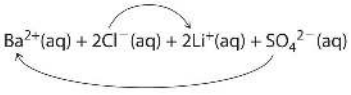
Asked for: reaction and net ionic equation

Strategy:

A. Identify the ions present in solution and write the products of each possible exchange reaction.
B. Refer to Table 10.6.2 to determine which, if any, of the products is insoluble and will therefore form a precipitate. If a precipitate forms, write the net ionic equation for the reaction.

Solution:

A Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains Ba^{2+}, Cl^-, Li^+, and SO_4^{2-} ions. The only possible exchange reaction is to form LiCl and $BaSO_4$:



B We now need to decide whether either of these products is insoluble. Table 10.6.2 shows that LiCl is soluble in water (rules 1 and 4), but $BaSO_4$ is not soluble in water (rule 5). Thus $BaSO_4$ will precipitate according to the net ionic equation

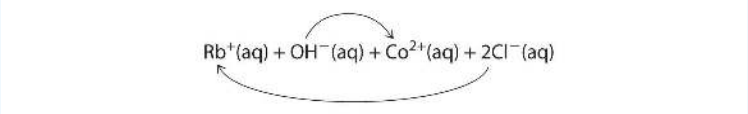
$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) \tag{10.6.9}$$

Although soluble barium salts are toxic, $BaSO_4$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a "barium milkshake" or a "barium enema"—a suspension of very fine $BaSO_4$ particles in water.



An x-ray of the digestive organs of a patient who has swallowed a "barium milkshake." A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. from Wikipedia.

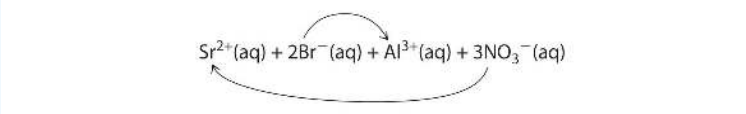
1. **A** Rubidium hydroxide and cobalt(II) chloride are strong electrolytes, so when aqueous solutions of these compounds are mixed, the resulting solution initially contains Rb^+, OH^-, Co^{2+}, and Cl^- ions. The possible products of an exchange reaction are rubidium chloride and cobalt(II) hydroxide):



B According to Table 10.6.2, RbCl is soluble (rules 1 and 4), but $Co(OH)_2$ is not soluble (rule 5). Hence $Co(OH)_2$ will precipitate according to the following net ionic equation:

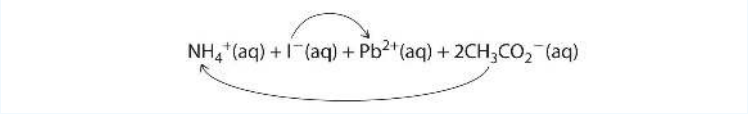
$$Co^{2+}(aq) + 2OH^-(aq) \rightarrow Co(OH)_2(s)$$

2. **A** When aqueous solutions of strontium bromide and aluminum nitrate are mixed, we initially obtain a solution that contains Sr^{2+}, Br^-, Al^{3+}, and NO_3^- ions. The two possible products from an exchange reaction are aluminum bromide and strontium nitrate:



B According to Table 10.6.2, both $AlBr_3$ (rule 4) and $Sr(NO_3)_2$ (rule 2) are soluble. Thus no net reaction will occur.

1. **A** According to Table 10.6.2, lead acetate is soluble (rule 3). Thus solid lead acetate dissolves in water to give Pb^{2+} and $CH_3CO_2^-$ ions. Because the solution also contains NH_4^+ and I^- ions, the possible products of an exchange reaction are ammonium acetate and lead(II) iodide:



B According to Table 10.6.2, ammonium acetate is soluble (rules 1 and 3), but PbI_2 is insoluble (rule 4). Thus $Pb(C_2H_3O_2)_2$ will dissolve, and PbI_2 will precipitate. The net ionic equation is as follows:

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

? Exercise 10.6.2

Using the information in Table 10.6.2, predict what will happen in each case involving strong electrolytes. Write the net ionic equation for any reaction that occurs.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer

a. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
b. $2PO_4^{3-}(aq) + 3Hg^{2+}(aq) \rightarrow Hg_3(PO_4)_2(s)$
c. $NaF(s)$ dissolves; no net reaction
d. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

Precipitation Reactions in Photography

Precipitation reactions can be used to recover silver from solutions used to develop conventional photographic film. Although largely supplanted by digital photography, conventional methods are often used for artistic purposes. Silver bromide is an off-white solid that turns black when exposed to light, which is due to the formation of small particles of silver metal. Black-and-white photography uses this reaction to capture images in shades of gray, with the darkest areas of the film corresponding to the areas that received the most light. The first step in film processing is to enhance the black/white contrast by using a developer to increase the amount of black. The developer is a reductant: because silver atoms catalyze the reduction reaction, grains of silver bromide that have already been partially reduced by exposure to light react with the reductant much more rapidly than unexposed grains.

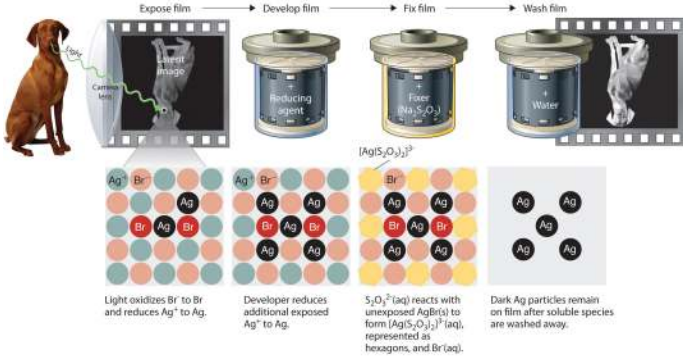


Figure 10.6.2: Outline of the Steps Involved in Producing a Black-and-White Photograph

After the film is developed, any unexposed silver bromide must be removed by a process called “fixing”; otherwise, the entire film would turn black with additional exposure to light. Although silver bromide is insoluble in water, it is soluble in a dilute solution of sodium thiosulfate ($Na_2S_2O_3$; *photographer’s hypo*) because of the formation of $[Ag(S_2O_3)_2]^{3-}$ ions. Thus washing the film with thiosulfate solution dissolves unexposed silver bromide and leaves a pattern of metallic silver granules that constitutes the negative. This procedure is summarized in Figure 10.6.2 The negative image is then projected onto paper coated with silver halides, and the developing and fixing processes are repeated to give a positive image. (Color photography works in much the same way, with a combination of silver halides and organic dyes superimposed in layers.) “Instant photo” operations can generate more than a hundred gallons of dilute silver waste solution per day. Recovery of silver from thiosulfate fixing solutions involves first removing the thiosulfate by oxidation and then precipitating Ag^+ ions with excess chloride ions.

Example 10.6.3

A silver recovery unit can process 1500 L of photographic silver waste solution per day. Adding excess solid sodium chloride to a 500 mL sample of the waste (after removing the thiosulfate as described previously) gives a white precipitate that, after filtration and drying, consists of 3.73 g of AgCl. What mass of NaCl must be added to the 1500 L of silver waste to ensure that all the Ag^+ ions precipitate?

Given: volume of solution of one reactant and mass of product from a sample of reactant solution

Asked for: mass of second reactant needed for complete reaction

Strategy:

A. Write the net ionic equation for the reaction. Calculate the number of moles of AgCl obtained from the 500 mL sample and then determine the concentration of Ag^+ in the sample by dividing the number of moles of AgCl formed by the volume of solution.
B. Determine the total number of moles of Ag^+ in the 1500 L solution by multiplying the Ag^+ concentration by the total volume.
C. Use mole ratios to calculate the number of moles of chloride needed to react with Ag^+. Obtain the mass of NaCl by multiplying the number of moles of NaCl needed by its molar mass.

Solution:

We can use the data provided to determine the concentration of Ag^+ ions in the waste, from which the number of moles of Ag^+ in the entire waste solution can be calculated. From the net ionic equation, we can determine how many moles of Cl^- are needed, which in turn will give us the mass of NaCl necessary.

A The first step is to write the net ionic equation for the reaction:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$$

We know that 500 mL of solution produced 3.73 g of AgCl. We can convert this value to the number of moles of AgCl as follows:

$$moles\ AgCl = \frac{grams\ AgCl}{molar\ mass\ AgCl} = 3.73\ \cancel{g\ AgCl}\left(\frac{1\ mol\ AgCl}{143.32\ \cancel{g\ AgCl}}\right) = 0.0260\ mol\ AgCl \tag{10.6.10}$$

Therefore, the 500 mL sample of the solution contained 0.0260 mol of Ag^+. The Ag^+ concentration is determined as follows:

$$[Ag^+] = \frac{moles\ Ag^+}{liters\ soln} = \frac{0.0260\ mol\ AgCl}{0.500\ L} = 0.0520\ M \tag{10.6.11}$$

B The total number of moles of Ag^+ present in 1500 L of solution is as follows:

$$moles\ Ag^+ = 1500\ \cancel{L}\left(\frac{0.520\ mol}{1\ \cancel{L}}\right) = 78.1\ mol\ Ag^+ \tag{10.6.12}$$

C According to the net ionic equation, one Cl^- ion is required for each Ag^+ ion. Thus 78.1 mol of NaCl are needed to precipitate the silver. The corresponding mass of NaCl is

$$mass\ NaCl = 78.1\ \cancel{mol\ NaCl}\left(\frac{58.44\ g\ NaCl}{1\ \cancel{mol\ NaCl}}\right) = 4560\ g\ NaCl = 4.56\ kg\ NaCl \tag{10.6.13}$$

Note that 78.1 mol of AgCl correspond to 8.43 kg of metallic silver, which is worth about $7983 at 2011 prices ($32.84 per troy ounce). Silver recovery may be economically attractive as well as ecologically sound, although the procedure outlined is becoming nearly obsolete for all but artistic purposes with the growth of digital photography.

? Exercise 10.6.3

Because of its toxicity, arsenic is the active ingredient in many pesticides. The arsenic content of a pesticide can be measured by oxidizing arsenic compounds to the arsenate ion (AsO_4^{3-}), which forms an insoluble silver salt (Ag_3AsO_4). Suppose you are asked to assess the purity of technical grade sodium arsenite ($NaAsO_2$), the active ingredient in a pesticide used against termites. You dissolve a 10.00 g sample in water, oxidize it to arsenate, and dilute it with water to a final volume of 500 mL. You then add excess $AgNO_3$ solution to a 50.0 mL sample of the arsenate solution. The resulting precipitate of Ag_3AsO_4 has a mass of 3.24 g after drying. What is the percentage by mass of $NaAsO_2$ in the original sample?

Answer: 91.0%

Note

Precipitation reactions are a subclass of double displacement reactions.



Determining the Net Ionic Equation for a Precipitation Reaction: https://youtu.be/AMJz1Sdz8IA

Summary

A complete ionic equation consists of the net ionic equation and spectator ions. Predicting the solubility of ionic compounds in water can give insight into whether or not a reaction will occur. The chemical equation for a reaction in solution can be written in three ways. The **overall chemical equation** shows all the substances present in their undissociated forms; the **complete ionic equation** shows all the substances present in the form in which they actually exist in solution; and the **net ionic equation** is derived from the complete ionic equation by omitting all **spectator ions**, ions that

occur on both sides of the equation with the same coefficients. Net ionic equations demonstrate that many different combinations of reactants can give the same net chemical reaction.

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when solutions of two substances are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)

10.6: Precipitation Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **5.2: Precipitation Reactions** is licensed CC BY-NC-SA 4.0.

10.7: Acid-Base Reactions

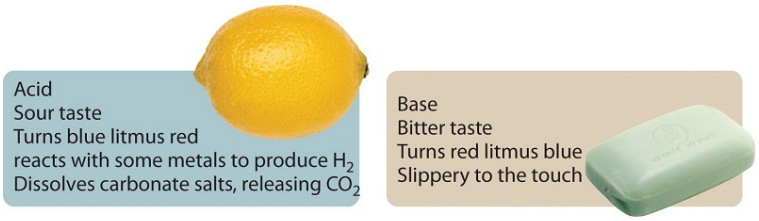
Learning Objectives

- To know the characteristic properties of acids and bases.

Acid–base reactions are essential in both biochemistry and industrial chemistry. Moreover, many of the substances we encounter in our homes, the supermarket, and the pharmacy are acids or bases. For example, aspirin is an acid (acetylsalicylic acid), and antacids are bases. In fact, every amateur chef who has prepared mayonnaise or squeezed a wedge of lemon to marinate a piece of fish has carried out an acid–base reaction. Before we discuss the characteristics of such reactions, let's first describe some of the properties of acids and bases.

Definitions of Acids and Bases

We defined **acids** as substances that dissolve in water to produce H^+ ions, whereas **bases** were defined as substances that dissolve in water to produce OH^- ions. In fact, this is only one possible set of definitions. Although the general properties of acids and bases have been known for more than a thousand years, the definitions of *acid* and *base* have changed dramatically as scientists have learned more about them. In ancient times, an acid was any substance that had a sour taste (e.g., vinegar or lemon juice), caused consistent color changes in dyes derived from plants (e.g., turning blue litmus paper red), reacted with certain metals to produce hydrogen gas and a solution of a salt containing a metal cation, and dissolved carbonate salts such as limestone ($CaCO_3$) with the evolution of carbon dioxide. In contrast, a base was any substance that had a bitter taste, felt slippery to the touch, and caused color changes in plant dyes that differed diametrically from the changes caused by acids (e.g., turning red litmus paper blue). Although these definitions were useful, they were entirely descriptive.



Although we include their tastes among the common characteristics of acids and bases, we never advocate tasting an unknown chemical!

The Arrhenius Definition of Acids and Bases

The first person to define acids and bases in detail was the Swedish chemist Svante Arrhenius (1859–1927; Nobel Prize in Chemistry, 1903). According to the *Arrhenius definition*, an acid is a substance like hydrochloric acid that dissolves in water to produce H^+ ions (protons; Equation 5.3.1), and a base is a substance like sodium hydroxide that dissolves in water to produce hydroxide (OH^-) ions (Equation 5.3.2):

$$\underset{an\ Arrhenius\ acid}{HCl_{(g)}} \xrightarrow{H_2O_{(l)}} H^+_{(aq)} + Cl^-_{(aq)} \tag{10.7.1}$$

$$\underset{an\ Arrhenius\ base}{NaOH_{(s)}} \xrightarrow{H_2O_{(l)}} Na^+_{(aq)} + OH^-_{(aq)} \tag{10.7.2}$$

According to Arrhenius, the characteristic properties of acids and bases are due exclusively to the presence of H^+ and OH^- ions, respectively, in solution. Although Arrhenius's ideas were widely accepted, his definition of acids and bases had two major limitations:

1. First, because acids and bases were defined in terms of ions obtained from water, the Arrhenius concept applied only to substances in aqueous solution.
2. Second, and more important, the Arrhenius definition predicted that *only* substances that dissolve in water to produce H^+ and OH^- ions should exhibit the properties of acids and bases, respectively. For example, according to the Arrhenius definition,

the reaction of ammonia (a base) with gaseous HCl (an acid) to give ammonium chloride (Equation 10.7.3) is not an acid–base reaction because it does not involve H^+ and OH^-:

$$NH_{3\ (g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)} \tag{10.7.3}$$

The Brønsted–Lowry Definition of Acids and Bases

Because of the limitations of the Arrhenius definition, a more general definition of acids and bases was needed. One was proposed independently in 1923 by the Danish chemist J. N. Brønsted (1879–1947) and the British chemist T. M. Lowry (1874–1936), who defined acid–base reactions in terms of the transfer of a proton (H^+ ion) from one substance to another.

According to Brønsted and Lowry, an acid (A substance with at least one hydrogen atom that can dissociate to form an anion and an H^+ ion (a proton) in aqueous solution, thereby forming an acidic solution) is any substance that can donate a proton, and a base (a substance that produces one or more hydroxide ions (OH^- and a cation when dissolved in aqueous solution, thereby forming a basic solution) is any substance that can accept a proton. The Brønsted–Lowry definition of an acid is essentially the same as the Arrhenius definition, except that it is not restricted to aqueous solutions. The Brønsted–Lowry definition of a base, however, is far more general because the hydroxide ion is just one of many substances that can accept a proton. Ammonia, for example, reacts with a proton to form NH_4^+, so in Equation 10.7.3, NH_3 is a Brønsted–Lowry base and HCl is a Brønsted–Lowry acid. Because of its more general nature, the Brønsted–Lowry definition is used throughout this text unless otherwise specified.



Definitions of Acids and Bases: https://youtu.be/r8reN0CSIHw

Polyprotic Acids

Acids differ in the number of protons they can donate. For example, monoprotic acids (a compound that is capable of donating one proton per molecule) are compounds that are capable of donating a single proton per molecule. Monoprotic acids include HF, HCl, HBr, HI, HNO_3, and HNO_2. All carboxylic acids that contain a single $-CO_2H$ group, such as acetic acid (CH_3CO_2H), are monoprotic acids, dissociating to form RCO_2^- and H^+. A compound that can donate more than one proton per molecule. can donate more than one proton per molecule. For example, H_2SO_4 can donate two H^+ ions in separate steps, so it is a diprotic acid (a compound that can donate two protons per molecule in separate steps) and H_3PO_4, which is capable of donating three protons in successive steps, is a triprotic acid (a compound that can donate three protons per molecule in separate steps), (Equations 10.7.4, 10.7.5, and 10.7.6):

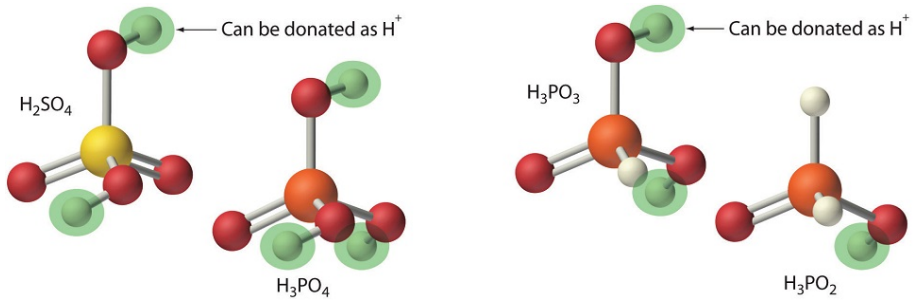
$$H_3PO_4(l) \overset{H_2O(l)}{\rightleftharpoons} H^+(aq) + H_2PO_4^-(aq) \tag{10.7.4}$$

$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq) \tag{10.7.5}$$

$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq) \tag{10.7.6}$$

In chemical equations such as these, a double arrow is used to indicate that both the forward and reverse reactions occur simultaneously, so the forward reaction does not go to completion. Instead, the solution contains significant amounts of both reactants and products. Over time, the reaction reaches a state in which the concentration of each species in solution remains

constant. The reaction is then said to be in equilibrium (the point at which the rates of the forward and reverse reactions become the same, so that the net composition of the system no longer changes with time).



Strengths of Acids and Bases

We will not discuss the strengths of acids and bases quantitatively until next semester. Qualitatively, however, we can state that strong acids (An acid that reacts essentially completely with water) to give H^+ and the corresponding anion. react essentially completely with water to give H^+ and the corresponding anion. Similarly, strong bases (A base that dissociates essentially completely in water) to give OH^- and the corresponding cation) dissociate essentially completely in water to give OH^- and the corresponding cation. Strong acids and strong bases are both strong electrolytes. In contrast, only a fraction of the molecules of weak acids (An acid in which only a fraction of the molecules react with water) to producee H^+ and the corresponding anion. and weak bases (A base in which only a fraction of the molecules react with water to produce OH^- and the corresponding cation) react with water to produce ions, so weak acids and weak bases are also weak electrolytes. Typically less than 5% of a weak electrolyte dissociates into ions in solution, whereas more than 95% is present in undissociated form.

In practice, only a few strong acids are commonly encountered: HCl, HBr, HI, HNO_3, $HClO_4$, and H_2SO_4 (H_3PO_4 is only moderately strong). The most common strong bases are ionic compounds that contain the hydroxide ion as the anion; three examples are NaOH, KOH, and $Ca(OH)_2$. Common weak acids include HCN, H_2S, HF, oxoacids such as HNO_2 and HClO, and carboxylic acids such as acetic acid. The ionization reaction of acetic acid is as follows:

$$CH_3CO_2H(l) \overset{H_2O(l)}{\rightleftharpoons} H^+(aq) + CH_3CO_2^-(aq) \tag{10.7.7}$$

Although acetic acid is *very* soluble in water, almost all of the acetic acid in solution exists in the form of neutral molecules (less than 1% dissociates), as we stated in section 8.1. Sulfuric acid is unusual in that it is a strong acid when it donates its first proton (Equation 5.3.8) but a weak acid when it donates its second proton (Equation 5.3.9) as indicated by the single and double arrows, respectively:

$$\underset{strong\ acid}{H_2SO_4(l)} \xrightarrow{H_2O(l)} H^+(aq) + HSO_4^-(aq) \tag{10.7.8}$$

$$\underset{weak\ acid}{HSO_4^-(aq)} \rightleftharpoons H^+(aq) + SO_4^{2-}(aq) \tag{10.7.9}$$

Consequently, an aqueous solution of sulfuric acid contains $H^+_{(aq)}$ ions and a mixture of $HSO^-_{4\ (aq)}$ and $SO^{2-}_{4\ (aq)}$ ions, but no H_2SO_4 molecules.

The most common weak base is ammonia, which reacts with water to form small amounts of hydroxide ion:

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \tag{10.7.10}$$

Most of the ammonia (>99%) is present in the form of NH_3(g). Amines, which are organic analogues of ammonia, are also weak bases, as are ionic compounds that contain anions derived from weak acids (such as S^{2-}).

> *There is no correlation between the solubility of a substance and whether it is a strong electrolyte, a weak electrolyte, or a nonelectrolyte.*

Table 10.7.1 lists some common strong acids and bases. Acids other than the six common strong acids are almost invariably weak acids. The only common strong bases are the hydroxides of the alkali metals and the heavier alkaline earths (Ca, Sr, and Ba); any

other bases you encounter are most likely weak. Remember that *there is no correlation between solubility and whether a substance is a strong or a weak electrolyte!* Many weak acids and bases are extremely soluble in water.

Table 10.7.1*: Common Strong Acids and Bases*

Strong Acids		Strong Bases	
Hydrogen Halides	Oxoacids	Group 1 Hydroxides	Hydroxides of the Heavier Group 2 Elements
HCl	HNO_3	LiOH	$Ca(OH)_2$
HBr	H_2SO_4	NaOH	$Sr(OH)_2$
HI	$HClO_4$	KOH	$Ba(OH)_2$
		RbOH	
		CsOH	

Example 10.7.1

Classify each compound as a strong acid, a weak acid, a strong base, a weak base, or none of these.

a. $CH_3CH_2CO_2H$
b. CH_3OH
c. $Sr(OH)_2$
d. $CH_3CH_2NH_2$
e. $HBrO_4$

Given: compound

Asked for: acid or base strength

Strategy:

A Determine whether the compound is organic or inorganic.

B If inorganic, determine whether the compound is acidic or basic by the presence of dissociable H^+ or OH^- ions, respectively. If organic, identify the compound as a weak base or a weak acid by the presence of an amine or a carboxylic acid group, respectively. Recall that all polyprotic acids except H_2SO_4 are weak acids.

Solution:

a. **A** This compound is propionic acid, which is organic. **B** It contains a carboxylic acid group analogous to that in acetic acid, so it must be a weak acid.
b. **A** CH_3OH is methanol, an organic compound that contains the −OH group. **B** As a covalent compound, it does not dissociate to form the OH^- ion. Because it does not contain a carboxylic acid ($-CO_2H$) group, methanol also cannot dissociate to form $H^+(aq)$ ions. Thus we predict that in aqueous solution methanol is neither an acid nor a base.
c. **A** $Sr(OH)_2$ is an inorganic compound that contains one Sr^{2+} and two OH^- ions per formula unit. **B** We therefore expect it to be a strong base, similar to $Ca(OH)_2$.
d. **A** $CH_3CH_2NH_2$ is an amine (ethylamine), an organic compound in which one hydrogen of ammonia has been replaced by an R group. **B** Consequently, we expect it to behave similarly to ammonia (Equation 5.3.7), reacting with water to produce small amounts of the OH^- ion. Ethylamine is therefore a weak base.
e. **A** $HBrO_4$ is perbromic acid, an inorganic compound. **B** It is not listed in Table 10.7.1 as one of the common strong acids, but that does not necessarily mean that it is a weak acid. If you examine the periodic table, you can see that Br lies directly below Cl in group 17. We might therefore expect that $HBrO_4$ is chemically similar to $HClO_4$, a strong acid—and, in fact, it is.

? Exercise 10.7.1

Classify each compound as a strong acid, a weak acid, a strong base, a weak base, or none of these.

a. $Ba(OH)_2$
b. HIO_4
c. $CH_3CH_2CH_2CO_2H$
d. $(CH_3)_2NH$
e. CH_2O

Answer

a. strong base
b. strong acid
c. weak acid
d. weak base
e. none of these; formaldehyde is a neutral molecule



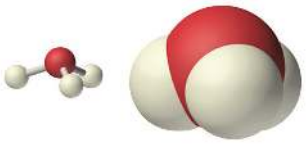
Definition of Strong/Weak Acids & Bases: https://youtu.be/gN2l8H_AWGU

The Hydronium Ion

Because isolated protons are *very* unstable and hence very reactive, an acid never simply "loses" an H^+ ion. Instead, the proton is always transferred to another substance, which acts as a base in the Brønsted–Lowry definition. Thus in every acid–base reaction, one species acts as an acid and one species acts as a base. Occasionally, the same substance performs both roles, as you will see later. When a strong acid dissolves in water, the proton that is released is transferred to a water molecule that acts as a proton acceptor or base, as shown for the dissociation of sulfuric acid:

$$\underset{acid\ (proton\ donor)}{H_2SO_4(l)} + \underset{base\ (proton\ acceptor)}{H_2O(l)} \rightarrow \underset{acid}{H_3O^+}(aq) + \underset{base}{HSO_4^-}(aq) \tag{10.7.11}$$

Technically, therefore, it is imprecise to describe the dissociation of a strong acid as producing $H^+_{(aq)}$ ions, as we have been doing. The resulting H_3O^+ ion, called the hydronium ion and is a more accurate representation of $H^+_{(aq)}$. For the sake of brevity, however, in discussing acid dissociation reactions, we often show the product as $H^+_{(aq)}$ (as in Equation 5.3.7) with the understanding that the product is actually the$H_3O^+_{(aq)}$ ion.



Hydronium ion, H_3O^+

Conversely, bases that do not contain the hydroxide ion accept a proton from water, so small amounts of OH^- are produced, as in the following:

$$\underset{base}{NH_3(g)} + \underset{acid}{H_2O(l)} \rightleftharpoons \underset{acid}{NH_4^+(aq)} + \underset{base}{OH^-(aq)} \tag{10.7.12}$$

Again, the double arrow indicates that the reaction does not go to completion but rather reaches a state of equilibrium. In this reaction, water acts as an acid by donating a proton to ammonia, and ammonia acts as a base by accepting a proton from water. Thus water can act as either an acid or a base by donating a proton to a base or by accepting a proton from an acid. Substances that can behave as both an acid and a base are said to be amphotericWhen substances can behave as both an acid and a base..

The products of an acid–base reaction are also an acid and a base. In Equation 10.7.11, for example, the products of the reaction are the hydronium ion, here an acid, and the hydrogen sulfate ion, here a weak base. In Equation 10.7.12, the products are NH_4^+, an acid, and OH^-, a base. The product NH_4^+ is called the conjugate acid of the base NH_3, and the product OH^- is called the conjugate baseThe substance formed when a Brønsted–Lowry acid donates a proton. of the acid H_2O. Thus all acid–base reactions actually involve two conjugate acid–base pairs. All acid–base reactions involve two conjugate acid–base pairs, the Brønsted–Lowry acid and the base it forms after donating its proton, and the Brønsted–Lowry base and the acid it forms after accepting a proton.; in Equation 10.7.12, they are NH_4^+/NH_3 and H_2O/OH^-.

Neutralization Reactions

A neutralization reaction (a chemical reaction in which an acid and a base react in stoichiometric amounts to produce water and a salt) is one in which an acid and a base react in stoichiometric amounts to produce water and a salt (the general term for any ionic substance that does not have OH− as the anion or H+ as the cation), the general term for any ionic substance that does not have OH^- as the anion or H^+ as the cation. If the base is a metal hydroxide, then the general formula for the reaction of an acid with a base is described as follows: *Acid plus base yields water plus salt*. For example, the reaction of equimolar amounts of HBr and NaOH to give water and a salt (NaBr) is a neutralization reaction:

$$\underset{acid}{HBr(aq)} + \underset{base}{NaOH(aq)} \rightarrow \underset{water}{H_2O(l)} + \underset{salt}{NaBr(aq)} \tag{10.7.13}$$

> *Acid plus base yields water plus salt.*

If we write the complete ionic equation for the reaction in Equation 10.7.13, we see that $Na^+_{(aq)}$ and $Br^-_{(aq)}$ are spectator ions and are not involved in the reaction:

$$H^+(aq) + \cancel{Br^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{Na^+(aq)} + \cancel{Br^-(aq)} \tag{10.7.14}$$

The overall reaction is therefore simply the combination of H^+(aq) and OH^-(aq) to produce H_2O, as shown in the net ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \tag{10.7.15}$$

The net ionic equation for the reaction of any strong acid with any strong base is identical to Equation 10.7.15.

The strengths of the acid and the base generally determine whether the reaction goes to completion. The reaction of *any* strong acid with *any* strong base goes essentially to completion, as does the reaction of a strong acid with a weak base, and a weak acid with a strong base. Examples of the last two are as follows:

$$\underset{strong\ acid}{HCl(aq)} + \underset{weak\ base}{NH_3(aq)} \rightarrow \underset{salt}{NH_4Cl(aq)} \tag{10.7.16}$$

$$\underset{weak\ acid}{CH_3CO_2H(aq)} + \underset{strong\ base}{NaOH(aq)} \rightarrow \underset{salt}{CH_3CO_2Na(aq)} + H_2O(l) \tag{10.7.17}$$

Sodium acetate is written with the organic component first followed by the cation, as is usual for organic salts. Most reactions of a weak acid with a weak base also go essentially to completion. One example is the reaction of acetic acid with ammonia:

$$\underset{weak\ acid}{CH_3CO_2H(aq)} + \underset{weak\ base}{NH_3(aq)} \rightarrow \underset{salt}{CH_3CO_2NH_4(aq)} \tag{10.7.18}$$

An example of an acid–base reaction that does not go to completion is the reaction of a weak acid or a weak base with water, which is both an extremely weak acid and an extremely weak base.

Note

Except for the reaction of a weak acid or a weak base with water, acid–base reactions essentially go to completion.

In some cases, the reaction of an acid with an anion derived from a weak acid (such as HS^-) produces a gas (in this case, H_2S). Because the gaseous product escapes from solution in the form of bubbles, the reverse reaction cannot occur. Therefore, these reactions tend to be forced, or driven, to completion. Examples include reactions in which an acid is added to ionic compounds that contain the HCO_3^-, CN^-, or S^{2-} anions, all of which are driven to completion:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq) \tag{10.7.19}$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l) \tag{10.7.20}$$

$$CN^-(aq) + H^+(aq) \rightarrow HCN(g) \tag{10.7.21}$$

$$S^{2-}(aq) + H^+(aq) \rightarrow HS^-(aq) \tag{10.7.22}$$

$$HS^-(aq) + H^+(aq) \rightarrow H_2S(g) \tag{10.7.23}$$

The reactions in Equations 10.7.19–10.7.22 are responsible for the rotten egg smell that is produced when metal sulfides come in contact with acids.



Acid/Base Neutralization Reactions & Net Ionic Equations: https://youtu.be/gDS93ySeF80

Example 10.7.2: Calcium Propionate

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$] to give calcium propionate. Do you expect this reaction to go to completion, making it a feasible method for the preparation of calcium propionate?

Given: reactants and product

Asked for: balanced chemical equation and whether the reaction will go to completion

Strategy

Write the balanced chemical equation for the reaction of propionic acid with calcium hydroxide. Based on their acid and base strengths, predict whether the reaction will go to completion.

Solution

Propionic acid is an organic compound that is a weak acid, and calcium hydroxide is an inorganic compound that is a strong base. The balanced chemical equation is as follows:

$$2CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + 2H_2O(l)$$

The reaction of a weak acid and a strong base will go to completion, so it is reasonable to prepare calcium propionate by mixing solutions of propionic acid and calcium hydroxide in a 2:1 mole ratio.

? Exercise 10.7.2

Write a balanced chemical equation for the reaction of solid sodium acetate with dilute sulfuric acid to give sodium sulfate.

Answer:

$$2CH_3CO_2Na(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2CH_3CO_2H(aq)$$

One of the most familiar and most heavily advertised applications of acid–base chemistry is *antacids*, which are bases that neutralize stomach acid. The human stomach contains an approximately 0.1 M solution of hydrochloric acid that helps digest foods. If the protective lining of the stomach breaks down, this acid can attack the stomach tissue, resulting in the formation of an *ulcer*. Because one factor that is believed to contribute to the formation of stomach ulcers is the production of excess acid in the stomach, many individuals routinely consume large quantities of antacids. The active ingredients in antacids include sodium bicarbonate and potassium bicarbonate ($NaHCO_3$ and $KHCO_3$; Alka-Seltzer); a mixture of magnesium hydroxide and aluminum hydroxide [$Mg(OH)_2$ and $Al(OH)_3$; Maalox, Mylanta]; calcium carbonate ($CaCO_3$; Tums); and a complex salt, dihydroxyaluminum sodium carbonate [$NaAl(OH)_2CO_3$; original Rolaids]. Each has certain advantages and disadvantages. For example, $Mg(OH)_2$ is a powerful laxative (it is the active ingredient in milk of magnesia), whereas $Al(OH)_3$ causes constipation. When mixed, each tends to counteract the unwanted effects of the other. Although all antacids contain both an anionic base (OH^-, CO_3^{2-}, or HCO_3^-) and an appropriate cation, they differ substantially in the amount of active ingredient in a given mass of product.



Stomach acid. An antacid tablet reacts with 0.1 M HCl (the approximate concentration found in the human stomach).

✓ Example 10.7.3

Assume that the stomach of someone suffering from acid indigestion contains 75 mL of 0.20 M HCl. How many Tums tablets are required to neutralize 90% of the stomach acid, if each tablet contains 500 mg of $CaCO_3$? (Neutralizing all of the stomach acid is not desirable because that would completely shut down digestion.)

Given: volume and molarity of acid and mass of base in an antacid tablet

Asked for: number of tablets required for 90% neutralization

Strategy:

A. Write the balanced chemical equation for the reaction and then decide whether the reaction will go to completion.
B. Calculate the number of moles of acid present. Multiply the number of moles by the percentage to obtain the quantity of acid that must be neutralized. Using mole ratios, calculate the number of moles of base required to neutralize the acid.
C. Calculate the number of moles of base contained in one tablet by dividing the mass of base by the corresponding molar mass. Calculate the number of tablets required by dividing the moles of base by the moles contained in one tablet.

Solution:

A We first write the balanced chemical equation for the reaction:

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2CO_3(aq)$$

Each carbonate ion can react with 2 mol of H^+ to produce H_2CO_3, which rapidly decomposes to H_2O and CO_2. Because HCl is a strong acid and CO_3^{2-} is a weak base, the reaction will go to completion.

B Next we need to determine the number of moles of HCl present:

$$75\ \cancel{mL}\left(\frac{1\ \cancel{L}}{1000\ \cancel{mL}}\right)\left(\frac{0.20\ mol\ HCl}{\cancel{L}}\right) = 0.015\ mol\ HCl$$

Because we want to neutralize only 90% of the acid present, we multiply the number of moles of HCl by 0.90:

$$(0.015\ mol\ HCl)(0.90) = 0.014\ mol\ HCl$$

We know from the stoichiometry of the reaction that each mole of $CaCO_3$ reacts with 2 mol of HCl, so we need

$$moles\ CaCO_3 = 0.014\ \cancel{mol\ HCl}\left(\frac{1\ mol\ CaCO_3}{2\ \cancel{mol\ HCl}}\right) = 0.0070\ mol\ CaCO_3$$

C Each Tums tablet contains

$$\left(\frac{500\ \cancel{mg\ CaCO_3}}{1\ Tums\ tablet}\right)\left(\frac{1\ \cancel{g}}{1000\ \cancel{mg\ CaCO_3}}\right)\left(\frac{1\ mol\ CaCO_3}{100.1\ \cancel{g}}\right) = 0.00500\ mol\ CaCO_3$$

Thus we need $\dfrac{0.0070\ \cancel{mol\ CaCO_3}}{0.00500\ \cancel{mol\ CaCO_3}} = 1.4$ Tums tablets.

? Exercise 10.7.3

Assume that as a result of overeating, a person's stomach contains 300 mL of 0.25 M HCl. How many Rolaids tablets must be consumed to neutralize 95% of the acid, if each tablet contains 400 mg of $NaAl(OH)_2CO_3$? The neutralization reaction can be written as follows:

$$NaAl(OH)_2CO_3(s) + 4HCl(aq) \rightarrow AlCl_3(aq) + NaCl(aq) + CO_2(g) + 3H_2O(l)$$

Answer: 6.4 tablets

The pH Scale

One of the key factors affecting reactions that occur in dilute solutions of acids and bases is the concentration of H^+ and OH^- ions. The pH scale provides a convenient way of expressing the hydrogen ion (H^+) concentration of a solution and enables us to describe acidity or basicity in quantitative terms.

Pure liquid water contains extremely low, but measurable concentrations of H_3O^+(aq) and OH^-(aq) ions produced via an *autoionization reaction*, in which water acts simultaneously as an acid and as a base:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \tag{10.7.24}$$

The concentration of hydrogen ions in pure water is only 1.0×10^{-7} M at 25°C. Because the autoionization reaction produces both a proton and a hydroxide ion, the OH^- concentration in pure water is also 1.0×10^{-7} M. Pure water is a neutral solution in which the total positive charge from all the cations is matched by an identical total negative charge from all the anions. At room temperature this is $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ M.

The pH scale describes the hydrogen ion concentration of a solution in a way that avoids the use of exponential notation. The negative base-10 logarithm of the hydrogen ion concentration: pH is often defined as the negative base-10 logarithm of the hydrogen ion concentration; pH is **actually** defined as the negative base-10 logarithm of hydrogen ion *activity*. As you will learn in a more advanced course, the activity of a substance in solution is related to its concentration. For dilute solutions such as those we are discussing, the activity and the concentration are approximately the same.

$$pH = -log[H^+] \tag{10.7.25}$$

Conversely,

$$[H^+] = 10^{-pH} \tag{10.7.26}$$

(If you are not familiar with logarithms or using a calculator to obtain logarithms and antilogarithms)

Because the hydrogen ion concentration is 1.0×10^{-7} M in pure water at 25°C, the pH of pure liquid water (and, by extension, of any neutral solution) is

$$pH = -\log[1.0 \times 10^{-7}] = 7.00 \tag{10.7.27}$$

Adding an acid to pure water increases the hydrogen ion concentration and decreases the hydroxide ion concentration because a neutralization reaction occurs, such as that shown in Equation 5.3.15. Because the negative exponent of $[H^+]$ becomes smaller as $[H^+]$ increases, the pH *decreases* with increasing $[H^+]$. For example, a 1.0 M solution of a strong monoprotic acid such as HCl or HNO_3 has a pH of 0.00:

$$pH = -\log[1.0] = 0.00 \tag{10.7.28}$$

Note

pH decreases with increasing $[H^+]$.

Conversely, adding a base to pure water increases the hydroxide ion concentration and decreases the hydrogen ion concentration. Because the autoionization reaction of water does not go to completion, neither does the neutralization reaction. Even a strongly basic solution contains a detectable amount of H^+ ions. For example, a 1.0 M OH^- solution has $[H^+] = 1.0 \times 10^{-14}$ M. The pH of a 1.0 M NaOH solution is therefore

$$pH = -log[1.0 \times 10^{-14}] = 14.00 \tag{10.7.29}$$

For practical purposes, the pH scale runs from pH = 0 (corresponding to 1 M H^+) to pH 14 (corresponding to 1 M OH^-), although pH values less than 0 or greater than 14 are possible.

We can summarize the relationships between acidity, basicity, and pH as follows:

- If pH = 7.0, the solution is neutral.
- If pH < 7.0, the solution is acidic.
- If pH > 7.0, the solution is basic.

Keep in mind that the pH scale is logarithmic, so a change of 1.0 in the pH of a solution corresponds to a *tenfold* change in the hydrogen ion concentration. The foods and consumer products we encounter daily represent a wide range of pH values, as shown in Figure 10.7.2

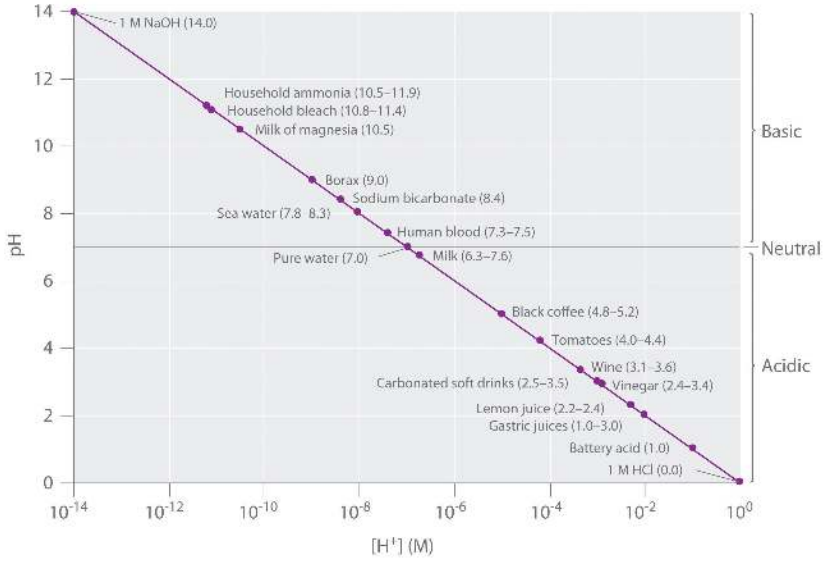


Figure 10.7.2 A Plot of pH versus $[H^+]$ for Some Common Aqueous Solutions. Although many substances exist in a range of pH values (indicated in parentheses), they are plotted using typical values.

Example 10.7.4

a. What is the pH of a 2.1×10^{-2} M aqueous solution of $HClO_4$?
b. The pH of a vinegar sample is 3.80. What is its hydrogen ion concentration?

Given: molarity of acid or pH

Asked for: pH or $[H^+]$

Strategy:

Using the balanced chemical equation for the acid dissociation reaction and Equation 5.3.25 or 5.3.24, determine $[H^+]$ and convert it to pH or vice versa.

Solution:

a. $HClO_4$ (perchloric acid) is a strong acid, so it dissociates completely into H^+ ions and ClO_4^- ions:

$$HClO_4(l) \rightarrow H^+(aq) + ClO_4^-(aq)$$

The H^+ ion concentration is therefore the same as the perchloric acid concentration. The pH of the perchloric acid solution is thus

$$pH = -log[H^+] = -log(2.1 \times 10^{-2}) = 1.68$$

The result makes sense: the H^+ ion concentration is between 10^{-1} M and 10^{-2} M, so the pH must be between 1 and 2.

Note: The assumption that $[H^+]$ is the same as the concentration of the acid is valid for *only* strong acids. Because weak acids do *not* dissociate completely in aqueous solution, a more complex procedure is needed to calculate the pH of their solutions.

b. We are given the pH and asked to calculate the hydrogen ion concentration. From Equation 5.3.24,

$$10^{-pH} = [H^+]$$

Thus $[H^+] = 10^{-3.80} = 1.6 \times 10^{-4}\ M$.

Exercise 10.7.4

a. What is the pH of a 3.0×10^{-5} M aqueous solution of HNO_3?
b. What is the hydrogen ion concentration of turnip juice, which has a pH of 5.41?

Answer

a. $pH = 4.52$
b. $[H^+] = 3.9 \times 10^{-6}\ M$

Tools have been developed that make the measurement of pH simple and convenient (Figures 5.3.3 and 5.3.4). For example, pH paper consists of strips of paper impregnated with one or more acid–base indicators, which are intensely colored organic molecules whose colors change dramatically depending on the pH of the solution. Placing a drop of a solution on a strip of pH paper and comparing its color with standards give the solution's approximate pH. A more accurate tool, the pH meter, uses a glass electrode, a device whose voltage depends on the H^+ ion concentration (Figure 10.7.4).



Figure 10.7.3: These pH strips can measure pH in a series of ranges by putting sample on a strip and comparing its color change with colors on the box that correspond to a certain pH. The strip on the left measures pH 0-7 and shows results of a strong acid sample; the center strip is pH range 5-10 and shows results of a 6.97 buffer solution sample; the strip on the right measures a broad range (pH 1-14) and shows results of a 10% bleach water solution sample. This image can be enlarged by clicking on it. Photo by Monica Bruckner (courtesy of the Science Education Resource Center (SERC) at Carleton College).



Figure 10.7.4: These three meters and probes can measure pH (left) and electrical conductivity (center and right). The image can be enlarged if you click on it. Photo by Monica Bruckner (courtesy of the Science Education Resource Center (SERC) at Carleton College).



Introduction to pH: https://youtu.be/pQOa3bb5YEE

Summary

An acidic solution and a basic solution react together in a neutralization reaction that also forms a salt. Acid–base reactions require both an acid and a base. In Brønsted–Lowry terms, an **acid** is a substance that can donate a proton (H^+), and a **base** is a substance that can accept a proton. All acid–base reactions contain two acid–base pairs: the reactants and the products. Acids can donate one proton (**monoprotic acids**), two protons (**diprotic acids**), or three protons (**triprotic acids**). Compounds that are capable of donating more than one proton are generally called **polyprotic acids**. Acids also differ in their tendency to donate a proton, a measure of their acid strength. **Strong acids** react completely with water to produce $H_3O^+(aq)$ (the **hydronium ion**), whereas **weak acids** dissociate only partially in water. Conversely, **strong bases** react completely with water to produce the hydroxide ion,

whereas **weak bases** react only partially with water to form hydroxide ions. The reaction of a strong acid with a strong base is a **neutralization reaction**, which produces water plus a **salt**.

The acidity or basicity of an aqueous solution is described quantitatively using the **pH scale**. The **pH** of a solution is the negative logarithm of the H^+ ion concentration and typically ranges from 0 for strongly acidic solutions to 14 for strongly basic ones. Because of the *autoionization reaction* of water, which produces small amounts of hydronium ions and hydroxide ions, a **neutral solution** of water contains 1×10^{-7} M H^+ ions and has a pH of 7.0. An **indicator** is an intensely colored organic substance whose color is pH dependent; it is used to determine the pH of a solution.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)

10.7: Acid-Base Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **5.3: Acid-Base Reactions** is licensed CC BY-NC-SA 4.0.

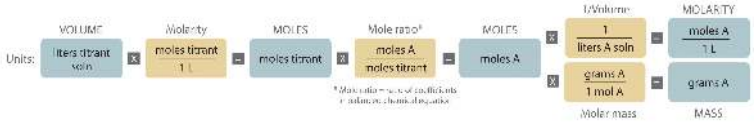
10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations

The Learning Objective of this section is to use titration methods to analyze solutions quantitatively.

To determine the amounts or concentrations of substances present in a sample, chemists use a combination of chemical reactions and stoichiometric calculations in a methodology called quantitative analysis (A methodology that combines chemical reactions and stoichiometric calculations to determine the amounts or concentrations of substances present in a sample). Suppose, for example, we know the identity of a certain compound in a solution but not its concentration. If the compound reacts rapidly and completely with another reactant, we may be able to use the reaction to determine the concentration of the compound of interest. A titration is an experimental procedure in which a carefully measured volume of a solution of known concentration is added to a measured volume of a solution containing a compound whose concentration is to be determined., a carefully measured volume of a solution of known concentration, called the titrant. The solution of known concentration that is reacted with a compound in a solution of unknown concentration in a titration., is added to a measured volume of a solution containing a compound whose concentration is to be determined (the *unknown*). The reaction used in a titration can be an acid–base reaction, a precipitation reaction, or an oxidation–reduction reaction. In all cases, the reaction chosen for the analysis must be fast, complete, and *specific*; that is, *only* the compound of interest should react with the titrant. The equivalence point (The point in a titration where a stoichiometric amount (i.e., the amount required to react completely with the unknown) of the titrant has been added) is reached when a stoichiometric amount of the titrant has been added—the amount required to react completely with the unknown.

Determining the Concentration of an Unknown Solution Using a Titration

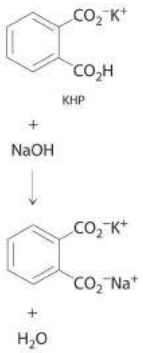
The chemical nature of the species present in the unknown dictates which type of reaction is most appropriate and also how to determine the equivalence point. The volume of titrant added, its concentration, and the coefficients from the balanced chemical equation for the reaction allow us to calculate the total number of moles of the unknown in the original solution. Because we have measured the volume of the solution that contains the unknown, we can calculate the molarity of the unknown substance. This procedure is summarized graphically here:



Graphical representation of the process used for determining molarity of an unknown substance.

Standard Solutions

The accuracy of any titration analysis depends on an accurate knowledge of the concentration of the titrant. Most titrants are first *standardized*; that is, their concentration is measured by titration with a standard solution, which is a solution whose concentration is known precisely. Only pure crystalline compounds that do not react with water or carbon dioxide are suitable for use in preparing a standard solution. One such compound is potassium hydrogen phthalate (KHP), a weak monoprotic acid suitable for standardizing solutions of bases such as sodium hydroxide. The reaction of KHP with NaOH is a simple acid–base reaction. If the concentration of the KHP solution is known accurately and the titration of a NaOH solution with the KHP solution is carried out carefully, then the concentration of the NaOH solution can be calculated precisely. The standardized NaOH solution can then be used to titrate a solution of an acid whose concentration is unknown.



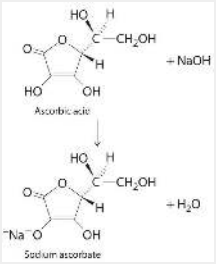
The reaction of KHP with NaOH. As with all acid-base reactions, a salt is formed.

Acid–Base Titrations

Because most common acids and bases are not intensely colored, a small amount of an acid–base **indicator** is usually added to detect the equivalence point in an acid–base titration. The point in the titration at which an indicator changes color is called the endpoint (the point in a titration at which an indicator changes color). The procedure is illustrated in Example 21.1.

Example 12.1: Vitamin C

The structure of vitamin C (ascorbic acid, a monoprotic acid) is as follows:



Ascorbic acid. The upper figure shows the three-dimensional representation of ascorbic acid. Hatched lines indicate bonds that are behind the plane of the paper, and wedged lines indicate bonds that are out of the plane of the paper.

An absence of vitamin C in the diet leads to the disease known as scurvy, a breakdown of connective tissue throughout the body and of dentin in the teeth. Because fresh fruits and vegetables rich in vitamin C are readily available in developed countries today, scurvy is not a major problem. In the days of slow voyages in wooden ships, however, scurvy was common. Ferdinand Magellan, the first person to sail around the world, lost more than 90% of his crew, many to scurvy. Although a diet rich in fruits and vegetables contains more than enough vitamin C to prevent scurvy, many people take supplemental doses of vitamin C, hoping that the extra amounts will help prevent colds and other illness.

Suppose a tablet advertised as containing 500 mg of vitamin C is dissolved in 100.0 mL of distilled water that contains a small amount of the acid–base indicator bromothymol blue, an indicator that is yellow in acid solution and blue in basic solution, to give a yellow solution. The addition of 53.5 mL of a 0.0520 M solution of NaOH results in a change to green at the endpoint, due to a mixture of the blue and yellow forms of the indicator (**Figure 8.9.3**). What is the actual mass of vitamin C in the tablet? (The molar mass of ascorbic acid is 176.13 g/mol.)

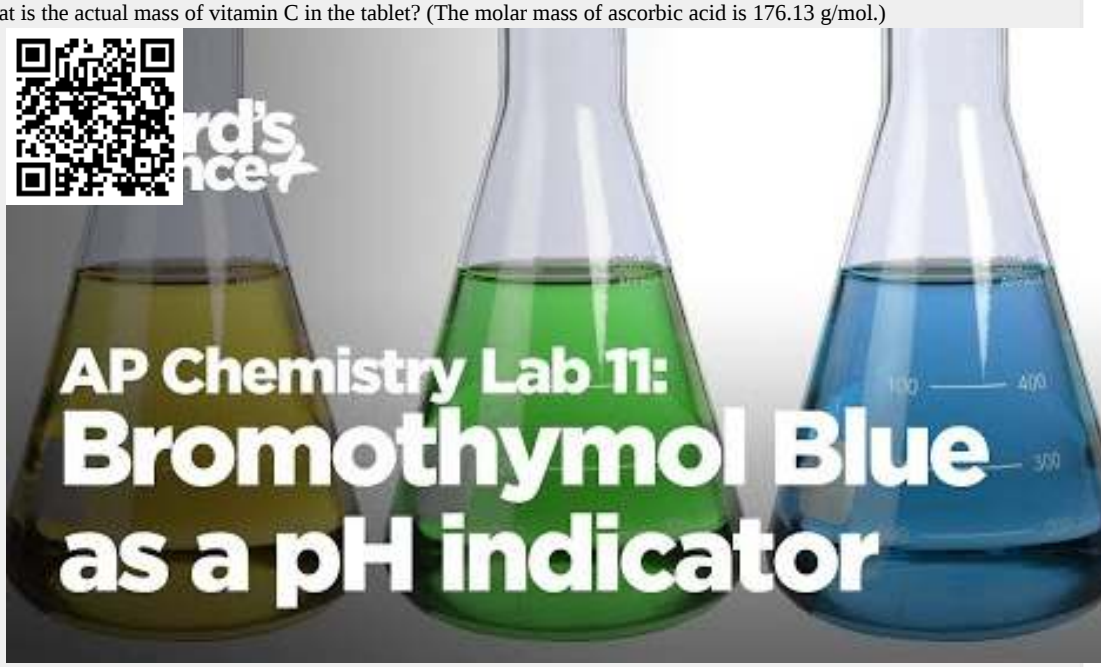


Figure 8.9.3: Bromothymol Blue is used as an indicator. If placed in a acidic solution it will turn the solution yellow, and if placed in a basic solution it will turn the solution blue. The equivalence point is seen when the solution turns green. pH is important for a variety of reasons. For example if you are trying to discard an acidic solution, you could add bromothymol blue to the solution and then titrate in some base until the solution turned green. This indicates that it is neutral

Given: reactant, volume of sample solution, and volume and molarity of titrant
Asked for: mass of unknown
Strategy:
A Write the balanced chemical equation for the reaction and calculate the number of moles of base needed to neutralize the ascorbic acid.
B Using mole ratios, determine the amount of ascorbic acid consumed. Calculate the mass of vitamin C by multiplying the number of moles of ascorbic acid by its molar mass.
Solution:
A Because ascorbic acid acts as a monoprotic acid, we can write the balanced chemical equation for the reaction as

$$HAsc(aq) + OH^-(aq) \rightarrow Asc^-(aq) + H_2O(l) \tag{10.8.1}$$

where HAsc is ascorbic acid and Asc^- is ascorbate. The number of moles of OH^- ions needed to neutralize the ascorbic acid is

$$moles\ OH^- = 53.5\ \cancel{mL}\left(\frac{\cancel{1\ L}}{1000\ \cancel{mL}}\right)\left(\frac{0.0520\ mol\ OH^-}{\cancel{1\ L}}\right) = 2.78 \times 10^{-3}\ mol\ OH^- \tag{10.8.2}$$

B The mole ratio of the base added to the acid consumed is 1:1, so the number of moles of OH^- added equals the number of moles of ascorbic acid present in the tablet:

$$mass\ ascorbic\ acid = 2.78 \times 10^{-3}\ \cancel{mol\ HAsc}\left(\frac{176.13\ g\ HAsc}{1\ \cancel{mol\ HAsc}}\right) = 0.490\ g\ HAsc \tag{10.8.3}$$

Because 0.490 g equals 490 mg, the tablet contains about 2% less vitamin C than advertised.

Exercise 12.1: Vinegar

Vinegar is essentially a dilute solution of acetic acid in water. Vinegar is usually produced in a concentrated form and then diluted with water to give a final concentration of 4%–7% acetic acid; that is, a 4% m/v solution contains 4.00 g of acetic acid per 100 mL of solution. If a drop of bromothymol blue indicator is added to 50.0 mL of concentrated vinegar stock and 31.0 mL of 2.51 M NaOH are needed to turn the solution from yellow to green, what is the percentage of acetic acid in the vinegar stock? (Assume that the density of the vinegar solution is 1.00 g/mL.)
Answer: 9.35%

Summary

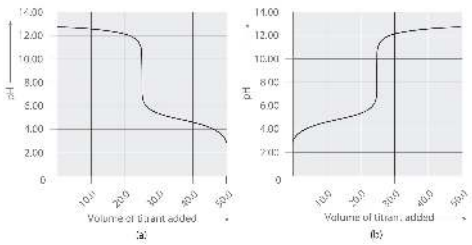
The concentration of a species in solution can be determined by **quantitative analysis**. One such method is a **titration**, in which a measured volume of a solution of one substance, the **titrant**, is added to a solution of another substance to determine its concentration. The **equivalence point** in a titration is the point at which exactly enough reactant has been added for the reaction to go to completion. A **standard solution**, a solution whose concentration is known precisely, is used to determine the concentration of the titrant. Many titrations, especially those that involve acid–base reactions, rely on an indicator. The point at which a color change is observed is the **endpoint**, which is close to the equivalence point if the indicator is chosen properly.

Key Takeaway

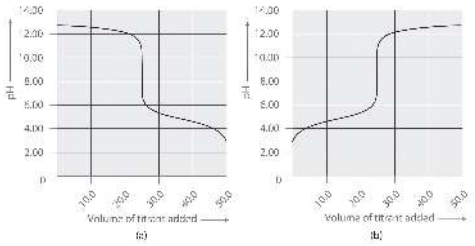
- Quantitative analysis of an unknown solution can be achieved using titration methods.

Conceptual Problems

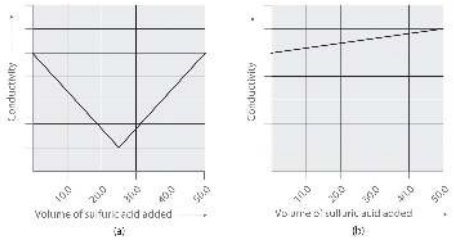
1. The titration procedure is an application of the use of limiting reactants. Explain why this is so.
2. Explain how to determine the concentration of a substance using a titration.
3. Following are two graphs that illustrate how the pH of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M acetic acid with 0.10 M NaOH, and the other corresponds to the titration of 100 mL 0.10 M NaOH with 0.10 M acetic acid. Which graph corresponds to which titration? Justify your answer.



4. Following are two graphs that illustrate how the pH of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M ammonia with 0.10 M HCl, and the other corresponds to the titration of 100 mL 0.10 M NH_4Cl with 0.10 M NaOH. Which graph corresponds to which titration? Justify your answer.



5. Following are two graphs that illustrate how the electrical conductivity of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M $Ba(OH)_2$ with 0.10 M H_2SO_4 , and the other corresponds to the titration of 100 mL of 0.10 M NaOH with 0.10 M H_2SO_4. Which graph corresponds to which titration? Justify your answer.



Answers

1.
2.
3. Figure a) corresponds to titration of NaOH with acetic acid, figure b) corresponds to titration of acetic acid with NaOH.
4. Figure a) corresponds to titration of ammonia with HCl, figure b) corresponds to titration of ammonium chloride with NaOH.
5. Figure a) corresponds to titration of $Ba(OH)_2$ with sulfuric acid, figure b) corresponds to titration of NaOH with sulfuric acid.

Numerical Problems

1. A 10.00 mL sample of a 1.07 M solution of potassium hydrogen phthalate (KHP, formula mass = 204.22 g/mol) is diluted to 250.0 mL. What is the molarity of the final solution? How many grams of KHP are in the 10.00 mL sample?
2. What volume of a 0.978 M solution of NaOH must be added to 25.0 mL of 0.583 M HCl to completely neutralize the acid? How many moles of NaOH are needed for the neutralization?
3. A student was titrating 25.00 mL of a basic solution with an HCl solution that was 0.281 M. The student ran out of the HCl solution after having added 32.46 mL, so she borrowed an HCl solution that was labeled as 0.317 M. An additional 11.5 mL of the second solution was needed to complete the titration. What was the concentration of the basic solution?

Contributors

- Anonymous

Modified by Joshua Halpern (Howard University)

10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.9: Oxidation-Reduction- Some General Principles

Learning Objectives

- To identify oxidation–reduction reactions in solution.

The term oxidation was first used to describe reactions in which metals react with oxygen in air to produce metal oxides. When iron is exposed to air in the presence of water, for example, the iron turns to rust—an iron oxide. When exposed to air, aluminum metal develops a continuous, coherent, transparent layer of aluminum oxide on its surface. In both cases, the metal acquires a positive charge by transferring electrons to the neutral oxygen atoms of an oxygen molecule. As a result, the oxygen atoms acquire a negative charge and form oxide ions (O^{2-}). Because the metals have lost electrons to oxygen, they have been oxidized; oxidation is therefore the loss of electrons. Conversely, because the oxygen atoms have gained electrons, they have been reduced, so reduction is the gain of electrons. For every oxidation, there must be an associated reduction.

Note

Any oxidation must **ALWAYS** be accompanied by a reduction and vice versa.

Originally, the term reduction referred to the decrease in mass observed when a metal oxide was heated with carbon monoxide, a reaction that was widely used to extract metals from their ores. When solid copper(I) oxide is heated with hydrogen, for example, its mass decreases because the formation of pure copper is accompanied by the loss of oxygen atoms as a volatile product (water). The reaction is as follows:

$$Cu_2O(s) + H_2(g) \rightarrow 2Cu(s) + H_2O(g) \tag{10.9.1}$$

Oxidation and reduction reactions are now defined as reactions that exhibit a change in the oxidation states of one or more elements in the reactants, which follows the mnemonic oxidation is loss reduction is gain, or oil rig. The oxidation state of each atom in a compound is the charge an atom would have if all its bonding electrons were transferred to the atom with the greater attraction for electrons. Atoms in their elemental form, such as O_2 or H_2, are assigned an oxidation state of zero. For example, the reaction of aluminum with oxygen to produce aluminum oxide is

$$4Al(s) + 3O_2 \rightarrow 2Al_2O_3(s) \tag{10.9.2}$$

Each neutral oxygen atom gains two electrons and becomes negatively charged, forming an oxide ion; thus, oxygen has an oxidation state of −2 in the product and has been reduced. Each neutral aluminum atom loses three electrons to produce an aluminum ion with an oxidation state of +3 in the product, so aluminum has been oxidized. In the formation of Al_2O_3, electrons are transferred as follows (the superscript 0 emphasizes the oxidation state of the elements):

$$4Al^0 + 3O_2^0 \rightarrow 4Al^{3+} + 6O^{2-} \tag{10.9.3}$$

Equation 5.4.1 and Equation 5.4.2 are examples of oxidation–reduction (redox) reactions. In redox reactions, there is a net transfer of electrons from one reactant to another. In any redox reaction, the total number of electrons lost must equal the total of electrons gained to preserve electrical neutrality. In Equation 5.4.3, for example, the total number of electrons lost by aluminum is equal to the total number gained by oxygen:

$$electrons\,lost = 4\,Al\,atoms \times \frac{3\,e^-\,lost}{Al\,atom} = 12\,e^-\,lost \tag{10.9.4}$$

$$electrons\,gained = 6\,O\,atoms \times \frac{2\,e^-\,gained}{O\,atom} = 12\,e^-\,gained \tag{10.9.5}$$

The same pattern is seen in all oxidation–reduction reactions: the number of electrons lost must equal the number of electrons gained.

An additional example of a redox reaction, the reaction of sodium metal with oxygen in air, is illustrated in Figure 10.9.1.

Note

In all oxidation–reduction (redox) reactions, the number of electrons lost equals the number of electrons gained.



Redox Reactions: https://youtu.be/1v3yaaR_nHc

Assigning Oxidation States

Assigning oxidation states to the elements in binary ionic compounds is straightforward: the oxidation states of the elements are identical to the charges on the monatomic ions. Previosuly, you learned how to predict the formulas of simple ionic compounds based on the sign and magnitude of the charge on monatomic ions formed by the neutral elements. Examples of such compounds are sodium chloride (NaCl; Figure 10.9.1), magnesium oxide (MgO), and calcium chloride ($CaCl_2$). In covalent compounds, in contrast, atoms share electrons. Oxidation states in covalent compounds are somewhat arbitrary, but they are useful bookkeeping devices to help you understand and predict many reactions.

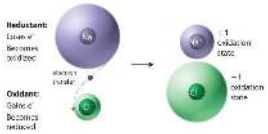


Figure 10.9.1: The Reaction of a Neutral Sodium Atom with a Neutral Chlorine Atom. The result is the transfer of one electron from sodium to chlorine, forming the ionic compound NaCl.

A set of rules for assigning oxidation states to atoms in chemical compounds follows.

Rules for Assigning Oxidation States

1. The oxidation state of an atom in any pure element, whether monatomic, diatomic, or polyatomic, is zero.
2. The oxidation state of a monatomic ion is the same as its charge—for example, $Na^+ = +1$, $Cl^- = -1$.
3. The oxidation state of fluorine in chemical compounds is always −1. Other halogens usually have oxidation states of −1 as well, except when combined with oxygen or other halogens.
4. Hydrogen is assigned an oxidation state of +1 in its compounds with nonmetals and −1 in its compounds with metals.
5. Oxygen is normally assigned an oxidation state of −2 in compounds, with two exceptions: in compounds that contain oxygen–fluorine or oxygen–oxygen bonds, the oxidation state of oxygen is determined by the oxidation states of the other elements present.
6. The sum of the oxidation states of all the atoms in a neutral molecule or ion must equal the charge on the molecule or ion.

Nonintegral oxidation states are encountered occasionally. They are usually due to the presence of two or more atoms of the same element with different oxidation states.

In any chemical reaction, the net charge must be conserved; that is, in a chemical reaction, the total number of electrons is constant, just like the total number of atoms. Consistent with this, rule 1 states that the sum of the individual oxidation states of the atoms in a molecule or ion must equal the net charge on that molecule or ion. In NaCl, for example, Na has an oxidation state of +1 and Cl is −1. The net charge is zero, as it must be for any compound.

Rule 3 is required because fluorine attracts electrons more strongly than any other element, for reasons you will discover in Chapter 6. Hence fluorine provides a reference for calculating the oxidation states of other atoms in chemical compounds. Rule 4 reflects the difference in chemistry observed for compounds of hydrogen with nonmetals (such as chlorine) as opposed to compounds of hydrogen with metals (such as sodium). For example, NaH contains the H^- ion, whereas HCl forms H^+ and Cl^- ions when dissolved in water. Rule 5 is necessary because fluorine has a greater attraction for electrons than oxygen does; this rule also prevents violations of rule 2. So the oxidation state of oxygen is +2 in OF_2 but −½ in KO_2. Note that an oxidation state of −½ for O in KO_2 is perfectly acceptable.

The reduction of copper(I) oxide shown in Equation 5.4.5 demonstrates how to apply these rules. Rule 1 states that atoms in their elemental form have an oxidation state of zero, which applies to H_2 and Cu. From rule 4, hydrogen in H_2O has an oxidation state of +1, and from rule 5, oxygen in both Cu_2O and H_2O has an oxidation state of −2. Rule 6 states that the sum of the oxidation states in a molecule or formula unit must equal the net charge on that compound. This means that each Cu atom in Cu_2O must have a charge of +1: 2(+1) + (−2) = 0. So the oxidation states are as follows:

$$\overset{+1}{Cu}_2\overset{-2}{O}(s) + \overset{0}{H}_2(g) \rightarrow 2\overset{0}{Cu}(s) + \overset{+1}{H}_2\overset{-2}{O}(g) \tag{10.9.6}$$

Assigning oxidation states allows us to see that there has been a net transfer of electrons from hydrogen (0 → +1) to copper (+1 → 0). So this is a redox reaction. Once again, the number of electrons lost equals the number of electrons gained, and there is a net conservation of charge:

$$electrons\,lost = 2\,H\,atoms \times \frac{1\,e^-\,lost}{H\,atom} = 2\,e^-\,lost \tag{10.9.7}$$

$$electrons\,gained = 2\,Cu\,atoms \times \frac{1\,e^-\,gained}{Cu\,atom} = 2\,e^-\,gained \tag{10.9.8}$$

Remember that oxidation states are useful for visualizing the transfer of electrons in oxidation–reduction reactions, but the oxidation state of an atom and its actual charge are the same only for simple ionic compounds. Oxidation states are a convenient way of assigning electrons to atoms, and they are useful for predicting the types of reactions that substances undergo.

Example 10.9.1

Assign oxidation states to all atoms in each compound.

a. sulfur hexafluoride (SF_6)
b. methanol (CH_3OH)
c. ammonium sulfate [$(NH_4)2SO_4$]
d. magnetite (Fe_3O_4)
e. ethanoic (acetic) acid (CH_3CO_2H)

Given: molecular or empirical formula

Asked for: oxidation states

Strategy:

Begin with atoms whose oxidation states can be determined unambiguously from the rules presented (such as fluorine, other halogens, oxygen, and monatomic ions). Then determine the oxidation states of other atoms present according to rule 1.

Solution:

a. We know from rule 3 that fluorine always has an oxidation state of −1 in its compounds. The six fluorine atoms in sulfur hexafluoride give a total negative charge of −6. Because rule 1 requires that the sum of the oxidation states of all atoms be zero in a neutral molecule (here SF_6), the oxidation state of sulfur must be +6:

[(6 F atoms)(−1)] + [(1 S atom) (+6)] = 0

b. According to rules 4 and 5, hydrogen and oxygen have oxidation states of +1 and −2, respectively. Because methanol has no net charge, carbon must have an oxidation state of −2:

[(4 H atoms)(+1)] + [(1 O atom)(−2)] + [(1 C atom)(−2)] = 0

c. Note that $(NH_4)_2SO_4$ is an ionic compound that consists of both a polyatomic cation (NH_4^+) and a polyatomic anion ($SO_4^{2−}$) (see Table 2.4). We assign oxidation states to the atoms in each polyatomic ion separately. For NH_4^+, hydrogen has an oxidation state of +1 (rule 4), so nitrogen must have an oxidation state of −3:

$$[(4\text{ H atoms})(+1)] + [(1\text{ N atom})(−3)] = +1, \text{ the charge on the } NH_4^+ \text{ ion}$$

For SO42−, oxygen has an oxidation state of −2 (rule 5), so sulfur must have an oxidation state of +6:

$$[(4\text{ O atoms})\ (−2)] + [(1\text{ S atom})(+6)] = −2, \text{ the charge on the sulfate ion}$$

d. Oxygen has an oxidation state of −2 (rule 5), giving an overall charge of −8 per formula unit. This must be balanced by the positive charge on three iron atoms, giving an oxidation state of +8/3 for iron:

$$[(4\text{ O atoms})(−2)]+[(3\text{ Fe atoms})\left(+\tfrac{8}{3}\right)]= 0$$

Fractional oxidation states are allowed because oxidation states are a somewhat arbitrary way of keeping track of electrons. In fact, Fe_3O_4 can be viewed as having two Fe^{3+} ions and one Fe^{2+} ion per formula unit, giving a net positive charge of +8 per formula unit. Fe_3O_4 is a magnetic iron ore commonly called magnetite. In ancient times, magnetite was known as lodestone because it could be used to make primitive compasses that pointed toward Polaris (the North Star), which was called the "lodestar."

e. Initially, we assign oxidation states to the components of CH_3CO_2H in the same way as any other compound. Hydrogen and oxygen have oxidation states of +1 and −2 (rules 4 and 5, respectively), resulting in a total charge for hydrogen and oxygen of

$$[(4\text{ H atoms})(+1)] + [(2\text{ O atoms})(−2)] = 0$$

So the oxidation state of carbon must also be zero (rule 6). This is, however, an average oxidation state for the two carbon atoms present. Because each carbon atom has a different set of atoms bonded to it, they are likely to have different oxidation states. To determine the oxidation states of the individual carbon atoms, we use the same rules as before but with the additional assumption that bonds between atoms of the same element do not affect the oxidation states of those atoms. The carbon atom of the methyl group ($−CH_3$) is bonded to three hydrogen atoms and one carbon atom. We know from rule 4 that hydrogen has an oxidation state of +1, and we have just said that the carbon–carbon bond can be ignored in calculating the oxidation state of the carbon atom. For the methyl group to be electrically neutral, its carbon atom must have an oxidation state of −3. Similarly, the carbon atom of the carboxylic acid group ($−CO_2H$) is bonded to one carbon atom and two oxygen atoms. Again ignoring the bonded carbon atom, we assign oxidation states of −2 and +1 to the oxygen and hydrogen atoms, respectively, leading to a net charge of

$$[(2\text{ O atoms})(−2)] + [(1\text{ H atom})(+1)] = −3$$

To obtain an electrically neutral carboxylic acid group, the charge on this carbon must be +3. The oxidation states of the individual atoms in acetic acid are thus

$$\overset{-3}{C}\overset{+1}{H_3}\overset{+3}{C}\overset{-2}{O_2}\overset{+1}{H} \tag{10.9.9}$$

Thus the sum of the oxidation states of the two carbon atoms is indeed zero.

? Exercise 10.9.1

Assign oxidation states to all atoms in each compound.

a. barium fluoride (BaF_2)
b. formaldehyde (CH_2O)
c. potassium dichromate ($K_2Cr_2O_7$)
d. cesium oxide (CsO_2)
e. ethanol (CH_3CH_2OH)

Answer:

a. Ba, +2; F, −1
b. C, 0; H, +1; O, −2
c. K, +1; Cr, +6; O, −2

d. Cs, +1; O, −½
e. C, −3; H, +1; C, −1; H, +1; O, −2; H, +1

Redox Reactions of Solid Metals in Aqueous Solution

A widely encountered class of oxidation–reduction reactions is the reaction of aqueous solutions of acids or metal salts with solid metals. An example is the corrosion of metal objects, such as the rusting of an automobile (Figure 10.9.1). Rust is formed from a complex oxidation–reduction reaction involving dilute acid solutions that contain Cl^- ions (effectively, dilute HCl), iron metal, and oxygen. When an object rusts, iron metal reacts with HCl(aq) to produce iron(II) chloride and hydrogen gas:

$$\boxed{Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)} \quad (5.4.5)$$

In subsequent steps, $FeCl_2$ undergoes oxidation to form a reddish-brown precipitate of $Fe(OH)_3$.



Figure 10.9.2: Rust Formation. The corrosion process involves an oxidation–reduction reaction in which metallic iron is converted to $Fe(OH)_3$, a reddish-brown solid.

Many metals dissolve through reactions of this type, which have the general form

$$metal + acid \rightarrow salt + hydrogen \quad (10.9.10)$$

Some of these reactions have important consequences. For example, it has been proposed that one factor that contributed to the fall of the Roman Empire was the widespread use of lead in cooking utensils and pipes that carried water. Rainwater, as we have seen, is slightly acidic, and foods such as fruits, wine, and vinegar contain organic acids. In the presence of these acids, lead dissolves:

$$Pb(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + H_2(g) \quad (10.9.11)$$

Consequently, it has been speculated that both the water and the food consumed by Romans contained toxic levels of lead, which resulted in widespread lead poisoning and eventual madness. Perhaps this explains why the Roman Emperor Caligula appointed his favorite horse as consul!

Single-Displacement Reactions

Certain metals are oxidized by aqueous acid, whereas others are oxidized by aqueous solutions of various metal salts. Both types of reactions are called single-displacement reactions, in which the ion in solution is displaced through oxidation of the metal. Two examples of single-displacement reactions are the reduction of iron salts by zinc (Equation 5.4.8) and the reduction of silver salts by copper (Equation 5.4.9 and Figure 10.9.2):

$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s) \quad (10.9.12)$$

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s) \quad (10.9.13)$$

The reaction in Equation 5.4.8 is widely used to prevent (or at least postpone) the corrosion of iron or steel objects, such as nails and sheet metal. The process of “galvanizing” consists of applying a thin coating of zinc to the iron or steel, thus protecting it from

oxidation as long as zinc remains on the object.



Figure used with permission of Wikipedia

The Activity Series

By observing what happens when samples of various metals are placed in contact with solutions of other metals, chemists have arranged the metals according to the relative ease or difficulty with which they can be oxidized in a single-displacement reaction. For example, metallic zinc reacts with iron salts, and metallic copper reacts with silver salts. Experimentally, it is found that zinc reacts with both copper salts and silver salts, producing Zn^{2+}. Zinc therefore has a greater tendency to be oxidized than does iron, copper, or silver. Although zinc will not react with magnesium salts to give magnesium metal, magnesium metal will react with zinc salts to give zinc metal:

$$Zn(s) + Mg^{2+}(aq) \not\rightarrow Zn^{2+}(aq) + Mg(s) \tag{10.9.14}$$

$$Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s) \tag{10.9.15}$$

Magnesium has a greater tendency to be oxidized than zinc does.

Pairwise reactions of this sort are the basis of the activity series (Figure 10.9.4), which lists metals and hydrogen in order of their relative tendency to be oxidized. The metals at the top of the series, which have the greatest tendency to lose electrons, are the alkali metals (group 1), the alkaline earth metals (group 2), and Al (group 13). In contrast, the metals at the bottom of the series, which have the lowest tendency to be oxidized, are the precious metals or coinage metals—platinum, gold, silver, and copper, and mercury, which are located in the lower right portion of the metals in the periodic table. You should be generally familiar with which kinds of metals are active metals, which have the greatest tendency to be oxidized. (located at the top of the series) and which are inert metals, which have the least tendency to be oxidized. (at the bottom of the series).

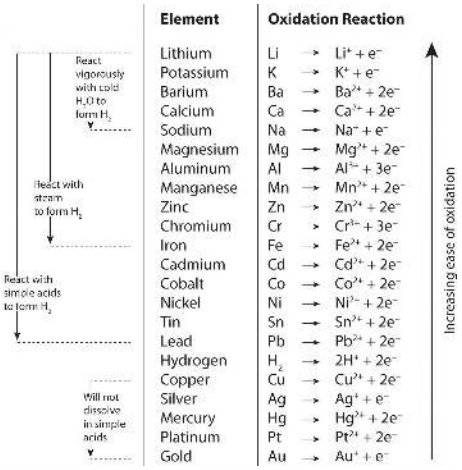


Figure 10.9.4: The Activity Series

When using the activity series to predict the outcome of a reaction, keep in mind that *any element will reduce compounds of the elements below it in the series*. Because magnesium is above zinc in Figure 10.9.4, magnesium metal will reduce zinc salts but not vice versa. Similarly, the precious metals are at the bottom of the activity series, so virtually any other metal will reduce precious metal salts to the pure precious metals. Hydrogen is included in the series, and the tendency of a metal to react with an acid is indicated by its position relative to hydrogen in the activity series. *Only those metals that lie above hydrogen in the activity series dissolve in acids to produce H_2.* Because the precious metals lie below hydrogen, they do not dissolve in dilute acid and therefore

do not corrode readily. Example 10.9.2 demonstrates how a familiarity with the activity series allows you to predict the products of many single-displacement reactions.

Example 10.9.2

Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation.

1. A strip of aluminum foil is placed in an aqueous solution of silver nitrate.
2. A few drops of liquid mercury are added to an aqueous solution of lead(II) acetate.
3. Some sulfuric acid from a car battery is accidentally spilled on the lead cable terminals.

Given: reactants

Asked for: overall reaction and net ionic equation

Strategy:

A. Locate the reactants in the activity series in Figure .5.4.4 and from their relative positions, predict whether a reaction will occur. If a reaction does occur, identify which metal is oxidized and which is reduced.
B. Write the net ionic equation for the redox reaction.

Solution:

1. **A** Aluminum is an active metal that lies above silver in the activity series, so we expect a reaction to occur. According to their relative positions, aluminum will be oxidized and dissolve, and silver ions will be reduced to silver metal. **B** The net ionic equation is as follows:

$$Al(s) + 3Ag^{+}(aq) \rightarrow Al^{3+}(aq) + 3Ag(s) \tag{10.9.16}$$

   Recall from our discussion of solubilities that most nitrate salts are soluble. In this case, the nitrate ions are spectator ions and are not involved in the reaction.

2. **A** Mercury lies below lead in the activity series, so no reaction will occur.
3. **A** Lead is above hydrogen in the activity series, so the lead terminals will be oxidized, and the acid will be reduced to form H_2. **B** From our discussion of solubilities, recall that Pb^{2+} and SO_4^{2-} form insoluble lead(II) sulfate. In this case, the sulfate ions are *not* spectator ions, and the reaction is as follows:

$$Pb(s) + 2H^{+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + H_2(g) \tag{10.9.17}$$

   Lead(II) sulfate is the white solid that forms on corroded battery terminals.



Corroded battery terminals. *The white solid is lead(II) sulfate, formed from the reaction of solid lead with a solution of sulfuric acid.*

Exercise 10.9.2

Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation.

1. A strip of chromium metal is placed in an aqueous solution of aluminum chloride.
2. A strip of zinc is placed in an aqueous solution of chromium(III) nitrate.
3. A piece of aluminum foil is dropped into a glass that contains vinegar (the active ingredient is acetic acid).

Answer

1. $no\ reaction$
2. $3Zn(s) + 2Cr^{3+}(aq) \rightarrow 3Zn^{2+}(aq) + 2Cr(s)$
3. $2Al(s) + 6CH_3CO_2H(aq) \rightarrow 2Al^{3+}(aq) + 6CH_3CO_2^-(aq) + 3H_2(g)$

Summary

- Oxidation–reduction reactions are balanced by separating the overall chemical equation into an oxidation equation and a reduction equation.

In oxidation–reduction reactions, electrons are transferred from one substance or atom to another. We can balance oxidation–reduction reactions in solution using the **oxidation state method** (Table 10.9.1), in which the overall reaction is separated into an oxidation equation and a reduction equation. **Single-displacement reactions** are reactions of metals with either acids or another metal salt that result in dissolution of the first metal and precipitation of a second (or evolution of hydrogen gas). The outcome of these reactions can be predicted using the **activity series** (Figure 10.9.4), which arranges metals and H_2 in decreasing order of their tendency to be oxidized. Any metal will reduce metal ions below it in the activity series. **Active metals** lie at the top of the activity series, whereas **inert metals** are at the bottom of the activity series.

10.9: Oxidation-Reduction- Some General Principles is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **5.4: Oxidation-Reduction: Some General Principles** is licensed CC BY-NC-SA 4.0.

10.10: Balancing Oxidation-Reduction Equations

The Learning Objectives of this Module are:

1. To identify oxidation–reduction reactions in solution.

We described the defining characteristics of oxidation–reduction, or redox, reactions already. Most of the reactions we considered there were relatively simple, and balancing them was straightforward. When oxidation–reduction reactions occur in aqueous solution, however, the equations are more complex and can be more difficult to balance by inspection. Because a balanced chemical equation is the most important prerequisite for solving any stoichiometry problem, we need a method for balancing oxidation–reduction reactions in aqueous solution that is generally applicable. One such method uses *oxidation states*, and a second is referred to as the *half-reaction* method. We show you how to balance redox equations using oxidation states in this section; the half-reaction method is discussed in the second semester

Balancing Redox Equations Using Oxidation States

To balance a redox equation using the oxidation state method, we conceptually separate the overall reaction into two parts: an oxidation—in which the atoms of one element lose electrons—and a reduction—in which the atoms of one element gain electrons. Consider, for example, the reaction of Cr^{2+}(aq) with manganese dioxide (MnO_2) in the presence of dilute acid. Equation 8.8.1 is the net ionic equation for this reaction before balancing; the oxidation state of each element in each species has been assigned using the procedure described in Section 7.5:

$$\overset{+2}{Cr^{2+}}(aq) + \overset{+4}{\underset{-2}{MnO_2}}(aq) + \overset{+1}{H^+}(aq) \rightarrow \overset{+3}{Cr^{3+}}(aq) + \overset{+2}{Mn^{2+}}(aq) + \overset{+1}{\underset{-2}{H_2O}}(l) \tag{8.8.1}$$

Notice that chromium is oxidized from the +2 to the +3 oxidation state, while manganese is reduced from the +4 to the +2 oxidation state. We can write an equation for this reaction that shows only the atoms that are oxidized and reduced:

$$Cr^{2+} + Mn^{4+} \rightarrow Cr^{3+} + Mn^{2+} \tag{8.8.2}$$

The oxidation can be written as

$$Cr^{2+} \rightarrow Cr^{3+} + e^- \tag{8.8.3}$$

and the reduction as

$$Mn^{4+} + 2e^- \rightarrow Mn^{2+} \tag{8.8.4}$$

For the overall chemical equation to be balanced, the number of electrons lost by the reductant must equal the number gained by the oxidant. We must therefore multiply the oxidation and the reduction equations by appropriate coefficients to give us the same number of electrons in both. In this example, we must multiply the oxidation equation by 2 to give

$$2Cr^{2+} \rightarrow 2Cr^{3+} + 2e^- \tag{8.8.5}$$

Note the Patterns
In a balanced redox reaction, the number of electrons lost by the reductant equals the number of electrons gained by the oxidant.

The number of electrons lost in the oxidation now equals the number of electrons gained in the reduction:

$$\begin{aligned} 2Cr^{2+} &\rightarrow 2Cr^{3+} + 2e^- \\ Mn^{4+} + 2e^- &\rightarrow Mn^{2+} \end{aligned} \tag{8.8.6}$$

We then add the equations for the oxidation and the reduction and cancel the electrons on both sides of the equation, using the actual chemical forms of the reactants and products:

$$2Cr^{2+}\,(aq) \rightarrow 2Cr^{3+}\,(aq) + 2\cancel{e^-}$$

$$2Cr^{2+}\,(aq) + 2\cancel{e^-} \rightarrow Mn^{2+}\,(aq) \tag{8.8.7}$$

$$2Cr^{2+}\,(aq) + 2Cr^{2+}\,(aq) \rightarrow 2Cr^{3+}\,(aq) + Mn^{2+}\,(aq)$$

Although the electrons cancel and the metal atoms are balanced, the total charge on the left side of the equation (+4) does not equal the charge on the right side (+8). Because the reaction is carried out in the presence of aqueous acid, we can add H^+ as necessary to either side of the equation to balance the charge. By the same token, if the reaction were carried out in the presence of aqueous base, we could balance the charge by adding OH^- as necessary to either side of the equation to balance the charges. In this case, adding four H^+ ions to the left side of the equation gives

$$2Cr^{2+}(aq) + MnO_2(s) + 4H^+(aq) \rightarrow 2Cr^{3+}(aq) + Mn^{2+}(aq) \tag{8.8.8}$$

Although the charges are now balanced, we have two oxygen atoms on the left side of the equation and none on the right. We can balance the oxygen atoms without affecting the overall charge balance by adding H_2O as necessary to either side of the equation. Here, we need to add two H_2O molecules to the right side:

$$2Cr^{2+}(aq) + MnO_2(s) + 4H^+(aq) \rightarrow 2Cr^{3+}(aq) + Mn^{2+}(aq) + 2H_2O(l) \tag{8.8.9}$$

Although we did not explicitly balance the hydrogen atoms, we can see by inspection that the overall chemical equation is now balanced. All that remains is to check to make sure that we have not made a mistake. This procedure for balancing reactions is summarized in Table 8.8.1 and illustrated in Example 17.

Table 8.8.1 Procedure for Balancing Oxidation–Reduction Reactions by the Oxidation State Method

1. Write the unbalanced chemical equation for the reaction, showing the reactants and the products.
2. Assign oxidation states to all atoms in the reactants and the products (see Section 3.5) and determine which atoms change oxidation state.
3. Write separate equations for oxidation and reduction, showing (a) the atom(s) that is (are) oxidized and reduced and (b) the number of electrons accepted or donated by each.
4. Multiply the oxidation and reduction equations by appropriate coefficients so that both contain the same number of electrons.
5. Write the oxidation and reduction equations showing the actual chemical forms of the reactants and the products, adjusting the coefficients as necessary to give the numbers of atoms in step 4.
6. Add the two equations and cancel the electrons.
7. Balance the charge by adding H^+ or OH^- ions as necessary for reactions in acidic or basic solution, respectively.
8. Balance the oxygen atoms by adding H_2O molecules to one side of the equation.
9. Check to make sure that the equation is balanced in both atoms and total charges.

Example 17

Arsenic acid (H_3AsO_4) is a highly poisonous substance that was once used as a pesticide. The reaction of elemental zinc with arsenic acid in acidic solution yields arsine (AsH_3, a highly toxic and unstable gas) and Zn^{2+}(aq). Balance the equation for this reaction using oxidation states:

$$H_3AsO_4(aq) + Zn(s) \rightarrow AsH_3(g) + Zn^{2+}(aq)$$

Given: reactants and products in acidic solution

Asked for: balanced chemical equation using oxidation states

Strategy:

Follow the procedure given in Table 8.8.1 for balancing a redox equation using oxidation states. When you are done, be certain to check that the equation is balanced.

Solution:

1. **Write a chemical equation showing the reactants and the products.** Because we are given this information, we can skip this step.

2. **Assign oxidation states using the procedure described in Section 3.5 and determine which atoms change oxidation state.** The oxidation state of arsenic in arsenic acid is +6, and the oxidation state of arsenic in arsine is −3. Conversely, the oxidation state of zinc in elemental zinc is 0, and the oxidation state of zinc in Zn^{2+}(aq) is +2:

$$[\mathrm{H_3\overset{+5}{As}O_4(aq) + \overset{0}{Zn}(s) \rightarrow \underset{-3}{As}H_3(g) + \overset{+2}{Zn^{2+}}(aq)}]$$

3. **Write separate equations for oxidation and reduction.** The arsenic atom in H_3AsO_4 is reduced from the +5 to the −3 oxidation state, which requires the addition of eight electrons:

$$Reduction\text{: } As^{5+} + 8e^- \rightarrow \underset{-3}{As^{3-}}$$

Each zinc atom in elemental zinc is oxidized from 0 to +2, which requires the loss of two electrons per zinc atom:

$$Oxidation\text{: } Zn \rightarrow Zn^{2+} + 2e^-$$

4. **Multiply the oxidation and reduction equations by appropriate coefficients so that both contain the same number of electrons.** The reduction equation has eight electrons, and the oxidation equation has two electrons, so we need to multiply the oxidation equation by 4 to obtain

$$Reduction\ (\times 1)\text{: } As^{5+} + 8e^- \rightarrow As^{3-} \quad (10.10.1)$$
$$Oxidation\ (\times 4)\text{: } 4Zn^0 \rightarrow 4Zn^{2+} + 8e^- \quad (10.10.2)$$

5. **Write the oxidation and reduction equations showing the actual chemical forms of the reactants and the products, adjusting coefficients as necessary to give the numbers of atoms shown in step 4.** Inserting the actual chemical forms of arsenic and zinc and adjusting the coefficients gives

$$Reduction\text{: } H_3AsO_4(aq) + 8e^- \rightarrow AsH_3(g) \quad (10.10.3)$$
$$Oxidation\text{: } 4Zn(s) \rightarrow 4Zn^{2+}(aq) + 8e^- \quad (10.10.4)$$

6. *Add the two equations and cancel the electrons.* The sum of the two equations in step 5 is

$$H_3AsO_4(aq) + 4Zn(s) + \cancel{8e^-} \rightarrow AsH_3(g) + 4Zn^{2+}(aq) + \cancel{8e^-}$$

which then yields

$$H_3AsO_4(aq) + 4Zn(s) \rightarrow AsH_3(g) + 4Zn^{2+}(aq)$$

7. *Balance the charge by adding H^+or OH^-ions as necessary for reactions in acidic or basic solution, respectively.* Because the reaction is carried out in acidic solution, we can add H^+ ions to whichever side of the equation requires them to balance the charge. The overall charge on the left side is zero, and the total charge on the right side is 4 × (+2) = +8. Adding eight H^+ ions to the left side gives a charge of +8 on both sides of the equation:

$$H_3AsO_4(aq) + 4Zn(s) + 8H^+(aq) \rightarrow AsH_3(g) + 4Zn^{2+}(aq)$$

8. *Balance the oxygen atoms by adding H_2O molecules to one side of the equation.* There are 4 O atoms on the left side of the equation. Adding 4 H_2O molecules to the right side balances the O atoms:

$$H_3AsO_4(aq) + 4Zn(s) + 8H^+(aq) \rightarrow AsH_3(g) + 4Zn^{2+}(aq) + 4H_2O(l)$$

Although we have not explicitly balanced H atoms, each side of the equation has 11 H atoms.

9. *Check to make sure that the equation is balanced in both atoms and total charges.* To guard against careless errors, it is important to check that both the total number of atoms of each element and the total charges are the same on both sides of the equation:

$$Atoms\text{: } 1As + 4Zn + 4O + 11H = 1As + 4Zn + 4O + 11H \quad (10.10.5)$$
$$Total\ charge\text{: } 8(+1) = 4(+2) = +8 \quad (10.10.6)$$

The balanced chemical equation for the reaction is therefore:

$$H_3AsO_4(aq) + 4Zn(s) + 8H^+(aq) \rightarrow AsH_3(g) + 4Zn^{2+}(aq) + 4H_2O(l)$$

Exercise

Copper commonly occurs as the sulfide mineral CuS. The first step in extracting copper from CuS is to dissolve the mineral in nitric acid, which oxidizes the sulfide to sulfate and reduces nitric acid to NO. Balance the equation for this reaction using

oxidation states:

$$CuS(s) + H^+(aq) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + NO(g) + SO_4^{2-}(aq)$$

Answer: $3CuS(s) + 8H^+(aq) + 8NO_3^-(aq) \rightarrow 3Cu^{2+}(aq) + 8NO(g) + 3SO_4^{2-}(aq) + 4H_2O(l)$

Reactions in basic solutions are balanced in exactly the same manner. To make sure you understand the procedure, consider Example 18.

Example 18

The commercial solid drain cleaner, Drano, contains a mixture of sodium hydroxide and powdered aluminum. The sodium hydroxide dissolves in standing water to form a strongly basic solution, capable of slowly dissolving organic substances, such as hair, that may be clogging the drain. The aluminum dissolves in the strongly basic solution to produce bubbles of hydrogen gas that agitate the solution to help break up the clogs. The reaction is as follows:

$$Al(s) + H_2O(aq) \rightarrow [Al(OH)_4]^-(aq) + H_2(g)$$

Balance this equation using oxidation states.

Given: reactants and products in a basic solution

Asked for: balanced chemical equation

Strategy:

Follow the procedure given in Table 8.8.1 for balancing a redox reaction using oxidation states. When you are done, be certain to check that the equation is balanced.

Solution:

We will apply the same procedure used in Example 17 but in a more abbreviated form.

1. The equation for the reaction is given, so we can skip this step.
2. The oxidation state of aluminum changes from 0 in metallic Al to +3 in $[Al(OH)_4]^-$. The oxidation state of hydrogen changes from +1 in H_2O to 0 in H_2. Aluminum is oxidized, while hydrogen is reduced:

$$\overset{0}{Al}(s) + \overset{+1}{H}_2O(aq) \rightarrow [\overset{+3}{Al}(OH)_4]^-(aq) + \overset{0}{H}_2(g)$$

3.
$$Reduction\text{: } Al^0 \rightarrow Al^{3+} + 3e^- \quad (10.10.7)$$
$$Oxidation\text{: } H^+ + e^- \rightarrow H^0\ (in\ H_2) \quad (10.10.8)$$

4. Multiply the reduction equation by 3 to obtain an equation with the same number of electrons as the oxidation equation:

$$Reduction\text{: } 3H^+ + 3e^- \rightarrow 3H^0\ (in\ H_2) \quad (10.10.9)$$
$$Oxidation\text{: } Al^0 \rightarrow Al^{3+} + 3e^- \quad (10.10.10)$$

5. Insert the actual chemical forms of the reactants and products, adjusting the coefficients as necessary to obtain the correct numbers of atoms as in step 4. Because a molecule of H_2O contains two protons, in this case, $3H^+$ corresponds to $3/2H_2O$. Similarly, each molecule of hydrogen gas contains two H atoms, so 3H corresponds to $3/2H_2$.

$$Reduction\text{: } \frac{3}{2}H_2O + 3e^- \rightarrow \frac{3}{2}H_2 \quad (10.10.11)$$
$$Oxidation\text{: } Al \rightarrow [Al(OH)_4]^- + 3e^- \quad (10.10.12)$$

6. Adding the equations and canceling the electrons gives

$$Al + \frac{3}{2}H_2O + \cancel{3e^-} \rightarrow [Al(OH)_4]^- + \frac{3}{2}H_2 + \cancel{3e^-}$$
$$Al + \frac{3}{2}H_2O \rightarrow [Al(OH)_4]^- + \frac{3}{2}H_2$$

   To remove the fractional coefficients, multiply both sides of the equation by 2:

$$2Al + 3H_2O \rightarrow 2[Al(OH)_4]^- + 3H_2$$

7. The right side of the equation has a total charge of −2, whereas the left side has a total charge of 0. Because the reaction is carried out in basic solution, we can balance the charge by adding two OH^- ions to the left side:

$$2Al + 2OH^- + 3H_2O \rightarrow 2[Al(OH)_4]^- + 3H_2$$

8. The left side of the equation contains five O atoms, and the right side contains eight O atoms. We can balance the O atoms by adding three H_2O molecules to the left side:

$$2Al + 2OH^- + 6H_2O \rightarrow 2[Al(OH)_4]^- + 3H_2$$

9. Be sure the equation is balanced:

$$Atoms: 2Al + 8O + 14H = 2Al + 8O + 14H \quad (10.10.13)$$
$$Total\ charge: (2)(0) + (2)(-1) + (6)(0) = (2)(-1) + (3)(0) \quad (10.10.14)$$

$$-2 = -2$$

10. The balanced chemical equation is therefore

$$2Al(s) + 2OH^-(aq) + 6H_2O(l) \rightarrow 2[Al(OH)_4]^-(aq) + 3H_2(g)$$

Thus 3 mol of H_2 gas are produced for every 2 mol of Al.

Exercise

The permanganate ion reacts with nitrite ion in basic solution to produce manganese(IV) oxide and nitrate ion. Write a balanced chemical equation for the reaction.

Answer: $2MnO_4^-(aq) + 3NO_2^-(aq) + H_2O(l) \rightarrow 2MnO_2(s) + 3NO_3^-(aq) + 2OH^-(aq)$

As suggested in Example 17 and Example 18, a wide variety of redox reactions are possible in aqueous solutions. The identity of the products obtained from a given set of reactants often depends on both the ratio of oxidant to reductant and whether the reaction is carried out in acidic or basic solution, which is one reason it can be difficult to predict the outcome of a reaction. Because oxidation–reduction reactions in solution are so common and so important, however, chemists have developed two general guidelines for predicting whether a redox reaction will occur and the identity of the products:

1. Compounds of elements in high oxidation states (such as ClO_4^-, NO_3^-, MnO_4^-, $Cr_2O_7^{2-}$, and UF_6) tend to act as *oxidants* and *become reduced* in chemical reactions.
2. Compounds of elements in low oxidation states (such as CH_4, NH_3, H_2S, and HI) tend to act as *reductants* and *become oxidized* in chemical reactions.

Note the Pattern
Species in high oxidation states act as oxidants, whereas species in low oxidation states act as reductants.

When an aqueous solution of a compound that contains an element in a high oxidation state is mixed with an aqueous solution of a compound that contains an element in a low oxidation state, an oxidation–reduction reaction is likely to occur.

Redox Reactions of Solid Metals in Aqueous Solution

A widely encountered class of oxidation–reduction reactions is the reaction of aqueous solutions of acids or metal salts with solid metals. An example is the corrosion of metal objects, such as the rusting of an automobile (Figure 8.8.1). Rust is formed from a complex oxidation–reduction reaction involving dilute acid solutions that contain Cl^- ions (effectively, dilute HCl), iron metal, and oxygen. When an object rusts, iron metal reacts with HCl(aq) to produce iron(II) chloride and hydrogen gas:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g) \quad (8.8.10)$$

In subsequent steps, $FeCl_2$ undergoes oxidation to form a reddish-brown precipitate of $Fe(OH)_3$.



Figure 8.8.1 Rust Formation. *The corrosion process involves an oxidation–reduction reaction in which metallic iron is converted to $Fe(OH)_3$, a reddish-brown solid.*

Many metals dissolve through reactions of this type, which have the general form

$$metal + acid \rightarrow salt + hydrogen \tag{8.8.11}$$

Some of these reactions have important consequences. For example, it has been proposed that one factor that contributed to the fall of the Roman Empire was the widespread use of lead in cooking utensils and pipes that carried water. Rainwater, as we have seen, is slightly acidic, and foods such as fruits, wine, and vinegar contain organic acids. In the presence of these acids, lead dissolves:

$$Pb(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + H_2(g) \tag{8.8.12}$$

Consequently, it has been speculated that both the water and the food consumed by Romans contained toxic levels of lead, which resulted in widespread lead poisoning and eventual madness. Perhaps this explains why the Roman Emperor Caligula appointed his favorite horse as consul!

Single-Displacement Reactions

Certain metals are oxidized by aqueous acid, whereas others are oxidized by aqueous solutions of various metal salts. Both types of reactions are called single-displacement reactions, in which the ion in solution is displaced through oxidation of the metal. Two examples of single-displacement reactions are the reduction of iron salts by zinc (Equation 8.8.13) and the reduction of silver salts by copper (Equation 8.8.14 and Figure 8.8.2):

$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s) \tag{8.8.13}$$

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s) \tag{8.8.14}$$

The reaction in Equation 8.8.13 is widely used to prevent (or at least postpone) the corrosion of iron or steel objects, such as nails and sheet metal. The process of “galvanizing” consists of applying a thin coating of zinc to the iron or steel, thus protecting it from oxidation as long as zinc remains on the object.



Figure 8.8.2 The Single-Displacement Reaction of Metallic Copper with a Solution of Silver Nitrate. *When a copper coil is placed in a solution of silver nitrate, silver ions are reduced to metallic silver on the copper surface, and some of the copper metal*

dissolves. Note the formation of a metallic silver precipitate on the copper coil and a blue color in the surrounding solution due to the presence of aqueous Cu^{2+} ions. Figure used with permission of Wikipedia

Summary

In oxidation–reduction reactions, electrons are transferred from one substance or atom to another. We can balance oxidation–reduction reactions in solution using the **oxidation state method** (Table 8.8.1), in which the overall reaction is separated into an oxidation equation and a reduction equation.

Key Takeaway

- Oxidation–reduction reactions are balanced by separating the overall chemical equation into an oxidation equation and a reduction equation.

Conceptual Problems

1. Which elements in the periodic table tend to be good oxidants? Which tend to be good reductants?
2. If two compounds are mixed, one containing an element that is a poor oxidant and one with an element that is a poor reductant, do you expect a redox reaction to occur? Explain your answer. What do you predict if one is a strong oxidant and the other is a weak reductant? Why?
3. In each redox reaction, determine which species is oxidized and which is reduced:
   a. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
   b. $Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$
   c. $BrO_3^-(aq) + 2MnO_2(s) + H_2O(l) \rightarrow Br^-(aq) + 2MnO_4^-(aq) + 2H^+(aq)$
4. Single-displacement reactions are a subset of redox reactions. In this subset, what is oxidized and what is reduced? Give an example of a redox reaction that is *not* a single-displacement reaction.

Numerical Problems

1. Balance each redox reaction under the conditions indicated.
   a. $CuS(s) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq) + NO(g)$; acidic solution
   b. $Ag(s) + HS^-(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2S(s) + Cr(OH)_3(s)$; basic solution
   c. $Zn(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + H_2(g)$; acidic solution
   d. $O_2(g) + Sb(s) \rightarrow H_2O_2(aq) + SbO_2^-(aq)$; basic solution
   e. $UO_2^{2+}(aq) + Te(s) \rightarrow U^{4+}(aq) + TeO_4^{2-}(aq)$; acidic solution
2. Balance each redox reaction under the conditions indicated.
   a. $MnO_4^-(aq) + S_2O_3^{2-}(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$; acidic solution
   b. $Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) \rightarrow Fe^{3+}(aq) + Cr^{3+}(aq)$; acidic solution
   c. $Fe(s) + CrO_4^{2-}(aq) \rightarrow Fe_2O_3(s) + Cr_2O_3(s)$; basic solution
   d. $Cl_2(aq) \rightarrow ClO_3^-(aq) + Cl^-(aq)$; acidic solution
   e. $CO_3^{2-}(aq) + N_2H_4(aq) \rightarrow CO(g) + N_2(g)$; basic solution
3. Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation; then write the complete ionic equation for the reaction.
   a. Platinum wire is dipped in hydrochloric acid.
   b. Manganese metal is added to a solution of iron(II) chloride.
   c. Tin is heated with steam.
   d. Hydrogen gas is bubbled through a solution of lead(II) nitrate.
4. Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation; then write the complete ionic equation for the reaction.
   a. A few drops of $NiBr_2$ are dropped onto a piece of iron.
   b. A strip of zinc is placed into a solution of HCl.
   c. Copper is dipped into a solution of $ZnCl_2$.
   d. A solution of silver nitrate is dropped onto an aluminum plate.

5. Dentists occasionally use metallic mixtures called *amalgams* for fillings. If an amalgam contains zinc, however, water can contaminate the amalgam as it is being manipulated, producing hydrogen gas under basic conditions. As the filling hardens, the gas can be released, causing pain and cracking the tooth. Write a balanced chemical equation for this reaction.
6. Copper metal readily dissolves in dilute aqueous nitric acid to form blue $Cu^{2+}(aq)$ and nitric oxide gas.
    a. What has been oxidized? What has been reduced?
    b. Balance the chemical equation.
7. Classify each reaction as an acid–base reaction, a precipitation reaction, or a redox reaction, or state if there is no reaction; then complete and balance the chemical equation:
    a. $Pt^{2+}(aq) + Ag(s) \rightarrow$
    b. $HCN(aq) + NaOH(aq) \rightarrow$
    c. $Fe(NO_3)_3(aq) + NaOH(aq) \rightarrow$
    d. $CH_4(g) + O_2(g) \rightarrow$
8. Classify each reaction as an acid–base reaction, a precipitation reaction, or a redox reaction, or state if there is no reaction; then complete and balance the chemical equation:
    a. $Zn(s) + HCl(aq) \rightarrow$
    b. $HNO_3(aq) + AlCl_3(aq) \rightarrow$
    c. $K_2CrO_4(aq) + Ba(NO_3)_2(aq) \rightarrow$
    d. $Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$

Contributors

- Anonymous

10.10: Balancing Oxidation-Reduction Equations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.11: Oxidizing and Reducing Agents

Learning Objectives

- To understand how an oxidant accept electrons (and are themselves reduced) while reductants donate electrons (and are themselves oxidized).
- Recognize redox reactions in chemcial reactions and everyday situtations such as in combustion.

Compounds that are capable of accepting electrons, such as O_2 or F_2, are called oxidants (or oxidizing agents) because they can oxidize other compounds. In the process of accepting electrons, an oxidant is reduced. Compounds that are capable of donating electrons, such as sodium metal or cyclohexane (C_6H_{12}), are called reductants (or reducing agents) because they can cause the reduction of another compound. In the process of donating electrons, a reductant is oxidized. These relationships are summarized in Equation 3.30:

$$\text{oxidant + reductant} \rightarrow \text{oxidation−reduction}$$

$$\underset{gains\ e^-\ (is\ reduced)}{O_2(g)} + \underset{loses\ e^-\ (is\ oxidized)}{4Na(s)} \rightarrow \underset{redox\ reaction}{2\ Na_2O(s)} \tag{3.30}$$

Some oxidants have a greater ability than others to remove electrons from other compounds. Oxidants can range from very powerful, capable of oxidizing most compounds with which they come in contact, to rather weak. Both F_2 and Cl_2 are powerful oxidants: for example, F_2 will oxidize H_2O in a vigorous, potentially explosive reaction. In contrast, S_8 is a rather weak oxidant, and O_2 falls somewhere in between. Conversely, reductants vary in their tendency to donate electrons to other compounds. Reductants can also range from very powerful, capable of giving up electrons to almost anything, to weak. The alkali metals are powerful reductants, so they must be kept away from atmospheric oxygen to avoid a potentially hazardous redox reaction.

A combustion reaction, first introduced in Section 3.2 "Determining Empirical and Molecular Formulas", is an oxidation–reduction reaction in which the oxidant is O_2. One example of a combustion reaction is the burning of a candle, shown in Figure 3.9 "An Example of a Combustion Reaction". Consider, for example, the combustion of cyclohexane, a typical hydrocarbon, in excess oxygen. The balanced chemical equation for the reaction, with the oxidation state shown for each atom, is as follows:

$$\overset{+1}{\underset{-2}{C_6H_{12}}} + 9\overset{0}{O_2} \rightarrow 6\overset{+4}{\underset{-2}{CO_2}} + 6\overset{+1}{\underset{-2}{H_2O}} \tag{3.31}$$

If we compare the oxidation state of each element in the products and the reactants, we see that hydrogen is the only element whose oxidation state does not change; it remains +1. Carbon, however, has an oxidation state of −2 in cyclohexane and +4 in CO_2; that is, each carbon atom changes its oxidation state by six electrons during the reaction. Oxygen has an oxidation state of 0 in the reactants, but it gains electrons to have an oxidation state of −2 in CO_2 and H_2O. Because carbon has been oxidized, cyclohexane is the reductant; because oxygen has been reduced, it is the oxidant. All combustion reactions are therefore oxidation–reduction reactions.

We described the defining characteristics of oxidation–reduction, or redox, reactions in Chapter 7. Most of the reactions we considered there were relatively simple, and balancing them was straightforward. When oxidation–reduction reactions occur in aqueous solution, however, the equations are more complex and can be more difficult to balance by inspection. Because a balanced chemical equation is the most important prerequisite for solving any stoichiometry problem, we need a method for balancing oxidation–reduction reactions in aqueous solution that is generally applicable. One such method uses *oxidation states*, and a second is referred to as the *half-reaction* method. We show you how to balance redox equations using oxidation states in this section; the half-reaction method is discussed in the second semester

Note the Pattern
Species in high oxidation states act as oxidants, whereas species in low oxidation states act as reductants.

When an aqueous solution of a compound that contains an element in a high oxidation state is mixed with an aqueous solution of a compound that contains an element in a low oxidation state, an oxidation–reduction reaction is likely to occur.

Contributors

- Anonymous

10.11: Oxidizing and Reducing Agents is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.12: Key Terms

10.12: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by LibreTexts.

10.13: Key Equations

10.13: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by LibreTexts.

10.E: Reactions in Aqueous Solution (Exercises)

4.1: General Properties of Aqueous Solutions

Conceptual Problems

1. What are the advantages to carrying out a reaction in solution rather than simply mixing the pure reactants?
2. What types of compounds dissolve in polar solvents?
3. Describe the charge distribution in liquid water. How does this distribution affect its physical properties?
4. Must a molecule have an asymmetric charge distribution to be polar? Explain your answer.
5. Why are many ionic substances soluble in water?
6. Explain the phrase *like dissolves like*.
7. What kinds of covalent compounds are soluble in water?
8. Why do most aromatic hydrocarbons have only limited solubility in water? Would you expect their solubility to be higher, lower, or the same in ethanol compared with water? Why?
9. Predict whether each compound will dissolve in water and explain why.
    a. toluene
    b. acetic acid
    c. sodium acetate
    d. butanol
    e. pentanoic acid
10. Predict whether each compound will dissolve in water and explain why.
    a. ammonium chloride
    b. 2-propanol
    c. heptane
    d. potassium dichromate
    e. 2-octanol
11. Given water and toluene, predict which is the better solvent for each compound and explain your reasoning.
    a. sodium cyanide
    b. benzene
    c. acetic acid
    d. sodium ethoxide (CH_3CH_2ONa)
12. Of water and toluene, predict which is the better solvent for each compound and explain your reasoning.
    a. *t*-butanol
    b. calcium chloride
    c. sucrose
    d. cyclohexene
13. Compound *A* is divided into three equal samples. The first sample does not dissolve in water, the second sample dissolves only slightly in ethanol, and the third sample dissolves completely in toluene. What does this suggest about the polarity of *A*?
14. You are given a mixture of three solid compounds—*A*, *B*, and *C*—and are told that *A* is a polar compound, *B* is slightly polar, and *C* is nonpolar. Suggest a method for separating these three compounds.
15. A laboratory technician is given a sample that contains only sodium chloride, sucrose, and cyclodecanone (a ketone). You must tell the technician how to separate these three compounds from the mixture. What would you suggest?
16. Many over-the-counter drugs are sold as ethanol/water solutions rather than as purely aqueous solutions. Give a plausible reason for this practice.

17. What distinguishes a weak electrolyte from a strong electrolyte?
18. Which organic groups result in aqueous solutions that conduct electricity?
19. It is considered highly dangerous to splash barefoot in puddles during a lightning storm. Why?
20. Which solution(s) would you expect to conduct electricity well? Explain your reasoning.
    a. an aqueous solution of sodium chloride
    b. a solution of ethanol in water
    c. a solution of calcium chloride in water
    d. a solution of sucrose in water
21. Which solution(s) would you expect to conduct electricity well? Explain your reasoning.
    a. an aqueous solution of acetic acid
    b. an aqueous solution of potassium hydroxide
    c. a solution of ethylene glycol in water
    d. a solution of ammonium chloride in water
22. Which of the following is a strong electrolyte, a weak electrolyte, or a nonelectrolyte in an aqueous solution? Explain your reasoning.
    a. potassium hydroxide
    b. ammonia
    c. calcium chloride
    d. butanoic acid
23. Which of the following is a strong electrolyte, a weak electrolyte, or a nonelectrolyte in an aqueous solution? Explain your reasoning.
    a. magnesium hydroxide
    b. butanol
    c. ammonium bromide
    d. pentanoic acid
24. Which of the following is a strong electrolyte, a weak electrolyte, or a nonelectrolyte in aqueous solution? Explain your reasoning.
    a. H_2SO_4
    b. diethylamine
    c. 2-propanol
    d. ammonium chloride
    e. propanoic acid

Conceptual Answers

1.
2.
3.
4.
5. Ionic compounds such as NaCl are held together by electrostatic interactions between oppositely charged ions in the highly ordered solid. When an ionic compound dissolves in water, the partially negatively charged oxygen atoms of the H_2O molecules surround the cations, and the partially positively charged hydrogen atoms in H_2O surround the anions. The favorable electrostatic interactions between water and the ions compensate for the loss of the electrostatic interactions between ions in the solid.
6.
7.
8.
9.

a. Because toluene is an aromatic hydrocarbon that lacks polar groups, it is unlikely to form a homogenous solution in water.
b. Acetic acid contains a carboxylic acid group attached to a small alkyl group (a methyl group). Consequently, the polar characteristics of the carboxylic acid group will be dominant, and acetic acid will form a homogenous solution with water.
c. Because most sodium salts are soluble, sodium acetate should form a homogenous solution with water.
d. Like all alcohols, butanol contains an −OH group that can interact well with water. The alkyl group is rather large, consisting of a 4-carbon chain. In this case, the nonpolar character of the alkyl group is likely to be as important as the polar character of the –OH, decreasing the likelihood that butanol will form a homogeneous solution with water.
e. Like acetic acid, pentanoic acid is a carboxylic acid. Unlike acetic acid, however, the alkyl group is rather large, consisting of a 4-carbon chain as in butanol. As with butanol, the nonpolar character of the alkyl group is likely to be as important as the polar character of the carboxylic acid group, making it unlikely that pentanoic acid will form a homogeneous solution with water. (In fact, the solubility of both butanol and pentanoic acid in water is quite low, only about 3 g per 100 g water at 25°C.)

10.
11.
12.
13.
14.
15.
16.

17. An *electrolyte* is any compound that can form ions when it dissolves in water. When a strong electrolyte dissolves in water, it dissociates completely to give the constituent ions. In contrast, when a weak electrolyte dissolves in water, it produces relatively few ions in solution.

4.2: Precipitation Reactions

Conceptual Problems

0. What information can be obtained from a complete ionic equation that cannot be obtained from the overall chemical equation?

1. Predict whether mixing each pair of solutions will result in the formation of a precipitate. If so, identify the precipitate.

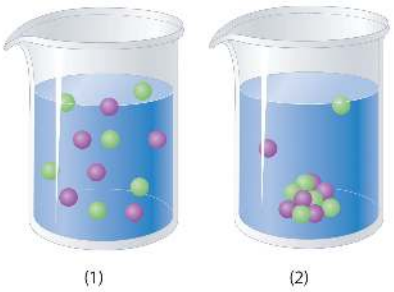
a. $FeCl_2(aq) + Na_2S(aq)$
b. $NaOH(aq) + H_3PO_4(aq)$
c. $ZnCl_2(aq) + (NH_4)_2S(aq)$

2. Predict whether mixing each pair of solutions will result in the formation of a precipitate. If so, identify the precipitate.

a. $KOH(aq) + H_3PO_4(aq)$
b. $K_2CO_3(aq) + BaCl_2(aq)$
c. $Ba(NO_3)_2(aq) + Na_2SO_4(aq)$

3. Which representation best corresponds to an aqueous solution originally containing each of the following?

a. 1 M NH_4Cl
b. 1 M NaO_2CCH_3
c. 1 M NaOH + 1 M HCl
d. 1 M $Ba(OH)_2$ + 1 M H_2SO_4



4. Which representation in Problem 3 best corresponds to an aqueous solution originally containing each of the following?

   a. 1 M CH_3CO_2H + 1 M NaOH
   b. 1 M NH_3 + 1 M HCl
   c. 1 M Na_2CO_3 + 1 M H_2SO_4
   d. 1 M $CaCl_2$ + 1 M H_3PO_4

Conceptual Answer

3

a. 1
b. 1
c. 1
d. 2

Numerical Problems

1. What mass of precipitate would you expect to obtain by mixing 250 mL of a solution containing 4.88 g of Na_2CrO_4 with 200 mL of a solution containing 3.84 g of $AgNO_3$? What is the final nitrate ion concentration?
2. Adding 10.0 mL of a dilute solution of zinc nitrate to 246 mL of 2.00 M sodium sulfide produced 0.279 g of a precipitate. How many grams of zinc(II) nitrate and sodium sulfide were consumed to produce this quantity of product? What was the concentration of each ion in the original solutions? What is the concentration of the sulfide ion in solution after the precipitation reaction, assuming no further reaction?

Numerical Answer

1. 3.75 g Ag_2CrO_4; 5.02×10^{-2} M nitrate

4.3: Acid-Base Reactions

Conceptual Problems

1. Why was it necessary to expand on the Arrhenius definition of an acid and a base? What specific point does the Brønsted–Lowry definition address?
2. State whether each compound is an acid, a base, or a salt.

   a. $CaCO_3$
   b. $NaHCO_3$
   c. H_2SO_4
   d. $CaCl_2$
   e. $Ba(OH)_2$

3. State whether each compound is an acid, a base, or a salt.

   a. NH_3
   b. NH_4Cl
   c. H_2CO_3
   d. CH_3COOH

e. NaOH

4. Classify each compound as a strong acid, a weak acid, a strong base, or a weak base in aqueous solution.

   a. sodium hydroxide
   b. acetic acid
   c. magnesium hydroxide
   d. tartaric acid
   e. sulfuric acid
   f. ammonia
   g. hydroxylamine (NH_2OH)
   h. hydrocyanic acid

5. Decide whether each compound forms an aqueous solution that is strongly acidic, weakly acidic, strongly basic, or weakly basic.

   a. propanoic acid
   b. hydrobromic acid
   c. methylamine
   d. lithium hydroxide
   e. citric acid
   f. sodium acetate
   g. ammonium chloride
   h. barium hydroxide

6. What is the relationship between the strength of an acid and the strength of the conjugate base derived from that acid? Would you expect the $CH_3CO_2^-$ ion to be a strong base or a weak base? Why? Is the hydronium ion a strong acid or a weak acid? Explain your answer.

7. What are the products of an acid–base reaction? Under what circumstances is one of the products a gas?

8. Explain how an aqueous solution that is strongly basic can have a pH, which is a measure of the *acidity* of a solution.

Answer

1.
2.
3.
4.
5.

   a. weakly acidic
   b. strongly acidic
   c. weakly basic
   d. strongly basic
   e. weakly acidic
   f. weakly basic
   g. weakly acidic
   h. strongly basic

6.
7.
8.

Numerical Problems

Please be sure you are familiar with the topics discussed in Essential Skills 3 (section 4.1 "Aqueous Solutions"0) before proceeding to the Numerical Problems.

1. Derive an equation to relate the hydrogen ion concentration to the molarity of a solution of a strong monoprotic acid.

2. Derive an equation to relate the hydroxide ion concentration to the molarity of a solution of
    a. a group I hydroxide.
    b. a group II hydroxide.
3. Given the following salts, identify the acid and the base in the neutralization reactions and then write the complete ionic equation:
    a. barium sulfate
    b. lithium nitrate
    c. sodium bromide
    d. calcium perchlorate
4. What is the pH of each solution?
    a. 5.8×10^{-3} mol of HNO_3 in 257 mL of water
    b. 0.0079 mol of HI in 750 mL of water
    c. 0.011 mol of $HClO_4$ in 500 mL of water
    d. 0.257 mol of HBr in 5.00 L of water
5. What is the hydrogen ion concentration of each substance in the indicated pH range?
    a. black coffee (pH 5.10)
    b. milk (pH 6.30–7.60)
    c. tomatoes (pH 4.00–4.40)
6. What is the hydrogen ion concentration of each substance in the indicated pH range?
    a. orange juice (pH 3–4)
    b. fresh egg white (pH 7.60–7.80)
    c. lemon juice (pH 2.20–2.40)
7. What is the pH of a solution prepared by diluting 25.00 mL of 0.879 M HCl to a volume of 555 mL?
8. Vinegar is primarily an aqueous solution of acetic acid. Commercial vinegar typically contains 5.0 g of acetic acid in 95.0 g of water. What is the concentration of commercial vinegar? If only 3.1% of the acetic acid dissociates to $CH_3CO_2^-$ and H^+, what is the pH of the solution? (Assume the density of the solution is 1.00 g/mL.)
9. If a typical household cleanser is 0.50 M in strong base, what volume of 0.998 M strong monoprotic acid is needed to neutralize 50.0 mL of the cleanser?
10. A 25.00 mL sample of a 0.9005 M solution of HCl is diluted to 500.0 mL. What is the molarity of the final solution? How many milliliters of 0.223 M NaOH are needed to neutralize 25.00 mL of this final solution?
11. If 20.0 mL of 0.10 M NaOH are needed to neutralize 15.0 mL of gastric fluid, what is the molarity of HCl in the fluid? (Assume all the acidity is due to the presence of HCl.) What other base might be used instead of NaOH?
12. Malonic acid ($C_3H_4O_4$) is a diprotic acid used in the manufacture of barbiturates. How many grams of malonic acid are in a 25.00 mL sample that requires 32.68 mL of 1.124 M KOH for complete neutralization to occur? Malonic acid is a dicarboxylic acid; propose a structure for malonic acid.
13. Describe how you would prepare 500 mL of a 1.00 M stock solution of HCl from an HCl solution that is 12.11 M. Using your stock solution, how would you prepare 500 mL of a solution that is 0.012 M in HCl?
14. Given a stock solution that is 8.52 M in HBr, describe how you would prepare a 500 mL solution with each concentration.
    a. 2.50 M
    b. 4.00×10^{-3} M
    c. 0.989 M
15. How many moles of solute are contained in each?
    a. 25.00 mL of 1.86 M NaOH
    b. 50.00 mL of 0.0898 M HCl
    c. 13.89 mL of 0.102 M HBr

16. A chemist needed a solution that was approximately 0.5 M in HCl but could measure only 10.00 mL samples into a 50.00 mL volumetric flask. Propose a method for preparing the solution. (Assume that concentrated HCl is 12.0 M.)

17. Write the balanced chemical equation for each reaction.

    a. perchloric acid with potassium hydroxide
    b. nitric acid with calcium hydroxide

18. Write the balanced chemical equation for each reaction.

    a. solid strontium hydroxide with hydrobromic acid
    b. aqueous sulfuric acid with solid sodium hydroxide

19. A neutralization reaction gives calcium nitrate as one of the two products. Identify the acid and the base in this reaction. What is the second product? If the product had been cesium iodide, what would have been the acid and the base? What is the complete ionic equation for each reaction?

Answers

1. $[H_3O^+] = [HA]$ M
2.
3.
    a. H_2SO_4 and $Ba(OH)_2$; $2H^+ + SO_4^{2-} + Ba^{2+} + 2OH^- \rightarrow 2H_2O + Ba^{2+} + SO_4^{2-}$
    b. HNO_3 and LiOH; $H^+ + NO_3^- + Li^+ + OH^- \rightarrow H_2O + Li^+ + NO_3^-$
    c. HBr and NaOH; $H^+ + Br^- + Na^+ + OH^- \rightarrow H_2O + Na^+ + Br^-$
    d. $HClO_4$ and $Ca(OH)_2$; $2H^+ + 2ClO_4^- + Ca^{2+} + 2OH^- \rightarrow 2H_2O + Ca^{2+} + 2ClO_4^-$
4.
5.
    a. 7.9×10^{-6} M H^+
    b. 5.0×10^{-7} to 2.5×10^{-8} M H^+
    c. 1.0×10^{-4} to 4.0×10^{-5} M H^+
6.
7. pH = 1.402
8.
9. 25 mL
10.
11. 0.13 M HCl; magnesium carbonate, $MgCO_3$, or aluminum hydroxide, $Al(OH)_3$
12.
13. 1.00 M solution: dilute 41.20 mL of the concentrated solution to a final volume of 500 mL. 0.012 M solution: dilute 12.0 mL of the 1.00 M stock solution to a final volume of 500 mL.
14.
15.
    a. 4.65×10^{-2} mol NaOH
    b. 4.49×10^{-3} mol HCl
    c. 1.42×10^{-3} mol HBr
16.
17.
    a. $HClO_4 + KOH \rightarrow KClO_4 + H_2O$
    b. $2HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2H_2O$

18.

19. The acid is nitric acid, and the base is calcium hydroxide. The other product is water.

$$2HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2H_2O$$

The acid is hydroiodic acid, and the base is cesium hydroxide. The other product is water.

$$HI + CsOH \rightarrow CsI + H_2O$$

The complete ionic equations are

$$2H^+ + 2NO_3^- + Ca^{2+} + 2OH^- \rightarrow Ca^{2+} + 2NO_3^- + H_2O$$
$$H^+ + I^- + Cs^+ + OH^- \rightarrow Cs^+ + I^- + H_2O$$

4.4: Oxidation-Reduction Reactions

Conceptual Problems

1. Which elements in the periodic table tend to be good oxidants? Which tend to be good reductants?
2. If two compounds are mixed, one containing an element that is a poor oxidant and one with an element that is a poor reductant, do you expect a redox reaction to occur? Explain your answer. What do you predict if one is a strong oxidant and the other is a weak reductant? Why?
3. In each redox reaction, determine which species is oxidized and which is reduced:
   a. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
   b. $Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$
   c. $BrO_3^-(aq) + 2MnO_2(s) + H_2O(l) \rightarrow Br^-(aq) + 2MnO_4^-(aq) + 2H^+(aq)$
4. Single-displacement reactions are a subset of redox reactions. In this subset, what is oxidized and what is reduced? Give an example of a redox reaction that is *not* a single-displacement reaction.

Conceptual Solutions

1. In the periodic table, the trend for oxidation increases towards the top-right of the table, while the trend for reduction increases towards the bottom-left. That is to say, the nonmetals and metalloids (with the exception of Noble Gases) tend to be better oxidants, and alkali and transition metals tend to be better reductants.
2. If the compounds mixed are both poor oxidants and poor reductants, no redox reaction will occur. The poor oxidizing agent has a weak ability to gain electrons, while the reducing agent has a weak ability to donate electrons. Due to this, no electrons (or very few electrons) will be transferred, and no reaction will occur. If one compound is a strong oxidant and the other is a weak reductant, the reaction will progress, but not as far as it would have with a strong reductant. Some reductants, such as gold, are so weak that only the strongest of oxidants can oxidize them.
3. Which is Oxidized and Reduced:
   1. Zn is oxidized, and H is reduced.
   2. Cu is oxidized, and N is reduced.
   3. Mn is oxidized, and Br is reduced.
4. A single-displacement reaction occurs when one "free" element replaces another similar element in a compound. What is oxidized and what is reduced depends on the elements in question. For example, in the single-displacement reaction $Zn(s)+2HCl(aq) \rightarrow ZnCl_2(aq)+H_2(g)$, the free element Zinc is oxidized while Hydrogen is reduced. In $Cl_2(g)+2NaBr(aq) \rightarrow 2NaCl(aq)+Br_2(l)$, on the other hand, the free element Chlorine is reduced while Bromine is oxidized.
   1. An example of a non-single-displacement reaction is a *double*-displacement reaction, such as: $2KI(aq)+Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq)+PbI_2(s)$

Numerical Problems

1. Balance each redox reaction under the conditions indicated.
   a. $CuS(s) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + SO_4^{2-}(aq) + NO(g)$; acidic solution
   b. $Ag(s) + HS^-(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2S(s) + Cr(OH)_3(s)$; basic solution

c. $Zn(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + H_2(g)$; acidic solution
d. $O_2(g) + Sb(s) \rightarrow H_2O_2(aq) + SbO_2^-(aq)$; basic solution
e. $UO_2^{2+}(aq) + Te(s) \rightarrow U^{4+}(aq) + TeO_4^{2-}(aq)$; acidic solution

2. Balance each redox reaction under the conditions indicated.

a. $MnO_4^-(aq) + S_2O_3^{2-}(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$; acidic solution
b. $Fe^{2+}(aq) + Cr_2O_7^{2-}(aq) \rightarrow Fe^{3+}(aq) + Cr^{3+}(aq)$; acidic solution
c. $Fe(s) + CrO_4^{2-}(aq) \rightarrow Fe_2O_3(s) + Cr_2O_3(s)$; basic solution
d. $Cl_2(aq) \rightarrow ClO_3^-(aq) + Cl^-(aq)$; acidic solution
e. $CO_3^{2-}(aq) + N_2H_4(aq) \rightarrow CO(g) + N_2(g)$; basic solution

3. Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation; then write the complete ionic equation for the reaction.

a. Platinum wire is dipped in hydrochloric acid.
b. Manganese metal is added to a solution of iron(II) chloride.
c. Tin is heated with steam.
d. Hydrogen gas is bubbled through a solution of lead(II) nitrate.

4. Using the activity series, predict what happens in each situation. If a reaction occurs, write the net ionic equation; then write the complete ionic equation for the reaction.

a. A few drops of $NiBr_2$ are dropped onto a piece of iron.
b. A strip of zinc is placed into a solution of HCl.
c. Copper is dipped into a solution of $ZnCl_2$.
d. A solution of silver nitrate is dropped onto an aluminum plate.

5. Dentists occasionally use metallic mixtures called *amalgams* for fillings. If an amalgam contains zinc, however, water can contaminate the amalgam as it is being manipulated, producing hydrogen gas under basic conditions. As the filling hardens, the gas can be released, causing pain and cracking the tooth. Write a balanced chemical equation for this reaction.

6. Copper metal readily dissolves in dilute aqueous nitric acid to form blue $Cu^{2+}(aq)$ and nitric oxide gas.

a. What has been oxidized? What has been reduced?
b. Balance the chemical equation.

7. Classify each reaction as an acid–base reaction, a precipitation reaction, or a redox reaction, or state if there is no reaction; then complete and balance the chemical equation:

a. $Pt^{2+}(aq) + Ag(s) \rightarrow$
b. $HCN(aq) + NaOH(aq) \rightarrow$
c. $Fe(NO_3)_3(aq) + NaOH(aq) \rightarrow$
d. $CH_4(g) + O_2(g) \rightarrow$

8. Classify each reaction as an acid–base reaction, a precipitation reaction, or a redox reaction, or state if there is no reaction; then complete and balance the chemical equation:

a. $Zn(s) + HCl(aq) \rightarrow$
b. $HNO_3(aq) + AlCl_3(aq) \rightarrow$
c. $K_2CrO_4(aq) + Ba(NO_3)_2(aq) \rightarrow$
d. $Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$Conceptual

Numerical Solutions

1. Balanced Redox Reactions:

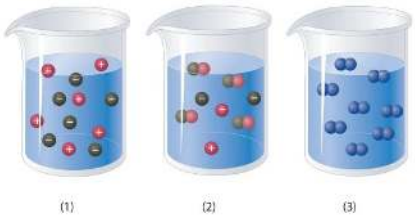
1. $3CuS(s) + 8NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 3SO_4^{2-}(aq) + 8NO(g) + 4H_2O(l)$
2. $6Ag(s) + 2CrO_4^{2-}(aq) + 3HS^-(aq) + 5H_2O(l) \rightarrow 3Ag_2S(s) + 2Cr(OH)_3(s) + 7OH^-(aq)$
3. $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
4. $2Sb(s) + 3O_2(g) + 2H_2O(l) + 2OH^-(aq) \rightarrow 2SbO_2^-(aq) + 3H_2O_2(aq)$
5. $Te(s) + 3UO_2^{2+}(aq) + 4H^+(aq) \rightarrow TeO_4^{2-}(aq) + 3U^{4+}(aq) + 2H_2O(l)$

4.5: Concentration of Solutions

Conceptual Problems

1. Which of the representations best corresponds to a 1 M aqueous solution of each compound? Justify your answers.

   a. NH_3
   b. HF
   c. $CH_3CH_2CH_2OH$
   d. Na_2SO_4



2. Which of the representations shown in Problem 1 best corresponds to a 1 M aqueous solution of each compound? Justify your answers.

   a. CH_3CO_2H
   b. NaCl
   c. Na_2S
   d. Na_3PO_4
   e. acetaldehyde

3. Would you expect a 1.0 M solution of $CaCl_2$ to be a better conductor of electricity than a 1.0 M solution of NaCl? Why or why not?

4. An alternative way to define the concentration of a solution is *molality*, abbreviated *m*. Molality is defined as the number of moles of solute in 1 kg of *solvent*. How is this different from molarity? Would you expect a 1 M solution of sucrose to be more or less concentrated than a 1 *m* solution of sucrose? Explain your answer.

5. What are the advantages of using solutions for quantitative calculations?

Conceptual Answer

1.

Compound	Corresponding Beaker	Reasoning
NH_3	2	Weak base dissociates into both ions and molecules.
HF	2	Weak acid dissociates into both ions and molecules.
$CH_3CH_2CH_2OH$	3	Covalent compound
Na_2SO_4	1	Soluble ionic compound

2.

CH_3CO_2H	2	Weak acid dissociates into both ions and molecules.
NaCl	1	Soluble ionic compound
Na_2S	1	Soluble ionic compound
Na_3PO_4	1	Soluble ionic compound

Acetaldehyde	1	Aldehydes with short chain lengths are miscible

3. $CaCl_2$ would be expected to be a better conductor than NaCl due to the fact that it contains more charges. The conductivity of a solution depends on the amount of mobile charges within it. While NaCl only has two charges (Na being + and Cl being -), $CaCl_2$ contains three (Ca being 2+ and the two Cls being -). With more free charges when dissolved in the solution, $CaCl_2$ would be a better conductor.
4. Assuming that the solvent is water, the solution with a Molarity (M) of 1 would be more concentrated than the solution with a Molality (m) of 1. Water is approximately 1 kg per liter, meaning that the solute is dissolved in 1 liter of water. This *increases* the amount of solution past 1 liter. Molarity, on the other hand, means that there is 1 liter of solution (that is, including both solute and solvent). Since in both cases there will be one mole, the 1M solution will be more concentrated than the 1m solution because there is less volume, but the same amount of solute.
5. If the amount of a substance required for a reaction is too small to be weighed accurately, the use of a solution of the substance, in which the solute is dispersed in a much larger mass of solvent, allows chemists to measure the quantity of the substance more accurately.

Numerical Problems

1. Calculate the number of grams of solute in 1.000 L of each solution.
   a. 0.2593 M $NaBrO_3$
   b. 1.592 M KNO_3
   c. 1.559 M acetic acid
   d. 0.943 M potassium iodate
2. Calculate the number of grams of solute in 1.000 L of each solution.
   a. 0.1065 M BaI_2
   b. 1.135 M Na_2SO_4
   c. 1.428 M NH_4Br
   d. 0.889 M sodium acetate
3. If all solutions contain the same solute, which solution contains the greater mass of solute?
   a. 1.40 L of a 0.334 M solution or 1.10 L of a 0.420 M solution
   b. 25.0 mL of a 0.134 M solution or 10.0 mL of a 0.295 M solution
   c. 250 mL of a 0.489 M solution or 150 mL of a 0.769 M solution
4. Complete the following table for 500 mL of solution.

Compound	Mass (g)	Moles	Concentration (M)
calcium sulfate	4.86		
acetic acid		3.62	
hydrogen iodide dihydrate			1.273
barium bromide	3.92		
glucose			0.983
sodium acetate		2.42	

5. What is the concentration of each species present in the following aqueous solutions?
   a. 0.489 mol of $NiSO_4$ in 600 mL of solution
   b. 1.045 mol of magnesium bromide in 500 mL of solution
   c. 0.146 mol of glucose in 800 mL of solution
   d. 0.479 mol of $CeCl_3$ in 700 mL of solution

6. What is the concentration of each species present in the following aqueous solutions?
    a. 0.324 mol of K_2MoO_4 in 250 mL of solution
    b. 0.528 mol of potassium formate in 300 mL of solution
    c. 0.477 mol of $KClO_3$ in 900 mL of solution
    d. 0.378 mol of potassium iodide in 750 mL of solution
7. What is the molar concentration of each solution?
    a. 8.7 g of calcium bromide in 250 mL of solution
    b. 9.8 g of lithium sulfate in 300 mL of solution
    c. 12.4 g of sucrose ($C_{12}H_{22}O_{11}$) in 750 mL of solution
    d. 14.2 g of iron(III) nitrate hexahydrate in 300 mL of solution
8. What is the molar concentration of each solution?
    a. 12.8 g of sodium hydrogen sulfate in 400 mL of solution
    b. 7.5 g of potassium hydrogen phosphate in 250 mL of solution
    c. 11.4 g of barium chloride in 350 mL of solution
    d. 4.3 g of tartaric acid ($C_4H_6O_6$) in 250 mL of solution
9. Give the concentration of each reactant in the following equations, assuming 20.0 g of each and a solution volume of 250 mL for each reactant.
    a. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow$
    b. $Ca(OH)_2(aq) + H_3PO_4(aq) \rightarrow$
    c. $Al(NO_3)_3(aq) + H_2SO_4(aq) \rightarrow$
    d. $Pb(NO_3)_2(aq) + CuSO_4(aq) \rightarrow$
    e. $Al(CH_3CO_2)_3(aq) + NaOH(aq) \rightarrow$
10. An experiment required 200.0 mL of a 0.330 M solution of Na_2CrO_4. A stock solution of Na_2CrO_4 containing 20.0% solute by mass with a density of 1.19 g/cm^3 was used to prepare this solution. Describe how to prepare 200.0 mL of a 0.330 M solution of Na_2CrO_4 using the stock solution.
11. Calcium hypochlorite [$Ca(OCl)_2$] is an effective disinfectant for clothing and bedding. If a solution has a $Ca(OCl)_2$ concentration of 3.4 g per 100 mL of solution, what is the molarity of hypochlorite?
12. Phenol (C_6H_5OH) is often used as an antiseptic in mouthwashes and throat lozenges. If a mouthwash has a phenol concentration of 1.5 g per 100 mL of solution, what is the molarity of phenol?
13. If a tablet containing 100 mg of caffeine ($C_8H_{10}N_4O_2$) is dissolved in water to give 10.0 oz of solution, what is the molar concentration of caffeine in the solution?
14. A certain drug label carries instructions to add 10.0 mL of sterile water, stating that each milliliter of the resulting solution will contain 0.500 g of medication. If a patient has a prescribed dose of 900.0 mg, how many milliliters of the solution should be administered?

Numerical Answers

1. Grams of solute in 1.000 L of each solution:
    1. 0.2593 M $NaBrO_3$
        1.
    2. 1.592 M KNO_3
        1.
    3. 1.559 M acetic acid
        1.
    4. 0.943 M potassium iodate
        1.
2.

3.
4.

Compound	Mass (g)	Moles	Concentration (M)
calcium sulfate	4.86	0.0357	0.07
acetic acid	217.39	3.62	7.24
hydrogen iodide dihydrate	103.79	0.637	1.273
barium bromide	3.92	0.013	0.026
glucose	88.55	0.491	0.983
sodium acetate	198.52	2.42	4.84

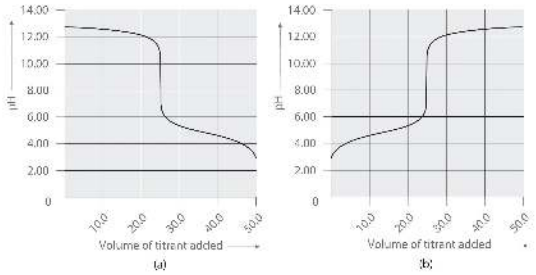
5. Molar Concentration in each solution:
    1. 0.815M $NiSO_4$
    2. 2.09M Magnesium Bromide
    3. 0.183M Glucose
    4. 0.684M $CeCl_3$
6. Molar Concentration in each solution:
    1. 1.296M K_2MoO_4
    2. 1.760M Potassium Formate
    3. 0.530M $KClO_3$
    4. 0.503M Potassium Iodide
7. Molar Concentration in each solution:
    1. 0.174M Calcium Bromide
    2. 0.297M Lithium Sulfate
    3. 0.048M Sucrose
    4. 0.135M Iron(III) Nitrate Hexahydrate
8. Molar Concentration in each solution:
    1. 0.267M Sodium Hydrogen Sulfate
    2. 0.172M Potassium Hydrogen Phosphate
    3. 0.156M Barium Chloride
    4. 0.115M Tartaric Acid
9. Concentration of Reactants in each solution:
    1. 0.192M of $BaCl_2$, 0.282M of Na_2SO_4
    2. 0.541M of $Ca(OH)_2$, 0.408M of H_3PO_4
    3. 0.188M of $Al(NO_3)_3$, 0.408M H_2SO_4(aq)
    4. 0.121M of $Pb(NO_3)_2$, 0.251M $CuSO_4$
    5. 0.196M of $Al(CH_3CO_2)_3$, 1M NaOH
10.
11. 0.4756 M ClO^-
12. 0.159M Phenol
13. 1.74×10^{-3} M caffeine
14. 1.8mL of solution

4.6: Solution Stoichiometry and Chemical Analysis

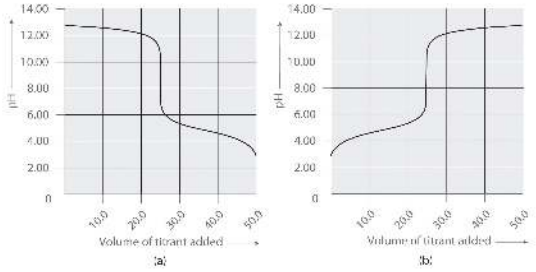
Conceptual Problems

1. The titration procedure is an application of the use of limiting reactants. Explain why this is so.
2. Explain how to determine the concentration of a substance using a titration.

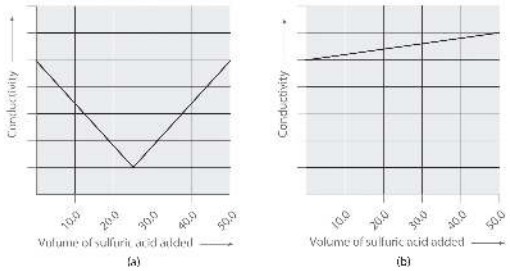
3. Following are two graphs that illustrate how the pH of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M acetic acid with 0.10 M NaOH, and the other corresponds to the titration of 100 mL 0.10 M NaOH with 0.10 M acetic acid. Which graph corresponds to which titration? Justify your answer.



4. Following are two graphs that illustrate how the pH of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M ammonia with 0.10 M HCl, and the other corresponds to the titration of 100 mL 0.10 M NH_4Cl with 0.10 M NaOH. Which graph corresponds to which titration? Justify your answer.



5. Following are two graphs that illustrate how the electrical conductivity of a solution varies during a titration. One graph corresponds to the titration of 100 mL 0.10 M $Ba(OH)_2$ with 0.10 M H_2SO_4 , and the other corresponds to the titration of 100 mL of 0.10 M NaOH with 0.10 M H_2SO_4. Which graph corresponds to which titration? Justify your answer.



Conceptual Answers

1. In a titration procedure, the analyte is considered to be the limiting reactant, as the amount of moles of it remain constant while more and more titrant can be added (making the titrant the excess reactant). By using the fact that the analyte is limited, we can continually add the excess reactant to find out just how many moles are in the analyte.
2. To determine the concentration of a solution, one must titrate it with an acid (if the solution is a base) or a base (if the solution is an acid). By titrating the unknown solution with a solution of known concentration, we can determine how many moles of the titrant were needed to fully neutralize the analyte. This point can be determined by finding the *equivalence point* during the titration, and from this knowledge as well as the mole ratio from the balanced neutralization equation, we can determine how many moles are in the analyte. With this, we can simply divide by the original volume to find its concentration.
3. Graph (a) corresponds to the titration of NaOH with acetic acid, while graph (b) corresponds to the titration of acetic acid with NaOH. This can be determined due to our knowledge that acids have a pH below 7 while bases have a pH above 7. Therefore, titrating a base with an acid would decrease its pH while titrating an acid with a base would increase its pH.
4. Graph (a) corresponds to the titration of ammonia with HCl, while graph (b) corresponds to the titration of NH_4Cl with NaOH. This can be determined due to our knowledge that acids have a pH below 7 while bases have a pH above 7. Therefore, titrating

a base with an acid would decrease its pH while titrating an acid with a base would increase its pH.

5.

1. titration of $Ba(OH)_2$ with sulfuric acid
2. titration of NaOH with sulfuric acid

Numerical Problems

1. A 10.00 mL sample of a 1.07 M solution of potassium hydrogen phthalate (KHP, formula mass = 204.22 g/mol) is diluted to 250.0 mL. What is the molarity of the final solution? How many grams of KHP are in the 10.00 mL sample?
2. What volume of a 0.978 M solution of NaOH must be added to 25.0 mL of 0.583 M HCl to completely neutralize the acid? How many moles of NaOH are needed for the neutralization?
3. A student was titrating 25.00 mL of a basic solution with an HCl solution that was 0.281 M. The student ran out of the HCl solution after having added 32.46 mL, so she borrowed an HCl solution that was labeled as 0.317 M. An additional 11.5 mL of the second solution was needed to complete the titration. What was the concentration of the basic solution?

Numerical Answers

1. Final Molarity = 0.0428M, and there are 0.76g of KHP in the initial sample.
2. 15mL of 0.978 M NaOH, or 0.0147 moles.
3. Concentration of basic solution = 0.536 M.

10.E: Reactions in Aqueous Solution (Exercises) is shared under a CC BY-NC-SA 3.0 license and was authored, remixed, and/or curated by LibreTexts.

10.S: Summary

aqueous solutions – solutions in which water is the dissolving medium

4.1: General Properties of Aqueous Solutions

- **electrolyte** – substance whose aqueous solution contains ions
- **nonelectrolyte** – substance that does not form ions in solution

4.2.1 Ionic Compounds in Water

- **dissociate** – when ions separate from a solid being dissolved

4.2.2 Molecular Compounds in Water

- the molecular structure is maintained

4.2.3 Strong and Weak Electrolytes

- strong electrolytes – ionic compounds that exists entirely of ions in solution
- weak electryolytes – molecular compounds that produce a small amound of ions
- chemical equilibrium – equilibrium of forming ions and recrystalizing ions

4.2: Precipitation Reactions

$$AX + BY \rightarrow AY + BX$$

- for methathesis to occur:

    1. the formation of an insoluble product
    2. the formation of either a weak electrolyte or a nonelectrolyte
    3. the formation of a gas that escapes from solution

4.5.1 Precipitation Reactions

- **precipitate** – insoluble solid formed by a reaction in solution
- **solubility** – amount of substance that can be dissolved in a given quantity

4.5.2 Solubility Guidelines for Ionic Compounds

- all common ionic compounds of the alkali metal ions and of the ammonium ion are soluble in water

4.5.3 Reactions in Which a Weak Electrolyte or Nonelectrolyte Forms

- hydrogen and hydroxide react to form water
- insoluble metal oxides react with acids

4.3: Acid-Base Reactions

4.3.1 Acids

- substances that ionize to form hydrogen ions
- proton donors

4.3.2 Bases

- substances that ionize to form hydroxide ions

4.3.3 Strong and Weak Acids and Bases

- **strong acid, strong base** – strong electrolyte
- **weak acid, weak base** – weak electrolyte

4.3.4 Neutralization Reactions and Salts

- **neutralization reaction** – when an acid and base are mixed
- produces water and a salt

4.4 Ionic Equations

- **molecular formula** – and equation written to show the complete chemical formulas of reactants and products
- **spectator ions** – ions that do not play a role in a reaction
- **net ionic equation** – equation where the spectator ions are removed
- only soluble strong electrolytes are written in ionic form

4.4: Oxidation-Reduction Reactions

4.6.1 Reactions in Which a Gas Forms

- carbonates and bicarbonates

4.6.2 Oxidation and Reduction

- **oxidation** – loss of electrons
- **reduction** – gain of electrons

4.6.3 Oxidation of Metals by Acids and Salts

- whenever one substance is oxidized, some other substance must be reduced
- metals react with acids to form salts and hydrogen gas

4.6.4 The Activity Series

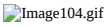
- **activity series** – list of metals arranged in order of decreasing ease of oxidation
- **active metals** – alkali metals and alkaline earth metals
- any metal on the list can be oxidized by ions of elements below it

4.5: Concentration of Solutions

- **solution** – homogeneous mixture of two or more substances
- **solvent** – component that is present in greatest quantity
- **solutes** – substances dissolved in the solvent

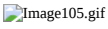
4.1.1 Molarity

- **concentration** – the amount of solute dissolved in a given quantity of solvent or solution
- **molarity** – number of moles of solute in a liter of solution



4.1.2 Dilution

- **dilution** - obtaining a lower concentration of a solution by adding water
- moles solute before dilution = moles solute after dilution



4.6: Solution Stoichiometry and Chemical Analysis

4.7.1 Titrations

- **statndard solution** – solution of known concentration
- **titration** – a known solution that undergoes a specific chemical reaction of known stoichiometry with the solution of unknown concentration
- **equivalence point** – stoichiometrically equivalent quantities of reactants are brought together
- **indicator** – used to show the endpoint of the titration

10.S: Summary is shared under a CC BY-NC-SA 3.0 license and was authored, remixed, and/or curated by LibreTexts.

- **4.S: Reactions in Aqueous Solution (Summary)** is licensed CC BY-NC-SA 3.0.

CHAPTER OVERVIEW

11: Nuclear Chemistry

The chemical reactions that we have considered in previous chapters involve changes in the *electronic* structure of the species involved, that is, the arrangement of the **electrons** around atoms, ions, or molecules. *Nuclear* structure, the numbers of protons and neutrons within the nuclei of the atoms involved, remains unchanged during chemical reactions. This chapter will introduce the topic of nuclear chemistry, which began with the discovery of radioactivity in 1896 by French physicist Antoine Becquerel and has become increasingly important during the twentieth and twenty-first centuries, providing the basis for various technologies related to energy, medicine, geology, and many other areas.

The realm of nuclear chemistry is a field essential for understanding the properties and reactions of atomic nuclei. Nuclear chemistry finds wide-ranging applications in medicine, energy production, and environmental science, making it a crucial area of study in modern scientific research.

Faculty Feature

Dr. Andre Simpson's research is focused on the development of Nuclear Magnetic Resonance (NMR) Spectroscopy techniques to unravel environmental processes at the molecular level. His research places a specific emphasis on the structural categories of carbon in terrestrial and aquatic environments. By employing advanced NMR techniques, he sheds light on the role of carbon in the global carbon cycle and its interactions with organic contaminants and heavy metals, providing invaluable insights into their associations, fate, and transport. Through the interpretation of metabolic fingerprints as indicators of environmental stress and exposure, his work contributes to a deeper understanding of ecosystem health and resilience in the face of changing environmental conditions. Dr. Simpson's interdisciplinary approach exemplifies the integral role of nuclear chemistry in addressing complex environmental challenges. His research not only advances our understanding of molecular interactions in the environment but also underscores the critical importance of nuclear chemistry in promoting sustainable stewardship of our planet.



Courses Taught:

CHMC11 - Principles of Analytical Instrumentation

CHMC16 - Analytical Instrumentation

Read more about Dr. Andre Simpson and his research here: https://www.utsc.utoronto.ca/labs/asimpson

11.1: Introduction
11.2: Nuclear Structure and Stability
11.3: Nuclear Equations
11.4: Radioactive Decay
11.5: Transmutation and Nuclear Energy
11.6: Uses of Radioisotopes
11.7: Biological Effects of Radiation
11.8: Key Terms
11.9: Key Equations
11.10: Summary
11.11: Exercises

This page titled 11: Nuclear Chemistry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

11.1: Introduction

The chemical reactions that we have considered in previous chapters involve changes in the *electronic* structure of the species involved, that is, the arrangement of the electrons around atoms, ions, or molecules. *Nuclear* structure, the numbers of protons and neutrons within the nuclei of the atoms involved, remains unchanged during chemical reactions.

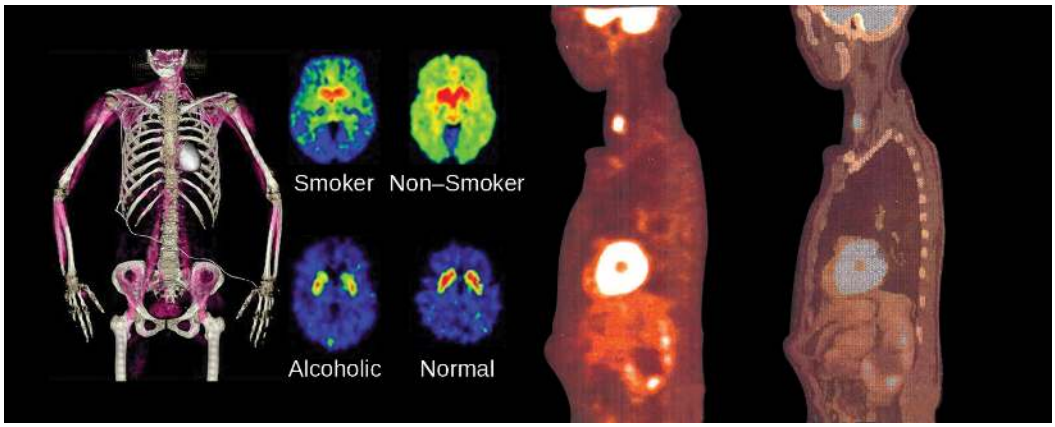


Figure 11.1.1: Nuclear chemistry provides the basis for many useful diagnostic and therapeutic methods in medicine, such as these positron emission tomography (PET) scans. The PET/computed tomography scan on the left shows muscle activity. The brain scans in the center show chemical differences in dopamine signaling in the brains of addicts and nonaddicts. The images on the right show an oncological application of PET scans to identify lymph node metastasis.

This chapter will introduce the topic of nuclear chemistry, which began with the discovery of radioactivity in 1896 by French physicist Antoine Becquerel and has become increasingly important during the twentieth and twenty-first centuries, providing the basis for various technologies related to energy, medicine, geology, and many other areas.

This page titled 11.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.0: Introduction** is licensed CC BY 4.0.

11.2: Nuclear Structure and Stability

Learning Objectives

By the end of this section, you will be able to:

- Describe nuclear structure in terms of protons, neutrons, and electrons
- Calculate mass defect and binding energy for nuclei
- Explain trends in the relative stability of nuclei

Nuclear chemistry is the study of reactions that involve changes in nuclear structure. The chapter on atoms, molecules, and ions introduced the basic idea of nuclear structure, that the nucleus of an atom is composed of protons and, with the exception of ${}^{1}_{1}\text{H}$.

Recall that the number of protons in the nucleus is called the atomic number (Z) of the element, and the sum of the number of protons and the number of neutrons is the mass number (A). Atoms with the same atomic number but different mass numbers are isotopes of the same element. When referring to a single type of nucleus, we often use the term **nuclide** and identify it by the notation ${}^{Z}_{A}\text{X}$, where X is the symbol for the element, A is the mass number, and Z is the atomic number (for example, ${}^{14}_{6}\text{C}$). Often a nuclide is referenced by the name of the element followed by a hyphen and the mass number. For example, ${}^{14}_{6}\text{C}$ is called "carbon-14."

Protons and neutrons, collectively called **nucleons**, are packed together tightly in a nucleus. With a radius of about 10^{-15} meters, a nucleus is quite small compared to the radius of the entire atom, which is about 10^{-10} meters. Nuclei are extremely dense compared to bulk matter, averaging 1.8×10^{14} grams per cubic centimeter. For example, water has a density of 1 gram per cubic centimeter, and iridium, one of the densest elements known, has a density of 22.6 g/cm3. If the earth's density were equal to the average nuclear density, the earth's radius would be only about 200 meters (earth's actual radius is approximately 6.4×10^{6} meters, 30,000 times larger). Example $11.2.1$ demonstrates just how great nuclear densities can be in the natural world.

Example 11.2.1: Density of a Neutron Star

Neutron stars form when the core of a very massive star undergoes gravitational collapse, causing the star's outer layers to explode in a supernova. Composed almost completely of neutrons, they are the densest-known stars in the universe, with densities comparable to the average density of an atomic nucleus. A neutron star in a faraway galaxy has a mass equal to 2.4 solar masses (1 solar mass = $M_{\odot}$ = mass of the sun = $1.99 \times 10^{30}\ kg$) and a diameter of 26 km.

a. What is the density of this neutron star?
b. How does this neutron star's density compare to the density of a uranium nucleus, which has a diameter of about 15 fm (1 fm = 10^{-15} m)?

Solution

We can treat both the neutron star and the U-235 nucleus as spheres. Then the density for both is given by:

$$d = \frac{m}{V} \quad \text{with} \quad V = \frac{4}{3}\pi r^3$$

a. The radius of the neutron star is so the density of the neutron star is:

$$d = \frac{m}{V} = \frac{m}{\frac{4}{3}\pi r^3} = \frac{2.4\left(1.99 \times 10^{30}\,\text{kg}\right)}{\frac{4}{3}\pi\left(1.3 \times 10^{4} m\right)^3} = 5.2 \times 10^{17}\,\text{kg/m}^3$$

b. The radius of the U-235 nucleus is so the density of the U-235 nucleus is:

$$d = \frac{m}{V} = \frac{m}{\frac{4}{3}\pi r^3} = \frac{235\,\text{amu}\left(\frac{1.66 \times 10^{-27}\,\text{kg}}{1\,\text{amu}}\right)}{\frac{4}{3}\pi\left(7.5 \times 10^{-15}\,\text{m}\right)^3} = 2.2 \times 10^{17}\,\text{kg/m}^3$$

These values are fairly similar (same order of magnitude), but the neutron star is more than twice as dense as the U-235 nucleus.

Exercise 11.2.1

Find the density of a neutron star with a mass of 1.97 solar masses and a diameter of 13 km, and compare it to the density of a hydrogen nucleus, which has a diameter of $1.75\ \text{fm}$ ($1\ \text{fm} = 1 \times 10^{-15}\ \text{m}$).

Answer

The density of the neutron star is $3.4 \times 10^{18}\ \text{kg/m}^3$. The density of a hydrogen nucleus is $\boxed{6.0 \times 10^{17}\ \text{kg/m}^3}$. The neutron star is 5.7 times denser than the hydrogen nucleus.

To hold positively charged protons together in the very small volume of a nucleus requires very strong attractive forces because the positively charged protons repel one another strongly at such short distances. The force of attraction that holds the nucleus together is the strong nuclear force. (The strong force is one of the four fundamental forces that are known to exist. The others are the electromagnetic force, the gravitational force, and the nuclear weak force.) This force acts between protons, between neutrons, and between protons and neutrons. It is very different from the electrostatic force that holds negatively charged electrons around a positively charged nucleus (the attraction between opposite charges). Over distances less than 10^{-15} meters and within the nucleus, the strong nuclear force is much stronger than electrostatic repulsions between protons; over larger distances and outside the nucleus, it is essentially nonexistent.

Link to Learning

Visit this website for more information about the four fundamental forces.

Nuclear Binding Energy

As a simple example of the energy associated with the strong nuclear force, consider the helium atom composed of two protons, two neutrons, and two electrons. The total mass of these six subatomic particles may be calculated as:

$$\underbrace{(2 \times 1.0073\ \text{amu})}_{\text{protons}} + \underbrace{(2 \times 1.0087\ \text{amu})}_{\text{neutrons}} + \underbrace{(2 \times 0.00055\ \text{amu})}_{\text{electrons}} = 4.0331\ \text{amu}$$

However, mass spectrometric measurements reveal that the mass of an atom is 4.0026 amu, less than the combined masses of its six constituent subatomic particles. This difference between the calculated and experimentally measured masses is known as the mass defect of the atom. In the case of helium, the **mass defect** indicates a "loss" in mass of 4.0331 amu – 4.0026 amu = 0.0305 amu. The loss in mass accompanying the formation of an atom from protons, neutrons, and electrons is due to the conversion of that mass into energy that is evolved as the atom forms. The **nuclear binding energy** is the energy produced when the atoms' nucleons are bound together; this is also the energy needed to break a nucleus into its constituent protons and neutrons. In comparison to chemical bond energies, nuclear binding energies are *vastly* greater, as we will learn in this section. Consequently, the energy changes associated with nuclear reactions are vastly greater than are those for chemical reactions.

The conversion between mass and energy is most identifiably represented by the **mass-energy equivalence equation** as stated by Albert Einstein:

$$E = mc^2 \tag{11.2.1}$$

where E is energy, m is mass of the matter being converted, and c is the speed of light in a vacuum. This equation can be used to find the amount of energy that results when matter is converted into energy. Using this mass-energy equivalence equation, the nuclear binding energy of a nucleus may be calculated from its mass defect, as demonstrated in Example 11.2.2 A variety of units are commonly used for nuclear binding energies, including electron volts (eV), with 1 eV equaling the amount of energy necessary to the move the charge of an electron across an electric potential difference of 1 volt, making 1 eV = 1.602 10^{-19} J.

Example 11.2.2: Calculation of Nuclear Binding Energy

Determine the binding energy for the nuclide in:

a. joules per mole of nuclei
b. joules per nucleus

c. MeV per nucleus

Solution

The mass defect for a ^4_2He nucleus is 0.0305 amu, as shown previously. Determine the binding energy in joules per nuclide using the mass-energy equivalence equation. To accommodate the requested energy units, the mass defect must be expressed in kilograms (recall that 1 J = 1 kg m^2/s^2).

a. First, express the mass defect in g/mol. This is easily done considering the *numerical equivalence* of atomic mass (amu) and molar mass (g/mol) that results from the definitions of the amu and mole units (refer to the previous discussion in the chapter on atoms, molecules, and ions if needed). The mass defect is therefore 0.0305 g/mol. To accommodate the units of the other terms in the mass-energy equation, the mass must be expressed in kg, since 1 J = 1 kg m^2/s^2. Converting grams into kilograms yields a mass defect of $3.05 \times 10^{-5}\ \text{kg/mol}$. Substituting this quantity into the mass-energy equivalence equation yields:

$$E = mc^2 = \frac{3.05 \times 10^{-5}\,kg}{mol} \times \left(\frac{2.998 \times 10^8 m}{s}\right)^2 = 2.74 \times 10^{12}\ \text{kg m}^2\text{s}^{-2}\text{mol}^{-1}$$

$$= 2.74 \times 10^{12}\ \text{J mol}^{-1} = 2.74\ \text{TJ mol}^{-1}$$

Note that this tremendous amount of energy is associated with the conversion of a very small amount of matter (about 30 mg, roughly the mass of typical drop of water).

b. The binding energy for a single nucleus is computed from the molar binding energy using Avogadro's number:

$$E = 2.74 \times 10^{12}\ \text{J mol}^{-1} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{nuclei}}$$

$$= 4.55 \times 10^{-12}\ \text{J} = 4.55\ \text{pJ}$$

c. Recall that $1\ \text{eV} = 1.602 \times 10^{-19}\ \text{J}$. Using the binding energy computed in part (b):

$$E = 4.55 \times 10^{-12}\ \text{J} \times \frac{1\ \text{eV}}{1.602 \times 10^{-19}\ \text{J}} = 2.84 \times 10^7\ \text{eV} = 28.4\ \text{MeV}$$

? Exercise 11.2.1

What is the binding energy for the $^{19}_9\text{F}$ nuclide (atomic mass: 18.9984 amu) in MeV per nucleus?

Answer

148.4 MeV

Because the energy changes for breaking and forming bonds are so small compared to the energy changes for breaking or forming nuclei, the changes in mass during all ordinary chemical reactions are virtually undetectable. As described in the chapter on thermochemistry, the most energetic chemical reactions exhibit enthalpies on the order of *thousands* of kJ/mol, which is equivalent to mass differences in the nanogram range (10^{-9} g). On the other hand, nuclear binding energies are typically on the order of *billions* of kJ/mol, corresponding to mass differences in the milligram range (10^{-3} g).

Nuclear Stability

A nucleus is stable if it cannot be transformed into another configuration without adding energy from the outside. Of the thousands of nuclides that exist, about 250 are stable. A plot of the number of neutrons versus the number of protons for stable nuclei reveals that the stable isotopes fall into a narrow band. This region is known as the **band of stability** (also called the belt, zone, or valley of stability). The straight line in Figure 11.2.1 represents nuclei that have a 1:1 ratio of protons to neutrons (n:p ratio). Note that the lighter stable nuclei, in general, have equal numbers of protons and neutrons. For example, nitrogen-14 has seven protons and seven neutrons. Heavier stable nuclei, however, have increasingly more neutrons than protons. For example: iron-56 has 30 neutrons and 26 protons, an n:p ratio of 1.15, whereas the stable nuclide lead-207 has 125 neutrons and 82 protons, an n:p ratio equal to 1.52. This is because larger nuclei have more proton-proton repulsions, and require larger numbers of neutrons to provide compensating strong forces to overcome these electrostatic repulsions and hold the nucleus together.

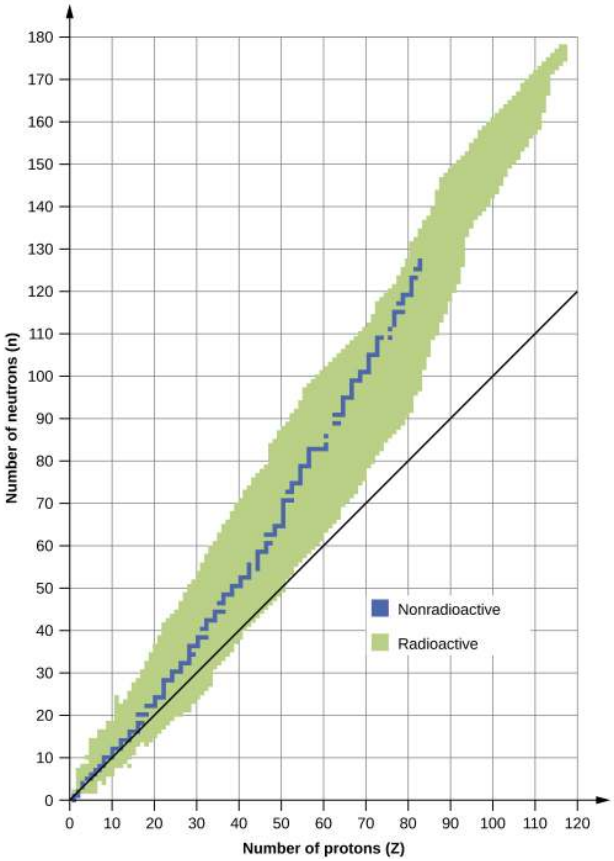


Figure 11.2.1: This plot shows the nuclides that are known to exist and those that are stable. The stable nuclides are indicated in blue, and the unstable nuclides are indicated in green. Note that all isotopes of elements with atomic numbers greater than 83 are unstable. The solid line is the line where n = Z.

The nuclei that are to the left or to the right of the band of stability are unstable and exhibit **radioactivity**. They change spontaneously (decay) into other nuclei that are either in, or closer to, the band of stability. These nuclear decay reactions convert one unstable isotope (or **radioisotope**) into another, more stable, isotope. We will discuss the nature and products of this radioactive decay in subsequent sections of this chapter.

Several observations may be made regarding the relationship between the stability of a nucleus and its structure. Nuclei with even numbers of protons, neutrons, or both are more likely to be stable (see Table 11.2.1). Nuclei with certain numbers of nucleons, known as **magic numbers**, are stable against nuclear decay. These numbers of protons or neutrons (2, 8, 20, 28, 50, 82, and 126) make complete shells in the nucleus. These are similar in concept to the stable electron shells observed for the noble gases. Nuclei that have magic numbers of both protons and neutrons, such as and are called “double magic” and are particularly stable. These trends in nuclear stability may be rationalized by considering a quantum mechanical model of nuclear energy states analogous to that used to describe electronic states earlier in this textbook. The details of this model are beyond the scope of this chapter.

Table 11.2.1: Stable Nuclear Isotopes

Number of Stable Isotopes	Proton Number	Neutron Number
157	even	even
53	even	odd
50	odd	even
5	odd	odd

The relative stability of a nucleus is correlated with its **binding energy per nucleon**, the total binding energy for the nucleus divided by the number or nucleons in the nucleus. For instance, we saw in Example 11.2.2 that the binding energy for a $^{4}_{2}\text{He}$

nucleus is 28.4 MeV. The *binding energy per nucleon* for a ^4_2He nucleus is therefore:

In Example 11.2.3, we learn how to calculate the binding energy per nucleon of a nuclide on the curve shown in Figure 11.2.2

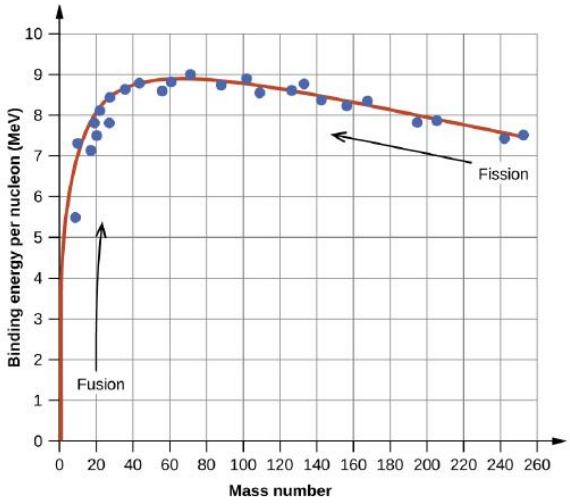


Figure 11.2.2: The binding energy per nucleon is largest for nuclides with mass number of approximately 56.

Example 11.2.3: Calculation of Binding Energy per Nucleon

The iron nuclide atom $^{56}_{26}\text{Fe}$ lies near the top of the binding energy curve (Figure 11.2.2) and is one of the most stable nuclides. What is the binding energy per nucleon (in MeV) for the nuclide $^{56}_{26}\text{Fe}$ (atomic mass of 55.9349 amu)?

Solution

As in Example 11.2.2, we first determine the mass defect of the nuclide, which is the difference between the mass of 26 protons, 30 neutrons, and 26 electrons, and the observed mass of an $^{56}_{26}\text{Fe}$ atom:

$$\begin{aligned}\text{Mass defect} &= [(26 \times 1.0073\ \text{amu}) + (30 \times 1.0087\ \text{amu}) + (26 \times 0.00055\ \text{amu})] - 55.9349 amu \\ &= 56.4651\ \text{amu} - 55.9349\ \text{amu} \\ &= 0.5302\ \text{amu}\end{aligned}$$

We next calculate the binding energy for one nucleus from the mass defect using the mass-energy equivalence equation:

$$\begin{aligned}E &= mc^2 = 0.5302\ \text{amu} \times \left(\frac{1.6605 \times 10^{-27}\ \text{kg}}{1\ \text{amu}}\right) \times \left(2.998 \times 10^8\ \text{m/s}\right)^2 \\ &= 7.913 \times 10^{-11}\ \text{kg} \cdot \text{m/s}^2 \\ &= 7.913 \times 10^{-11}\ \text{J}\end{aligned}$$

We then convert the binding energy in joules per nucleus into units of MeV per nuclide:

$$7.913 \times 10^{-11}\ \text{J} \times \frac{1\ \text{MeV}}{1.602 \times 10^{-13}\ \text{J}} = 493.9\ \text{MeV}$$

Finally, we determine the binding energy per nucleon by dividing the total nuclear binding energy by the number of nucleons in the atom:

Note that this is almost 25% larger than the binding energy per nucleon for ^4_2He.

$$\text{Binding energy per nucleon} = \frac{493.9\ \text{MeV}}{56} = 8.820\ \text{MeV/ nucleon}$$

(Note also that this is the same process as in Example 11.2.1, but with the additional step of dividing the total nuclear binding energy by the number of nucleons.)

Exercise 11.2.1

What is the binding energy per nucleon in $^{19}_{9}\mathrm{F}$ (atomic mass, 18.9984 amu)?

Answer

$7.810\ \mathrm{MeV/nucleon}$

This page titled 11.2: Nuclear Structure and Stability is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **21.1: Nuclear Structure and Stability** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.3: Nuclear Equations

Learning Objectives

By the end of this section, you will be able to:

- Identify common particles and energies involved in nuclear reactions
- Write and balance nuclear equations

Changes of nuclei that result in changes in their atomic numbers, mass numbers, or energy states are nuclear reactions. To describe a nuclear reaction, we use an equation that identifies the nuclides involved in the reaction, their mass numbers and atomic numbers, and the other particles involved in the reaction.

Types of Particles in Nuclear Reactions

Many entities can be involved in nuclear reactions. The most common are protons, neutrons, alpha particles, beta particles, positrons, and gamma rays, as shown in Figure 11.3.1. Protons (^1_1p, also represented by the symbol ^1_1H) and neutrons (^1_0n) are the constituents of atomic nuclei, and have been described previously. **Alpha particles** (^4_2He, also represented by the symbol $^4_2\alpha$) are high-energy helium nuclei. **Beta particles** ($^0_{-1}\beta$, also represented by the symbol ^0_1e) are high-energy electrons, and gamma rays are photons of very high-energy electromagnetic radiation. **Positrons** ($^0_{+1}\text{e}$, also represented by the symbol $^0_1\beta$) are positively charged electrons ("anti-electrons"). The subscripts and superscripts are necessary for balancing nuclear equations, but are usually optional in other circumstances. For example, an alpha particle is a helium nucleus (He) with a charge of +2 and a mass number of 4, so it is symbolized ^4_2He. This works because, in general, the ion charge is not important in the balancing of nuclear equations.

Name	Symbol(s)	Representation	Description
Alpha particle	^4_2He or $^4_2\alpha$		(High-energy) helium nuclei consisting of two protons and two neutrons
Beta particle	$^0_{-1}\text{e}$ or $^0_{-1}\beta$		(High-energy) electrons
Positron	$^0_{+1}\text{e}$ or $^0_{+1}\beta$		Particles with the same mass as an electron but with 1 unit of positive charge
Proton	^1_1H or ^1_1p		Nuclei of hydrogen atoms
Neutron	^1_0n		Particles with a mass approximately equal to that of a proton but with no charge
Gamma ray	γ		Very high-energy electromagnetic radiation

Figure 11.3.1: Although many species are encountered in nuclear reactions, this table summarizes the names, symbols, representations, and descriptions of the most common of these.

Note that positrons are exactly like electrons, except they have the opposite charge. They are the most common example of **antimatter**, particles with the same mass but the opposite state of another property (for example, charge) than ordinary matter. When antimatter encounters ordinary matter, both are annihilated and their mass is converted into energy in the form of **gamma rays** (γ)—and other much smaller subnuclear particles, which are beyond the scope of this chapter—according to the mass-energy equivalence equation $E = mc^2$, seen in the preceding section. For example, when a positron and an electron collide, both are annihilated and two gamma ray photons are created:

$$^0_{-1}\text{e} + ^0_{+1}\text{e} \longrightarrow \gamma + \gamma$$

As seen in the chapter discussing light and electromagnetic radiation, gamma rays compose short wavelength, high-energy electromagnetic radiation and are (much) more energetic than better-known X-rays that can behave as particles in the wave-particle duality sense. Gamma rays are a type of high energy electromagnetic radiation produced when a nucleus undergoes a transition from a higher to a lower energy state, similar to how a photon is produced by an electronic transition from a higher to a lower

energy level. Due to the much larger energy differences between nuclear energy shells, gamma rays emanating from a nucleus have energies that are typically millions of times larger than electromagnetic radiation emanating from electronic transitions.

Balancing Nuclear Reactions

A balanced chemical reaction equation reflects the fact that during a chemical reaction, bonds break and form, and atoms are rearranged, but the total numbers of atoms of each element are conserved and do not change. A balanced nuclear reaction equation indicates that there is a rearrangement during a nuclear reaction, but of nucleons (subatomic particles within the atoms' nuclei) rather than atoms. Nuclear reactions also follow conservation laws, and they are balanced in two ways:

1. The sum of the mass numbers of the reactants equals the sum of the mass numbers of the products.
2. The sum of the charges of the reactants equals the sum of the charges of the products.

If the atomic number and the mass number of all but one of the particles in a nuclear reaction are known, we can identify the particle by balancing the reaction. For instance, we could determine that ${}^{17}_{8}\mathrm{O}$ is a product of the nuclear reaction of ${}^{14}_{7}\mathrm{N}$ and ${}^{4}_{2}\mathrm{He}$ if we knew that a proton, ${}^{1}_{1}\mathrm{H}$, were one of the two products. Example 11.3.1 shows how we can identify a nuclide by balancing the nuclear reaction.

Example 11.3.1: Balancing Equations for Nuclear Reactions

The reaction of an α particle with magnesium-25 (${}^{25}_{12}\mathrm{Mg}$) produces a proton and a nuclide of another element. Identify the new nuclide produced.

Solution

The nuclear reaction can be written as:

$$ {}^{25}_{12}\mathrm{Mg} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{1}_{1}\mathrm{H} + {}^{\mathrm{A}}_{\mathrm{Z}}\mathrm{X} $$

where A is the mass number and Z is the atomic number of the new nuclide, X. Because the sum of the mass numbers of the reactants must equal the sum of the mass numbers of the products:

$$ 25 + 4 = A + 1 $$

or $A = 28$.

Similarly, the charges must balance, so:

$$ 12 + 2 = Z + 1 $$

and $Z = 13$.

Check the periodic table: The element with nuclear charge = +13 is aluminum. Thus, the product is ${}^{28}_{13}\mathrm{Al}$

Exercise 11.3.1

The nuclide ${}^{125}_{53}\mathrm{I}$ combines with an electron and produces a new nucleus and no other massive particles. What is the equation for this reaction?

Answer

$$ {}^{125}_{53}\mathrm{I} + {}^{0}_{-1}\mathrm{e} \longrightarrow {}^{125}_{52}\mathrm{Te} $$

Following are the equations of several nuclear reactions that have important roles in the history of nuclear chemistry:

- The first naturally occurring unstable element that was isolated, polonium, was discovered by the Polish scientist Marie Curie and her husband Pierre in 1898. It decays, emitting α particles:

$$ {}^{212}_{84}\mathrm{Po} \longrightarrow {}^{208}_{82}\mathrm{Pb} + {}^{4}_{2}\mathrm{He} $$

- The first nuclide to be prepared by artificial means was an isotope of oxygen, ${}^{17}\mathrm{O}$. It was made by Ernest Rutherford in 1919 by bombarding nitrogen atoms with α particles:

$$^{14}_{7}\mathrm{N} + ^{4}_{2}\mathrm{He} \longrightarrow ^{17}_{8}\mathrm{O} + ^{1}_{1}\mathrm{H}$$

- James Chadwick discovered the neutron in 1932, as a previously unknown neutral particle produced along with ^{12}C by the nuclear reaction between ^{9}Be and ^{4}He:

$$^{9}_{4}\mathrm{Be} + ^{4}_{2}\mathrm{He} \longrightarrow ^{12}_{6}\mathrm{C} + ^{1}_{0}\mathrm{n}$$

- The first element to be prepared that does not occur naturally on the earth, technetium, was created by bombardment of molybdenum by deuterons (heavy hydrogen, $^{2}_{1}H$), by Emilio Segre and Carlo Perrier in 1937:

$$^{2}_{1}\mathrm{H} + ^{97}_{42}\mathrm{Mo} \longrightarrow 2\,^{1}_{0}\mathrm{n} + ^{97}_{43}\mathrm{Tc}$$

- The first controlled nuclear chain reaction was carried out in a reactor at the University of Chicago in 1942. One of the many reactions involved was:

$$^{235}_{92}\mathrm{U} + ^{1}_{0}\mathrm{n} \longrightarrow ^{87}_{35}\mathrm{Br} + ^{146}_{57}\mathrm{La} + 3\,^{1}_{0}\mathrm{n}$$

This page titled 11.3: Nuclear Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.2: Nuclear Equations** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.4: Radioactive Decay

Learning Objectives

By the end of this section, you will be able to:

- Recognize common modes of radioactive decay
- Identify common particles and energies involved in nuclear decay reactions
- Write and balance nuclear decay equations
- Calculate kinetic parameters for decay processes, including half-life
- Describe common radiometric dating techniques

Following the somewhat serendipitous discovery of radioactivity by Becquerel, many prominent scientists began to investigate this new, intriguing phenomenon. Among them were Marie Curie (the first woman to win a Nobel Prize, and the only person to win two Nobel Prizes in different sciences—chemistry and physics), who was the first to coin the term "radioactivity," and Ernest Rutherford (of gold foil experiment fame), who investigated and named three of the most common types of radiation. During the beginning of the twentieth century, many radioactive substances were discovered, the properties of radiation were investigated and quantified, and a solid understanding of radiation and nuclear decay was developed.

The spontaneous change of an unstable nuclide into another is **radioactive decay**. The unstable nuclide is called the **parent nuclide**; the nuclide that results from the decay is known as the **daughter nuclide**. The daughter nuclide may be stable, or it may decay itself. The radiation produced during radioactive decay is such that the daughter nuclide lies closer to the band of stability than the parent nuclide, so the location of a nuclide relative to the band of stability can serve as a guide to the kind of decay it will undergo (Figure 11.4.1).

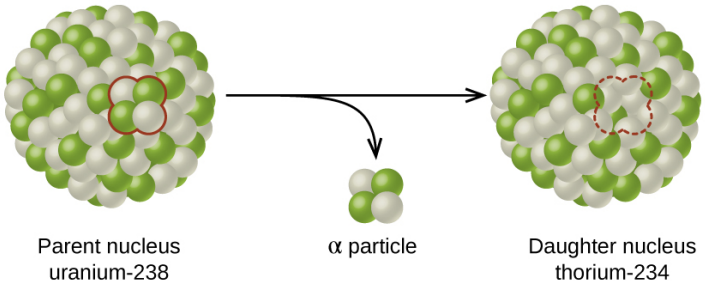


Figure 11.4.1: A nucleus of uranium-238 (the parent nuclide) undergoes α decay to form thorium-234 (the daughter nuclide). The alpha particle removes two protons (green) and two neutrons (gray) from the uranium-238 nucleus.

Link to Learning

Although the radioactive decay of a nucleus is too small to see with the naked eye, we can indirectly view radioactive decay in an environment called a cloud chamber. Click here to learn about cloud chambers and to view an interesting Cloud Chamber Demonstration from the Jefferson Lab.

Types of Radioactive Decay

Ernest Rutherford's experiments involving the interaction of radiation with a magnetic or electric field (Figure 11.4.2) helped him determine that one type of radiation consisted of positively charged and relatively massive α particles; a second type was made up of negatively charged and much less massive β particles; and a third was uncharged electromagnetic waves, γ rays. We now know that α particles are high-energy helium nuclei, β particles are high-energy electrons, and γ radiation compose high-energy electromagnetic radiation. We classify different types of radioactive decay by the radiation produced.

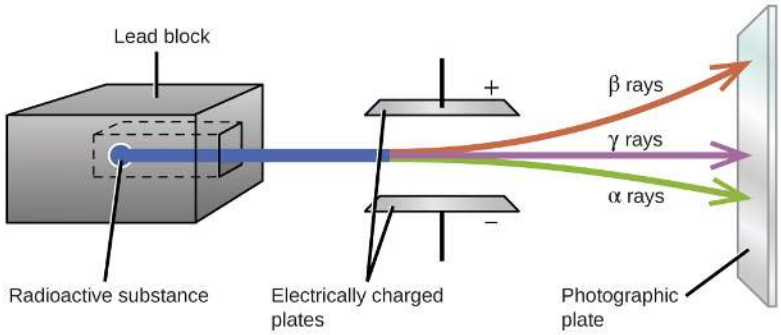


Figure 11.4.2: Alpha particles, which are attracted to the negative plate and deflected by a relatively small amount, must be positively charged and relatively massive. Beta particles, which are attracted to the positive plate and deflected a relatively large amount, must be negatively charged and relatively light. Gamma rays, which are unaffected by the electric field, must be uncharged.

Alpha (α) decay is the emission of an α particle from the nucleus. For example, polonium-210 undergoes α decay:

$$^{210}_{84}\mathrm{Po} \longrightarrow ^{4}_{2}\mathrm{He} + ^{206}_{82}\mathrm{Pb}$$

or

$$^{210}_{84}\mathrm{Po} \longrightarrow ^{4}_{2}\alpha + ^{206}_{82}\mathrm{Pb}$$

Alpha decay occurs primarily in heavy nuclei ($A > 200$, $Z > 83$). Because the loss of an α particle gives a daughter nuclide with a mass number four units smaller and an atomic number two units smaller than those of the parent nuclide, the daughter nuclide has a larger n:p ratio than the parent nuclide. If the parent nuclide undergoing α decay lies below the band of stability (refer to Figure 21.2), the daughter nuclide will lie closer to the band.

Beta (β) decay is the emission of an electron (i.e., a β particle) from a nucleus. Iodine-131 is an example of a nuclide that undergoes β decay:

$$^{131}_{53}\mathrm{I} \longrightarrow ^{0}_{-1}\mathrm{e} + ^{131}_{54}\mathrm{Xe}$$

or

$$^{131}_{53}\mathrm{I} \longrightarrow ^{0}_{-1}\beta + ^{131}_{54}\mathrm{Xe}$$

Beta decay, which can be thought of as the conversion of a neutron into a proton and a β particle, is observed in nuclides with a large n:p ratio. The beta particle (electron) emitted is from the atomic nucleus and is not one of the electrons surrounding the nucleus. Such nuclei lie above the band of stability. Emission of an electron does not change the mass number of the nuclide but does increase the number of its protons and decrease the number of its neutrons. Consequently, the n:p ratio is decreased, and the daughter nuclide lies closer to the band of stability than did the parent nuclide.

Gamma emission (γ emission) is observed when a nuclide is formed in an excited state and then decays to its ground state with the emission of a γ ray, a quantum of high-energy electromagnetic radiation. The presence of a nucleus in an excited state is often indicated by an asterisk (*). Cobalt-60 emits γ radiation and is used in many applications including cancer treatment:

$$^{60}_{27}\mathrm{Co}^{\cdot} \longrightarrow ^{0}_{0}\gamma + ^{60}_{27}\mathrm{Co}$$

There is no change in mass number or atomic number during the emission of a γ ray unless the γ emission accompanies one of the other modes of decay.

Positron emission (β^{+} decay) is the emission of a positron from the nucleus. Oxygen-15 is an example of a nuclide that undergoes positron emission:

$$^{15}_{8}\mathrm{O} \longrightarrow ^{0}_{+1}\mathrm{e} + ^{15}_{7}\mathrm{N}$$

or

$$^{15}_{8}\mathrm{O} \longrightarrow ^{0}_{+1}\beta + ^{15}_{7}\mathrm{N}$$

Positron emission is observed for nuclides in which the n:p ratio is low. These nuclides lie below the band of stability. Positron decay is the conversion of a proton into a neutron with the emission of a positron. The n:p ratio increases, and the daughter nuclide lies closer to the band of stability than did the parent nuclide.

Electron capture occurs when one of the inner electrons in an atom is captured by the atom's nucleus. For example, potassium-40 undergoes electron capture:

$$^{40}_{19}\mathrm{K} + ^{0}_{-1}\mathrm{e} \longrightarrow ^{40}_{18}\mathrm{Ar}$$

Electron capture occurs when an inner shell electron combines with a proton and is converted into a neutron. The loss of an inner shell electron leaves a vacancy that will be filled by one of the outer electrons. As the outer electron drops into the vacancy, it will emit energy. In most cases, the energy emitted will be in the form of an X-ray. Like positron emission, electron capture occurs for "proton-rich" nuclei that lie below the band of stability. Electron capture has the same effect on the nucleus as does positron emission: The atomic number is decreased by one and the mass number does not change. This increases the n:p ratio, and the daughter nuclide lies closer to the band of stability than did the parent nuclide. Whether electron capture or positron emission occurs is difficult to predict. The choice is primarily due to kinetic factors, with the one requiring the smaller activation energy being the one more likely to occur.

Figure 11.4.3 summarizes these types of decay, along with their equations and changes in atomic and mass numbers.

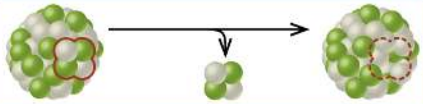
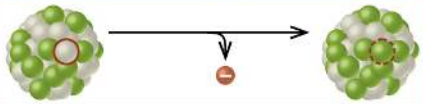
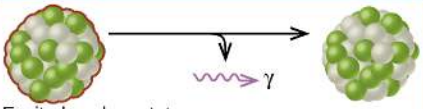
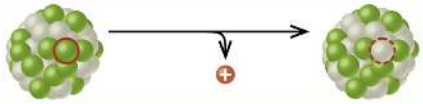
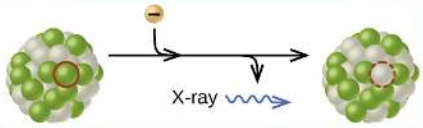
Type	Nuclear equation	Representation	Change in mass/atomic numbers
Alpha decay	$^{A}_{Z}\mathrm{X} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{A-4}_{Z-2}\mathrm{Y}$		A: decrease by 4 Z: decrease by 2
Beta decay	$^{A}_{Z}\mathrm{X} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{A}_{Z+1}\mathrm{Y}$		A: unchanged Z: increase by 1
Gamma decay	$^{A}_{Z}\mathrm{X} \rightarrow {}^{0}_{0}\gamma + {}^{A}_{Z}\mathrm{Y}$		A: unchanged Z: unchanged
Positron emission	$^{A}_{Z}\mathrm{X} \rightarrow {}^{0}_{+1}\mathrm{e} + {}^{A}_{Y-1}\mathrm{Y}$		A: unchanged Z: decrease by 1
Electron capture	$^{A}_{Z}\mathrm{X} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{A}_{Y-1}\mathrm{Y}$		A: unchanged Z: decrease by 1

Figure 11.4.3: This table summarizes the type, nuclear equation, representation, and any changes in the mass or atomic numbers for various types of decay.

Chemistry in Everyday Life: PET Scan

Positron emission tomography (PET) scans use radiation to diagnose and track health conditions and monitor medical treatments by revealing how parts of a patient's body function (Figure 11.4.4). To perform a PET scan, a positron-emitting radioisotope is produced in a cyclotron and then attached to a substance that is used by the part of the body being investigated. This "tagged" compound, or radiotracer, is then put into the patient (injected via IV or breathed in as a gas), and how it is used by the tissue reveals how that organ or other area of the body functions.

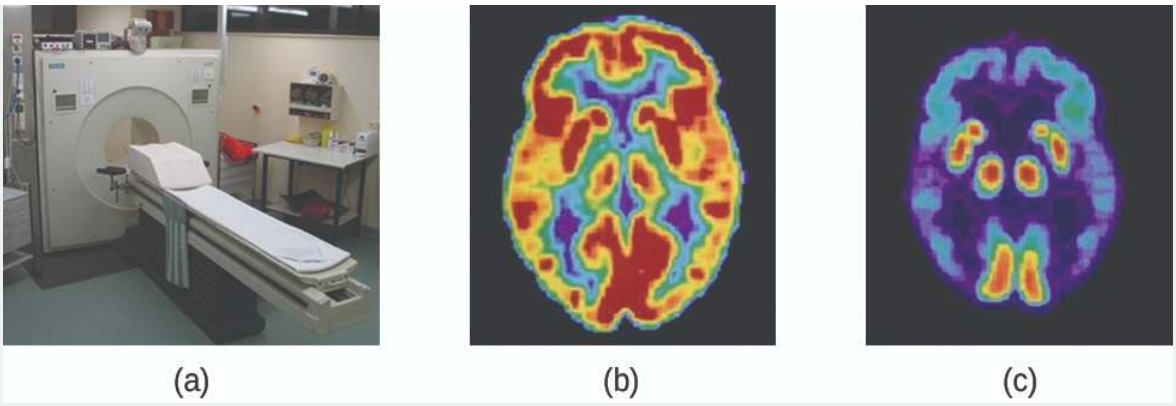


Figure 11.4.4: A PET scanner (a) uses radiation to provide an image of how part of a patient's body functions. The scans it produces can be used to image a healthy brain (b) or can be used for diagnosing medical conditions such as Alzheimer's disease (c). (credit a: modification of work by Jens Maus)

For example, F-18 is produced by proton bombardment of \(\ce{^{18}O}\)

$$^{18}_{8}\mathrm{O} + ^{1}_{1}\mathrm{p} \longrightarrow ^{18}_{9}\mathrm{F} + ^{1}_{0}\mathrm{n}$$

and incorporated into a glucose analog called fludeoxyglucose (FDG). How FDG is used by the body provides critical diagnostic information; for example, since cancers use glucose differently than normal tissues, FDG can reveal cancers. The $^{18}\mathrm{F}$ emits positrons that interact with nearby electrons, producing a burst of gamma radiation. This energy is detected by the scanner and converted into a detailed, three-dimensional, color image that shows how that part of the patient's body functions. Different levels of gamma radiation produce different amounts of brightness and colors in the image, which can then be interpreted by a radiologist to reveal what is going on. PET scans can detect heart damage and heart disease, help diagnose Alzheimer's disease, indicate the part of a brain that is affected by epilepsy, reveal cancer, show what stage it is, and how much it has spread, and whether treatments are effective. Unlike magnetic resonance imaging and X-rays, which only show how something looks, the big advantage of PET scans is that they show how something functions. PET scans are now usually performed in conjunction with a computed tomography scan.

Radioactive Decay Series

The naturally occurring radioactive isotopes of the heaviest elements fall into chains of successive disintegrations, or decays, and all the species in one chain constitute a radioactive family, or **radioactive decay series**. Three of these series include most of the naturally radioactive elements of the periodic table. They are the uranium series, the actinide series, and the thorium series. The neptunium series is a fourth series, which is no longer significant on the earth because of the short half-lives of the species involved. Each series is characterized by a parent (first member) that has a long half-life and a series of daughter nuclides that ultimately lead to a stable end-product—that is, a nuclide on the band of stability (Figure 11.4.5). In all three series, the end-product is a stable isotope of lead. The neptunium series, previously thought to terminate with bismuth-209, terminates with thallium-205.

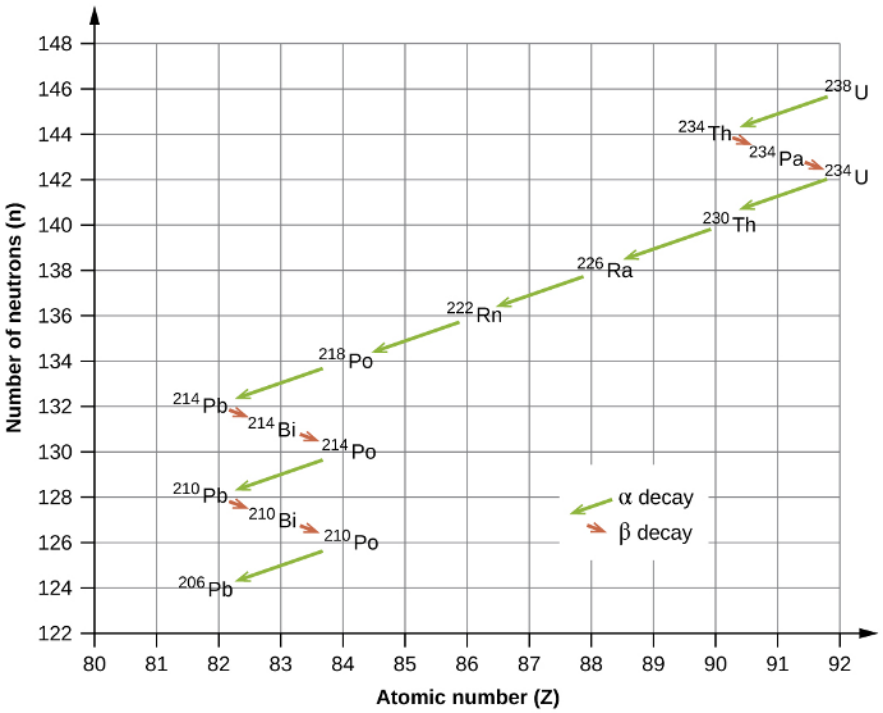


Figure 11.4.5: Uranium-238 undergoes a radioactive decay series consisting of 14 separate steps before producing stable lead-206. This series consists of eight α decays and six β decays.

Radioactive Half-Lives

Radioactive decay follows first-order kinetics. Since first-order reactions have already been covered in detail in the kinetics chapter, we will now apply those concepts to nuclear decay reactions. Each radioactive nuclide has a characteristic, constant **half-life** ($t_{1/2}$), the time required for half of the atoms in a sample to decay. An isotope's half-life allows us to determine how long a sample of a useful isotope will be available, and how long a sample of an undesirable or dangerous isotope must be stored before it decays to a low-enough radiation level that is no longer a problem.

For example, coba source, since half of the nuclei decay every 5.27 years, both the amount of material and the intensity of the radiation emitted is cut in half every 5.27 years. (Note that for a given substance, the intensity of radiation that it produces is directly proportional to the rate of decay of the substance and the amount of the substance.) This is as expected for a process following first-order kinetics. Thus, a cobalt-60 source that is used for cancer treatment must be replaced regularly to continue to be effective.

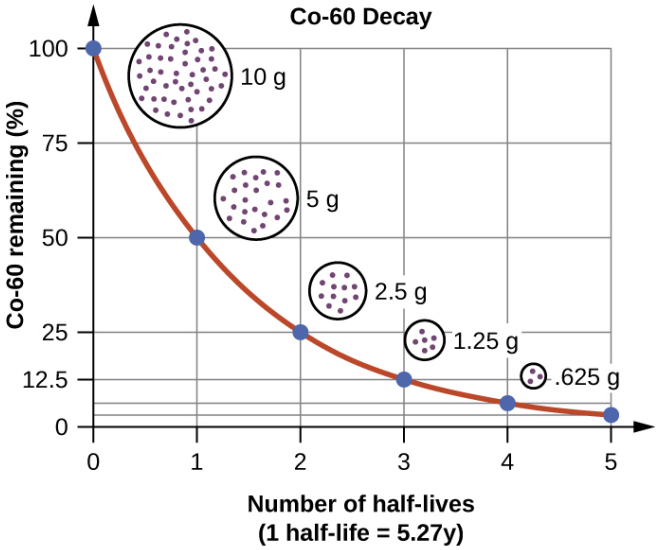


Figure 11.4.6: For cobalt-60, which has a half-life of 5.27 years, 50% remains after 5.27 years (one half-life), 25% remains after 10.54 years (two half-lives), 12.5% remains after 15.81 years (three half-lives), and so on.

Since nuclear decay follows first-order kinetics, we can adapt the mathematical relationships used for first-order chemical reactions. We generally substitute the number of nuclei, N, for the concentration. If the rate is stated in nuclear decays per second, we refer to it as the **activity** of the radioactive sample. The rate for radioactive decay is:

$$\text{decay rate} = \lambda N$$

with λ = the decay constant for the particular radioisotope.

The decay constant, λ, which is the same as a rate constant discussed in the kinetics chapter. It is possible to express the decay constant in terms of the half-life, $t_{1/2}$:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{t_{1/2}}$$

or

$$t_{1/2} = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda}$$

The first-order equations relating amount, N, and time are:

$$N_t = N_0 e^{-\lambda t} \tag{11.4.1}$$

or

$$t = -\frac{1}{\lambda} \ln\left(\frac{N_t}{N_0}\right)$$

Example 11.4.1: Rates of Radioactive Decay

$^{60}_{27}\text{Co}$ decays with a half-life of 5.27 years to produce $^{60}_{28}\text{Co}$

a. What is the decay constant for the radioactive disintegration of cobalt-60?
b. Calculate the fraction of a sample of the $^{60}_{27}\text{Co}$ isotope that will remain after 15 years.
c. How long does it take for a sample of $^{60}_{27}\text{Co}$ to disintegrate to the extent that only 2.0% of the original amount remains?

Solution

(a) The value of the rate constant is given by:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{5.27y} = 0.132y^{-1}$$

(b) The fraction of $^{60}_{27}\text{Co}$ that is left after time t is given by $\frac{N_t}{N_0}$. Rearranging the first-order relationship in Equation 11.4.1 ($N_t = N_0 e^{-\lambda t}$ to solve for this ratio yields:

$$\frac{N_t}{N_0} = e^{-\lambda t} = e^{-(0.132/\text{y})(15\ \text{y})} = 0.138$$

The fraction of $^{60}_{27}\text{Co}$ that will remain after 15.0 years is 0.138. Or put another way, 13.8% of the original $^{60}_{27}\text{Co}$ present will remain after 15 years.

(c) 2.00% of the original amount of $^{60}_{27}\text{Co}$ is equal to 0.0200 N_0. Substituting this into the equation for time for first-order kinetics, we have:

$$t = -\frac{1}{\lambda}\ln\left(\frac{N_t}{N_0}\right) = -\frac{1}{0.132\ \text{y}^{-1}}\ln\left(\frac{0.0200 \times N_0}{N_0}\right) = 29.6y$$

? Exercise 11.4.1

Radon-222, $^{222}_{86}\text{Rn}$ has a half-life of 3.823 days. How long will it take a sample of radon-222 with a mass of 0.750 g to decay into other elements, leaving only 0.100 g of radon-222?

Answer

11.1 days

Because each nuclide has a specific number of nucleons, a particular balance of repulsion and attraction, and its own degree of stability, the half-lives of radioactive nuclides vary widely. For example: the half-life of $^{209}_{83}\text{Bi}$ is 1.9×10^{19} years; $^{239}_{94}\text{Ra}$ is 24,000 years; $^{222}_{86}\text{Rn}$ is 3.82 days; and element-111 (Rg for roentgenium) is 1.5×10^{-3} seconds. The half-lives of a number of radioactive isotopes important to medicine are shown in Table 11.4.1, and others are listed in Appendix M.

Table 11.4.1: Half-lives of Radioactive Isotopes Important to Medicine

Type[1]	Decay Mode	Half-Life	Uses
F-18	β^+ decay	110. minutes	PET scans
Co-60	β decay, γ decay	5.27 years	cancer treatment
Tc-99m	γ decay	8.01 hours	scans of brain, lung, heart, bone
I-131	β decay	8.02 days	thyroid scans and treatment
Tl-201	electron capture	73 hours	heart and arteries scans; cardiac stress tests

Radiometric Dating

Several radioisotopes have half-lives and other properties that make them useful for purposes of “dating” the origin of objects such as archaeological artifacts, formerly living organisms, or geological formations. This process is **radiometric dating** and has been responsible for many breakthrough scientific discoveries about the geological history of the earth, the evolution of life, and the history of human civilization. We will explore some of the most common types of radioactive dating and how the particular isotopes work for each type.

Radioactive Dating Using Carbon-14

The radioactivity of carbon-14 provides a method for dating objects that were a part of a living organism. This method of radiometric dating, which is also called **radiocarbon dating** or carbon-14 dating, is accurate for dating carbon-containing substances that are up to about 30,000 years old, and can provide reasonably accurate dates up to a maximum of about 50,000 years old.

Naturally occurring carbon consists of three isotopes: $^{12}_{6}\text{C}$ which constitutes about 99% of the carbon on earth; $^{13}_{6}\text{C}$, about 1% of the total; and trace amounts of $^{14}_{6}\text{C}$. Carbon-14 forms in the upper atmosphere by the reaction of nitrogen atoms with neutrons from

cosmic rays in space:

$$^{14}_{7}\mathrm{N} + ^{1}_{0}\mathrm{n} \longrightarrow ^{14}_{6}\mathrm{C} + ^{1}_{1}\mathrm{H}$$

All isotopes of carbon react with oxygen to produce $\mathrm{CO_2}$ molecules. The ratio of $^{14}_{6}\mathrm{CO_2}$ to $^{12}_{6}\mathrm{CO_2}$ depends on the ratio of $^{14}_{6}\mathrm{CO}$ to $^{12}_{6}\mathrm{CO}$ in the atmosphere. The natural abundance of $^{14}_{6}\mathrm{CO}$ in the atmosphere is approximately 1 part per trillion; until recently, this has generally been constant over time, as seen is gas samples found trapped in ice. The incorporation of $^{14}_{6}\mathrm{CO_2}$ and $^{14}_{6}\mathrm{CO_2}$ into plants is a regular part of the photosynthesis process, which means that the $^{14}_{6}\mathrm{C} : ^{12}_{6}\mathrm{C}$ ratio found in a living plant is the **same** as the $^{14}_{6}\mathrm{C} : ^{12}_{6}\mathrm{C}$ ratio in the atmosphere. But when the plant dies, it no longer traps carbon through photosynthesis. Because $^{12}_{6}\mathrm{C}$ is a stable isotope and does not undergo radioactive decay, its concentration in the plant does not change. However, carbon-14 decays by β emission with a half-life of 5730 years:

$$^{14}_{6}\mathrm{C} \longrightarrow ^{14}_{7}\mathrm{N} + ^{0}_{-1}\mathrm{e}$$

Thus, the ratio gradually decreases after the plant dies. The decrease in the ratio with time provides a measure of the time that has elapsed since the death of the plant (or other organism that ate the plant). Figure 11.4.7 visually depicts this process.

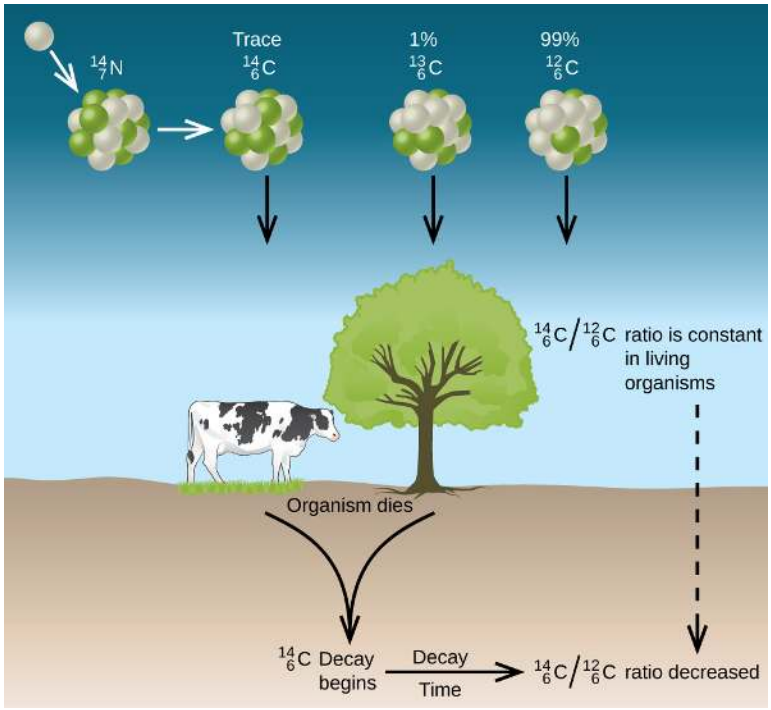


Figure 11.4.7: Along with stable carbon-12, radioactive carbon-14 is taken in by plants and animals, and remains at a constant level within them while they are alive. After death, the C-14 decays and the C-14:C-12 ratio in the remains decreases. Comparing this ratio to the C-14:C-12 ratio in living organisms allows us to determine how long ago the organism lived (and died).

For example, with the half-life of being 5730 years, if the ratio in a wooden object found in an archaeological dig is half what it is in a living tree, this indicates that the wooden object is 5730 years old. Highly accurate determinations of ratios can be obtained from very small samples (as little as a milligram) by the use of a mass spectrometer.

Link to Learning

Visit this website to perform simulations of radiometric dating.

Example 11.4.2: Radiocarbon Dating

A tiny piece of paper (produced from formerly living plant matter) taken from the Dead Sea Scrolls has an activity of 10.8 disintegrations per minute per gram of carbon. If the initial C-14 activity was 13.6 disintegrations/min/g of C, estimate the age of the Dead Sea Scrolls.

Solution

The rate of decay (number of disintegrations/minute/gram of carbon) is proportional to the amount of radioactive C-14 left in the paper, so we can substitute the rates for the amounts, N, in the relationship:

$$t = -\frac{1}{\lambda}\ln\left(\frac{N_t}{N_0}\right) \longrightarrow t = -\frac{1}{\lambda}\ln\left(\frac{\text{Rate}_t}{\text{Rate}_0}\right)$$

where the subscript 0 represents the time when the plants were cut to make the paper, and the subscript t represents the current time.

The decay constant can be determined from the half-life of C-14, 5730 years:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{5730\ \text{y}} = 1.21 \times 10^{-4}\ \text{y}^{-1}$$

Substituting and solving, we have:

$$t = -\frac{1}{\lambda}\ln\left(\frac{\text{Rate}_t}{\text{Rate}_0}\right) = -\frac{1}{1.21 \times 10^{-4}\ \text{y}^{-1}}\ln\left(\frac{10.8\ \text{dis/min/g C}}{13.6\ \text{dis/min/g C}}\right) = 1910\ \text{y}$$

Therefore, the Dead Sea Scrolls are approximately 1900 years old (Figure 11.4.8).



Figure 11.4.8: Carbon-14 dating has shown that these pages from the Dead Sea Scrolls were written or copied on paper made from plants that died between 100 BC and AD 50.

? Exercise 11.4.1

More accurate dates of the reigns of ancient Egyptian pharaohs have been determined recently using plants that were preserved in their tombs. Samples of seeds and plant matter from King Tutankhamun's tomb have a C-14 decay rate of 9.07 disintegrations/min/g of C. How long ago did King Tut's reign come to an end?

Answer

about 3350 years ago, or approximately 1340 BC

There have been some significant, well-documented changes to the ${}^{14}_{6}\text{C} : {}^{12}_{6}\text{C}$ ratio. The accuracy of a straightforward application of this technique depends on the ${}^{14}_{6}\text{C} : {}^{12}_{6}\text{C}$ ratio in a living plant being the same now as it was in an earlier era, but this is not always valid. Due to the increasing accumulation of CO_2 molecules (largely ${}^{12}_{6}\text{C}$) in the atmosphere caused by combustion of fossil fuels (in which essentially all of the ${}^{14}_{6}\text{C}$ has decayed), the ratio of ${}^{14}_{6}\text{C} : {}^{12}_{6}\text{C}$ in the atmosphere may be changing. This manmade increase in ${}^{12}_{6}\text{C}$ in the atmosphere causes the ${}^{14}_{6}\text{C} : {}^{12}_{6}\text{C}$ ratio to decrease, and this in turn affects the ratio in currently living organisms on the earth. Fortunately, however, we can use other data, such as tree dating via examination of annual growth rings, to calculate correction factors. With these correction factors, accurate dates can be determined. In general, radioactive dating only works for about 10 half-lives; therefore, the limit for carbon-14 dating is about 57,000 years.

Radioactive Dating Using Nuclides Other than Carbon-14

Radioactive dating can also use other radioactive nuclides with longer half-lives to date older events. For example, uranium-238 (which decays in a series of steps into lead-206) can be used for establishing the age of rocks (and the approximate age of the oldest rocks on earth). Since U-238 has a half-life of 4.5 billion years, it takes that amount of time for half of the original U-238 to decay into Pb-206. In a sample of rock that does not contain appreciable amounts of Pb-208, the most abundant isotope of lead, we can

assume that lead was not present when the rock was formed. Therefore, by measuring and analyzing the ratio of U-238:Pb-206, we can determine the age of the rock. This assumes that all of the lead-206 present came from the decay of uranium-238. If there is additional lead-206 present, which is indicated by the presence of other lead isotopes in the sample, it is necessary to make an adjustment. Potassium-argon dating uses a similar method. K-40 decays by positron emission and electron capture to form Ar-40 with a half-life of 1.25 billion years. If a rock sample is crushed and the amount of Ar-40 gas that escapes is measured, determination of the Ar-40:K-40 ratio yields the age of the rock. Other methods, such as rubidium-strontium dating (Rb-87 decays into Sr-87 with a half-life of 48.8 billion years), operate on the same principle. To estimate the lower limit for the earth's age, scientists determine the age of various rocks and minerals, making the assumption that the earth is older than the oldest rocks and minerals in its crust. As of 2014, the oldest known rocks on earth are the Jack Hills zircons from Australia, found by uranium-lead dating to be almost 4.4 billion years old.

Example 11.4.3: Radioactive Dating of Rocks

An igneous rock contains 9.58×10^{-5} g of U-238 and 2.51×10^{-5} g of Pb-206, and much, much smaller amounts of Pb-208. Determine the approximate time at which the rock formed.

Solution

The sample of rock contains very little Pb-208, the most common isotope of lead, so we can safely assume that all the Pb-206 in the rock was produced by the radioactive decay of U-238. When the rock formed, it contained all of the U-238 currently in it, plus some U-238 that has since undergone radioactive decay.

The amount of U-238 currently in the rock is:

$$9.58 \times 10^{-5} \ \cancel{\text{g U}} \times \left(\frac{1 \text{ mol U}}{238 \ \cancel{\text{g U}}} \right) = 4.03 \times 10^{-7} \text{ mol U}$$

Because when one mole of U-238 decays, it produces one mole of Pb-206, the amount of U-238 that has undergone radioactive decay since the rock was formed is:

$$2.51 \times 10^{-5} \ \cancel{\text{g Pb}} \times \left(\frac{1 \ \cancel{\text{mol Pb}}}{206 \ \cancel{\text{g Pb}}} \right) \times \left(\frac{1 \text{ mol U}}{1 \ \cancel{\text{mol Pb}}} \right) = 1.22 \times 10^{-7} \text{ mol U}$$

The total amount of U-238 originally present in the rock is therefore:

$$4.03 \times 10^{-7} \text{ mol U} + 1.22 \times 10^{-7} \text{ mol U} = 5.25 \times 10^{-7} \text{ mol U}$$

The amount of time that has passed since the formation of the rock is given by:

$$t = -\frac{1}{\lambda} \ln \left(\frac{N_t}{N_0} \right)$$

with N_0 representing the original amount of U-238 and N_t representing the present amount of U-238.

U-238 decays into Pb-206 with a half-life of 4.5×10^9 y, so the decay constant λ is:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{4.5 \times 10^9 \text{ y}} = 1.54 \times 10^{-10} \text{ y}^{-1}$$

Substituting and solving, we have:

$$t = -\frac{1}{1.54 \times 10^{-10} \text{ y}^{-1}} \ln \left(\frac{4.03 \times 10^{-7} \ \cancel{\text{mol U}}}{5.25 \times 10^{-7} \ \cancel{\text{mol U}}} \right) = 1.7 \times 10^9 \text{ y}$$

Therefore, the rock is approximately 1.7 billion years old.

Exercise 11.4.1

A sample of rock contains 6.14×10^{-4} g of Rb-87 and $3.51 \times 10^{-5} g$ of Sr-87. Calculate the age of the rock. (The half-life of the β decay of Rb-87 is 4.7×10^{10} y.)

Answer

3.7 10^9 y

Footnotes

- 1The “m” in Tc-99m stands for “metastable,” indicating that this is an unstable, high-energy state of Tc-99. Metastable isotopes emit γ radiation to rid themselves of excess energy and become (more) stable.

This page titled 11.4: Radioactive Decay is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **21.3: Radioactive Decay** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.5: Transmutation and Nuclear Energy

Learning Objectives

By the end of this section, you will be able to:

- Describe the synthesis of transuranium nuclides
- Explain nuclear fission and fusion processes
- Relate the concepts of critical mass and nuclear chain reactions
- Summarize basic requirements for nuclear fission and fusion reactors

After the discovery of radioactivity, the field of nuclear chemistry was created and developed rapidly during the early twentieth century. A slew of new discoveries in the 1930s and 1940s, along with World War II, combined to usher in the Nuclear Age in the mid-twentieth century. Scientists learned how to create new substances, and certain isotopes of certain elements were found to possess the capacity to produce unprecedented amounts of energy, with the potential to cause tremendous damage during war, as well as produce enormous amounts of power for society's needs during peace.

Synthesis of Nuclides

Nuclear transmutation is the conversion of one nuclide into another. It can occur by the radioactive decay of a nucleus, or the reaction of a nucleus with another particle. The first manmade nucleus was produced in Ernest Rutherford's laboratory in 1919 by a transmutation reaction, the bombardment of one type of nuclei with other nuclei or with neutrons. Rutherford bombarded nitrogen atoms with high-speed α particles from a natural radioactive isotope of radium and observed protons resulting from the reaction:

$$^{14}_{7}\mathrm{N} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{17}_{8}\mathrm{O} + {}^{1}_{1}\mathrm{H}$$

The $^{17}_{8}\mathrm{O}$ and $^{14}_{7}\mathrm{N}$ nuclei that are produced are stable, so no further (nuclear) changes occur.

To reach the kinetic energies necessary to produce transmutation reactions, devices called **particle accelerators** are used. These devices use magnetic and electric fields to increase the speeds of nuclear particles. In all accelerators, the particles move in a vacuum to avoid collisions with gas molecules.

When neutrons are required for transmutation reactions, they are usually obtained from radioactive decay reactions or from various nuclear reactions occurring in nuclear reactors. The Chemistry in Everyday Life feature that follows discusses a famous particle accelerator that made worldwide news.

Chemistry in Everyday Life: CERN Particle Accelerator

Located near Geneva, the CERN ("Conseil Européen pour la Recherche Nucléaire," or European Council for Nuclear Research) Laboratory is the world's premier center for the investigations of the fundamental particles that make up matter. It contains the 27-kilometer (17 mile) long, circular Large Hadron Collider (LHC), the largest particle accelerator in the world (Figure 11.5.1). In the LHC, particles are boosted to high energies and are then made to collide with each other or with stationary targets at nearly the speed of light. Superconducting electromagnets are used to produce a strong magnetic field that guides the particles around the ring. Specialized, purpose-built detectors observe and record the results of these collisions, which are then analyzed by CERN scientists using powerful computers.



Figure 11.5.1: A small section of the LHC is shown with workers traveling along it. (credit: Christophe Delaere)

In 2012, CERN announced that experiments at the LHC showed the first observations of the Higgs boson, an elementary particle that helps explain the origin of mass in fundamental particles. This long-anticipated discovery made worldwide news and resulted in the awarding of the 2013 Nobel Prize in Physics to François Englert and Peter Higgs, who had predicted the existence of this particle almost 50 years previously.

Link to Learning

Famous physicist Brian Cox talks about his work on the Large Hadron Collider at CERN, providing an entertaining and engaging tour of this massive project and the physics behind it.

View a short video from CERN, describing the basics of how its particle accelerators work.

Prior to 1940, the heaviest-known element was uranium, whose atomic number is 92. Now, many artificial elements have been synthesized and isolated, including several on such a large scale that they have had a profound effect on society. One of these—element 93, neptunium (Np)—was first made in 1940 by McMillan and Abelson by bombarding uranium-238 with neutrons. The reaction creates unstable uranium-239, with a half-life of 23.5 minutes, which then decays into neptunium-239. Neptunium-239 is also radioactive, with a half-life of 2.36 days, and it decays into plutonium-239. The nuclear reactions are:

$$^{238}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{239}_{92}\mathrm{U}$$

$$^{239}_{92}\mathrm{U} \longrightarrow {}^{239}_{93}\mathrm{Np} + {}^{0}_{-1}\mathrm{e} \quad \text{half-life } = 23.5 \text{ min}$$

$$^{239}_{93}\mathrm{Np} \longrightarrow {}^{239}_{94}\mathrm{Pu} + {}^{0}_{-1}\mathrm{e} \quad \text{half-life } = 2.36 \text{ days}$$

Plutonium is now mostly formed in nuclear reactors as a byproduct during the fission of U-235. Additional neutrons are released during this fission process (see the next section), some of which combine with U-238 nuclei to form uranium-239; this undergoes β decay to form neptunium-239, which in turn undergoes β decay to form plutonium-239 as illustrated in the preceding three equations. These processes are summarized in the equation:

$$^{238}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{239}_{92}\mathrm{U} \xrightarrow{\beta^-} {}^{239}_{93}\mathrm{Np} \xrightarrow{\beta^-} {}^{239}_{94}\mathrm{Pu} \tag{11.5.1}$$

Heavier isotopes of plutonium—Pu-240, Pu-241, and Pu-242—are also produced when lighter plutonium nuclei capture neutrons. Some of this highly radioactive plutonium is used to produce military weapons, and the rest presents a serious storage problem because they have half-lives from thousands to hundreds of thousands of years.

Although they have not been prepared in the same quantity as plutonium, many other synthetic nuclei have been produced. Nuclear medicine has developed from the ability to convert atoms of one type into other types of atoms. Radioactive isotopes of several dozen elements are currently used for medical applications. The radiation produced by their decay is used to image or treat various organs or portions of the body, among other uses.

The elements beyond element 92 (uranium) are called **transuranium elements**. As of this writing, 22 transuranium elements have been produced and officially recognized by IUPAC; several other elements have formation claims that are waiting for approval. Some of these elements are shown in Table 21.3.

Table 21.3: Preparation of Some of the Transuranium Elements

Name	Symbol	Atomic Number	Reaction
americium	Am	95	$^{239}_{94}\text{Pu} + {}^{1}_{0}\text{n} \longrightarrow {}^{240}_{95}\text{Am} + {}^{0}_{-1}\text{e}$
curium	Cm	96	$^{239}_{94}\text{Pu} + {}^{4}_{2}\text{He} \longrightarrow {}^{242}_{96}\text{Cm} + {}^{1}_{0}\text{n}$
californium	Cf	98	$^{242}_{96}\text{Cm} + {}^{4}_{2}\text{He} \longrightarrow {}^{245}_{98}\text{Cf} + {}^{1}_{0}\text{n}$
einsteinium	Es	99	$^{242}_{96}\text{Cm} + {}^{4}_{2}\text{He} \longrightarrow {}^{245}_{98}\text{Cf} + {}^{1}_{0}\text{n}$
mendelevium	Md	101	$^{253}_{99}\text{Es} + {}^{4}_{2}\text{He} \longrightarrow {}^{256}_{101}\text{Md} + {}^{1}_{0}\text{n}$
nobelium	No	102	$^{246}_{96}\text{Cm} + {}^{12}_{6}\text{C} \longrightarrow {}^{254}_{102}\text{No} + 4{}^{1}_{0}\text{n}$
rutherfordium	Rf	104	$^{249}_{98}\text{Cf} + {}^{12}_{6}\text{C} \longrightarrow {}^{257}_{104}\text{Rf} + 4{}^{1}_{0}\text{n}$ $^{249}_{98}\text{Cf} + {}^{18}_{8}\text{O} \longrightarrow {}^{263}_{106}\text{Sg} + 4{}^{1}_{0}\text{n}$
seaborgium	Sg	106	$^{206}_{82}\text{Pb} + {}^{54}_{24}\text{Cr} \longrightarrow {}^{257}_{106}\text{Sg} + 3{}^{1}_{0}\text{n}$
meitnerium	Mt	107	$^{209}_{83}\text{Bi} + {}^{58}_{26}\text{Fe} \longrightarrow {}^{266}_{109}\text{Mt} + {}^{1}_{0}\text{n}$

Nuclear Fission

Many heavier elements with smaller binding energies per nucleon can decompose into more stable elements that have intermediate mass numbers and larger binding energies per nucleon—that is, mass numbers and binding energies per nucleon that are closer to the "peak" of the binding energy graph near 56 (see Figure 21.3). Sometimes neutrons are also produced. This decomposition is called fission, the breaking of a large nucleus into smaller pieces. The breaking is rather random with the formation of a large number of different products. Fission usually does not occur naturally, but is induced by bombardment with neutrons. The first reported nuclear fission occurred in 1939 when three German scientists, Lise Meitner, Otto Hahn, and Fritz Strassman, bombarded uranium-235 atoms with slow-moving neutrons that split the U-238 nuclei into smaller fragments that consisted of several neutrons and elements near the middle of the periodic table. Since then, fission has been observed in many other isotopes, including most actinide isotopes that have an odd number of neutrons. A typical nuclear fission reaction is shown in Figure 11.5.2

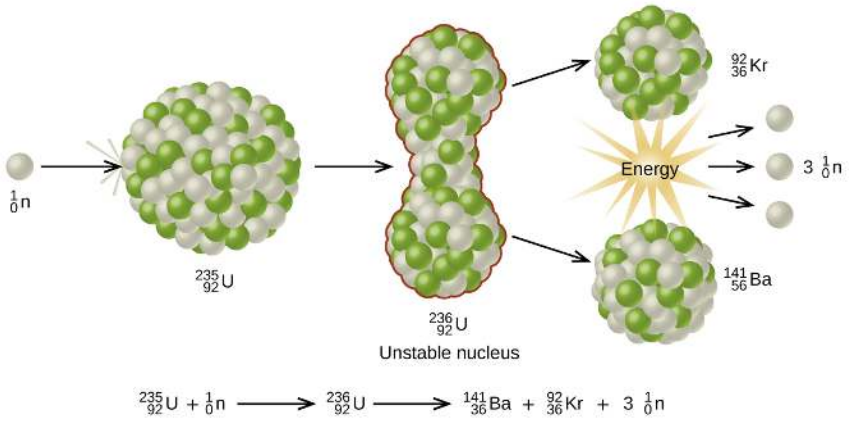


Figure 11.5.2: When a slow neutron hits a fissionable U-235 nucleus, it is absorbed and forms an unstable U-236 nucleus. The U-236 nucleus then rapidly breaks apart into two smaller nuclei (in this case, Ba-141 and Kr-92) along with several neutrons (usually two or three), and releases a very large amount of energy.

Among the products of Meitner, Hahn, and Strassman's fission reaction were barium, krypton, lanthanum, and cerium, all of which have nuclei that are more stable than uranium-235. Since then, hundreds of different isotopes have been observed among the products of fissionable substances. A few of the many reactions that occur for U-235, and a graph showing the distribution of its fission products and their yields, are shown in Figure 11.5.3 Similar fission reactions have been observed with other uranium isotopes, as well as with a variety of other isotopes such as those of plutonium.

$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow {}^{90}_{38}Sr + {}^{144}_{54}Xe + 2\ {}^{1}_{0}n$$

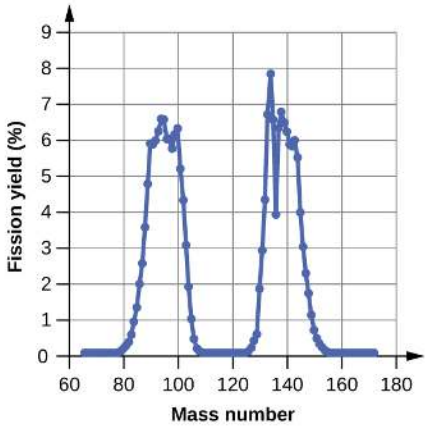
$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow {}^{87}_{35}Br + {}^{146}_{57}La + 3\ {}^{1}_{0}n$$

$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow {}^{96}_{37}Rb + {}^{137}_{55}Cs + 3\ {}^{1}_{0}n$$

$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow {}^{137}_{52}Te + {}^{97}_{40}Zr + 2\ {}^{1}_{0}n$$

$$^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{236}_{92}U \longrightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + 3\ {}^{1}_{0}n$$

(a)



(b)

Figure 11.5.3: (a) Nuclear fission of U-235 produces a range of fission products. (b) The larger fission products of U-235 are typically one isotope with a mass number around 85–105, and another isotope with a mass number that is about 50% larger, that is, about 130–150.

Link to Learning

View this link to see a simulation of nuclear fission.

A tremendous amount of energy is produced by the fission of heavy elements. For instance, when one mole of U-235 undergoes fission, the products weigh about 0.2 grams less than the reactants; this "lost" mass is converted into a very large amount of energy, about $1.8\ 10^{10}$ kJ per mole of U-235. Nuclear fission reactions produce incredibly large amounts of energy compared to chemical reactions. The fission of 1 kilogram of uranium-235, for example, produces about 2.5 million times as much energy as is produced by burning 1 kilogram of coal.

As described earlier, when undergoing fission U-235 produces two "medium-sized" nuclei, and two or three neutrons. These neutrons may then cause the fission of other uranium-235 atoms, which in turn provide more neutrons that can cause fission of even more nuclei, and so on. If this occurs, we have a nuclear **chain reaction** (see Figure 11.5.4). On the other hand, if too many neutrons escape the bulk material without interacting with a nucleus, then no chain reaction will occur.

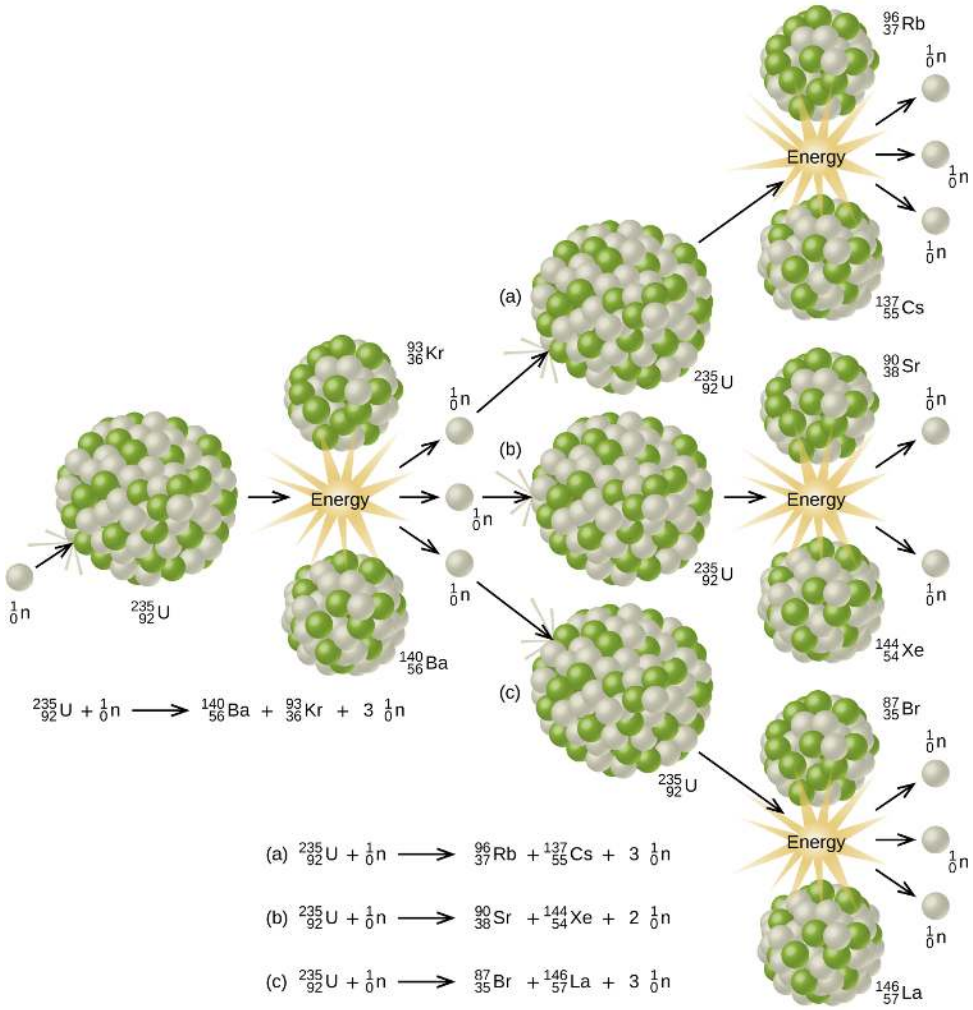


Figure 11.5.4: The fission of a large nucleus, such as U-235, produces two or three neutrons, each of which is capable of causing fission of another nucleus by the reactions shown. If this process continues, a nuclear chain reaction occurs.

Material that can sustain a nuclear fission chain reaction is said to be **fissile** or **fissionable**. (Technically, fissile material can undergo fission with neutrons of any energy, whereas fissionable material requires high-energy neutrons.) Nuclear fission becomes self-sustaining when the number of neutrons produced by fission equals or exceeds the number of neutrons absorbed by splitting nuclei plus the number that escape into the surroundings. The amount of a fissionable material that will support a self-sustaining chain reaction is a **critical mass**. An amount of fissionable material that cannot sustain a chain reaction is a **subcritical mass**. An amount of material in which there is an increasing rate of fission is known as a **supercritical mass**. The critical mass depends on the type of material: its purity, the temperature, the shape of the sample, and how the neutron reactions are controlled (Figure 11.5.5).

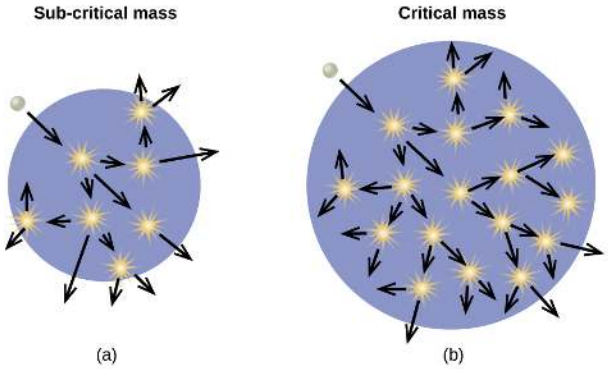


Figure 11.5.5: (a) In a subcritical mass, the fissile material is too small and allows too many neutrons to escape the material, so a chain reaction does not occur. (b) In a critical mass, a large enough number of neutrons in the fissile material induce fission to create a chain reaction.

An atomic bomb (Figure 11.5.6) contains several pounds of fissionable material, $^{235}_{92}\text{U}$ or $^{239}_{94}\text{Pu}$ a source of neutrons, and an explosive device for compressing it quickly into a small volume. When fissionable material is in small pieces, the proportion of neutrons that escape through the relatively large surface area is great, and a chain reaction does not take place. When the small pieces of fissionable material are brought together quickly to form a body with a mass larger than the critical mass, the relative number of escaping neutrons decreases, and a chain reaction and explosion result.

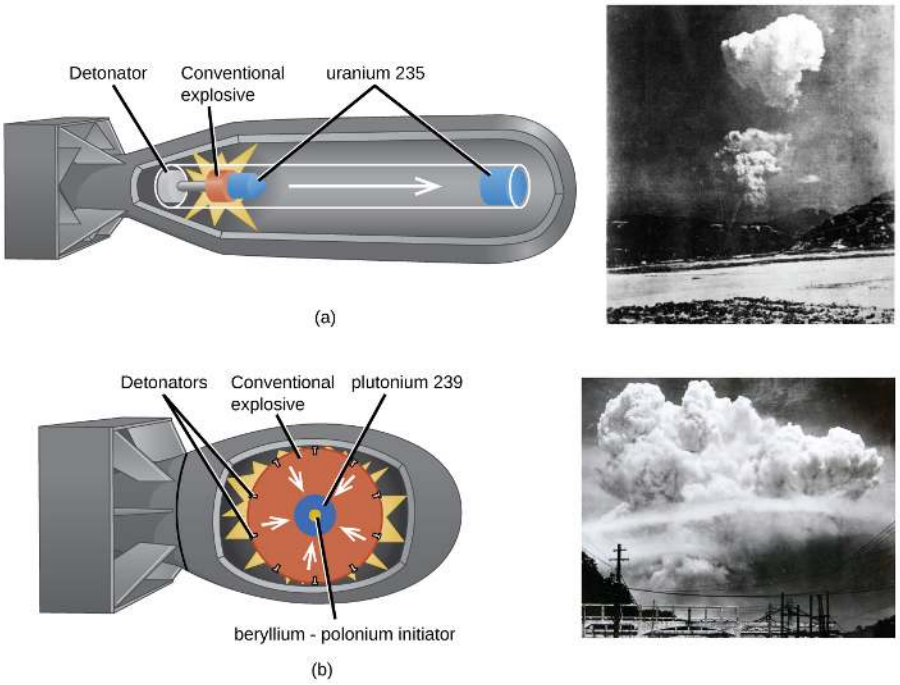


Figure 11.5.6: (a) The nuclear fission bomb that destroyed Hiroshima on August 6, 1945, consisted of two subcritical masses of U-235, where conventional explosives were used to fire one of the subcritical masses into the other, creating the critical mass for the nuclear explosion. (b) The plutonium bomb that destroyed Nagasaki on August 9, 1945, consisted of a hollow sphere of plutonium that was rapidly compressed by conventional explosives. This led to a concentration of plutonium in the center that was greater than the critical mass necessary for the nuclear explosion.

Fission Reactors

Chain reactions of fissionable materials can be controlled and sustained without an explosion in a **nuclear reactor** (Figure 11.5.7). Any nuclear reactor that produces power via the fission of uranium or plutonium by bombardment with neutrons must have at least five components: nuclear fuel consisting of fissionable material, a nuclear moderator, reactor coolant, control rods, and a shield and containment system. We will discuss these components in greater detail later in the section. The reactor works by separating the fissionable nuclear material such that a critical mass cannot be formed, controlling both the flux and absorption of neutrons to allow shutting down the fission reactions. In a nuclear reactor used for the production of electricity, the energy released by fission

reactions is trapped as thermal energy and used to boil water and produce steam. The steam is used to turn a turbine, which powers a generator for the production of electricity.

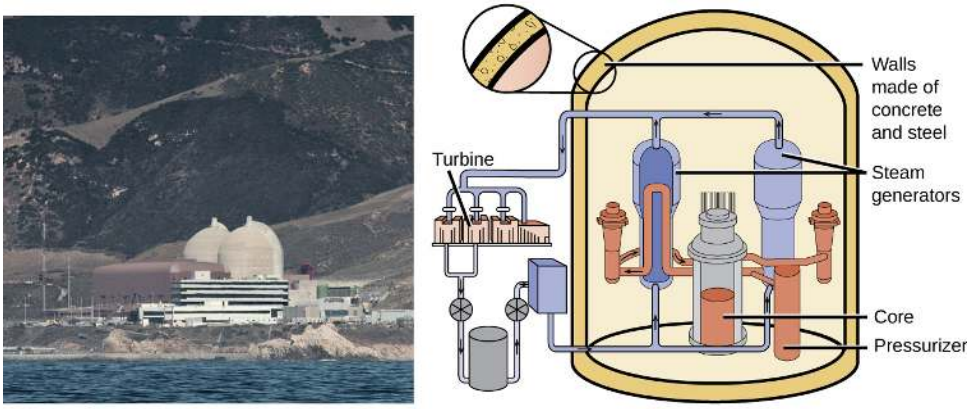


Figure 11.5.7: (a) The Diablo Canyon Nuclear Power Plant near San Luis Obispo is the only nuclear power plant currently in operation in California. The domes are the containment structures for the nuclear reactors, and the brown building houses the turbine where electricity is generated. Ocean water is used for cooling. (b) The Diablo Canyon uses a pressurized water reactor, one of a few different fission reactor designs in use around the world, to produce electricity. Energy from the nuclear fission reactions in the core heats water in a closed, pressurized system. Heat from this system produces steam that drives a turbine, which in turn produces electricity. (credit a: modification of work by “Mike” Michael L. Baird; credit b: modification of work by the Nuclear Regulatory Commission)

Nuclear Fuels

Nuclear fuel consists of a fissionable isotope, such as uranium-235, which must be present in sufficient quantity to provide a self-sustaining chain reaction. In the United States, uranium ores contain from 0.05–0.3% of the uranium oxide U_3O_8; the uranium in the ore is about 99.3% nonfissionable U-238 with only 0.7% fissionable U-235. Nuclear reactors require a fuel with a higher concentration of U-235 than is found in nature; it is normally enriched to have about 5% of uranium mass as U-235. At this concentration, it is not possible to achieve the supercritical mass necessary for a nuclear explosion. Uranium can be enriched by gaseous diffusion (the only method currently used in the US), using a gas centrifuge, or by laser separation.

In the gaseous diffusion enrichment plant where U-235 fuel is prepared, UF_6 (uranium hexafluoride) gas at low pressure moves through barriers that have holes just barely large enough for UF_6 to pass through. The slightly lighter $^{235}UF_6$ molecules diffuse through the barrier slightly faster than the heavier $^{238}UF_6$ molecules. This process is repeated through hundreds of barriers, gradually increasing the concentration of $^{235}UF_6$ to the level needed by the nuclear reactor. The basis for this process, Graham’s law, is described in the chapter on gases. The enriched UF_6 gas is collected, cooled until it solidifies, and then taken to a fabrication facility where it is made into fuel assemblies. Each fuel assembly consists of fuel rods that contain many thimble-sized, ceramic-encased, enriched uranium (usually UO_2) fuel pellets. Modern nuclear reactors may contain as many as 10 million fuel pellets. The amount of energy in each of these pellets is equal to that in almost a ton of coal or 150 gallons of oil.

Nuclear Moderators

Neutrons produced by nuclear reactions move too fast to cause fission (refer back to Figure 11.5.5). They must first be slowed to be absorbed by the fuel and produce additional nuclear reactions. A **nuclear moderator** is a substance that slows the neutrons to a speed that is low enough to cause fission. Early reactors used high-purity graphite as a moderator. Modern reactors in the US exclusively use heavy water or light water (ordinary H_2O), whereas some reactors in other countries use other materials, such as carbon dioxide, beryllium, or graphite.

Reactor Coolants

A nuclear reactor coolant is used to carry the heat produced by the fission reaction to an external boiler and turbine, where it is transformed into electricity. Two overlapping coolant loops are often used; this counteracts the transfer of radioactivity from the reactor to the primary coolant loop. All nuclear power plants in the US use water as a coolant. Other coolants include molten sodium, lead, a lead-bismuth mixture, or molten salts.

Control Rods

Nuclear reactors use control rods (Figure 11.5.8) to control the fission rate of the nuclear fuel by adjusting the number of slow neutrons present to keep the rate of the chain reaction at a safe level. Control rods are made of boron, cadmium, hafnium, or other elements that are able to absorb neutrons. Boron-10, for example, absorbs neutrons by a reaction that produces lithium-7 and alpha particles:

$$^{10}_{5}\mathrm{B} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{7}_{3}\mathrm{Li} + {}^{4}_{2}\mathrm{He} \tag{11.5.2}$$

When control rod assemblies are inserted into the fuel element in the reactor core, they absorb a larger fraction of the slow neutrons, thereby slowing the rate of the fission reaction and decreasing the power produced. Conversely, if the control rods are removed, fewer neutrons are absorbed, and the fission rate and energy production increase. In an emergency, the chain reaction can be shut down by fully inserting all of the control rods into the nuclear core between the fuel rods.

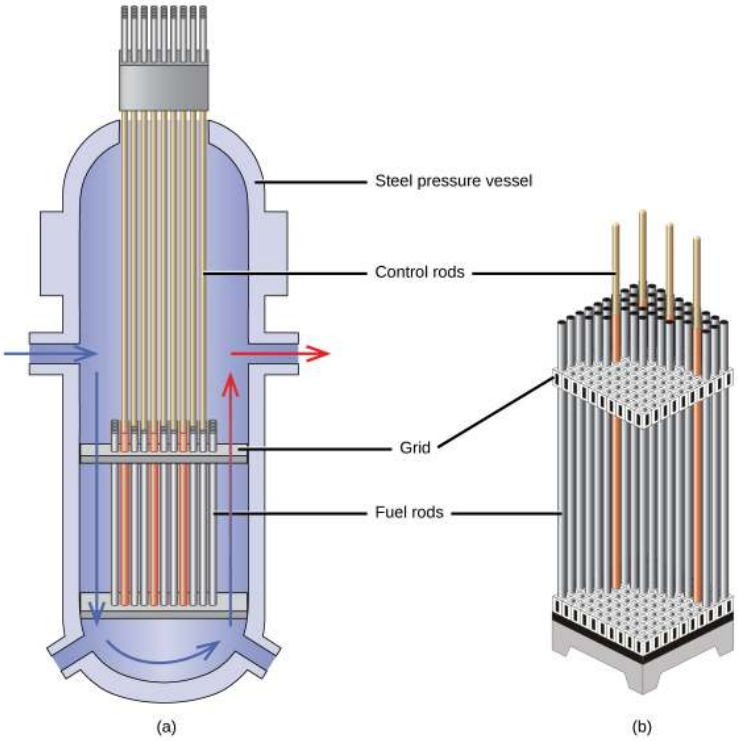


Figure 11.5.8: The nuclear reactor core shown in (a) contains the fuel and control rod assembly shown in (b). (credit: modification of work by E. Generalic, http://glossary.periodni.com/glossar...en=control+rod)

Shield and Containment System

During its operation, a nuclear reactor produces neutrons and other radiation. Even when shut down, the decay products are radioactive. In addition, an operating reactor is thermally very hot, and high pressures result from the circulation of water or another coolant through it. Thus, a reactor must withstand high temperatures and pressures, and must protect operating personnel from the radiation. Reactors are equipped with a **containment system** (or shield) that consists of three parts:

1. The reactor vessel, a steel shell that is 3–20-centimeters thick and, with the moderator, absorbs much of the radiation produced by the reactor
2. A main shield of 1–3 meters of high-density concrete
3. A personnel shield of lighter materials that protects operators from γ rays and X-rays

In addition, reactors are often covered with a steel or concrete dome that is designed to contain any radioactive materials might be released by a reactor accident.

Link to Learning

Click here to watch a 3-minute video from the Nuclear Energy Institute on how nuclear reactors work.

Nuclear power plants are designed in such a way that they cannot form a supercritical mass of fissionable material and therefore cannot create a nuclear explosion. But as history has shown, failures of systems and safeguards can cause catastrophic accidents, including chemical explosions and nuclear meltdowns (damage to the reactor core from overheating). The following Chemistry in Everyday Life feature explores three infamous meltdown incidents.

Chemistry in Everyday Life: Nuclear Accidents

The importance of cooling and containment are amply illustrated by three major accidents that occurred with the nuclear reactors at nuclear power generating stations in the United States (Three Mile Island), the former Soviet Union (Chernobyl), and Japan (Fukushima).

In March 1979, the cooling system of the Unit 2 reactor at Three Mile Island Nuclear Generating Station in Pennsylvania failed, and the cooling water spilled from the reactor onto the floor of the containment building. After the pumps stopped, the reactors overheated due to the high radioactive decay heat produced in the first few days after the nuclear reactor shut down. The temperature of the core climbed to at least 2200 °C, and the upper portion of the core began to melt. In addition, the zirconium alloy cladding of the fuel rods began to react with steam and produced hydrogen:

$$\mathrm{Zr(s) + 2\,H_2O(g) \longrightarrow ZrO_2(s) + 2\,H_2(g)}$$

The hydrogen accumulated in the confinement building, and it was feared that there was danger of an explosion of the mixture of hydrogen and air in the building. Consequently, hydrogen gas and radioactive gases (primarily krypton and xenon) were vented from the building. Within a week, cooling water circulation was restored and the core began to cool. The plant was closed for nearly 10 years during the cleanup process.

Although zero discharge of radioactive material is desirable, the discharge of radioactive krypton and xenon, such as occurred at the Three Mile Island plant, is among the most tolerable. These gases readily disperse in the atmosphere and thus do not produce highly radioactive areas. Moreover, they are noble gases and are not incorporated into plant and animal matter in the food chain. Effectively none of the heavy elements of the core of the reactor were released into the environment, and no cleanup of the area outside of the containment building was necessary (Figure 11.5.9).

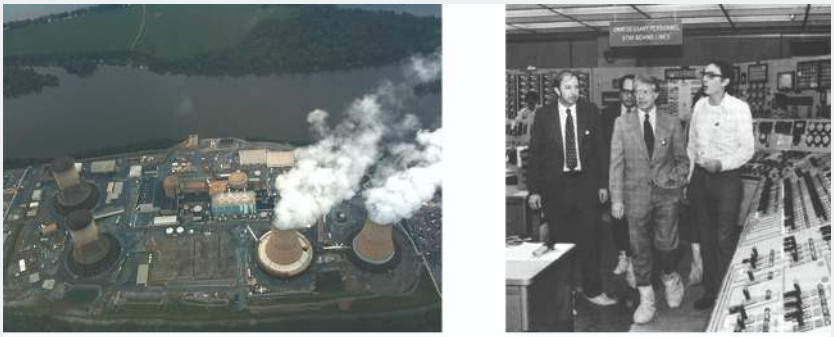


Figure 11.5.9: (a) In this 2010 photo of Three Mile Island, the remaining structures from the damaged Unit 2 reactor are seen on the left, whereas the separate Unit 1 reactor, unaffected by the accident, continues generating power to this day (right). (b) President Jimmy Carter visited the Unit 2 control room a few days after the accident in 1979.

Another major nuclear accident involving a reactor occurred in April 1986, at the Chernobyl Nuclear Power Plant in Ukraine, which was still a part of the former Soviet Union. While operating at low power during an unauthorized experiment with some of its safety devices shut off, one of the reactors at the plant became unstable. Its chain reaction became uncontrollable and increased to a level far beyond what the reactor was designed for. The steam pressure in the reactor rose to between 100 and 500 times the full power pressure and ruptured the reactor. Because the reactor was not enclosed in a containment building, a large amount of radioactive material spewed out, and additional fission products were released, as the graphite (carbon) moderator of the core ignited and burned. The fire was controlled, but over 200 plant workers and firefighters developed acute radiation sickness and at least 32 soon died from the effects of the radiation. It is predicted that about 4000 more deaths will

occur among emergency workers and former Chernobyl residents from radiation-induced cancer and leukemia. The reactor has since been encapsulated in steel and concrete, a now-decaying structure known as the sarcophagus. Almost 30 years later, significant radiation problems still persist in the area, and Chernobyl largely remains a wasteland.

In 2011, the Fukushima Daiichi Nuclear Power Plant in Japan was badly damaged by a 9.0-magnitude earthquake and resulting tsunami. Three reactors up and running at the time were shut down automatically, and emergency generators came online to power electronics and coolant systems. However, the tsunami quickly flooded the emergency generators and cut power to the pumps that circulated coolant water through the reactors. High-temperature steam in the reactors reacted with zirconium alloy to produce hydrogen gas. The gas escaped into the containment building, and the mixture of hydrogen and air exploded. Radioactive material was released from the containment vessels as the result of deliberate venting to reduce the hydrogen pressure, deliberate discharge of coolant water into the sea, and accidental or uncontrolled events.

An evacuation zone around the damaged plant extended over 12.4 miles away, and an estimated 200,000 people were evacuated from the area. All 48 of Japan's nuclear power plants were subsequently shut down, remaining shuttered as of December 2014. Since the disaster, public opinion has shifted from largely favoring to largely opposing increasing the use of nuclear power plants, and a restart of Japan's atomic energy program is still stalled (Figure 11.5.10).



Figure 11.5.10: (a) After the accident, contaminated waste had to be removed, and (b) an evacuation zone was set up around the plant in areas that received heavy doses of radioactive fallout. (credit a: modification of work by "Live Action Hero"/Flickr)

The energy produced by a reactor fueled with enriched uranium results from the fission of uranium as well as from the fission of plutonium produced as the reactor operates. As discussed previously, the plutonium forms from the combination of neutrons and the uranium in the fuel. In any nuclear reactor, only about 0.1% of the mass of the fuel is converted into energy. The other 99.9% remains in the fuel rods as fission products and unused fuel. All of the fission products absorb neutrons, and after a period of several months to a few years, depending on the reactor, the fission products must be removed by changing the fuel rods. Otherwise, the concentration of these fission products would increase and absorb more neutrons until the reactor could no longer operate.

Spent fuel rods contain a variety of products, consisting of unstable nuclei ranging in atomic number from 25 to 60, some transuranium elements, including plutonium and americium, and unreacted uranium isotopes. The unstable nuclei and the transuranium isotopes give the spent fuel a dangerously high level of radioactivity. The long-lived isotopes require thousands of years to decay to a safe level. The ultimate fate of the nuclear reactor as a significant source of energy in the United States probably rests on whether or not a politically and scientifically satisfactory technique for processing and storing the components of spent fuel rods can be developed.

Link to Learning

Explore the information in this link to learn about the approaches to nuclear waste management.

Nuclear Fusion and Fusion Reactors

The process of converting very light nuclei into heavier nuclei is also accompanied by the conversion of mass into large amounts of energy, a process called fusion. The principal source of energy in the sun is a net fusion reaction in which four hydrogen nuclei fuse

and produce one helium nucleus and two positrons. This is a net reaction of a more complicated series of events:

$$4\,{}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{0}_{+1}\mathrm{e}^{+}$$

A helium nucleus has a mass that is 0.7% less than that of four hydrogen nuclei; this lost mass is converted into energy during the fusion. This reaction produces about 3.6 10^{11} kJ of energy per mole of \({ }_2^4 \mathrm{He}\) produced. This is somewhat larger than the energy produced by the nuclear fission of one mole of U-235 (1.8 10^{10} kJ), and over 3 million times larger than the energy produced by the (chemical) combustion of one mole of octane (5471 kJ).

It has been determined that the nuclei of the heavy isotopes of hydrogen, a deuteron, ${}^{2}_{1}\mathrm{H}$, and a triton, ${}^{3}_{1}\mathrm{H}$, undergo fusion at extremely high temperatures (thermonuclear fusion). They form a helium nucleus and a neutron:

$${}^{2}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{1}_{0}\mathrm{n} \tag{11.5.3}$$

This change proceeds with a mass loss of 0.0188 amu, corresponding to the release of 1.69 10^{9} kilojoules per mole of \({ }_2^4 \mathrm{He}\) formed. The very high temperature is necessary to give the nuclei enough kinetic energy to overcome the very strong repulsive forces resulting from the positive charges on their nuclei so they can collide.

Useful fusion reactions require very high temperatures for their initiation—about 15,000,000 K or more. At these temperatures, all molecules dissociate into atoms, and the atoms ionize, forming plasma. These conditions occur in an extremely large number of locations throughout the universe—stars are powered by fusion. Humans have already figured out how to create temperatures high enough to achieve fusion on a large scale in thermonuclear weapons. A thermonuclear weapon such as a hydrogen bomb contains a nuclear fission bomb that, when exploded, gives off enough energy to produce the extremely high temperatures necessary for fusion to occur.

Another much more beneficial way to create fusion reactions is in a **fusion reactor**, a nuclear reactor in which fusion reactions of light nuclei are controlled. Because no solid materials are stable at such high temperatures, mechanical devices cannot contain the plasma in which fusion reactions occur. Two techniques to contain plasma at the density and temperature necessary for a fusion reaction are currently the focus of intensive research efforts: containment by a magnetic field and by the use of focused laser beams (Figure 11.5.11). A number of large projects are working to attain one of the biggest goals in science: getting hydrogen fuel to ignite and produce more energy than the amount supplied to achieve the extremely high temperatures and pressures that are required for fusion. At the time of this writing, there are no self-sustaining fusion reactors operating in the world, although small-scale controlled fusion reactions have been run for very brief periods.

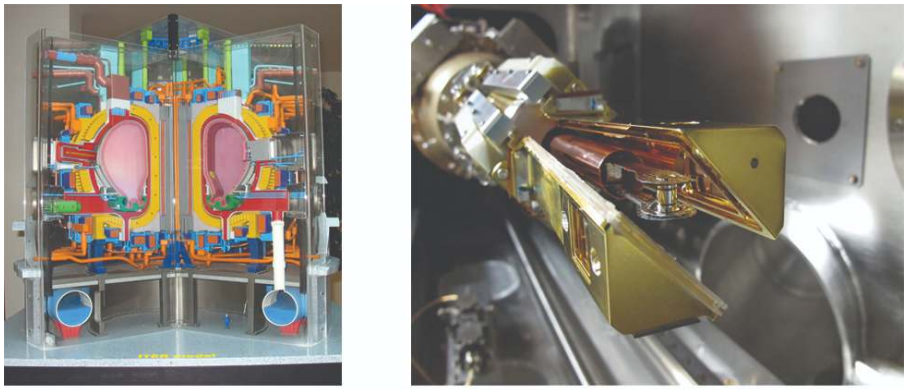


Figure 11.5.11: (a) This model is of the International Thermonuclear Experimental Reactor (ITER) reactor. Currently under construction in the south of France with an expected completion date of 2027, the ITER will be the world's largest experimental Tokamak nuclear fusion reactor with a goal of achieving large-scale sustained energy production. (b) In 2012, the National Ignition Facility at Lawrence Livermore National Laboratory briefly produced over 500,000,000,000 watts (500 terawatts, or 500 TW) of peak power and delivered 1,850,000 joules (1.85 MJ) of energy, the largest laser energy ever produced and 1000 times the power usage of the entire United States in any given moment. Although lasting only a few billionths of a second, the 192 lasers attained the conditions needed for nuclear fusion ignition. This image shows the target prior to the laser shot. (credit a: modification of work by Stephan Mosel)

This page titled 11.5: Transmutation and Nuclear Energy is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **21.4: Transmutation and Nuclear Energy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.6: Uses of Radioisotopes

Learning Objectives

By the end of this section, you will be able to:

- List common applications of radioactive isotopes

Radioactive isotopes have the same chemical properties as stable isotopes of the same element, but they emit radiation, which can be detected. If we replace one (or more) atom(s) with radioisotope(s) in a compound, we can track them by monitoring their radioactive emissions. This type of compound is called a radioactive tracer (or radioactive label). Radioisotopes are used to follow the paths of biochemical reactions or to determine how a substance is distributed within an organism. Radioactive tracers are also used in many medical applications, including both diagnosis and treatment. They are used to measure engine wear, analyze the geological formation around oil wells, and much more.

Radioimmunossays (RIA), for example, rely on radioisotopes to detect the presence and/or concentration of certain antigens. Developed by Rosalyn Sussman Yalow and Solomon Berson in the 1950s, the technique is known for extreme sensitivity, meaning that it can detect and measure very small quantities of a substance. Prior to its discovery, most similar detection relied on large enough quantities to produce visible outcomes. RIA revolutionized and expanded entire fields of study, most notably endocrinology, and is commonly used in narcotics detection, blood bank screening, early cancer screening, hormone measurement, and allergy diagnosis. Based on her significant contribution to medicine, Yalow received a Nobel Prize, making her the second woman to be awarded the prize for medicine.

Radioisotopes have revolutionized medical practice (see Appendix M), where they are used extensively. Over 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States. Four typical examples of radioactive tracers used in medicine are technetium-99 ($^{99}_{43}\mathrm{Tc}$), thallium-201 ($^{201}_{81}\mathrm{Tl}$), iodine-131 ($^{131}_{53}\mathrm{I}$), and sodium-24 ($^{24}_{11}\mathrm{Na}$). Damaged tissues in the heart, liver, and lungs absorb certain compounds of technetium-99 preferentially. After it is injected, the location of the technetium compound, and hence the damaged tissue, can be determined by detecting the γ rays emitted by the Tc-99 isotope. Thallium-201 (Figure 11.6.1) becomes concentrated in healthy heart tissue, so the two isotopes, Tc-99 and Tl-201, are used together to study heart tissue. Iodine-131 concentrates in the thyroid gland, the liver, and some parts of the brain. It can therefore be used to monitor goiter and treat thyroid conditions, such as Grave's disease, as well as liver and brain tumors. Salt solutions containing compounds of sodium-24 are injected into the bloodstream to help locate obstructions to the flow of blood.

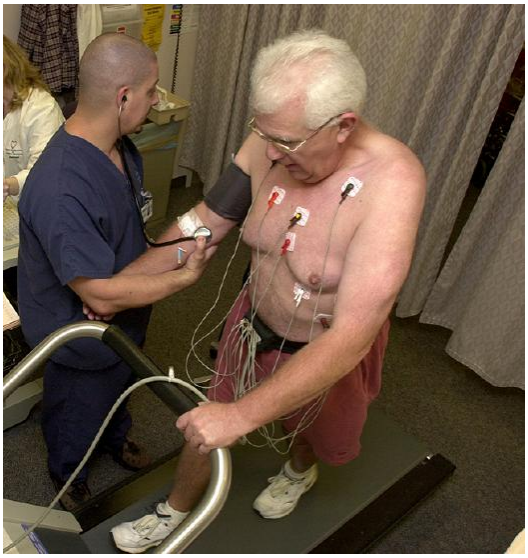


Figure 11.6.1: Administering thallium-201 to a patient and subsequently performing a stress test offer medical professionals an opportunity to visually analyze heart function and blood flow. (credit: modification of work by "Blue0ctane"/Wikimedia Commons)

Radioisotopes used in medicine typically have short half-lives—for example, the ubiquitous Tc-99m has a half-life of 6.01 hours. This makes Tc-99m essentially impossible to store and prohibitively expensive to transport, so it is made on-site instead. Hospitals and other medical facilities use Mo-99 (which is primarily extracted from U-235 fission products) to generate Tc-99. Mo-99 undergoes β decay with a half-life of 66 hours, and the Tc-99 is then chemically extracted (Figure 11.6.2). The parent nuclide Mo-99 is part of a molybdate ion, MoO_4^{2-}; when it decays, it forms the pertechnetate ion, TcO_4^-. These two water-soluble ions are separated by column chromatography, with the higher charge molybdate ion adsorbing onto the alumina in the column, and the lower charge pertechnetate ion passing through the column in the solution. A few micrograms of Mo-99 can produce enough Tc-99 to perform as many as 10,000 tests.

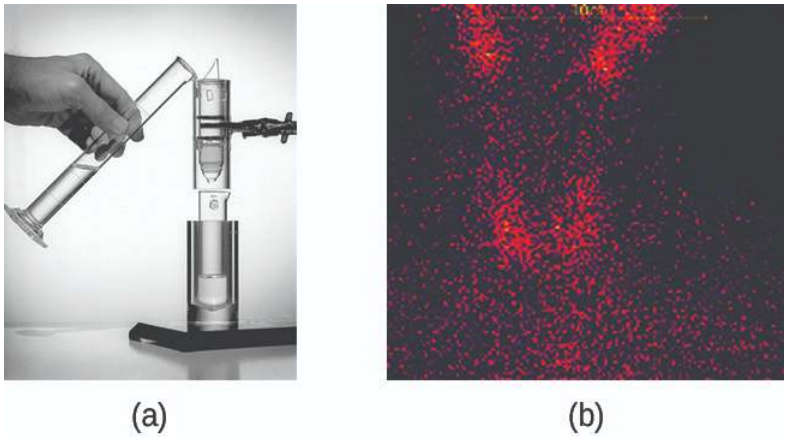


Figure 11.6.2: (a) The first Tc-99m generator (circa 1958) is used to separate Tc-99 from Mo-99. The is retained by the matrix in the column, whereas the passes through and is collected. (b) Tc-99 was used in this scan of the neck of a patient with Grave's disease. The scan shows the location of high concentrations of Tc-99. (credit a: modification of work by the Department of Energy; credit b: modification of work by "MBq"/Wikimedia Commons)

Radioisotopes can also be used, typically in higher doses than as a tracer, as treatment. Radiation therapy is the use of high-energy radiation to damage the DNA of cancer cells, which kills them or keeps them from dividing (Figure 11.6.3). A cancer patient may receive external beam radiation therapy delivered by a machine outside the body, or internal radiation therapy (brachytherapy) from a radioactive substance that has been introduced into the body. Note that chemotherapy is similar to internal radiation therapy in that the cancer treatment is injected into the body, but differs in that chemotherapy uses chemical rather than radioactive substances to kill the cancer cells.

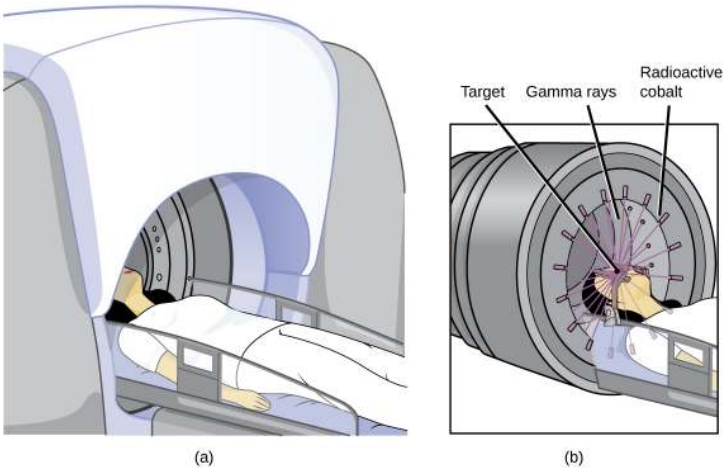


Figure 11.6.3: The cartoon in (a) shows a coba machine swings through an arc, focusing radiation on the targeted region (tumor) and minimizing the amount of radiation that passes through nearby regions.

Cobalt-60 is a synthetic radioisotope produced by the neutron activation of Co-59, which then undergoes β decay to form Ni-60, along with the emission of γ radiation. The overall process is:

$$ {}^{59}_{27}\mathrm{Co} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{60}_{27}\mathrm{Co} \rightarrow {}^{60}_{28}\mathrm{Ni} + {}^{0}_{-1}\beta + 2{}^{0}_{0}\gamma \qquad (11.6.1) $$

The overall decay scheme for this is shown graphically in Figure 11.6.4

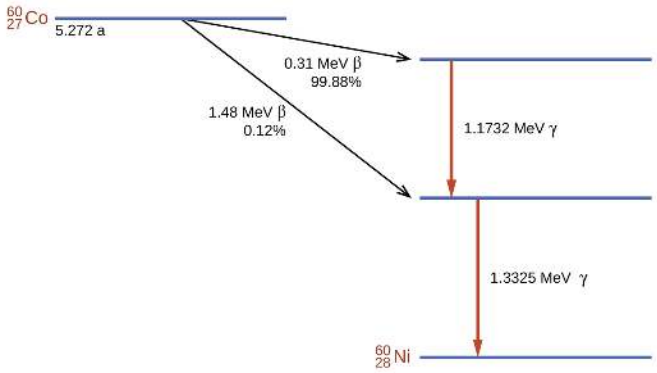


Figure 11.6.4: Co-60 undergoes a series of radioactive decays. The γ emissions are used for radiation therapy.

Radioisotopes are used in diverse ways to study the mechanisms of chemical reactions in plants and animals. These include labeling fertilizers in studies of nutrient uptake by plants and crop growth, investigations of digestive and milk-producing processes in cows, and studies on the growth and metabolism of animals and plants.

For example, the radioisotope C-14 was used to elucidate the details of how photosynthesis occurs. The overall reaction is:

$$6CO_2(g) + 6H_2O\ (l) \rightarrow C_6H_{12}O_6\ (s) + 6O_2\ (g), \tag{11.6.2}$$

but the process is much more complex, proceeding through a series of steps in which various organic compounds are produced. In studies of the pathway of this reaction, plants were exposed to CO_2 containing a high concentration of ($^{14}_6C$). At regular intervals, the plants were analyzed to determine which organic compounds contained carbon-14 and how much of each compound was present. From the time sequence in which the compounds appeared and the amount of each present at given time intervals, scientists learned more about the pathway of the reaction.

Commercial applications of radioactive materials are equally diverse (Figure 11.6.5). They include determining the thickness of films and thin metal sheets by exploiting the penetration power of various types of radiation. Flaws in metals used for structural purposes can be detected using high-energy gamma rays from cobalt-60 in a fashion similar to the way X-rays are used to examine the human body. In one form of pest control, flies are controlled by sterilizing male flies with γ radiation so that females breeding with them do not produce offspring. Many foods are preserved by radiation that kills microorganisms that cause the foods to spoil.



Figure 11.6.5: Common commercial uses of radiation include (a) X-ray examination of luggage at an airport and (b) preservation of food. (credit a: modification of work by the Department of the Navy; credit b: modification of work by the US Department of Agriculture)

Americium-241, an α emitter with a half-life of 458 years, is used in tiny amounts in ionization-type smoke detectors (Figure 11.6.6). The α emissions from Am-241 ionize the air between two electrode plates in the ionizing chamber. A battery supplies a potential that causes movement of the ions, thus creating a small electric current. When smoke enters the chamber, the movement of the ions is impeded, reducing the conductivity of the air. This causes a marked drop in the current, triggering an alarm.

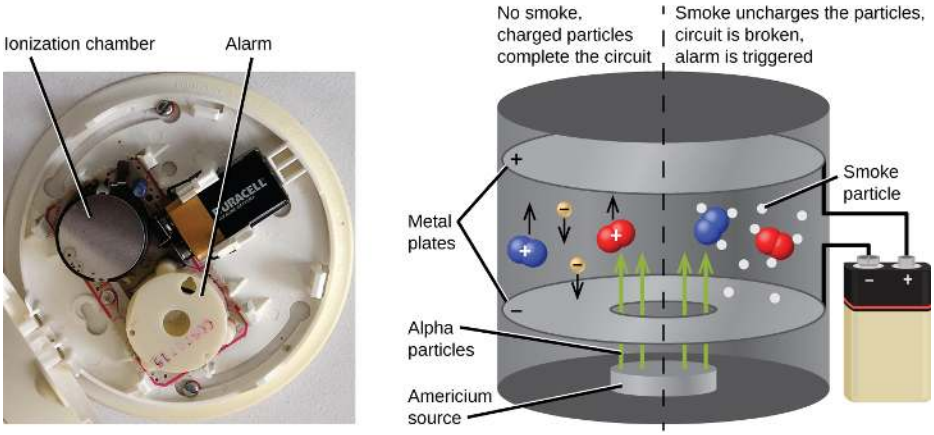


Figure 11.6.6: Inside a smoke detector, Am-241 emits α particles that ionize the air, creating a small electric current. During a fire, smoke particles impede the flow of ions, reducing the current and triggering an alarm. (credit a: modification of work by "Muffet"/Wikimedia Commons)

This page titled 11.6: Uses of Radioisotopes is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **21.5: Uses of Radioisotopes** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.7: Biological Effects of Radiation

Learning Objectives

By the end of this section, you will be able to:

- Describe the biological impact of ionizing radiation
- Define units for measuring radiation exposure
- Explain the operation of common tools for detecting radioactivity
- List common sources of radiation exposure in the US

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All radioactive nuclides emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause heating, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure 11.7.1).

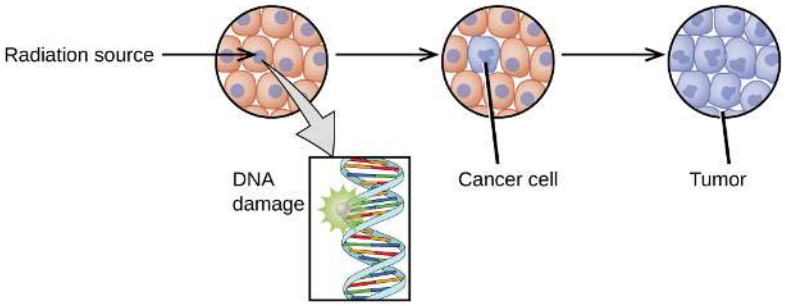


Figure 11.7.1: Radiation can harm biological systems by damaging the DNA of cells. If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer.

Ionizing and Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules (for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation) (Figure 11.7.2).

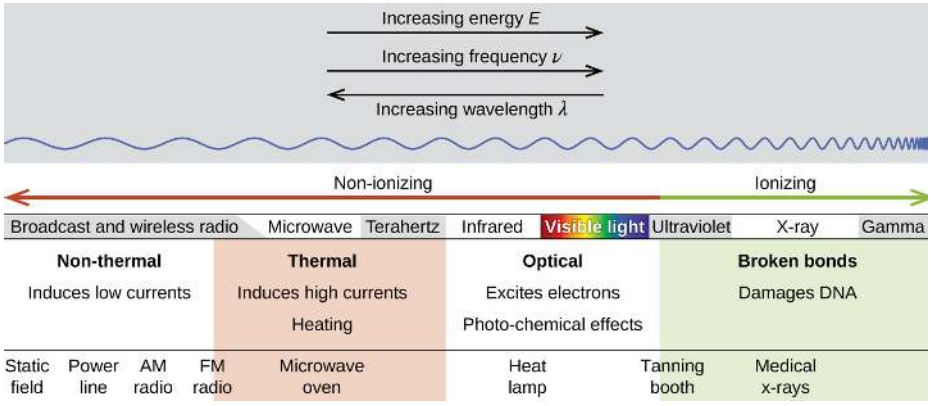


Figure 11.7.2: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing.

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day at the beach in the sun), a large amount of nonionizing radiation is necessary before dangerous levels are reached. Ionizing radiation,

however, may cause much more severe damage by breaking bonds or removing electrons in biological molecules, disrupting their structure and function. The damage can also be done indirectly, by first ionizing H_2O (the most abundant molecule in living organisms), which forms a H_2O^+ ion that reacts with water, forming a hydronium ion and a hydroxyl radical:

$$H_2O + \text{radiation} \xrightarrow{e^-} H_2O^+ + H_2O \longrightarrow H_3O^+ + OH^\bullet$$

Because the hydroxyl radical has an unpaired electron, it is highly reactive. (This is true of any substance with unpaired electrons, known as a free radical.) This hydroxyl radical can react with all kinds of biological molecules (DNA, proteins, enzymes, and so on), causing damage to the molecules and disrupting physiological processes. Examples of direct and indirect damage are shown in Figure 21.32.

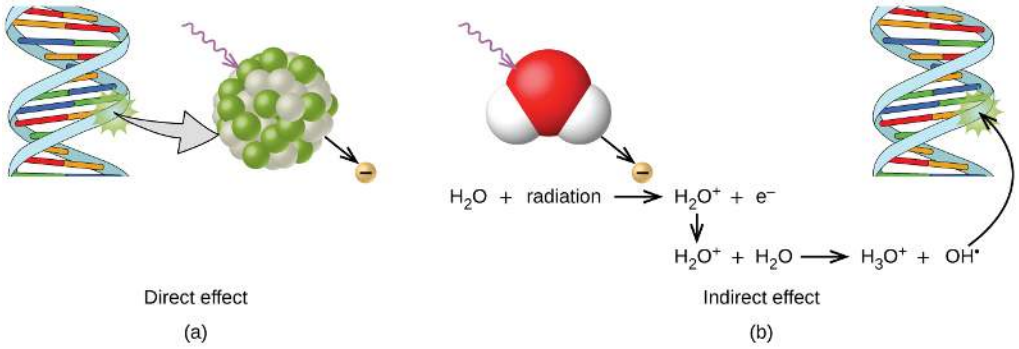


Figure 21.32 Ionizing radiation can (a) directly damage a biomolecule by ionizing it or breaking its bonds, or (b) create an H_2O^+ ion, which reacts with H_2O to form a hydroxyl radical, which in turn reacts with the biomolecule, causing damage indirectly.

Biological Effects of Exposure to Radiation

Radiation can harm either the whole body (somatic damage) or eggs and sperm (genetic damage). Its effects are more pronounced in cells that reproduce rapidly, such as the stomach lining, hair follicles, bone marrow, and embryos. This is why patients undergoing radiation therapy often feel nauseous or sick to their stomach, lose hair, have bone aches, and so on, and why particular care must be taken when undergoing radiation therapy during pregnancy.

Different types of radiation have differing abilities to pass through material (Figure 11.7.3). A very thin barrier, such as a sheet or two of paper, or the top layer of skin cells, usually stops alpha particles. Because of this, alpha particle sources are usually not dangerous if outside the body, but are quite hazardous if ingested or inhaled (see the Chemistry in Everyday Life feature on Radon Exposure). Beta particles will pass through a hand, or a thin layer of material like paper or wood, but are stopped by a thin layer of metal. Gamma radiation is very penetrating and can pass through a thick layer of most materials. Some high-energy gamma radiation is able to pass through a few feet of concrete. Certain dense, high atomic number elements (such as lead) can effectively attenuate gamma radiation with thinner material and are used for shielding. The ability of various kinds of emissions to cause ionization varies greatly, and some particles have almost no tendency to produce ionization. Alpha particles have about twice the ionizing power of fast-moving neutrons, about 10 times that of β particles, and about 20 times that of γ rays and X-rays.

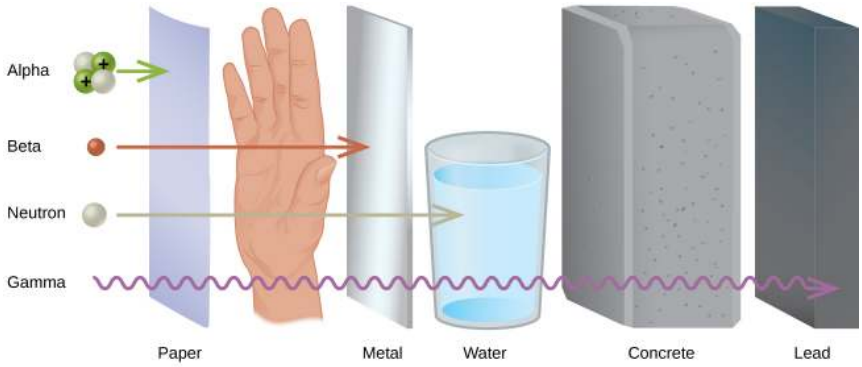


Figure 11.7.3: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma.

https://chem.libretexts.org/@go/page/479008

Chemistry in Everyday Life: Radon Exposure

For many people, one of the largest sources of exposure to radiation is from radon gas (Rn-222). Radon-222 is an α emitter with a half–life of 3.82 days. It is one of the products of the radioactive decay series of U-238 (Figure 21.9), which is found in trace amounts in soil and rocks. The radon gas that is produced slowly escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 11.7.4).

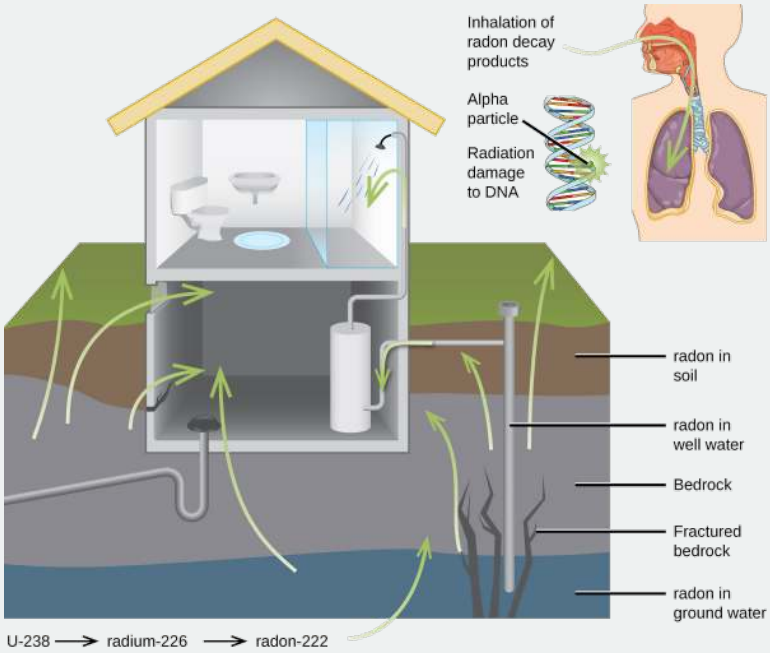


Figure 11.7.4: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinderblock foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts depending on where you live. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the levels found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

Measuring Radiation Exposure

Several different devices are used to detect and measure radiation, including Geiger counters, scintillation counters (scintillators), and radiation dosimeters (Figure 11.7.5). Probably the best-known radiation instrument, the Geiger counter (also called the Geiger-Müller counter) detects and measures radiation. Radiation causes the ionization of the gas in a Geiger-Müller tube. The rate of ionization is proportional to the amount of radiation. A scintillation counter contains a scintillator—a material that emits light (luminesces) when excited by ionizing radiation—and a sensor that converts the light into an electric signal. Radiation dosimeters also measure ionizing radiation and are often used to determine personal radiation exposure. Commonly used types are electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

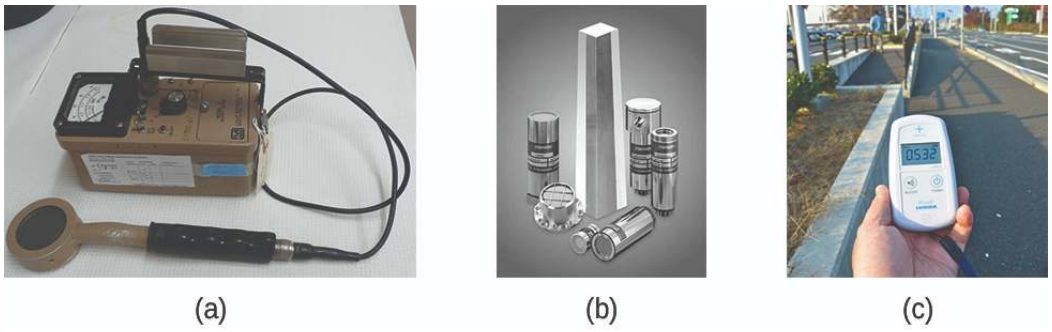


Figure 11.7.5: Devices such as (a) Geiger counters, (b) scintillators, and (c) dosimeters can be used to measure radiation. (credit c: modification of work by "osaMu"/Wikimedia commons)

A variety of units are used to measure various aspects of radiation (Figure 11.7.6). The SI unit for rate of radioactive decay is the becquerel (Bq), with 1 Bq = 1 disintegration per second. The curie (Ci) and millicurie (mCi) are much larger units and are frequently used in medicine (1 curie = 1 Ci = 3.7 10^{10} disintegrations per second). The SI unit for measuring radiation dose is the gray (Gy), with 1 Gy = 1 J of energy absorbed per kilogram of tissue. In medical applications, the radiation absorbed dose (rad) is more often used (1 rad = 0.01 Gy; 1 rad results in the absorption of 0.01 J/kg of tissue). The SI unit measuring tissue damage caused by radiation is the sievert (Sv). This takes into account both the energy and the biological effects of the type of radiation involved in the radiation dose. The roentgen equivalent for man (rem) is the unit for radiation damage that is used most frequently in medicine (100 rem = 1 Sv). Note that the tissue damage units (rem or Sv) includes the energy of the radiation dose (rad or Gy) along with a biological factor referred to as the RBE (for relative biological effectiveness) that is an approximate measure of the relative damage done by the radiation. These are related by:

$$\text{number of rems} = \text{RBE} \times \text{number of rads} \tag{11.7.1}$$

with RBE approximately 10 for α radiation, 2(+) for protons and neutrons, and 1 for β and γ radiation.

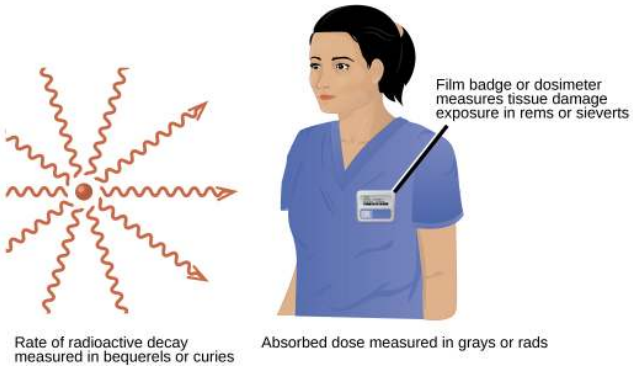


Figure 11.7.6: Different units are used to measure the rate of emission from a radioactive source, the energy that is absorbed from the source, and the amount of damage the absorbed radiation does.

Units of Radiation Measurement

Table 11.7.1 summarizes the units used for measuring radiation.

Table 11.7.1: Units Used for Measuring Radiation

Measurement Purpose	Unit	Quantity Measured	Description
activity of source	becquerel (Bq)	radioactive decays or emissions	amount of sample that undergoes 1 decay/second
	curie (Ci)		amount of sample that undergoes 3.7 10^{10} decays/second
absorbed dose	gray (Gy)	energy absorbed per kg of tissue	1 Gy = 1 J/kg tissue
	radiation absorbed dose (rad)		1 rad = 0.01 J/kg tissue

Measurement Purpose	Unit	Quantity Measured	Description
biologically effective dose	sievert (Sv)	tissue damage	Sv = RBE Gy
	roentgen equivalent for man (rem)		Rem = RBE rad

Example 11.7.1: Amount of Radiation

Cobalt-60 ($t_{1/2} = 5.26\ y$) is used in cancer therapy since the γ rays it emits can be focused in small areas where the cancer is located. A 5.00-g sample of Co-60 is available for cancer treatment.

a. What is its activity in Bq?
b. What is its activity in Ci?

Solution

The activity is given by:

$$\text{Activity } = \lambda N = \left(\frac{\ln 2}{t_{1/2}}\right) N = \left(\frac{\ln 2}{5.26\text{y}}\right) \times 5.00\text{ g} = 0.659 \frac{\text{g}}{\text{y}} \text{ of Co } -60 \text{ that decay}$$

And to convert this to decays per second:

$$0.659 \frac{\text{g}}{\text{y}} \times \frac{1\text{y}}{365\text{ d}} \times \frac{1\text{ d}}{24\text{ h}} \times \frac{1\text{ h}}{3600\text{ s}} \times \frac{1\text{ mol}}{59.9\text{ g}} \times \frac{6.02 \times 10^{23}\text{ atoms}}{1\text{ mol}} \times \frac{1\text{ decay}}{1\text{ atom}}$$
$$= 2.10 \times 10^{14} \frac{\text{decay}}{\text{s}}$$

a. Since $1\text{ Bq} = \frac{1\text{ decay}}{\text{s}}$ the activity in Becquerel (Bq) is:

$$2.10 \times 10^{14} \frac{\text{decay}}{\text{s}} \times \left(\frac{1\text{ Bq}}{1 \frac{\text{decay}}{\text{s}}}\right) = 2.10 \times 10^{14}\text{ Bq}$$

b. Since $1\text{Ci} = \frac{3.7 \times 10^{11}\text{ decay}}{\text{s}}$ the activity in curie (Ci) is:

$$2.10 \times 10^{14} \frac{\text{decay}}{\text{s}} \times \left(\frac{1\text{Ci}}{\frac{3.7 \times 10^{11}\text{ decay}}{\text{s}}}\right) = 5.7 \times 10^{2}\text{Ci}$$

Exercise 11.7.1

Tritium is a radioactive isotope of hydrogen ($t_{1/2}$ = 12.32 y) that has several uses, including self-powered lighting, in which electrons emitted in tritium radioactive decay cause phosphorus to glow. Its nucleus contains one proton and two neutrons, and the atomic mass of tritium is 3.016 amu. What is the activity of a sample containing 1.00mg of tritium (a) in Bq and (b) in Ci?

Answer

a. 3.56 10^{11} Bq;
b. 0.962 Ci

Effects of Long-term Radiation Exposure on the Human Body

The effects of radiation depend on the type, energy, and location of the radiation source, and the length of exposure. As shown in Figure 11.7.7 , the average person is exposed to background radiation, including cosmic rays from the sun and radon from uranium in the ground (see the Chemistry in Everyday Life feature on Radon Exposure); radiation from medical exposure, including CAT scans, radioisotope tests, X-rays, and so on; and small amounts of radiation from other human activities, such as airplane flights (which are bombarded by increased numbers of cosmic rays in the upper atmosphere), radioactivity from consumer

products, and a variety of radionuclides that enter our bodies when we breathe (for example, carbon-14) or through the food chain (for example, potassium-40, strontium-90, and iodine-131).

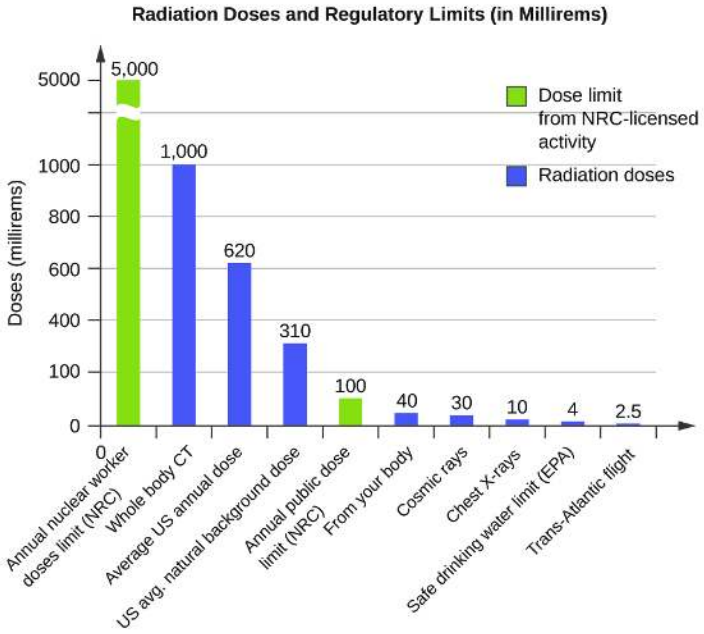


Figure 11.7.7: The total annual radiation exposure for a person in the US is about 620 mrem. The various sources and their relative amounts are shown in this bar graph. (source: U.S. Nuclear Regulatory Commission)

A short-term, sudden dose of a large amount of radiation can cause a wide range of health effects, from changes in blood chemistry to death. Short-term exposure to tens of rems of radiation will likely cause very noticeable symptoms or illness; a dose of about 500 rems is estimated to have a 50% probability of causing the death of the victim within 30 days of exposure. Exposure to radioactive emissions has a cumulative effect on the body during a person's lifetime, which is another reason why it is important to avoid any unnecessary exposure to radiation. Health effects of short-term exposure to radiation are shown in Table 11.7.2

Table 11.7.2: Health Effects of Radiation[2]

Exposure (rem)	Health Effect	Time to Onset (without treatment)
5–10	changes in blood chemistry	—
50	nausea	hours
55	fatigue	—
70	vomiting	—
75	hair loss	2–3 weeks
90	diarrhea	—
100	hemorrhage	—
400	possible death	within 2 months
1000	destruction of intestinal lining	—
	internal bleeding	—
	death	1–2 weeks
2000	damage to central nervous system	—
	loss of consciousness;	minutes
	death	hours to days

It is impossible to avoid some exposure to ionizing radiation. We are constantly exposed to background radiation from a variety of natural sources, including cosmic radiation, rocks, medical procedures, consumer products, and even our own atoms. We can minimize our exposure by blocking or shielding the radiation, moving farther from the source, and limiting the time of exposure.

Footnotes

- 2Source: US Environmental Protection Agency

This page titled 11.7: Biological Effects of Radiation is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.6: Biological Effects of Radiation** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

11.8: Key Terms

alpha (α) decay | loss of an alpha particle during radioactive decay

alpha particle | (α or or high-energy helium nucleus; a helium atom that has lost two electrons and contains two protons and two neutrons

antimatter | particles with the same mass but opposite properties (such as charge) of ordinary particles

band of stability | (also, belt of stability, zone of stability, or valley of stability) region of graph of number of protons versus number of neutrons containing stable (nonradioactive) nuclides

becquerel (Bq) | SI unit for rate of radioactive decay; 1 Bq = 1 disintegration/s

beta (β) decay | breakdown of a neutron into a proton, which remains in the nucleus, and an electron, which is emitted as a beta particle

beta particle | or or high-energy electron

binding energy per nucleon | total binding energy for the nucleus divided by the number of nucleons in the nucleus

chain reaction | repeated fission caused when the neutrons released in fission bombard other atoms

chemotherapy | similar to internal radiation therapy, but chemical rather than radioactive substances are introduced into the body to kill cancer cells

containment system | (also, shield) a three-part structure of materials that protects the exterior of a nuclear fission reactor and operating personnel from the high temperatures, pressures, and radiation levels inside the reactor

control rod | material inserted into the fuel assembly that absorbs neutrons and can be raised or lowered to adjust the rate of a fission reaction

critical mass | amount of fissionable material that will support a self-sustaining (nuclear fission) chain reaction

curie (Ci) | larger unit for rate of radioactive decay frequently used in medicine; 1 Ci = 3.7 10^{10} disintegrations/s

daughter nuclide | nuclide produced by the radioactive decay of another nuclide; may be stable or may decay further

electron capture | combination of a core electron with a proton to yield a neutron within the nucleus

electron volt (eV) | measurement unit of nuclear binding energies, with 1 eV equaling the amount energy due to the moving an electron across an electric potential difference of 1 volt

external beam radiation therapy | radiation delivered by a machine outside the body

fissile (or fissionable) | when a material is capable of sustaining a nuclear fission reaction

fission | splitting of a heavier nucleus into two or more lighter nuclei, usually accompanied by the conversion of mass into large amounts of energy

fusion | combination of very light nuclei into heavier nuclei, accompanied by the conversion of mass into large amounts of energy

fusion reactor | nuclear reactor in which fusion reactions of light nuclei are controlled

gamma (γ) emission | decay of an excited-state nuclide accompanied by emission of a gamma ray

gamma ray | (γ or short wavelength, high-energy electromagnetic radiation that exhibits wave-particle duality

Geiger counter | instrument that detects and measures radiation via the ionization produced in a Geiger-Müller tube

gray (Gy) | SI unit for measuring radiation dose; 1 Gy = 1 J absorbed/kg tissue

half-life ($t_{1/2}$) | time required for half of the atoms in a radioactive sample to decay

internal radiation therapy | (also, brachytherapy) radiation from a radioactive substance introduced into the body to kill cancer cells

ionizing radiation | radiation that can cause a molecule to lose an electron and form an ion

magic number | nuclei with specific numbers of nucleons that are within the band of stability

mass defect | difference between the mass of an atom and the summed mass of its constituent subatomic particles (or the mass “lost” when nucleons are brought together to form a nucleus)

mass-energy equivalence equation | Albert Einstein’s relationship showing that mass and energy are equivalent

millicurie (mCi) | larger unit for rate of radioactive decay frequently used in medicine; 1 Ci = 3.7 10^{10} disintegrations/s

nonionizing radiation | radiation that speeds up the movement of atoms and molecules; it is equivalent to heating a sample, but is not energetic enough to cause the ionization of molecules

nuclear binding energy | energy lost when an atom’s nucleons are bound together (or the energy needed to break a nucleus into its constituent protons and neutrons)

nuclear chemistry | study of the structure of atomic nuclei and processes that change nuclear structure

nuclear fuel | fissionable isotope present in sufficient quantities to provide a self-sustaining chain reaction in a nuclear reactor

nuclear moderator | substance that slows neutrons to a speed low enough to cause fission

nuclear reaction | change to a nucleus resulting in changes in the atomic number, mass number, or energy state

nuclear reactor | environment that produces energy via nuclear fission in which the chain reaction is controlled and sustained without explosion

nuclear transmutation | conversion of one nuclide into another nuclide

nucleon | collective term for protons and neutrons in a nucleus

nuclide | nucleus of a particular isotope

parent nuclide | unstable nuclide that changes spontaneously into another (daughter) nuclide

particle accelerator | device that uses electric and magnetic fields to increase the kinetic energy of nuclei used in transmutation reactions

positron or | antiparticle to the electron; it has identical properties to an electron, except for having the opposite (positive) charge

positron emission | (also, β^+ decay) conversion of a proton into a neutron, which remains in the nucleus, and a positron, which is emitted

radiation absorbed dose (rad) | SI unit for measuring radiation dose, frequently used in medical applications; 1 rad = 0.01 Gy

radiation dosimeter | device that measures ionizing radiation and is used to determine personal radiation exposure

radiation therapy | use of high-energy radiation to damage the DNA of cancer cells, which kills them or keeps them from dividing

radioactive decay | spontaneous decay of an unstable nuclide into another nuclide

radioactive decay series | chains of successive disintegrations (radioactive decays) that ultimately lead to a stable end-product

radioactive tracer | (also, radioactive label) radioisotope used to track or follow a substance by monitoring its radioactive emissions

radioactivity | phenomenon exhibited by an unstable nucleon that spontaneously undergoes change into a nucleon that is more stable; an unstable nucleon is said to be radioactive

radiocarbon dating | highly accurate means of dating objects 30,000–50,000 years old that were derived from once-living matter; achieved by calculating the ratio of in the object vs. the ratio of in the present-day atmosphere

radioisotope | isotope that is unstable and undergoes conversion into a different, more stable isotope

radiometric dating | use of radioisotopes and their properties to date the formation of objects such as archeological artifacts, formerly living organisms, or geological formations

reactor coolant | assembly used to carry the heat produced by fission in a reactor to an external boiler and turbine where it is transformed into electricity

relative biological effectiveness (RBE) | measure of the relative damage done by radiation

roentgen equivalent man (rem) | unit for radiation damage, frequently used in medicine; 100 rem = 1 Sv

scintillation counter | instrument that uses a scintillator—a material that emits light when excited by ionizing radiation—to detect and measure radiation

sievert (Sv) | SI unit measuring tissue damage caused by radiation; takes into account energy and biological effects of radiation

strong nuclear force | force of attraction between nucleons that holds a nucleus together

subcritical mass | amount of fissionable material that cannot sustain a chain reaction; less than a critical mass

supercritical mass | amount of material in which there is an increasing rate of fission

transmutation reaction | bombardment of one type of nuclei with other nuclei or neutrons

transuranium element | element with an atomic number greater than 92; these elements do not occur in nature

---

This page titled 11.8: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **21.7: Key Terms** is licensed CC BY 4.0.

11.9: Key Equations

$E = mc^2$

decay rate = λN

rem = RBE rad
Sv = RBE Gy

This page titled 11.9: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.8: Key Equations** is licensed CC BY 4.0.

11.10: Summary

An atomic nucleus consists of protons and neutrons, collectively called nucleons. Although protons repel each other, the nucleus is held tightly together by a short-range, but very strong, force called the strong nuclear force. A nucleus has less mass than the total mass of its constituent nucleons. This “missing” mass is the mass defect, which has been converted into the binding energy that holds the nucleus together according to Einstein’s mass-energy equivalence equation, $E = mc^2$. Of the many nuclides that exist, only a small number are stable. Nuclides with even numbers of protons or neutrons, or those with magic numbers of nucleons, are especially likely to be stable. These stable nuclides occupy a narrow band of stability on a graph of number of protons versus number of neutrons. The binding energy per nucleon is largest for the elements with mass numbers near 56; these are the most stable nuclei.

Nuclei can undergo reactions that change their number of protons, number of neutrons, or energy state. Many different particles can be involved in nuclear reactions. The most common are protons, neutrons, positrons (which are positively charged electrons), alpha (α) particles (which are high-energy helium nuclei), beta (β) particles (which are high-energy electrons), and gamma (γ) rays (which compose high-energy electromagnetic radiation). As with chemical reactions, nuclear reactions are always balanced. When a nuclear reaction occurs, the total mass (number) and the total charge remain unchanged.

Nuclei that have unstable n:p ratios undergo spontaneous radioactive decay. The most common types of radioactivity are α decay, β decay, γ emission, positron emission, and electron capture. Nuclear reactions also often involve γ rays, and some nuclei decay by electron capture. Each of these modes of decay leads to the formation of a new nucleus with a more stable n:p ratio. Some substances undergo radioactive decay series, proceeding through multiple decays before ending in a stable isotope. All nuclear decay processes follow first-order kinetics, and each radioisotope has its own characteristic half-life, the time that is required for half of its atoms to decay. Because of the large differences in stability among nuclides, there is a very wide range of half-lives of radioactive substances. Many of these substances have found useful applications in medical diagnosis and treatment, determining the age of archaeological and geological objects, and more.

It is possible to produce new atoms by bombarding other atoms with nuclei or high-speed particles. The products of these transmutation reactions can be stable or radioactive. A number of artificial elements, including technetium, astatine, and the transuranium elements, have been produced in this way.

Nuclear power as well as nuclear weapon detonations can be generated through fission (reactions in which a heavy nucleus is split into two or more lighter nuclei and several neutrons). Because the neutrons may induce additional fission reactions when they combine with other heavy nuclei, a chain reaction can result. Useful power is obtained if the fission process is carried out in a nuclear reactor. The conversion of light nuclei into heavier nuclei (fusion) also produces energy. At present, this energy has not been contained adequately and is too expensive to be feasible for commercial energy production.

Compounds known as radioactive tracers can be used to follow reactions, track the distribution of a substance, diagnose and treat medical conditions, and much more. Other radioactive substances are helpful for controlling pests, visualizing structures, providing fire warnings, and for many other applications. Hundreds of millions of nuclear medicine tests and procedures, using a wide variety of radioisotopes with relatively short half-lives, are performed every year in the US. Most of these radioisotopes have relatively short half-lives; some are short enough that the radioisotope must be made on-site at medical facilities. Radiation therapy uses high-energy radiation to kill cancer cells by damaging their DNA. The radiation used for this treatment may be delivered externally or internally.

We are constantly exposed to radiation from a variety of naturally occurring and human-produced sources. This radiation can affect living organisms. Ionizing radiation is the most harmful because it can ionize molecules or break chemical bonds, which damages the molecule and causes malfunctions in cell processes. It can also create reactive hydroxyl radicals that damage biological molecules and disrupt physiological processes. Radiation can cause somatic or genetic damage, and is most harmful to rapidly reproducing cells. Types of radiation differ in their ability to penetrate material and damage tissue, with alpha particles the least penetrating but potentially most damaging and gamma rays the most penetrating.

Various devices, including Geiger counters, scintillators, and dosimeters, are used to detect and measure radiation, and monitor radiation exposure. We use several units to measure radiation: becquerels or curies for rates of radioactive decay; gray or rads for energy absorbed; and rems or sieverts for biological effects of radiation. Exposure to radiation can cause a wide range of health effects, from minor to severe, and including death. We can minimize the effects of radiation by shielding with dense materials such as lead, moving away from the source, and limiting time of exposure.

This page titled 11.10: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.9: Summary** is licensed CC BY 4.0.

11.11: Exercises

1. Write the following isotopes in hyphenated form (e.g., “carbon-14”)

(a) $^{24}_{11}\text{Na}$
(b) $^{29}_{13}\text{Al}$
(c) $^{73}_{36}\text{Kr}$
(d) $^{194}_{77}\text{Ir}$

2. Write the following isotopes in nuclide notation (e.g.,

a. oxygen-14
b. copper-70
c. tantalum-175
d. francium-217

3. For the following isotopes that have missing information, fill in the missing information to complete the notation

(a) $^{34}_{14}\text{X}$
(b) $^{36}_{\text{X}}\text{P}$
(c) $^{57}_{\text{X}}\text{Mn}$
(d) $^{121}_{56}\text{X}$

4. For each of the isotopes in Exercise 21.1, determine the numbers of protons, neutrons, and electrons in a neutral atom of the isotope.

5. Write the nuclide notation, including charge if applicable, for atoms with the following characteristics:

a. 25 protons, 20 neutrons, 24 electrons
b. 45 protons, 24 neutrons, 43 electrons
c. 53 protons, 89 neutrons, 54 electrons
d. 97 protons, 146 neutrons, 97 electrons

6. Calculate the density of the nucleus in g/mL, assuming that it has the typical nuclear diameter of 1 10^{-13} cm and is spherical in shape. 7. What are the two principal differences between nuclear reactions and ordinary chemical changes?

8. The mass of the atom is 22.9898 amu.

a. Calculate its binding energy per atom in millions of electron volts.
b. Calculate its binding energy per nucleon.

9. Which of the following nuclei lie within the band of stability shown in Figure 21.2?

a. chlorine-37
b. calcium-40
c. ^{204}Bi
d. ^{56}Fe
e. ^{206}Pb
f. ^{211}Pb
g. ^{222}Rn
h. carbon-14

10. Which of the following nuclei lie within the band of stability shown in Figure 21.2?

a. argon-40
b. oxygen-16
c. ^{122}Ba
d. ^{58}Ni
e. ^{205}Tl
f. ^{210}Tl
g. ^{226}Ra

h. magnesium-24

11. Write a brief description or definition of each of the following:

a. nucleon
b. α particle
c. β particle
d. positron
e. γ ray
f. nuclide
g. mass number
h. atomic number

12. Which of the various particles (α particles, β particles, and so on) that may be produced in a nuclear reaction are actually nuclei?

13. Complete each of the following equations by adding the missing species:

(a) $^{27}_{13}\mathrm{Al} + {}^{4}_{2}\mathrm{He} \rightarrow \ ? + {}^{1}_{0}\mathrm{n}$
(b) $^{239}_{94}\mathrm{Pu} + ? \rightarrow {}^{242}_{96}\mathrm{Cm} + {}^{1}_{0}\mathrm{n}$
(c) $^{14}_{7}\mathrm{N} + {}^{4}_{2}\mathrm{He} \rightarrow ? + {}^{1}_{1}\mathrm{H}$
(d) $^{235}_{92}\mathrm{U} \rightarrow ? + {}^{135}_{55}\mathrm{Cs} + 4{}^{1}_{0}\mathrm{n}$

14. Complete each of the following equations:

(a) $^{7}_{3}\mathrm{Li} + ? \rightarrow 2{}^{4}_{2}\mathrm{He}$
(b) $^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + \ ?$
(c) $^{27}_{13}\mathrm{Al} + {}^{4}_{2}\mathrm{He} \rightarrow ? + {}^{1}_{0}\mathrm{n}$
(d) $^{250}_{96}\mathrm{Cm} \rightarrow ? + {}^{98}_{38}\mathrm{Sr} + 4{}^{1}_{0}\mathrm{n}$

15. Write a balanced equation for each of the following nuclear reactions:

a. the production of ^{17}O from ^{14}N by α particle bombardment
b. the production of ^{14}C from ^{14}N by neutron bombardment
c. the production of ^{233}Th from ^{232}Th by neutron bombardment
d. the production of ^{239}U from ^{238}U by bombardment

16. Technetium-99 is prepared from ^{98}Mo. Molybdenum-98 combines with a neutron to give molybdenum-99, an unstable isotope that emits a β particle to yield an excited form of technetium-99, represented as ^{99}Tc*. This excited nucleus relaxes to the ground state, represented as ^{99}Tc, by emitting a γ ray. The ground state of ^{99}Tc then emits a β particle. Write the equations for each of these nuclear reactions.

17. The mass of the atom is 18.99840 amu.

a. Calculate its binding energy per atom in millions of electron volts.
b. Calculate its binding energy per nucleon.

18. For the reaction if 100.0 g of carbon reacts, what volume of nitrogen gas (N_2) is produced at 273K and 1 atm?

19. What are the types of radiation emitted by the nuclei of radioactive elements?

20. What changes occur to the atomic number and mass of a nucleus during each of the following decay scenarios?

a. an α particle is emitted
b. a β particle is emitted
c. γ radiation is emitted
d. a positron is emitted
e. an electron is captured

21. What is the change in the nucleus that results from the following decay scenarios?

a. emission of a β particle
b. emission of a β^{+} particle
c. capture of an electron

22. Many nuclides with atomic numbers greater than 83 decay by processes such as electron emission. Explain the observation that the emissions from these unstable nuclides also normally include α particles.

23. Why is electron capture accompanied by the emission of an X-ray?

24. Explain, in terms of Figure 21.2, how unstable heavy nuclides (atomic number > 83) may decompose to form nuclides of greater stability (a) if they are below the band of stability and (b) if they are above the band of stability.

25. Which of the following nuclei is most likely to decay by positron emission? Explain your choice.

a. chromium-53
b. manganese-51
c. iron-59

26. The following nuclei do not lie in the band of stability. How would they be expected to decay? Explain your answer.

(a) $^{34}_{15}\mathrm{P}$
(b) $^{239}_{92}U$
(c) $^{38}_{20}\mathrm{Ca}$
(d) $^{3}_{1}\mathrm{H}$
(e) $^{245}_{94}\mathrm{Pu}$

27. The following nuclei do not lie in the band of stability. How would they be expected to decay?

(a) $^{28}_{15}\mathrm{P}$
(b) $^{235}_{92}\mathrm{U}$
(c) $^{37}_{20}\mathrm{Ca}$
(d) $^{9}_{3}\mathrm{Li}$
(e) $^{245}_{96}\mathrm{Cm}$

28. Predict by what mode(s) of spontaneous radioactive decay each of the following unstable isotopes might proceed:

(a) $^{6}_{2}\mathrm{He}$
(b) $^{60}_{30}\mathrm{Zn}$
(c) $^{235}_{91}\ \mathrm{Pa}$
(d) $^{241}_{94}\ \mathrm{Np}$
(e) $^{18}\ \mathrm{F}$
(f) $^{129}\mathrm{Ba}$
(g) $^{237}\mathrm{Pu}$

29. Write a nuclear reaction for each step in the formation of from which proceeds by a series of decay reactions involving the step-wise emission of α, β, β, α, α, α particles, in that order.

30. Write a nuclear reaction for each step in the formation of from which proceeds by a series of decay reactions involving the step-wise emission of α, α, α, α, β, β, α particles, in that order.

31. Define the term half-life and illustrate it with an example.

32. A 1.00 10^{-6}-g sample of nobelium, has a half-life of 55 seconds after it is formed. What is the percentage of remaining at the following times?

a. 5.0 min after it forms
b. 1.0 h after it forms

33. ^{239}Pu is a nuclear waste byproduct with a half-life of 24,000 y. What fraction of the ^{239}Pu present today will be present in 1000 y?

34. The isotope ^{208}Tl undergoes β decay with a half-life of 3.1 min.

a. What isotope is produced by the decay?
b. How long will it take for 99.0% of a sample of pure ^{208}Tl to decay?
c. What percentage of a sample of pure ^{208}Tl remains un-decayed after 1.0 h?

35. If 1.000 g of produces 0.0001 mL of the gas at STP (standard temperature and pressure) in 24 h, what is the half-life of ^{226}Ra in years?

36. The isotope is one of the extremely hazardous species in the residues from nuclear power generation. The strontium in a 0.500-g sample diminishes to 0.393 g in 10.0 y. Calculate the half-life.

37. Technetium-99 is often used for assessing heart, liver, and lung damage because certain technetium compounds are absorbed by damaged tissues. It has a half-life of 6.0 h. Calculate the rate constant for the decay of

38. What is the age of mummified primate skin that contains 8.25% of the original quantity of ^{14}C?

39. A sample of rock was found to contain 8.23 mg of rubidium-87 and 0.47 mg of strontium-87.

a. Calculate the age of the rock if the half-life of the decay of rubidium by β emission is $4.7\ 10^{10}$ y.
b. If some was initially present in the rock, would the rock be younger, older, or the same age as the age calculated in (a)? Explain your answer.

40. A laboratory investigation shows that a sample of uranium ore contains 5.37 mg of and 2.52 mg of Calculate the age of the ore. The half-life of is $4.5\ 10^{9}$ yr.

41. Plutonium was detected in trace amounts in natural uranium deposits by Glenn Seaborg and his associates in 1941. They proposed that the source of this ^{239}Pu was the capture of neutrons by ^{238}U nuclei. Why is this plutonium not likely to have been trapped at the time the solar system formed $4.7\ 10^{9}$ years ago?

42. A atom (mass = 7.0169 amu) decays into a atom (mass = 7.0160 amu) by electron capture. How much energy (in millions of electron volts, MeV) is produced by this reaction?

43. A atom (mass = 8.0246 amu) decays into a atom (mass = 8.0053 amu) by loss of a β^+ particle (mass = 0.00055 amu) or by electron capture. How much energy (in millions of electron volts) is produced by this reaction?

44. Isotopes such as ^{26}Al (half-life: $7.2\ 10^{5}$ years) are believed to have been present in our solar system as it formed, but have since decayed and are now called extinct nuclides.

a. ^{26}Al decays by β^+ emission or electron capture. Write the equations for these two nuclear transformations.
b. The earth was formed about $4.7\ 10^{9}$ (4.7 billion) years ago. How old was the earth when 99.999999% of the ^{26}Al originally present had decayed?

45. Write a balanced equation for each of the following nuclear reactions:

a. bismuth-212 decays into polonium-212
b. beryllium-8 and a positron are produced by the decay of an unstable nucleus
c. neptunium-239 forms from the reaction of uranium-238 with a neutron and then spontaneously converts into plutonium-239
d. strontium-90 decays into yttrium-90

46. Write a balanced equation for each of the following nuclear reactions:

a. mercury-180 decays into platinum-176
b. zirconium-90 and an electron are produced by the decay of an unstable nucleus
c. thorium-232 decays and produces an alpha particle and a radium-228 nucleus, which decays into actinium-228 by beta decay
d. neon-19 decays into fluorine-19

47. Write the balanced nuclear equation for the production of the following transuranium elements:

a. berkelium-244, made by the reaction of Am-241 and He-4
b. fermium-254, made by the reaction of Pu-239 with a large number of neutrons
c. lawrencium-257, made by the reaction of Cf-250 and B-11
d. dubnium-260, made by the reaction of Cf-249 and N-15

48. How does nuclear fission differ from nuclear fusion? Why are both of these processes exothermic?

49. Both fusion and fission are nuclear reactions. Why is a very high temperature required for fusion, but not for fission?

50.Cite the conditions necessary for a nuclear chain reaction to take place. Explain how it can be controlled to produce energy, but not produce an explosion.

51. Describe the components of a nuclear reactor.

52. In usual practice, both a moderator and control rods are necessary to operate a nuclear chain reaction safely for the purpose of energy production. Cite the function of each and explain why both are necessary.

53. Describe how the potential energy of uranium is converted into electrical energy in a nuclear power plant.

54. The mass of a hydrogen atom is 1.007825 amu; that of a tritium atom is 3.01605 amu; and that of an α particle is 4.00150 amu. How much energy in kilojoules per mole of produced is released by the following fusion reaction:

55. How can a radioactive nuclide be used to show that the equilibrium is a dynamic equilibrium?

56. Technetium-99m has a half-life of 6.01 hours. If a patient injected with technetium-99m is safe to leave the hospital once 75% of the dose has decayed, when is the patient allowed to leave?

57. Iodine that enters the body is stored in the thyroid gland from which it is released to control growth and metabolism. The thyroid can be imaged if iodine-131 is injected into the body. In larger doses, I-133 is also used as a means of treating cancer of the thyroid. I-131 has a half-life of 8.70 days and decays by β^- emission.

a. Write an equation for the decay.
b. How long will it take for 95.0% of a dose of I-131 to decay?

58. If a hospital were storing radioisotopes, what is the minimum containment needed to protect against:

a. cobalt-60 (a strong γ emitter used for irradiation)
b. molybdenum-99 (a beta emitter used to produce technetium-99 for imaging)

59. Based on what is known about Radon-222's primary decay method, why is inhalation so dangerous?

60. Given specimens uranium-232 ($t_{1/2}$ = 68.9 y) and uranium-233 ($t_{1/2}$ = 159,200 y) of equal mass, which one would have greater activity and why?

61. A scientist is studying a 2.234 g sample of thorium-229 ($t_{1/2}$ = 7340 y) in a laboratory.

a. What is its activity in Bq?
b. What is its activity in Ci?

62. Given specimens neon-24 ($t_{1/2}$ = 3.38 min) and bismuth-211 ($t_{1/2}$ = 2.14 min) of equal mass, which one would have greater activity and why?

This page titled 11.11: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **21.10: Exercises** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

CHAPTER OVERVIEW

12: Kinetics

Chemical kinetics is the study of rates of chemical processes and includes investigations of how different experimental conditions can influence the speed of a chemical reaction and yield information about the reaction's mechanism and transition states, as well as the construction of mathematical models that can describe the characteristics of a chemical reaction.

Faculty Feature

Dr. Maria Dittrich's research explores the processes at the biosphere-geosphere interface, focusing on the geochemical cycling of carbon, nutrients, and metals in aquatic ecosystems. Her work examines how biogeochemical and physical processes, particularly at the sediment-water interface, influence nutrient release, pollutant dynamics, and carbon burial. Using field studies, lab experiments, and modeling, she investigates the impacts of seasonal variations, climate change, and human activities. Dr. Dittrich's findings provide critical insights into aquatic ecosystem responses to environmental changes, highlighting the relevance of chemical kinetics and reaction mechanisms in understanding complex natural systems.

Read more about Dr. Maria Dittrich and her research here: https://mdbiogeochemistry.com/



12.1: Introduction
12.2: Chemical Reaction Rates
12.3: Factors Affecting Reaction Rates
12.4: Rate Laws
12.5: Integrated Rate Laws
12.6: Collision Theory
12.7: Reaction Mechanisms
12.8: Catalysis
12.9: Key Terms
12.10: Key Equations
12.11: Summary
12.12: Exercises

This page titled 12: Kinetics is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

12.1: Introduction

The lizard in the photograph is not simply enjoying the sunshine or working on its tan. The heat from the sun's rays is critical to the lizard's survival. A warm lizard can move faster than a cold one because the chemical reactions that allow its muscles to move occur more rapidly at higher temperatures. A cold lizard is a slower lizard and an easier meal for predators.



Figure 12.1.1: An agama lizard basks in the sun. As its body warms, the chemical reactions of its metabolism speed up.

From baking a cake to determining the useful lifespan of a bridge, rates of chemical reactions play important roles in our understanding of processes that involve chemical changes. Two questions are typically posed when planning to carry out a chemical reaction. The first is: "Will the reaction produce the desired products in useful quantities?" The second question is: "How rapidly will the reaction occur?" A third question is often asked when investigating reactions in greater detail: "What specific molecular-level processes take place as the reaction occurs?" Knowing the answer to this question is of practical importance when the yield or rate of a reaction needs to be controlled.

The study of chemical kinetics concerns the second and third questions—that is, the rate at which a reaction yields products and the molecular-scale means by which a reaction occurs. This chapter examines the factors that influence the rates of chemical reactions, the mechanisms by which reactions proceed, and the quantitative techniques used to describe the rates at which reactions occur.

This page titled 12.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **12.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.2: Chemical Reaction Rates

Learning Objectives

By the end of this section, you will be able to:

- Define chemical reaction rate
- Derive rate expressions from the balanced equation for a given chemical reaction
- Calculate reaction rates from experimental data

A *rate* is a measure of how some property varies with time. Speed is a familiar rate that expresses the distance traveled by an object in a given amount of time. Wage is a rate that represents the amount of money earned by a person working for a given amount of time. Likewise, the rate of a chemical reaction is a measure of how much reactant is consumed, or how much product is produced, by the reaction in a given amount of time.

The **rate of reaction** is the change in the amount of a reactant or product per unit time. Reaction rates are therefore determined by measuring the time dependence of some property that can be related to reactant or product amounts. Rates of reactions that consume or produce gaseous substances, for example, are conveniently determined by measuring changes in volume or pressure. For reactions involving one or more colored substances, rates may be monitored via measurements of light absorption. For reactions involving aqueous electrolytes, rates may be measured via changes in a solution's conductivity.

For reactants and products in solution, their relative amounts (concentrations) are conveniently used for purposes of expressing reaction rates. For example, the concentration of hydrogen peroxide, H_2O_2, in an aqueous solution changes slowly over time as it decomposes according to the equation:

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

The rate at which the hydrogen peroxide decomposes can be expressed in terms of the rate of change of its concentration, as shown here:

$$\begin{aligned} H_2O_2 &= -\frac{\text{change in concentration of reactant}}{\text{time interval}} \\ \text{rate of decomposition of} \qquad &= -\frac{[H_2O_2]_{t_2} - [H_2O_2]_{t_1}}{t_2 - t_1} \\ &= -\frac{\Delta[H_2O_2]}{\Delta t} \end{aligned}$$

This mathematical representation of the change in species concentration over time is the **rate expression** for the reaction. The brackets indicate molar concentrations, and the symbol delta (Δ) indicates "change in." Thus, $[\mathrm{H_2O_2}]_{t_1}$ represents the molar concentration of hydrogen peroxide at some time t_1; likewise, $[\mathrm{H_2O_2}]_{t_2}$ represents the molar concentration of hydrogen peroxide at a later time t2; and Δ[H2O2] represents the change in molar concentration of hydrogen peroxide during the time interval Δt (that is, $t_2 - t_1$). Since the reactant concentration decreases as the reaction proceeds, $\Delta[\mathrm{H_2O_2}]$ is a negative quantity. Reaction rates are, by convention, positive quantities, and so this negative change in concentration is multiplied by −1. Figure 12.2.1provides an example of data collected during the decomposition of $\mathrm{H_2O_2}$.

Time (h)	$[H_2O_2]$ (mol L^{-1})	$\Delta[H_2O_2]$ (mol L^{-1})	Δt (h)	Rate of Decomposition, (mol L^{-1} h^{-1})
0.00	1.000			
		−0.500	6.00	0.0833
6.00	0.500			
		−0.250	6.00	0.0417
12.00	0.250			
		−0.125	6.00	0.0208
18.00	0.125			
		−0.062	6.00	0.010
24.00	0.0625			

Figure 12.2.1: The rate of decomposition of H_2O_2 in an aqueous solution decreases as the concentration of H_2O_2 decreases.

To obtain the tabulated results for this decomposition, the concentration of hydrogen peroxide was measured every 6 hours over the course of a day at a constant temperature of 40 °C. Reaction rates were computed for each time interval by dividing the change in concentration by the corresponding time increment, as shown here for the first 6-hour period:

$$\frac{-\Delta\left[H_2O_2\right]}{\Delta t}=\frac{-(0.500mol/L-1.000mol/L)}{(6.00h-0.00h)}=0.0833molL^{-1}h^{-1}$$

Notice that the reaction rates vary with time, decreasing as the reaction proceeds. Results for the last 6-hour period yield a reaction rate of:

$$\frac{-\Delta\left[H_2O_2\right]}{\Delta t}=\frac{-(0.0625mol/L-0.125mol/L)}{(24.00h-18.00h)}=0.010molL^{-1}h^{-1}$$

This behavior indicates the reaction continually slows with time. Using the concentrations at the beginning and end of a time period over which the reaction rate is changing results in the calculation of an **average rate** for the reaction over this time interval. At any specific time, the rate at which a reaction is proceeding is known as its **instantaneous rate**. The instantaneous rate of a reaction at "time zero," when the reaction commences, is its **initial rate**. Consider the analogy of a car slowing down as it approaches a stop sign. The vehicle's initial rate—analogous to the beginning of a chemical reaction—would be the speedometer reading at the moment the driver begins pressing the brakes (t_0). A few moments later, the instantaneous rate at a specific moment—call it t_1—would be somewhat slower, as indicated by the speedometer reading at that point in time. As time passes, the instantaneous rate will continue to fall until it reaches zero, when the car (or reaction) stops. Unlike instantaneous speed, the car's average speed is not indicated by the speedometer; but it can be calculated as the ratio of the distance traveled to the time required to bring the vehicle to a complete stop (Δt). Like the decelerating car, the average rate of a chemical reaction will fall somewhere between its initial and final rates.

The instantaneous rate of a reaction may be determined one of two ways. If experimental conditions permit the measurement of concentration changes over very short time intervals, then average rates computed as described earlier provide reasonably good approximations of instantaneous rates. Alternatively, a graphical procedure may be used that, in effect, yields the results that would be obtained if short time interval measurements were possible. In a plot of the concentration of hydrogen peroxide against time, the instantaneous rate of decomposition of H_2O_2 at any time t is given by the slope of a straight line that is tangent to the curve at that time (Figure 12.2.2). These tangent line slopes may be evaluated using calculus, but the procedure for doing so is beyond the scope of this chapter.

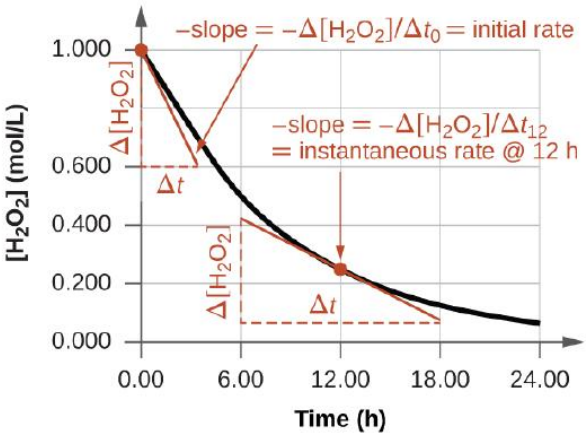


Figure 12.2.2: This graph shows a plot of concentration versus time for a 1.000 *M* solution of H_2O_2. The rate at any time is equal to the negative of the slope of a line tangent to the curve at that time. Tangents are shown at t = 0 h ("initial rate") and at t = 12 h ("instantaneous rate" at 12 h).

Chemistry in Everyday Life: Reaction Rates in Analysis: Test Strips for Urinalysis

Physicians often use disposable test strips to measure the amounts of various substances in a patient's urine (Figure 12.2.3). These test strips contain various chemical reagents, embedded in small pads at various locations along the strip, which undergo changes in color upon exposure to sufficient concentrations of specific substances. The usage instructions for test strips often stress that proper read time is critical for optimal results. This emphasis on read time suggests that kinetic aspects of the chemical reactions occurring on the test strip are important considerations.

The test for urinary glucose relies on a two-step process represented by the chemical equations shown here:

$$C_6H_{12}O_6 + O_2 \xrightarrow{\text{catalyst}} C_6H_{10}O_6 + H_2O_2$$

$$2H_2O_2 + 2I^- \xrightarrow{\text{catalyst}} I_2 + 2H_2O + O_2$$

The first equation depicts the oxidation of glucose in the urine to yield glucolactone and hydrogen peroxide. The hydrogen peroxide produced subsequently oxidizes colorless iodide ion to yield brown iodine, which may be visually detected. Some strips include an additional substance that reacts with iodine to produce a more distinct color change.

The two test reactions shown above are inherently very slow, but their rates are increased by special enzymes embedded in the test strip pad. This is an example of *catalysis*, a topic discussed later in this chapter. A typical glucose test strip for use with urine requires approximately 30 seconds for completion of the color-forming reactions. Reading the result too soon might lead one to conclude that the glucose concentration of the urine sample is lower than it actually is (a *false-negative* result). Waiting too long to assess the color change can lead to a *false positive* due to the slower (not catalyzed) oxidation of iodide ion by other substances found in urine.

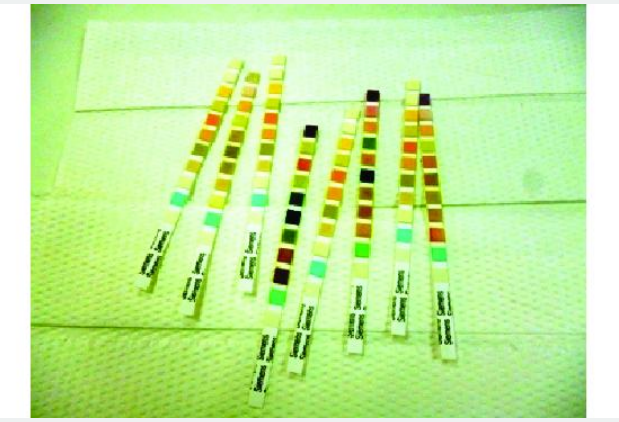


Figure 12.2.3: Test strips are commonly used to detect the presence of specific substances in a person's urine. Many test strips have several pads containing various reagents to permit the detection of multiple substances on a single strip. (credit: Iqbal Osman)

Relative Rates of Reaction

The rate of a reaction may be expressed as the change in concentration of any reactant or product. For any given reaction, these rate expressions are all related simply to one another according to the reaction stoichiometry. The rate of the general reaction

$$A \longrightarrow B \tag{12.2.1}$$

can be expressed in terms of the decrease in the concentration of A or the increase in the concentration of B. These two rate expressions are related by the stoichiometry of the reaction:

$$\text{rate} = -\left(\frac{1}{a}\right)\left(\frac{\Delta A}{\Delta t}\right) = \left(\frac{1}{b}\right)\left(\frac{\Delta B}{\Delta t}\right)$$

Consider the reaction represented by the following equation:

$$2NH_3(g) \longrightarrow N_2(g) + 3H_2(g)$$

The relation between the reaction rates expressed in terms of nitrogen production and ammonia consumption, for example, is:

$$-\frac{\Delta mol\ NH_3}{\Delta t} \times \frac{1 mol\ N_2}{2 mol\ NH_3} = \frac{\Delta mol\ N_2}{\Delta t}$$

This may be represented in an abbreviated format by omitting the units of the stoichiometric factor:

$$-\frac{1}{2}\frac{\Delta mol\ NH_3}{\Delta t} = \frac{\Delta mol\ N_2}{\Delta t}$$

Note that a negative sign has been included as a factor to account for the opposite signs of the two amount changes (the reactant amount is decreasing while the product amount is increasing). For homogeneous reactions, both the reactants and products are present in the same solution and thus occupy the same volume, so the molar amounts may be replaced with molar concentrations:

$$-\frac{1}{2}\frac{\Delta\left[NH_3\right]}{\Delta t}=\frac{\Delta\left[N_2\right]}{\Delta t}$$

Similarly, the rate of formation of H_2 is three times the rate of formation of N_2 because three moles of H_2 are produced for each mole of N_2 produced.

$$\frac{1}{3}\frac{\Delta\left[H_2\right]}{\Delta t}=\frac{\Delta\left[N_2\right]}{\Delta t}$$

Figure 12.2.4 illustrates the change in concentrations over time for the decomposition of ammonia into nitrogen and hydrogen at 1100 °C. Slopes of the tangent lines at t = 500 s show that the instantaneous rates derived from all three species involved in the reaction are related by their stoichiometric factors. The rate of hydrogen production, for example, is observed to be three times greater than that for nitrogen production:

$$\frac{2.91\times10^{-6}M/s}{9.70\times10^{-7}M/s}\approx3$$

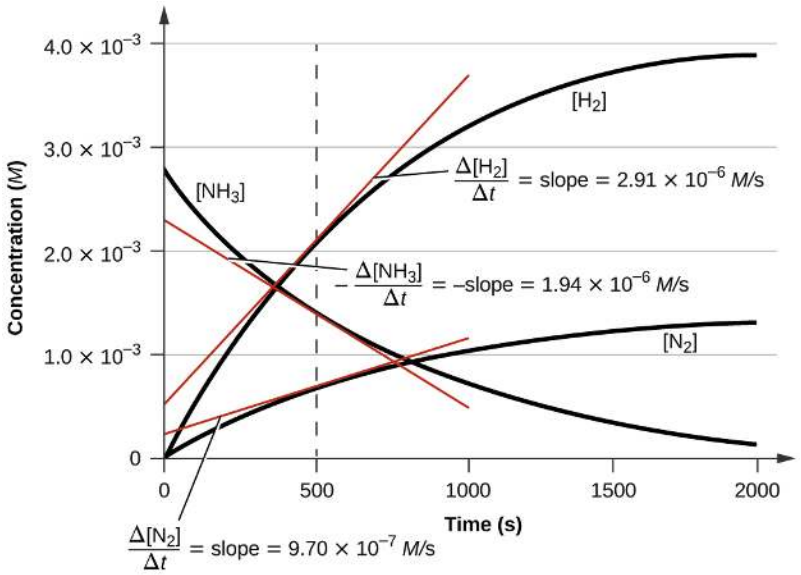


Figure 12.2.4: Changes in concentrations of the reactant and products for the reaction $2\,\mathrm{NH_3} \longrightarrow \mathrm{N_2} + 3\,\mathrm{H_2}$. The rates of change of the three concentrations are related by the reaction stoichiometry, as shown by the different slopes of the tangents at t = 500 s.

Example 12.2.1: Expressions for Relative Reaction Rates

The first step in the production of nitric acid is the combustion of ammonia:

$$4NH_3(g)+5O_2(g)\longrightarrow4NO(g)+6H_2O(g) \tag{12.2.2}$$

Write the equations that relate the rates of consumption of the reactants and the rates of formation of the products.

Solution

Considering the stoichiometry of this homogeneous reaction, the rates for the consumption of reactants and formation of products are:

$$-\frac{1}{4}\frac{\Delta\left[NH_3\right]}{\Delta t}=-\frac{1}{5}\frac{\Delta\left[O_2\right]}{\Delta t}=\frac{1}{4}\frac{\Delta[NO]}{\Delta t}=\frac{1}{6}\frac{\Delta\left[H_2O\right]}{\Delta t}$$

Exercise 12.2.1

The rate of formation of Br_2 is 6.0 10^{-6} mol/L/s in a reaction described by the following net ionic equation:

$$5Br^- + BrO_3^- + 6H^+ \longrightarrow 3Br_2 + 3H_2O$$

Write the equations that relate the rates of consumption of the reactants and the rates of formation of the products.

Answer

$$-\frac{1}{5}\frac{\Delta\left[Br^-\right]}{\Delta t} = -\frac{\Delta\left[BrO_3^-\right]}{\Delta t} = -\frac{1}{6}\frac{\Delta\left[H^+\right]}{\Delta t} = \frac{1}{3}\frac{\Delta\left[Br_2\right]}{\Delta t} = \frac{1}{3}\frac{\Delta\left[H_2O\right]}{\Delta t}$$

Example 12.2.2: Reaction Rate Expressions for Decomposition of H_2O_2

The graph in Figure 12.2.2 shows the rate of the decomposition of H_2O_2 over time:

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

Based on these data, the instantaneous rate of decomposition of H_2O_2 at t = 11.1 h is determined to be 3.20 10^{-2} mol/L/h, that is:

$$-\frac{1}{2}\frac{\Delta\left[\mathrm{H_2O_2}\right]}{\Delta t} = \frac{1}{2}\frac{\Delta\left[\mathrm{H_2O}\right]}{\Delta t} = \frac{\Delta\left[\mathrm{O_2}\right]}{\Delta t} \tag{12.2.3}$$

What is the instantaneous rate of production of $\mathrm{H_2O}$ and $\mathrm{O_2}$?

Solution

The reaction stoichiometry shows that

$$-\frac{1}{2}\frac{\Delta\left[\mathrm{H_2O_2}\right]}{\Delta t} = \frac{1}{2}\frac{\Delta\left[\mathrm{H_2O}\right]}{\Delta t} = \frac{\Delta\left[\mathrm{O_2}\right]}{\Delta t} \tag{12.2.4}$$

Therefore:

$$\frac{1}{2} \times 3.20 \times 10^{-2} molL^{-1}h^{-1} = \frac{\Delta\left[O_2\right]}{\Delta t}$$

and

$$\frac{\Delta\left[O_2\right]}{\Delta t} = 1.60 \times 10^{-2} molL^{-1}h^{-1}$$

Exercise 12.2.2

If the rate of decomposition of ammonia, NH_3, at 1150 K is 2.10 10^{-6} mol/L/s, what is the rate of production of nitrogen and hydrogen?

Answer

1.05 10^{-6} mol/L/s, N_2 and 3.15 10^{-6} mol/L/s, H_2.

This page titled 12.2: Chemical Reaction Rates is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **12.1: Chemical Reaction Rates** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.3: Factors Affecting Reaction Rates

Learning Objectives

By the end of this section, you will be able to:

- Describe the effects of chemical nature, physical state, temperature, concentration, and catalysis on reaction rates

The rates at which reactants are consumed and products are formed during chemical reactions vary greatly. Five factors typically affecting the rates of chemical reactions will be explored in this section: the chemical nature of the reacting substances, the state of subdivision (one large lump versus many small particles) of the reactants, the temperature of the reactants, the concentration of the reactants, and the presence of a catalyst.

The Chemical Nature of the Reacting Substances

The rate of a reaction depends on the nature of the participating substances. Reactions that appear similar may have different rates under the same conditions, depending on the identity of the reactants. For example, when small pieces of the metals iron and sodium are exposed to air, the sodium reacts completely with air overnight, whereas the iron is barely affected. The active metals calcium and sodium both react with water to form hydrogen gas and a base. Yet calcium reacts at a moderate rate, whereas sodium reacts so rapidly that the reaction is almost explosive.

The Physical States of the Reactants

A chemical reaction between two or more substances requires intimate contact between the reactants. When reactants are in different physical states, or phases (solid, liquid, gaseous, dissolved), the reaction takes place only at the interface between the phases. Consider the heterogeneous reaction between a solid phase and either a liquid or gaseous phase. Compared with the reaction rate for large solid particles, the rate for smaller particles will be greater because the surface area in contact with the other reactant phase is greater. For example, large pieces of iron react more slowly with acids than they do with finely divided iron powder (Figure 12.3.1). Large pieces of wood smolder, smaller pieces burn rapidly, and saw dust burns explosively.



Figure 12.3.1: (a) Iron powder reacts rapidly with dilute hydrochloric acid and produces bubbles of hydrogen gas: $2Fe(s) + 6HCl(aq)\ 2FeCl_3(aq) + 3H_2(g)$. (b) An iron nail reacts more slowly because the surface area exposed to the acid is much less.

Link to Learning

Watch this video to see the reaction of cesium with water in slow motion and a discussion of how the state of reactants and particle size affect reaction rates.

Temperature of the Reactants

Chemical reactions typically occur faster at higher temperatures. Food can spoil quickly when left on the kitchen counter. However, the lower temperature inside of a refrigerator slows that process so that the same food remains fresh for days. Gas burners, hot plates, and ovens are often used in the laboratory to increase the speed of reactions that proceed slowly at ordinary temperatures. For many chemical processes, reaction rates are approximately doubled when the temperature is raised by 10 °C.

Concentrations of the Reactants

The rates of many reactions depend on the concentrations of the reactants. Rates usually increase when the concentration of one or more of the reactants increases. For example, calcium carbonate ($CaCO_3$) deteriorates as a result of its reaction with the pollutant sulfur dioxide. The rate of this reaction depends on the amount of sulfur dioxide in the air (Figure 12.3.2). An acidic oxide, sulfur dioxide combines with water vapor in the air to produce sulfurous acid in the following reaction:

$$\mathrm{SO_2(g) + H_2O(g) \longrightarrow H_2SO_3(aq)}$$

Calcium carbonate reacts with sulfurous acid as follows:

$$\mathrm{CaCO_3(s) + H_2SO_3(aq) \longrightarrow CaSO_3(aq) + CO_2(g) + H_2O(l)}$$

In a polluted atmosphere where the concentration of sulfur dioxide is high, calcium carbonate deteriorates more rapidly than in less polluted air. Similarly, phosphorus burns much more rapidly in an atmosphere of pure oxygen than in air, which is only about 20% oxygen.



Figure 12.3.2: Statues made from carbonate compounds such as limestone and marble typically weather slowly over time due to the actions of water, and thermal expansion and contraction. However, pollutants like sulfur dioxide can accelerate weathering. As the concentration of air pollutants increases, deterioration of limestone occurs more rapidly. (credit: James P Fisher III)

Link to Learning

Phosphorus burns rapidly in air, but it will burn even more rapidly if the concentration of oxygen is higher. Watch this video to see an example.

The Presence of a Catalyst

Relatively dilute aqueous solutions of hydrogen peroxide, $\mathrm{H_2O_2}$, are commonly used as topical antiseptics. Hydrogen peroxide decomposes to yield water and oxygen gas according to the equation:

$$\mathrm{2\,H_2O_2(aq) \longrightarrow 2\,H_2O(l) + O_2(g)}$$

Under typical conditions, this decomposition occurs very slowly. When dilute $\mathrm{H_2O_2(aq)}$ is poured onto an open wound, however, the reaction occurs rapidly and the solution foams because of the vigorous production of oxygen gas. This dramatic difference is caused by the presence of substances within the wound's exposed tissues that accelerate the decomposition process. Substances that function to increase the rate of a reaction are called **catalysts**, a topic treated in greater detail later in this chapter.

Link to Learning

Chemical reactions occur when molecules collide with each other and undergo a chemical transformation. Before physically performing a reaction in a laboratory, scientists can use molecular modeling simulations to predict how the parameters discussed earlier will influence the rate of a reaction. Use the PhET Reactions & Rates interactive to explore how temperature, concentration, and the nature of the reactants affect reaction rates.

This page titled 12.3: Factors Affecting Reaction Rates is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **12.2: Factors Affecting Reaction Rates** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.4: Rate Laws

Learning Objectives

By the end of this section, you will be able to:

- Explain the form and function of a rate law
- Use rate laws to calculate reaction rates
- Use rate and concentration data to identify reaction orders and derive rate laws

As described in the previous module, the rate of a reaction is often affected by the concentrations of reactants. Rate laws (sometimes called *differential rate laws*) or rate equations are mathematical expressions that describe the relationship between the rate of a chemical reaction and the concentration of its reactants. As an example, consider the reaction described by the chemical equation

$$aA + bB \longrightarrow \text{ products}$$

where a and b are stoichiometric coefficients. The rate law for this reaction is written as:

$$\text{rate } = k[A]^m[B]^n$$

in which [*A*] and [*B*] represent the molar concentrations of reactants, and *k* is the **rate constant**, which is specific for a particular reaction at a particular temperature. The exponents *m* and *n* are the **reaction orders** and are typically positive integers, though they can be fractions, negative, or zero. The rate constant *k* and the reaction orders *m* and *n* must be determined experimentally by observing how the rate of a reaction changes as the concentrations of the reactants are changed. The rate constant *k* is independent of the reactant concentrations, but it does vary with temperature.

The reaction orders in a rate law describe the mathematical dependence of the rate on reactant concentrations. Referring to the generic rate law above, the reaction is *m* order with respect to *A* and *n* order with respect to *B*. For example, if *m* = 1 and *n* = 2, the reaction is first order in *A* and second order in *B*. The overall reaction order is simply the sum of orders for each reactant. For the example rate law here, the reaction is third order overall (1 + 2 = 3). A few specific examples are shown below to further illustrate this concept.

The rate law:

$$\text{rate } = k\left[H_2O_2\right]$$

describes a reaction that is first order in hydrogen peroxide and first order overall. The rate law:

$$\text{rate } = k\left[\mathrm{C_4H_6}\right]^2$$

describes a reaction that is second order in $\mathrm{C_4H_6}$ and second order overall. The rate law:

$$\text{rate } = k\left[\mathrm{H^+}\right]\left[\mathrm{OH^-}\right]$$

describes a reaction that is first order in $\mathrm{H^+}$, first order in $\mathrm{OH^-}$, and second order overall.

Example 12.4.1: Writing Rate Laws from Reaction Orders

An experiment shows that the reaction of nitrogen dioxide with carbon monoxide:

$$\mathrm{NO_2(g) + CO(g) \longrightarrow NO(g) + CO_2(g)}$$

is second order in $\mathrm{NO_2}$ and zero order in CO at 100 °C. What is the rate law for the reaction?

Solution

The reaction will have the form:

$$\text{rate } = k[\mathrm{NO_2}]^m[\mathrm{CO}]^n$$

The reaction is second order in $\mathrm{NO_2}$; thus $m = 2$. The reaction is zero order in CO; thus $n = 0$. The rate law is:

$$\text{rate} = k[\text{NO}_2]^2[\text{CO}]^0 = k[\text{NO}_2]^2$$

Remember that a number raised to the zero power is equal to 1, thus $[\text{CO}]^0 = 1$, which is why the CO concentration term may be omitted from the rate law: the rate of reaction is solely dependent on the concentration of NO_2. A later chapter section on reaction mechanisms will explain how a reactant's concentration can have no effect on a reaction rate despite being involved in the reaction.

Exercise 12.4.1A

The rate law for the reaction:

$$\text{H}_2(\text{g}) + 2\,\text{NO}(\text{g}) \longrightarrow \text{N}_2\text{O}(\text{g}) + \text{H}_2\text{O}(\text{g})$$

has been determined to be rate = $k[\text{NO}]^2[\text{H}_2]$. What are the orders with respect to each reactant, and what is the overall order of the reaction?

Answer

order in NO = 2; order in H_2 = 1; overall order = 3

Exercise 12.4.1B

In a transesterification reaction, a triglyceride reacts with an alcohol to form an ester and glycerol. Many students learn about the reaction between methanol (CH_3OH) and ethyl acetate ($\text{CH}_3\text{CH}_2\text{OCOCH}_3$) as a sample reaction before studying the chemical reactions that produce biodiesel:

$$\text{CH}_3\text{OH} + \text{CH}_3\text{CH}_2\text{OCOCH}_3 \longrightarrow \text{CH}_3\text{OCOCH}_3 + \text{CH}_3\text{CH}_2\text{OH}$$

The rate law for the reaction between methanol and ethyl acetate is, under certain conditions, determined to be:

$$\text{rate} = k\left[\text{CH}_3\text{OH}\right]$$

What is the order of reaction with respect to methanol and ethyl acetate, and what is the overall order of reaction?

Answer

order in CH_3OH = 1; order in $\text{CH}_3\text{CH}_2\text{OCOCH}_3$ = 0; overall order = 1

A common experimental approach to the determination of rate laws is the **method of initial rates**. This method involves measuring reaction rates for multiple experimental trials carried out using different initial reactant concentrations. Comparing the measured rates for these trials permits determination of the reaction orders and, subsequently, the rate constant, which together are used to formulate a rate law. This approach is illustrated in the next two example exercises.

Example 12.4.2: Determining a Rate Law from Initial Rates

Ozone in the upper atmosphere is depleted when it reacts with nitrogen oxides. The rates of the reactions of nitrogen oxides with ozone are important factors in deciding how significant these reactions are in the formation of the ozone hole over Antarctica (Figure 12.4.1). One such reaction is the combination of nitric oxide, NO, with ozone, O_3:

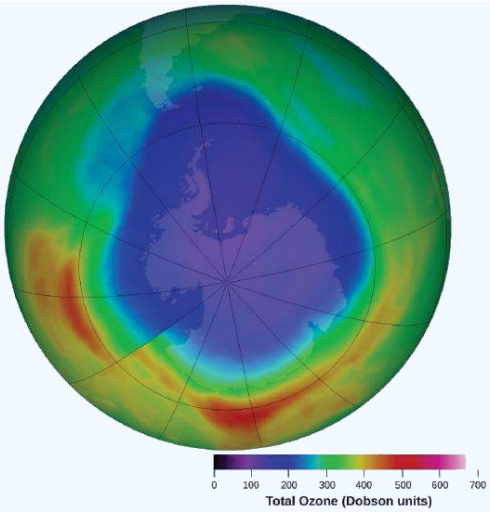


Figure 12.4.1: A contour map showing stratospheric ozone concentration and the “ozone hole” that occurs over Antarctica during its spring months. (credit: modification of work by NASA)

$$\mathrm{NO(g)} + \mathrm{O_3(g)} \longrightarrow \mathrm{NO_2(g)} + \mathrm{O_2(g)}$$

This reaction has been studied in the laboratory, and the following rate data were determined at 25 °C.

Trial	[NO] (mol/L)	[O_3] (mol/L)	$\frac{\Delta[\mathrm{NO_2}]}{\Delta t}(\mathrm{molL^{-1}s^{-1}})$
1	1.00×10^{-6}	3.00×10^{-6}	6.60×10^{-5}
2	1.00×10^{-6}	6.00×10^{-6}	1.32×10^{-4}
3	1.00×10^{-6}	9.00×10^{-6}	1.98×10^{-4}
4	2.00×10^{-6}	9.00×10^{-6}	3.96×10^{-4}
5	3.00×10^{-6}	9.00×10^{-6}	5.94×10^{-4}

Determine the rate law and the rate constant for the reaction at 25 °C.

Solution

The rate law will have the form:

$$\text{rate } = k[NO]^m[O_3]^n$$

Determine the values of *m*, *n*, and *k* from the experimental data using the following three-part process:

Step 1:

Determine the value of m *from the data in which [NO] varies and [O_3] is constant.* In the last three experiments, [NO] varies while [O_3] remains constant. When [NO] doubles from trial 3 to 4, the rate doubles, and when [NO] triples from trial 3 to 5, the rate also triples. Thus, the rate is also directly proportional to [NO], and *m* in the rate law is equal to 1.

Step 2:

Determine the value of n *from data in which [O_3] varies and [NO] is constant.* In the first three experiments, [NO] is constant and [O_3] varies. The reaction rate changes in direct proportion to the change in [O_3]. When [O_3] doubles from trial 1 to 2, the rate doubles; when [O_3] triples from trial 1 to 3, the rate increases also triples. Thus, the rate is directly proportional to [O_3], and *n* is equal to 1.The rate law is thus:

$$\text{rate } = k[NO]^1[O_3]^1 = k[NO]\,[O_3]$$

Step 3:

Determine the value of k *from one set of concentrations and the corresponding rate*. The data from trial 1 are used below:

$$k = \frac{\text{rate}}{[\text{NO}]\,[\text{O}_3]}$$
$$= \frac{6.60 \times 10^{-5}\ \text{mol L}^{-1}\text{s}^{-1}}{\left(1.00 \times 10^{-6}\ \text{mol L}^{-1}\right)\left(3.00 \times 10^{-6}\ \text{mol L}^{-1}\right)}$$
$$= 2.20 \times 10^{7}\ \text{L mol}^{-1}\text{s}^{-1}$$

Exercise 12.4.2

Acetaldehyde decomposes when heated to yield methane and carbon monoxide according to the equation:

$$\text{CH}_3\text{CHO(g)} \longrightarrow \text{CH}_4\text{(g)} + \text{CO(g)}$$

Determine the rate law and the rate constant for the reaction from the following experimental data:

Trial	[CH3CHO] (mol/L)	$-\frac{\Delta[\text{CH}_3\text{CHO}]}{\Delta t}(\text{mol L}^{-1}\text{s}^{-1})$
1	1.75×10^{-3}	2.06×10^{-11}
2	3.50×10^{-3}	8.24×10^{-11}
3	7.00×10^{-3}	3.30×10^{-10}

Answer

$\text{rate} = k\left[\text{CH}_3\text{CHO}\right]^2$ with k = 6.7310^{-6} L/mol/s

Example 12.4.3: Determining Rate Laws from Initial Rates

Using the initial rates method and the experimental data, determine the rate law and the value of the rate constant for this reaction:

$$2NO(g) + Cl_2(g) \longrightarrow 2NOCl(g)$$

Trial	[NO] (mol/L)	[Cl2] (mol/L)	$-\frac{\Delta[\text{CO}]}{\Delta t}(\text{mol } L^{-1}\text{s}^{-1})$
1	0.10	0.10	0.00300
2	0.10	0.15	0.00450
3	0.15	0.10	0.00675

Solution

The rate law for this reaction will have the form:

$$\text{rate} = k[NO]^m[Cl_2]^n$$

As in Example 12.4.2, approach this problem in a stepwise fashion, determining the values of *m* and *n* from the experimental data and then using these values to determine the value of *k*. In this example, however, an explicit algebraic approach (vs. the implicit approach of the previous example) will be used to determine the values of *m* and *n*:

Step 1

Determine the value of m from the data in which $[\text{NO}]$ varies and $[\text{Cl}_2]$ is constant. Write the ratios with the subscripts x and y to indicate data from two different trials:

$$\frac{\text{rate}_x}{\text{rate}_y} = \frac{k[NO]_x^m[Cl_2]_x^n}{k[NO]_y^m[Cl_2]_y^n}$$

Using the third trial and the first trial, in which $[Cl_2]$ does not vary, gives:

$$\frac{\text{rate } 3}{\text{rate } 1} = \frac{0.00675}{0.00300} = \frac{k(0.15)^m(0.10)^n}{k(0.10)^m(0.10)^n}$$

Canceling equivalent terms in the numerator and denominator leaves:

$$\frac{0.00675}{0.00300} = \frac{(0.15)^m}{(0.10)^m}$$

which simplifies to:

$$2.25 = (1.5)^m$$

Use logarithms to determine the value of the exponent *m*:

$$\ln(2.25) = m\ln(1.5)$$

$$\frac{\ln(2.25)}{\ln(1.5)} = m$$

$$2 = m$$

Confirm the result

$$1.5^2 = 2.25$$

Step 2

Determine the value of n from data in which $[Cl_2]$ varies and $[NO]$ is constant.

$$\frac{\text{rate } 2}{\text{rate } 1} = \frac{0.00450}{0.00300} = \frac{k0.10)^m(0.15)^n}{k(0.10)^m(0.10)^n}$$

Cancelation gives:

$$\frac{0.0045}{0.0030} = \frac{(0.15)^n}{(0.10)^n}$$

which simplifies to:

$$1.5 = (1.5)^n$$

Thus n must be 1, and the form of the rate law is:

$$\text{rate } = k[NO]^m[Cl_2]^n = k[NO]^2\,[Cl_2]$$

Step 3.

Determine the numerical value of the constant with appropriate units. The units for the rate of a reaction are mol/L/s. The units for are whatever is needed so that substituting into the rate law expression affords the appropriate units for the rate. In this example, the concentration units are mol^3/L^3. The units for *k* should be mol^- L/s so that the rate is in terms of mol/L/s.

To determine the value of *k* once the rate law expression has been solved, simply plug in values from the first experimental trial and solve for *k*:

$$0.00300 \text{ mol L}^{-1}\text{s}^{-1} = k\left(0.10 \text{ mol L}^{-1}\right)^2\left(0.10 \text{ mol L}^{-1}\right)^1$$

$$k = 3.0 \text{ mol}^{-2}\text{L}^2\text{s}^{-1}$$

Exercise 12.4.3

Use the provided initial rate data to derive the rate law for the reaction whose equation is:

$$\mathrm{OCl^-(aq) + I^-(aq) \longrightarrow OI^-(aq) + Cl^-(aq)}$$

Trial	[OCl^-] (mol/L)	[I^-] (mol/L)	Initial Rate (mol/L/s)
1	0.0040	0.0020	0.00184
2	0.0020	0.0040	0.00092
3	0.0020	0.0020	0.00046

Determine the rate law expression and the value of the rate constant k with appropriate units for this reaction.

Answer

$$\frac{\text{rate 2}}{\text{rate 3}} = \frac{0.00092}{0.00046} = \frac{k(0.0020)^x(0.0040)^y}{k(0.0020)^x(0.0020)^y}$$

$$2.00 = 2.00^y$$

$$y = 1$$

$$\frac{\text{rate 1}}{\text{rate 2}} = \frac{0.00184}{0.00092} = \frac{k(0.0040)^x(0.0020)^y}{k(0.0020)^x(0.0040)^y}$$

$$2.00 = \frac{2^x}{2^y}$$

$$2.00 = \frac{2^x}{2^1}$$

$$4.00 = 2^x$$

$$x = 2$$

Substituting the concentration data from trial 1 and solving for k yields:

$$\text{rate } = k\left[OCl^-\right]^2\left[I^-\right]^1$$

$$0.00184 = k(0.0040)^2(0.0020)^1$$

$$k = 5.75 \times 10^4 mol^{-2}L^2s^{-1}$$

Reaction Order and Rate Constant Units

In some of our examples, the reaction orders in the rate law happen to be the same as the coefficients in the chemical equation for the reaction. This is merely a coincidence and very often not the case.

Rate laws may exhibit fractional orders for some reactants, and negative reaction orders are sometimes observed when an increase in the concentration of one reactant causes a decrease in reaction rate. A few examples illustrating these points are provided:

$$\mathrm{NO_2 + CO \longrightarrow NO + CO_2} \qquad \text{rate } = k[NO_2]^2$$

$$\mathrm{CH_3CHO \longrightarrow CH_4 + CO} \qquad \text{rate } = k[CH_3CHO]^2$$

$$\mathrm{2\,N_2O_5 \longrightarrow NO_2 + O_2} \qquad \text{rate } = k\left[N_2O_5\right]$$

$$\mathrm{2\,NO_2 + F_2 \longrightarrow 2\,NO_2F} \qquad \text{rate } = k\left[NO_2\right]\left[F_2\right]$$

$$\mathrm{2\,NO_2Cl \longrightarrow 2\,NO_2 + Cl_2} \qquad \text{rate } = k\left[NO_2Cl\right]$$

Note

It is important to note that *rate laws are determined by experiment only and are not reliably predicted by reaction stoichiometry.*

The units for a rate constant will vary as appropriate to accommodate the overall order of the reaction. The unit of the rate constant for the second-order reaction described in Example 12.4.2 was determined to be k was derived to be $\mathrm{L\,mol^{-1}s^{-1}}$. For the third-order reaction, k was derived to be $\mathrm{L^2mol^{-2}s^{-1}}$. Dimensional analysis requires the rate constant unit for a reaction whose overall order is x to be $\mathrm{L^{x-1}\ mol^{1-x}s^{-1}}$. Table 12.4.1 summarizes the rate constant units for common reaction orders.

Table 12.4.1: Rate Constant Units for Common Reaction Orders

Overall Reaction Order (x)	Rate Constant Unit ($\mathrm{L^{x-1}\ mol^{1-x}\ s^{-1}}$)
0 (zero)	$\mathrm{mol\ L^{-1}\ s^{-1}}$
1 (first)	$\mathrm{s^{-1}}$
2 (second)	$\mathrm{L\ mol^{-1}\ s^{-1}}$
3 (third)	$\mathrm{L^2\ mol^{-2}\ s^{-1}}$

Note that the units in this table were derived using specific units for concentration (mol/L) and time (s), though any valid units for these two properties may be used.

This page titled 12.4: Rate Laws is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **12.3: Rate Laws** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.5: Integrated Rate Laws

Learning Objectives

By the end of this section, you will be able to:

- Explain the form and function of an integrated rate law
- Perform integrated rate law calculations for zero-, first-, and second-order reactions
- Define half-life and carry out related calculations
- Identify the order of a reaction from concentration/time data

The rate laws discussed thus far relate the rate and the concentrations of reactants. We can also determine a second form of each rate law that relates the concentrations of reactants and time. These are called **integrated rate laws**. We can use an integrated rate law to determine the amount of reactant or product present after a period of time or to estimate the time required for a reaction to proceed to a certain extent. For example, an integrated rate law is used to determine the length of time a radioactive material must be stored for its radioactivity to decay to a safe level.

Using calculus, the differential rate law for a chemical reaction can be integrated with respect to time to give an equation that relates the amount of reactant or product present in a reaction mixture to the elapsed time of the reaction. This process can either be very straightforward or very complex, depending on the complexity of the differential rate law. For purposes of discussion, we will focus on the resulting integrated rate laws for first-, second-, and zero-order reactions.

First-Order Reactions

Integration of the rate law for a simple first-order reaction (rate = $k[A]$) results in an equation describing how the reactant concentration varies with time:

$$[A]_t = [A]_0 e^{-kt}$$

where $[A]t$ is the concentration of A at any time t, $[A]_0$ is the initial concentration of A, and k is the first-order rate constant.

For mathematical convenience, this equation may be rearranged to other formats, including direct and indirect proportionalities:

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt$$

or

$$\ln\left(\frac{[A]_0}{[A]_t}\right) = kt$$

and a format showing a linear dependence of concentration in time:

$$\ln[A]_t = \ln[A]_0 - kt$$

Example 12.5.1: The Integrated Rate Law for a First-Order Reaction

The rate constant for the first-order decomposition of cyclobutane, C_4H_8 at 500 °C is 9.2 10^{-3} s^{-1}:

$$C_4H_8 \longrightarrow 2\,C_2H_4$$

How long will it take for 80.0% of a sample of C_4H_8 to decompose?

Solution

Since the relative change in reactant concentration is provided, a convenient format for the integrated rate law is:

$$\ln\left(\frac{[A]_0}{[A]_t}\right) = kt$$

The initial concentration of C_4H_8, $[A]_0$, is not provided, but the provision that 80.0% of the sample has decomposed is enough information to solve this problem. Let x be the initial concentration, in which case the concentration after 80.0% decomposition

is 20.0% of *x* or 0.200*x*. Rearranging the rate law to isolate *t* and substituting the provided quantities yields:

$$\begin{aligned} t &= \ln \frac{[x]}{[0.200x]} \times \frac{1}{k} \\ &= \ln 5 \times \frac{1}{9.2 \times 10^{-3}\ \text{s}^{-1}} \\ &= 1.609 \times \frac{1}{9.2 \times 10^{-3}\ \text{s}^{-1}} \\ &= 1.7 \times 10^{2}\ \text{s} \end{aligned}$$

Exercise 12.5.1

Iodine-131 is a radioactive isotope that is used to diagnose and treat some forms of thyroid cancer. Iodine-131 decays to xenon-131 according to the equation:

$$\text{I-131} \longrightarrow \text{Xe-131} + \text{electron}$$

The decay is first-order with a rate constant of 0.138 d^{-1}. How many days will it take for 90% of the iodine−131 in a 0.500 *M* solution of this substance to decay to Xe-131?

Answer

16.7 days

In the next example exercise, a linear format for the integrated rate law will be convenient:

$$\begin{aligned} \ln[A]_t &= (-k)(t) + \ln[A]_0 \\ y &= mx + b \end{aligned}$$

A plot of $\ln[A]_t$ versus *t* for a first-order reaction is a straight line with a slope of $-k$ and a *y*-intercept of $\ln[A]_0$. If a set of rate data are plotted in this fashion but do *not* result in a straight line, the reaction is not first order in *A*.

Example 12.5.2: Graphical Determination of Reaction Order and Rate Constant

Show that the data in Figure 12.5.1 can be represented by a first-order rate law by graphing $\ln[H_2O_2]$ versus time. Determine the rate constant for the decomposition of H_2O_2 from these data.

Solution

The data from Figure 12.5.1 are tabulated below, and a plot of $\ln[H_2O_2]$ is shown in Figure 12.5.1.

Time (h)	$[H_2O_2]$ (*M*)	$\ln[H_2O_2]$
0.00	1.000	0.000
6.00	0.500	−0.693
12.00	0.250	−1.386
18.00	0.125	−2.079
24.00	0.0625	−2.772

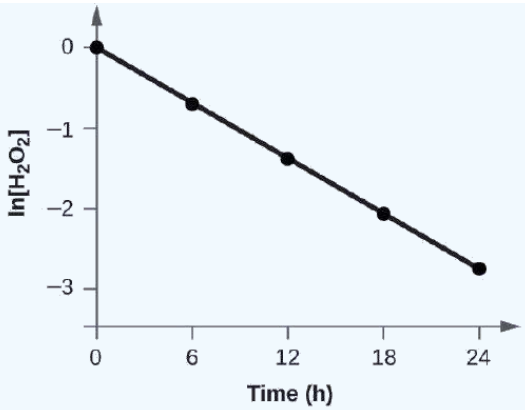


Figure 12.5.1: A linear relationship between $\ln[H_2O_2]$ and time suggests the decomposition of hydrogen peroxide is a first-order reaction.

The plot of $\ln[H_2O_2]$ versus time is linear, indicating that the reaction may be described by a first-order rate law.

According to the linear format of the first-order integrated rate law, the rate constant is given by the negative of this plot's slope.

$$\text{slope} = \frac{\text{change in } y}{\text{change in } x} = \frac{\Delta y}{\Delta x} = \frac{\Delta \ln[H_2O_2]}{\Delta t}$$

The slope of this line may be derived from two values of $\ln[H_2O_2]$ at different values of t (one near each end of the line is preferable). For example, the value of $\ln[H_2O_2]$ when t is 0.00 h is 0.000; the value when t = 24.00 h is −2.772

$$\begin{aligned}\text{slope} &= \frac{-2.772 - 0.000}{24.00 - 0.00h} \\ &= \frac{-2.772}{24.00h} \\ &= -0.116h^{-1} \\ k &= -\text{ slope} = -\left(-0.116h^{-1}\right) = 0.116h^{-1}\end{aligned}$$

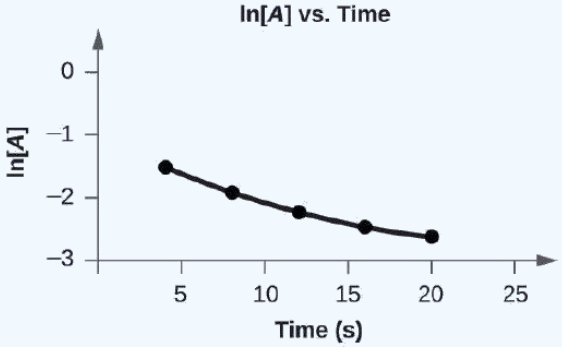
Exercise 12.5.2

Graph the following data to determine whether the reaction $A \longrightarrow B + C$ is first order.

Time (s)	[A]
4.0	0.220
8.0	0.144
12.0	0.110
16.0	0.088
20.0	0.074

Answer

The plot of $\ln[A]_t$ vs. t is not linear, indicating the reaction is not first order:



Second-Order Reactions

The equations that relate the concentrations of reactants and the rate constant of second-order reactions can be fairly complicated. To illustrate the point with minimal complexity, only the simplest second-order reactions will be described here, namely, those whose rates depend on the concentration of just one reactant. For these types of reactions, the differential rate law is written as:

$$\text{rate } = k[A]^2$$

For these second-order reactions, the integrated rate law is:

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

where the terms in the equation have their usual meanings as defined earlier.

Example 12.5.3: The Integrated Rate Law for a Second-Order Reaction

The reaction of butadiene gas (C_4H_6) to yield C_8H_{12} gas is described by the equation:

$$2C_4H_6(g) \longrightarrow C_8H_{12}(g)$$

This "dimerization" reaction is second order with a rate constant equal to 5.76 10^{-2} L mol^{-1} min^{-1} under certain conditions. If the initial concentration of butadiene is 0.200 *M*, what is the concentration after 10.0 min?

Solution

For a second-order reaction, the integrated rate law is written

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

We know three variables in this equation: $[A]_0$ = 0.200 mol/L, k = 5.76 10^{-2} L/mol/min, and t = 10.0 min. Therefore, we can solve for $[A]$, the fourth variable:

$$\begin{aligned}\frac{1}{[A]_t} &= \left(5.76 \times 10^{-2} L mol^{-1} min^{-1}\right)(10 min) + \frac{1}{0.200 mol^{-1}} \\ &= \left(5.76 \times 10^{-1} L mol^{-1}\right) + 5.00 L mol^{-1} \\ &= 5.58 L mol^{-1} \\ [A]_t &= 1.79 \times 10^{-1} mol L^{-1}\end{aligned}$$

Therefore 0.179 mol/L of butadiene remain at the end of 10.0 min, compared to the 0.200 mol/L that was originally present.

Exercise 12.5.3

If the initial concentration of butadiene is 0.0200 *M*, what is the concentration remaining after 20.0 min?

Answer

0.0195 mol/L

The integrated rate law for second-order reactions has the form of the equation of a straight line:

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

$$y = mx + b$$

A plot of $\frac{1}{[A]_t}$ versus t for a second-order reaction is a straight line with a slope of k and a y-intercept of $\frac{1}{[A]_0}$. If the plot is not a straight line, then the reaction is not second order.

Example 12.5.4: Graphical Determination of Reaction Order and Rate Constant

The data below are for the same reaction described in Example 12.5.3. Prepare and compare two appropriate data plots to identify the reaction as being either first or second order. After identifying the reaction order, estimate a value for the rate constant.

Solution

Time (s)	$[C_4H_6]$ (M)
0	1.00×10^{-2}
1600	5.04×10^{-3}
3200	3.37×10^{-3}
4800	2.53×10^{-3}
6200	2.08×10^{-3}

In order to distinguish a first-order reaction from a second-order reaction, prepare a plot of $\ln[C_4H_6]_t$ versus t and compare it to a plot of $\frac{1}{[C_4H_6]_t}$ versus t. The values needed for these plots follow.

Time (s)	$\frac{1}{[C_4H_6]}(M^{-1})$	$\ln[C_4H_6]$
0	100	−4.605
1600	198	−5.289
3200	296	−5.692
4800	395	−5.978
6200	481	−6.175

The plots are shown in Figure 12.5.2, which clearly shows the plot of $\ln[C_4H_6]_t$ versus t is not linear, therefore the reaction is not first order. The plot of $\frac{1}{[C_4H_6]_t}$ versus t is linear, indicating that the reaction is second order.

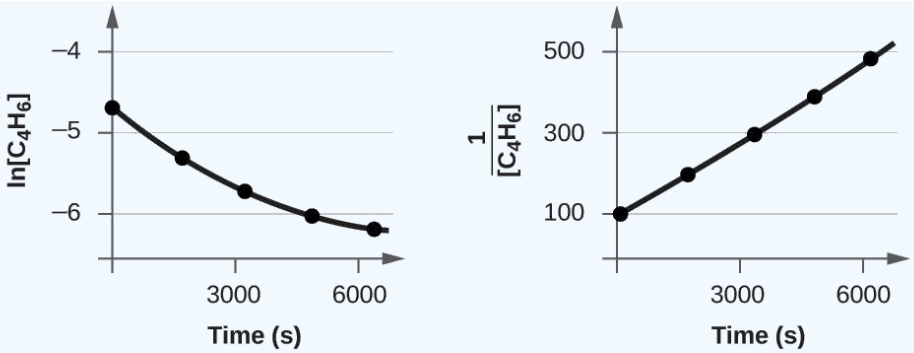


Figure 12.5.2: These two graphs show first- and second-order plots for the dimerization of C_4H_6. The linear trend in the second-order plot (right) indicates that the reaction follows second-order kinetics.

According to the second-order integrated rate law, the rate constant is equal to the slope of the versus *t* plot. Using the data for *t* = 0 *s* and *t* = 6200 *s*, the rate constant is estimated as follows:

$$k = \text{ slope } = \frac{\left(481M^{-1} - 100M^{-1}\right)}{(6200s - 0s)} = 0.0614M^{-1}s^{-1}$$

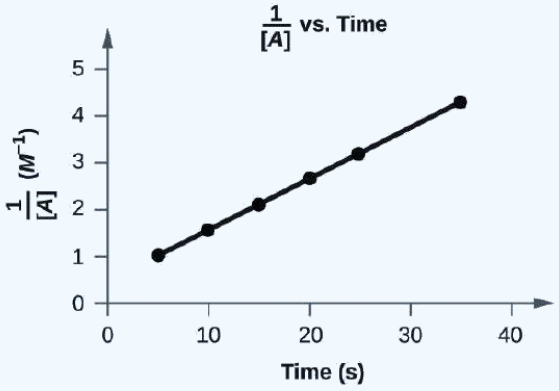
Exercise 12.5.4

Do the following data fit a second-order rate law?

Time (s)	[*A*] (*M*)
5	0.952
10	0.625
15	0.465
20	0.370
25	0.308
35	0.230

Answer

Yes. The plot of $\frac{1}{[A]_t}$ vs. *t* is linear:



Zero-Order Reactions

For zero-order reactions, the differential rate law is:

$$\text{rate} = k$$

A zero-order reaction thus exhibits a constant reaction rate, regardless of the concentration of its reactant(s). This may seem counterintuitive, since the reaction rate certainly can't be finite when the reactant concentration is zero. For purposes of this introductory text, it will suffice to note that zero-order kinetics are observed for some reactions only under certain specific conditions. These same reactions exhibit different kinetic behaviors when the specific conditions aren't met, and for this reason the more prudent term *pseudo-zero-order* is sometimes used.

The integrated rate law for a zero-order reaction is a linear function:

$$[A]_t = -kt + [A]_0$$
$$y = mx + b$$

A plot of [*A*] versus *t* for a zero-order reaction is a straight line with a slope of $-k$ and a *y*-intercept of $[A]_0$. Figure 12.5.3 shows a plot of [NH_3] versus *t* for the thermal decomposition of ammonia at the surface of two different heated solids. The decomposition reaction exhibits first-order behavior at a quartz (SiO_2) surface, as suggested by the exponentially decaying plot of concentration versus time. On a tungsten surface, however, the plot is linear, indicating zero-order kinetics.

Example 12.5.5: Graphical Determination of Zero-Order Rate Constant

Use the data plot in Figure 12.5.3 to graphically estimate the zero-order rate constant for ammonia decomposition at a tungsten surface.

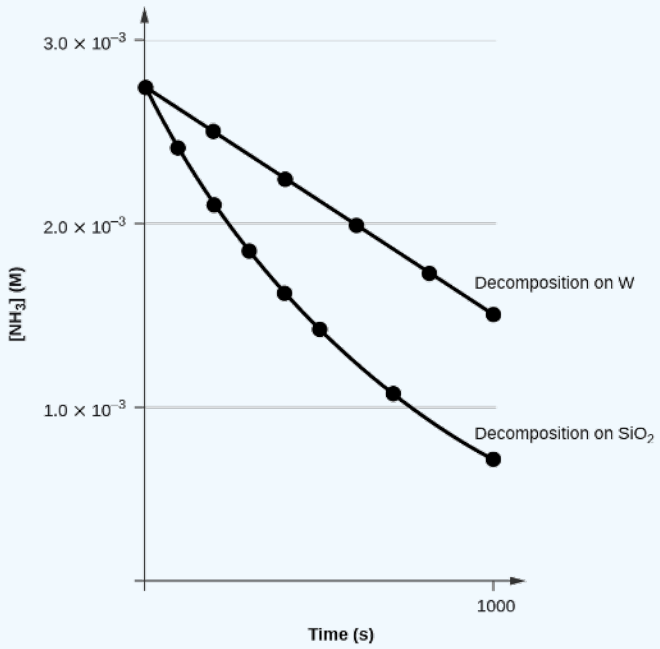


Figure 12.5.3: The decomposition of NH_3 on a tungsten (W) surface is a zero-order reaction, whereas on a quartz (SiO_2) surface, the reaction is first order.

Solution

The integrated rate law for zero-order kinetics describes a linear plot of reactant concentration, $[A]_t$, versus time, *t*, with a slope equal to the negative of the rate constant, $-k$. Following the mathematical approach of previous examples, the slope of the linear data plot (for decomposition on W) is estimated from the graph. Using the ammonia concentrations at *t* = 0 and *t* = 1000 s:

$$k = -\text{ slope } = -\frac{\left(0.0015 molL^{-1} - 0.0028 molL^{-1}\right)}{(1000s - 0s)} = 1.3 \times 10^{-6} molL^{-1}s^{-1}$$

Exercise 12.5.5

The zero-order plot in Figure 12.5.3 shows an initial ammonia concentration of 0.0028 mol L^{-1} decreasing linearly with time for 1000 s. Assuming no change in this zero-order behavior, at what time (min) will the concentration reach 0.0001 mol L^{-1}?

Answer

35 min

The Half-Life of a Reaction

The half-life of a reaction ($t_{1/2}$) is the time required for one-half of a given amount of reactant to be consumed. In each succeeding half-life, half of the remaining concentration of the reactant is consumed. Using the decomposition of hydrogen peroxide (Figure 12.5.1) as an example, we find that during the first half-life (from 0.00 hours to 6.00 hours), the concentration of H_2O_2 decreases from 1.000 *M* to 0.500 *M*. During the second half-life (from 6.00 hours to 12.00 hours), it decreases from 0.500 *M* to 0.250 *M*; during the third half-life, it decreases from 0.250 *M* to 0.125 *M*. The concentration of H_2O_2 decreases by half during each successive period of 6.00 hours. The decomposition of hydrogen peroxide is a first-order reaction, and, as can be shown, the half-life of a first-order reaction is independent of the concentration of the reactant. However, half-lives of reactions with other orders depend on the concentrations of the reactants.

First-Order Reactions

An equation relating the half-life of a first-order reaction to its rate constant may be derived from the integrated rate law as follows:

$$\ln \frac{[A]_0}{[A]_t} = kt$$

$$t = \ln \frac{[A]_0}{[A]_t} \times \frac{1}{k}$$

Invoking the definition of half-life, symbolized requires that the concentration of *A* at this point is one-half its initial concentration: $t = t_{1/2}$, $[A]_t = \frac{1}{2}[A]_0$

Substituting these terms into the rearranged integrated rate law and simplifying yields the equation for half-life:

$$t_{1/2} = \ln \frac{[A]_0}{\frac{1}{2}[A]_0} \times \frac{1}{k}$$

$$= \ln 2 \times \frac{1}{k} = 0.693 \times \frac{1}{k}$$

$$t_{1/2} = \frac{0.693}{k}$$

This equation describes an expected inverse relation between the half-life of the reaction and its rate constant, *k*. Faster reactions exhibit larger rate constants and correspondingly shorter half-lives. Slower reactions exhibit smaller rate constants and longer half-lives.

Example 12.5.6: Calculation of a First-order Rate Constant using Half-Life

Calculate the rate constant for the first-order decomposition of hydrogen peroxide in water at 40 °C, using the data given in Figure 12.5.4



Figure 12.5.4: The decomposition of H_2O_2 at 40 °C is illustrated. The intensity of the color symbolizes the concentration of H_2O_2 at the indicated times; H_2O_2 is actually colorless.

Solution

Inspecting the concentration/time data in Figure 12.5.4 shows the half-life for the decomposition of H_2O_2 is 2.16×10^4 s:

$$t_{1/2} = \frac{0.693}{k}$$

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{2.16 \times 10^4 s} = 3.21 \times 10^{-5} s^{-1}$$

? Exercise 12.5.1

The first-order radioactive decay of iodine-131 exhibits a rate constant of 0.138 d^{-1}. What is the half-life for this decay?

Answer

5.02 d.

Second-Order Reactions

Following the same approach as used for first-order reactions, an equation relating the half-life of a second-order reaction to its rate constant and initial concentration may be derived from its integrated rate law:

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

or

$$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$$

Restrict t to $t_{1/2}$

$$t = t_{1/2}$$

define $[A]_t$ as one-half $[A]_0$

$$[A]_t = \frac{1}{2}[A]_0$$

and then substitute into the integrated rate law and simplify:

$$\frac{1}{\frac{1}{2}[A]_0} - \frac{1}{[A]_0} = kt_{1/2}$$

$$\frac{2}{[A]_0} - \frac{1}{[A]_0} = kt_{1/2}$$

$$\frac{1}{[A]_0} = kt_{1/2}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

For a second-order reaction, $t_{1/2}$ is inversely proportional to the concentration of the reactant, and the half-life increases as the reaction proceeds because the concentration of reactant decreases. Unlike with first-order reactions, the rate constant of a second-order reaction cannot be calculated directly from the half-life unless the initial concentration is known.

Zero-Order Reactions

As for other reaction orders, an equation for zero-order half-life may be derived from the integrated rate law:

$$[A] = -kt + [A]_0$$

Restricting the time and concentrations to those defined by half-life: and Substituting these terms into the zero-order integrated rate law yields:

$$\frac{[A]_0}{2} = -kt_{1/2} + [A]_0$$

$$kt_{1/2} = \frac{[A]_0}{2}$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

As for all reaction orders, the half-life for a zero-order reaction is inversely proportional to its rate constant. However, the half-life of a zero-order reaction increases as the initial concentration increases.

Equations for both differential and integrated rate laws and the corresponding half-lives for zero-, first-, and second-order reactions are summarized in Table 12.5.1.

Table 12.5.1: Summary of Rate Laws for Zero-, First-, and Second-Order Reactions

	Zero-Order	First-Order	Second-Order
rate law	rate = k	rate = $k[A]$	rate = $k[A]^2$
units of rate constant	M s^{-1}	s^{-1}	M^{-1} s^{-1}
integrated rate law	$[A] = -kt + [A]_0$	$\ln[A] = -kt + \ln[A]_0$	$\frac{1}{[A]} = kt + \left(\frac{1}{[A]_0}\right)$
plot needed for linear fit of rate data	$[A]$ vs. t	ln$[A]$ vs. t	vs. t
relationship between slope of linear plot and rate constant	k = −slope	k = −slope	k = slope
half-life	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{0.693}{k}$	$t_{1/2} = \frac{1}{[A]_0 k}$

Example 12.5.7: Half-Life for Zero-Order and Second-Order Reactions

What is the half-life for the butadiene dimerization reaction described in Example 12.5.3?

Solution

The reaction in question is second order, is initiated with a 0.200 mol L^{-1} reactant solution, and exhibits a rate constant of 0.0576 L mol^{-1} min^{-1}. Substituting these quantities into the second-order half-life equation:

$$t_{1/2} = \frac{1}{\left[\left(0.0576 L mol^{-1} min^{-1}\right)\left(0.200 mol L^{-1}\right)\right]} = 86.8 \text{ min}$$

Exercise 12.5.7

What is the half-life (min) for the thermal decomposition of ammonia on tungsten (see Figure 12.5.3)?

Answer

87 min

This page titled 12.5: Integrated Rate Laws is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **12.4: Integrated Rate Laws** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.6: Collision Theory

Learning Objectives

By the end of this section, you will be able to:

- Use the postulates of collision theory to explain the effects of physical state, temperature, and concentration on reaction rates
- Define the concepts of activation energy and transition state
- Use the Arrhenius equation in calculations relating rate constants to temperature

We should not be surprised that atoms, molecules, or ions must collide before they can react with each other. Atoms must be close together to form chemical bonds. This simple premise is the basis for a very powerful theory that explains many observations regarding chemical kinetics, including factors affecting reaction rates.

Collision theory is based on the following postulates:

1. The rate of a reaction is proportional to the rate of reactant collisions:

$$\text{reaction rate} \propto \frac{\#\ \text{collisions}}{\text{time}}$$

2. The reacting species must collide in an orientation that allows contact between the atoms that will become bonded together in the product.
3. The collision must occur with adequate energy to permit mutual penetration of the reacting species' valence shells so that the electrons can rearrange and form new bonds (and new chemical species).

We can see the importance of the two physical factors noted in postulates 2 and 3, the orientation and energy of collisions, when we consider the reaction of carbon monoxide with oxygen:

$$2\,CO(g) + O_2(g) \longrightarrow 2\,CO_2(g)$$

Carbon monoxide is a pollutant produced by the combustion of hydrocarbon fuels. To reduce this pollutant, automobiles have catalytic converters that use a catalyst to carry out this reaction. It is also a side reaction of the combustion of gunpowder that results in muzzle flash for many firearms. If carbon monoxide and oxygen are present in sufficient amounts, the reaction will occur at high temperature and pressure.

The first step in the gas-phase reaction between carbon monoxide and oxygen is a collision between the two molecules:

$$CO(g) + O_2(g) \longrightarrow CO_2(g) + O(g)$$

Although there are many different possible orientations the two molecules can have relative to each other, consider the two presented in Figure 12.6.1. In the first case, the oxygen side of the carbon monoxide molecule collides with the oxygen molecule. In the second case, the carbon side of the carbon monoxide molecule collides with the oxygen molecule. The second case is clearly more likely to result in the formation of carbon dioxide, which has a central carbon atom bonded to two oxygen atoms (\ce{O=C=O}\). This is a rather simple example of how important the orientation of the collision is in terms of creating the desired product of the reaction.

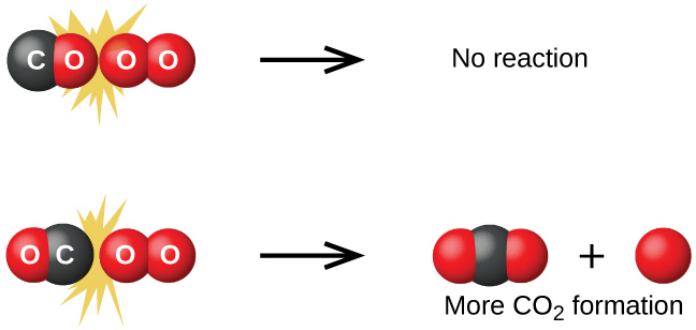


Figure 12.6.1: Illustrated are two collisions that might take place between carbon monoxide and oxygen molecules. The orientation of the colliding molecules partially determines whether a reaction between the two molecules will occur.

If the collision does take place with the correct orientation, there is still no guarantee that the reaction will proceed to form carbon dioxide. In addition to a proper orientation, the collision must also occur with sufficient energy to result in product formation. When reactant species collide with both proper orientation and adequate energy, they combine to form an unstable species called an **activated complex** or a **transition state**. These species are very short lived and usually undetectable by most analytical instruments. In some cases, sophisticated spectral measurements have been used to observe transition states.

Collision theory explains why most reaction rates increase as concentrations increase. With an increase in the concentration of any reacting substance, the chances for collisions between molecules are increased because there are more molecules per unit of volume. More collisions mean a faster reaction rate, assuming the energy of the collisions is adequate.

Activation Energy and the Arrhenius Equation

The minimum energy necessary to form a product during a collision between reactants is called the **activation energy** (E_a). How this energy compares to the kinetic energy provided by colliding reactant molecules is a primary factor affecting the rate of a chemical reaction. If the activation energy is much larger than the average kinetic energy of the molecules, the reaction will occur slowly since only a few fast-moving molecules will have enough energy to react. If the activation energy is much smaller than the average kinetic energy of the molecules, a large fraction of molecules will be adequately energetic and the reaction will proceed rapidly.

Figure $12.6.2$ shows how the energy of a chemical system changes as it undergoes a reaction converting reactants to products according to the equation

$$\mathrm{A^+B \longrightarrow C^+D}$$

These reaction diagrams are widely used in chemical kinetics to illustrate various properties of the reaction of interest. Viewing the diagram from left to right, the system initially comprises reactants only, $\mathrm{A+B}$. Reactant molecules with sufficient energy can collide to form a high-energy activated complex or transition state. The unstable transition state can then subsequently decay to yield stable products, *C* + *D*. The diagram depicts the reaction's activation energy, E_a, as the energy difference between the reactants and the transition state. Using a specific energy, the *enthalpy* (see chapter on thermochemistry), the enthalpy change of the reaction, ΔH, is estimated as the energy difference between the reactants and products. In this case, the reaction is exothermic ($\Delta H < 0$) since it yields a decrease in system enthalpy.

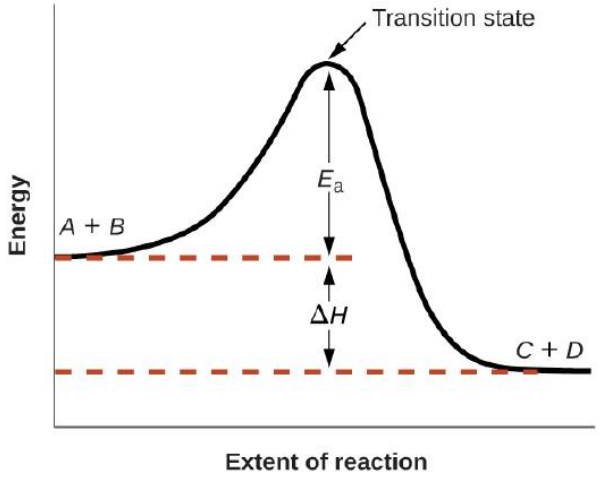


Figure 12.6.2: Reaction diagram for the exothermic reaction

The **Arrhenius equation** relates the activation energy and the rate constant, *k*, for many chemical reactions:

$$k = Ae^{-E_a/RT}$$

In this equation, R is the ideal gas constant, which has a value 8.314 J/mol/K, T is temperature on the Kelvin scale, E_a is the activation energy in joules per mole, *e* is the constant 2.7183, and *A* is a constant called the **frequency factor**, which is related to the frequency of collisions and the orientation of the reacting molecules.

Postulates of collision theory are nicely accommodated by the Arrhenius equation. The frequency factor, *A*, reflects how well the reaction conditions favor properly oriented collisions between reactant molecules. An increased probability of effectively oriented collisions results in larger values for *A* and faster reaction rates.

The exponential term, $e^{-Ea/RT}$, describes the effect of activation energy on reaction rate. According to kinetic molecular theory (see chapter on gases), the temperature of matter is a measure of the average kinetic energy of its constituent atoms or molecules. The distribution of energies among the molecules composing a sample of matter at any given temperature is described by the plot shown in Figure $12.6.3a$. Two shaded areas under the curve represent the numbers of molecules possessing adequate energy (*RT*) to overcome the activation barriers (E_a). A lower activation energy results in a greater fraction of adequately energized molecules and a faster reaction.

The exponential term also describes the effect of temperature on reaction rate. A higher temperature represents a correspondingly greater fraction of molecules possessing sufficient energy (*RT*) to overcome the activation barrier (E_a), as shown in Figure $12.6.3b$ This yields a greater value for the rate constant and a correspondingly faster reaction rate.

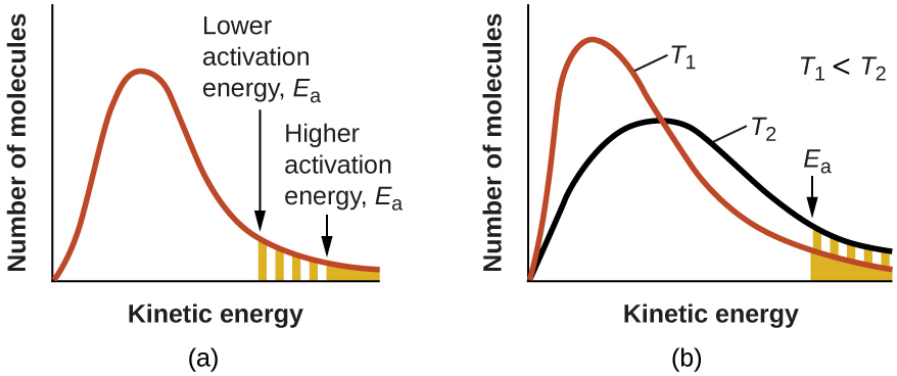


Figure 12.6.3: Molecular energy distributions showing numbers of molecules with energies exceeding (a) two different activation energies at a given temperature, and (b) a given activation energy at two different temperatures.

A convenient approach for determining E_a for a reaction involves the measurement of *k* at two or more different temperatures and using an alternate version of the Arrhenius equation that takes the form of a linear equation

$$\ln k = \left(\frac{-E_a}{R}\right)\left(\frac{1}{T}\right) + \ln A$$

$$y = mx + b$$

A plot of $\ln k$ versus $\frac{1}{T}$ is linear with a slope equal to $\frac{-E_a}{R}$ and a y-intercept equal to $\ln A$.

Example 12.6.1: Determination of E_a

The variation of the rate constant with temperature for the decomposition of $HI(g)$ to $H_2(g)$ and $I_2(g)$ is given here. What is the activation energy for the reaction?

$$2\,HI(g) \longrightarrow H_2(g) + I_2(g)$$

T (K)	k (L/mol/s)
555	3.52×10^{-7}
575	1.22×10^{-6}
645	8.59×10^{-5}
700	1.16×10^{-3}
781	3.95×10^{-2}

Solution

Use the provided data to derive values of $\frac{1}{T}$ and $\ln k$:

$\frac{1}{T}$	$\ln k$
1.80×10^{-3}	−14.860
1.74×10^{-3}	−13.617
1.55×10^{-3}	−9.362
1.43×10^{-3}	−6.759
1.28×10^{-3}	−3.231

Figure 12.6.4is a graph of $\ln k$ versus $\frac{1}{T}$. In practice, the equation of the line (slope and *y*-intercept) that best fits these plotted data points would be derived using a statistical process called regression. This is helpful for most experimental data because a perfect fit of each data point with the line is rarely encountered. For the data here, the fit is nearly perfect and the slope may be estimated using any two of the provided data pairs. Using the first and last data points permits estimation of the slope.

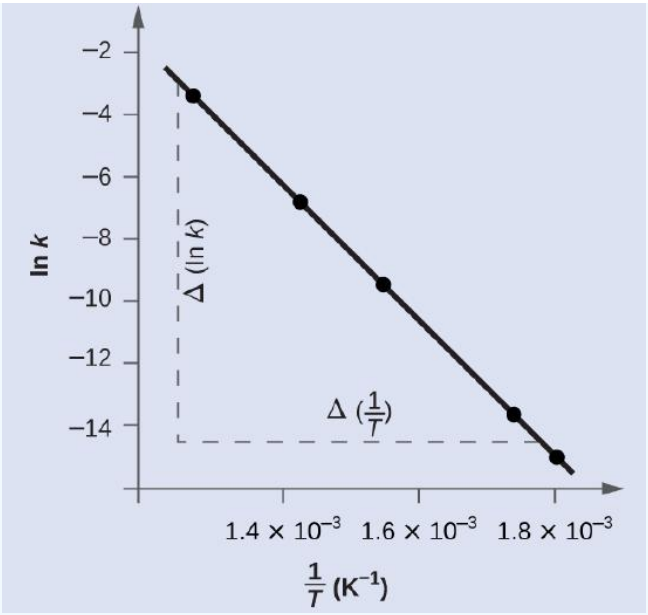


Figure 12.6.4: This graph shows the linear relationship between ln k versus $\frac{1}{T}$ for the reaction $2HI \longrightarrow H_2 + I_2$ according to the Arrhenius equation.

$$\begin{aligned}\text{Slope} &= \frac{\Delta(\ln k)}{\Delta\left(\frac{1}{T}\right)} \\ &= \frac{(-14.860)-(-3.231)}{\left(1.80\times10^{-3}K^{-1}\right)-\left(1.28\times10^{-3}K^{-1}\right)} \\ &= \frac{-11.629}{0.52\times10^{-3}K^{-1}} = -2.2\times10^{4}K \\ &= -\frac{E_a}{R} \\ E_a &= -\text{ slope }\times R = -\left(-2.2\times10^{4}K\times8.314Jmol^{-1}K^{-1}\right) \\ &\quad 1.8\times10^{5}Jmol^{-1} \text{ or } 180kJmol^{-1}\end{aligned}$$

Alternative approach: A more expedient approach involves deriving activation energy from measurements of the rate constant at just two temperatures. In this approach, the Arrhenius equation is rearranged to a convenient two-point form:

$$\ln\frac{k_1}{k_2} = \frac{E_a}{R}\left(\frac{1}{T_2}-\frac{1}{T_1}\right)$$

Rearranging this equation to isolate activation energy yields:

$$E_a = -R\left(\frac{\ln k_2 - \ln k_1}{\left(\frac{1}{T_2}\right)-\left(\frac{1}{T_1}\right)}\right)$$

Any two data pairs may be substituted into this equation—for example, the first and last entries from the above data table:

$$E_a = -8.314Jmol^{-1}K^{-1}\left(\frac{-3.231-(-14.860)}{1.28\times10^{-3}K^{-1}-1.80\times10^{-3}K^{-1}}\right)$$

and the result is E_a = 1.8 10^5 J mol^{-1} or 180 kJ mol^{-1}

This approach yields the same result as the more rigorous graphical approach used above, as expected. In practice, the graphical approach typically provides more reliable results when working with actual experimental data.

Exercise 12.6.1

The rate constant for the rate of decomposition of N_2O_5 to NO and O_2 in the gas phase is 1.66 L/mol/s at 650 K and 7.39 L/mol/s at 700 K:

$$2\,N_2O_5(g) \longrightarrow 4\,NO(g) + 3\,O_2(g)$$

Assuming the kinetics of this reaction are consistent with the Arrhenius equation, calculate the activation energy for this decomposition.

Answer

$1.1\ 10^5$ J mol^{-1} or 110 kJ mol^{-1}

This page titled 12.6: Collision Theory is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **12.5: Collision Theory** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.7: Reaction Mechanisms

Learning Objectives

By the end of this section, you will be able to:

- Distinguish net reactions from elementary reactions (steps)
- Identify the molecularity of elementary reactions
- Write a balanced chemical equation for a process given its reaction mechanism
- Derive the rate law consistent with a given reaction mechanism

Chemical reactions very often occur in a step-wise fashion, involving two or more distinct reactions taking place in sequence. A balanced equation indicates what is reacting and what is produced, but it reveals no details about how the reaction actually takes place. The reaction mechanism (or reaction path) provides details regarding the precise, step-by-step process by which a reaction occurs.

The decomposition of ozone, for example, appears to follow a mechanism with two steps:

$$O_3(g) \longrightarrow O_2(g) + O$$
$$O + O_3(g) \longrightarrow 2O_2(g)$$

Each of the steps in a reaction mechanism is an **elementary reaction**. These elementary reactions occur precisely as represented in the step equations, and they must sum to yield the balanced chemical equation representing the overall reaction:

$$2O_3(g) \longrightarrow 3O_2(g)$$

Notice that the oxygen atom produced in the first step of this mechanism is consumed in the second step and therefore does not appear as a product in the overall reaction. Species that are produced in one step and consumed in a subsequent step are called intermediates.

While the overall reaction equation for the decomposition of ozone indicates that two molecules of ozone react to give three molecules of oxygen, the mechanism of the reaction *does not involve the direct collision and reaction of two ozone molecules*. Instead, one O_3 decomposes to yield O_2 and an oxygen atom, and a second O_3 molecule subsequently reacts with the oxygen atom to yield two additional O_2 molecules.

Unlike balanced equations representing an overall reaction, the equations for elementary reactions are explicit representations of the chemical change taking place. The reactant(s) in an elementary reaction's equation undergo only the bond-breaking and/or making events depicted to yield the product(s). For this reason, *the rate law for an elementary reaction may be derived directly from the balanced chemical equation describing the reaction*. This is not the case for typical chemical reactions, for which rate laws may be reliably determined only via experimentation.

Unimolecular Elementary Reactions

The molecularity of an elementary reaction is the number of reactant species (atoms, molecules, or ions). For example, a unimolecular reaction involves the reaction of a *single* reactant species to produce one or more molecules of product:

$$A \longrightarrow \text{ products}$$

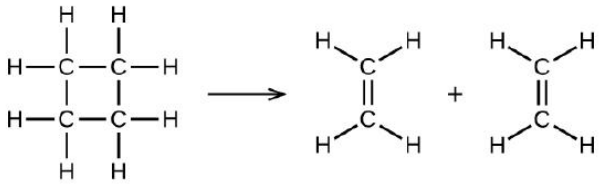
The rate law for a unimolecular reaction is first order:

$$\text{rate } = k[A]$$

A unimolecular reaction may be one of several elementary reactions in a complex mechanism. For example, the reaction:

$$O_3 \longrightarrow O_2 + O$$

illustrates a unimolecular elementary reaction that occurs as one part of a two-step reaction mechanism as described above. However, some unimolecular reactions may be the only step of a single-step reaction mechanism. (In other words, an "overall" reaction may also be an elementary reaction in some cases.) For example, the gas-phase decomposition of cyclobutane, C_4H_8, to ethylene, C_2H_4, is represented by the following chemical equation:



This equation represents the overall reaction observed, and it might also represent a legitimate unimolecular elementary reaction. The rate law predicted from this equation, assuming it is an elementary reaction, turns out to be the same as the rate law derived experimentally for the overall reaction, namely, one showing first-order behavior:

$$\text{rate } = -\frac{\Delta\left[C_4H_8\right]}{\Delta t} = k\left[C_4H_8\right]$$

This agreement between observed and predicted rate laws is interpreted to mean that the proposed unimolecular, single-step process is a reasonable mechanism for the butadiene reaction.

Bimolecular Elementary Reactions

A bimolecular reaction involves two reactant species, for example:

$$A + B \longrightarrow \text{ products}$$

and

$$2A \longrightarrow \text{ products}$$

For the first type, in which the two reactant molecules are different, the rate law is first-order in *A* and first order in *B* (second-order overall):

$$\text{rate } = k[A][B]$$

For the second type, in which two identical molecules collide and react, the rate law is second order in *A*:

$$\text{rate } = k[A][A] = k[A]^2$$

Some chemical reactions occur by mechanisms that consist of a single bimolecular elementary reaction. One example is the reaction of nitrogen dioxide with carbon monoxide:

$$NO_2(g) + CO(g) \longrightarrow NO(g) + CO_2(g)$$

(see Figure 12.7.1)



Figure 12.7.1: The probable mechanism for the reaction between NO_2 and CO to yield NO and CO_2.

Bimolecular elementary reactions may also be involved as steps in a multistep reaction mechanism. The reaction of atomic oxygen with ozone is the second step of the two-step ozone decomposition mechanism discussed earlier in this section:

$$O(g) + O_3(g) \longrightarrow 2O_2(g)$$

Termolecular Elementary Reactions

An elementary termolecular reaction involves the simultaneous collision of three atoms, molecules, or ions. Termolecular elementary reactions are uncommon because the probability of three particles colliding simultaneously is less than one one-thousandth of the probability of two particles colliding. There are, however, a few established termolecular elementary reactions. The reaction of nitric oxide with oxygen appears to involve termolecular steps:

$$2NO + O_2 \longrightarrow 2NO_2$$

$$\text{rate } = k[NO]^2\,[O_2]$$

Likewise, the reaction of nitric oxide with chlorine appears to involve termolecular steps:

$$2NO + Cl_2 \longrightarrow 2NOCl$$

$$\text{rate } = k[NO]^2\,[Cl_2]$$

Relating Reaction Mechanisms to Rate Laws

It's often the case that one step in a multistep reaction mechanism is significantly slower than the others. Because a reaction cannot proceed faster than its slowest step, this step will limit the rate at which the overall reaction occurs. The slowest step is therefore called the rate-limiting step (or rate-determining step) of the reaction Figure 12.7.2



Figure 12.7.2: A cattle chute is a nonchemical example of a rate-determining step. Cattle can only be moved from one holding pen to another as quickly as one animal can make its way through the chute. (credit: Loren Kerns)

As described earlier, rate laws may be derived directly from the chemical equations for elementary reactions. This is not the case, however, for ordinary chemical reactions. The balanced equations most often encountered represent the overall change for some chemical system, and very often this is the result of some multistep reaction mechanisms. In every case, the rate law must be determined from experimental data and the reaction mechanism subsequently deduced from the rate law (and sometimes from other data). The reaction of NO_2 and CO provides an illustrative example:

$$NO_2(g) + CO(g) \longrightarrow CO_2(g) + NO(g)$$

For temperatures above 225 °C, the rate law has been found to be:

$$\text{rate } = k\,[NO_2]\,[CO]$$

The reaction is first order with respect to NO_2 and first-order with respect to CO. This is consistent with a single-step bimolecular mechanism and it is *possible* that this is the mechanism for this reaction at high temperatures.

At temperatures below 225 °C, the reaction is described by a rate law that is second order with respect to NO_2:

$$\text{rate } = k[NO_2]^2$$

This rate law is not consistent with the single-step mechanism, but is consistent with the following two-step mechanism:

$$NO_2(g) + NO_2(g) \longrightarrow NO_3(g) + NO(g)\ (\text{slow})$$

$$NO_3(g) + CO(g) \longrightarrow NO_2(g) + CO_2(g)\ (\text{fast})$$

The rate-determining (slower) step gives a rate law showing second-order dependence on the NO_2 concentration, and the sum of the two equations gives the net overall reaction.

In general, when the rate-determining (slower) step is the first step in a mechanism, the rate law for the overall reaction is the same as the rate law for this step. However, when the rate-determining step is preceded by a step involving a rapidly reversible reaction the rate law for the overall reaction may be more difficult to derive.

As discussed in several chapters of this text, a reversible reaction is at *equilibrium* when the rates of the forward and reverse processes are equal. Consider the reversible elementary reaction in which NO dimerizes to yield an intermediate species N_2O_2.

$$NO + NO \rightleftharpoons N_2O_2$$

When this reaction is at equilibrium:

$$\text{rate}_{\text{forward}} = \text{rate}_{\text{reverse}}$$
$$k_1[NO]^2 = k_{-1}\,[N_2O_2]$$

This expression may be rearranged to express the concentration of the intermediate in terms of the reactant NO:

$$\left(\frac{k_1[NO]^2}{k_{-1}}\right) = [N_2O_2]$$

Since intermediate species concentrations are not used in formulating rate laws for overall reactions, this approach is sometimes necessary, as illustrated in the following example exercise.

Example 12.7.1: Deriving a Rate Law from a Reaction Mechanism

The two-step mechanism below has been proposed for a reaction between nitrogen monoxide and molecular chlorine:

$$\text{Step 1:} \quad \mathrm{NO(g) + Cl_2(g) \rightleftharpoons NOCl_2(g)} \qquad \text{(fast)}$$
$$\text{Step 2:} \quad \mathrm{NOCl_2(g) + NO(g) \rightleftharpoons 2\,NOCl(g)} \qquad \text{(slow)}$$

Use this mechanism to derive the equation and predicted rate law for the overall reaction.

Solution

The equation for the overall reaction is obtained by adding the two elementary reactions:

$$2NO(g) + Cl_2(g) \longrightarrow 2NOCl(g)$$

To derive a rate law from this mechanism, first write rates laws for each of the two steps.

$$\text{rate}_{\,1} = k_1[NO]\,[Cl_2] \text{ for the forward reaction of step 1}$$
$$\text{rate}_{\,-1} = k_{-1}\,[NOCl_2] \text{ for the reverse reaction of step 1}$$
$$\text{rate}_{\,2} = k_2\,[NOCl_2]\,[NO] \text{ for step 2}$$

Step 2 is the rate-determining step, and so the rate law for the overall reaction should be the same as for this step. However, the step 2 rate law, as written, contains an intermediate species concentration, [$NOCl_2$]. To remedy this, use the first step's rate laws to derive an expression for the intermediate concentration in terms of the reactant concentrations.

Assuming step 1 is at equilibrium:

$$\text{rate}_{\,1} = \text{rate}_{\,-1}$$
$$k_1[NO]\,[Cl_2] = k_{-1}\,[NOCl_2]$$
$$[NOCl_2] = \left(\frac{k_1}{k_{-1}}\right)[NO]\,[Cl_2]$$

Substituting this expression into the rate law for step 2 yields:

$$\text{rate}_{\,2} = \text{rate}_{\,\text{overall}} = \left(\frac{k_2k_1}{k_{-1}}\right)[NO]^2\,[Cl_2]$$

Exercise 12.7.1

The first step of a proposed multistep mechanism is:

$$\boxed{\ce{F2(g) <=> 2 F(g) \tag{fast}}}$$

Derive the equation relating atomic fluorine concentration to molecular fluorine concentration.

Answer

$$[\mathrm{F}] = \left(\frac{k_1\,[F_2]}{k_{-1}}\right)^{1/2}$$

This page titled 12.7: Reaction Mechanisms is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **12.6: Reaction Mechanisms** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

12.8: Catalysis

Learning Objectives

By the end of this section, you will be able to:

- Explain the function of a catalyst in terms of reaction mechanisms and potential energy diagrams
- List examples of catalysis in natural and industrial processes

Catalysts Do Not Affect Equilibrium

A catalyst can speed up the rate of a reaction. Though this increase in reaction rate may cause a system to reach equilibrium more quickly (by speeding up the forward and reverse reactions), a catalyst has no effect on the value of an equilibrium constant nor on equilibrium concentrations.

The interplay of changes in concentration or pressure, temperature, and the lack of an influence of a catalyst on a chemical equilibrium is illustrated in the industrial synthesis of ammonia from nitrogen and hydrogen according to the equation

$$\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g)$$

A large quantity of ammonia is manufactured by this reaction. Each year, ammonia is among the top 10 chemicals, by mass, manufactured in the world. About 2 billion pounds are manufactured in the United States each year.

Ammonia plays a vital role in our global economy. It is used in the production of fertilizers and is, itself, an important fertilizer for the growth of corn, cotton, and other crops. Large quantities of ammonia are converted to nitric acid, which plays an important role in the production of fertilizers, explosives, plastics, dyes, and fibers, and is also used in the steel industry.

Portrait of a Chemist

Fritz Haber

In the early 20th century, German chemist Fritz Haber (Figure 17.19) developed a practical process for converting diatomic nitrogen, which cannot be used by plants as a nutrient, to ammonia, a form of nitrogen that is easiest for plants to absorb.

$$\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g)$$

The availability of nitrogen is a strong limiting factor to the growth of plants. Despite accounting for 78% of air, diatomic nitrogen ($\mathrm{N_2}$) is nutritionally unavailable due the tremendous stability of the nitrogen-nitrogen triple bond. For plants to use atmospheric nitrogen, the nitrogen must be converted to a more bioavailable form (this conversion is called nitrogen fixation).

Haber was born in Breslau, Prussia (presently Wroclaw, Poland) in December 1868. He went on to study chemistry and, while at the University of Karlsruhe, he developed what would later be known as the Haber process: the catalytic formation of ammonia from hydrogen and atmospheric nitrogen under high temperatures and pressures. For this work, Haber was awarded the 1918 Nobel Prize in Chemistry for synthesis of ammonia from its elements. The Haber process was a boon to agriculture, as it allowed the production of fertilizers to no longer be dependent on mined feed stocks such as sodium nitrate. Currently, the annual production of synthetic nitrogen fertilizers exceeds 100 million tons and synthetic fertilizer production has increased the number of humans that arable land can support from 1.9 persons per hectare in 1908 to 4.3 in 2008.



Figure 17.19 The work of Nobel Prize recipient Fritz Haber revolutionized agricultural practices in the early 20th century. His work also affected wartime strategies, adding chemical weapons to the artillery.

In addition to his work in ammonia production, Haber is also remembered by history as one of the fathers of chemical warfare. During World War I, he played a major role in the development of poisonous gases used for trench warfare. Regarding his role in these developments, Haber said, “During peace time a scientist belongs to the World, but during war time he belongs to his country.”[1] Haber defended the use of gas warfare against accusations that it was inhumane, saying that death was death, by whatever means it was inflicted. He stands as an example of the ethical dilemmas that face scientists in times of war and the double-edged nature of the sword of science.

Like Haber, the products made from ammonia can be multifaceted. In addition to their value for agriculture, nitrogen compounds can also be used to achieve destructive ends. Ammonium nitrate has also been used in explosives, including improvised explosive devices. Ammonium nitrate was one of the components of the bomb used in the attack on the Alfred P. Murrah Federal Building in downtown Oklahoma City on April 19, 1995.

It has long been known that nitrogen and hydrogen react to form ammonia. However, it became possible to manufacture ammonia in useful quantities by the reaction of nitrogen and hydrogen only in the early 20th century after the factors that influence its equilibrium were understood.

To be practical, an industrial process must give a large yield of product relatively quickly. One way to increase the yield of ammonia is to increase the pressure on the system in which N_2, H_2, and NH_3 are at equilibrium or are coming to equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The formation of additional amounts of ammonia reduces the total pressure exerted by the system and somewhat reduces the stress of the increased pressure.

Although increasing the pressure of a mixture of N_2, H_2, and NH_3 will increase the yield of ammonia, at low temperatures, the rate of formation of ammonia is slow. At room temperature, for example, the reaction is so slow that if we prepared a mixture of N_2 and H_2, no detectable amount of ammonia would form during our lifetime. The formation of ammonia from hydrogen and nitrogen is an exothermic process:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) \qquad \Delta H = -92.2\ \text{kJ}$$

Thus, increasing the temperature to increase the rate lowers the yield. If we lower the temperature to shift the equilibrium to favor the formation of more ammonia, equilibrium is reached more slowly because of the large decrease of reaction rate with decreasing temperature.

Part of the rate of formation lost by operating at lower temperatures can be recovered by using a catalyst. The net effect of the catalyst on the reaction is to cause equilibrium to be reached more rapidly.

In the commercial production of ammonia, conditions of about 500 °C, 150–900 atm, and the presence of a catalyst are used to give the best compromise among rate, yield, and the cost of the equipment necessary to produce and contain high-pressure gases at high temperatures (Figure 17.20).

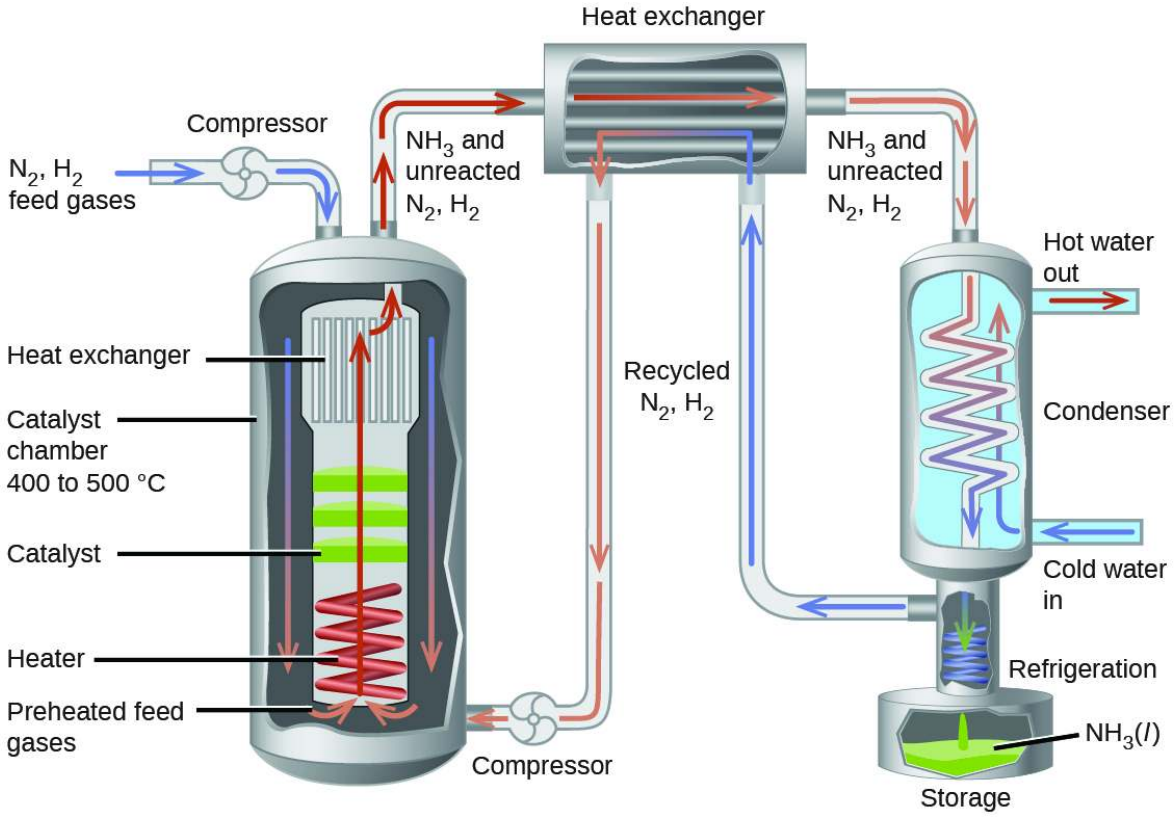


Figure 17.20 Commercial production of ammonia requires heavy equipment to handle the high temperatures and pressures required. This schematic outlines the design of an ammonia plant.

Among the factors affecting chemical reaction rates discussed earlier in this chapter was the presence of a *catalyst*, a substance that can increase the reaction rate without being consumed in the reaction. The concepts introduced in the previous section on reaction mechanisms provide the basis for understanding how catalysts are able to accomplish this very important function.

Figure 17.21 shows reaction diagrams for a chemical process in the absence and presence of a catalyst. Inspection of the diagrams reveals several traits of these reactions. Consistent with the fact that the two diagrams represent the same overall reaction, both curves begin and end at the same energies (in this case, because products are more energetic than reactants, the reaction is

endothermic). The reaction mechanisms, however, are clearly different. The uncatalyzed reaction proceeds via a one-step mechanism (one transition state observed), whereas the catalyzed reaction follows a two-step mechanism (two transition states observed) with *a notably lesser activation energy*. This difference illustrates the means by which a catalyst functions to accelerate reactions, namely, by providing an alternative reaction mechanism with a lower activation energy. Although the catalyzed reaction mechanism for a reaction needn't necessarily involve a different number of steps than the uncatalyzed mechanism, it must provide a reaction path whose rate determining step is faster (lower E_a).

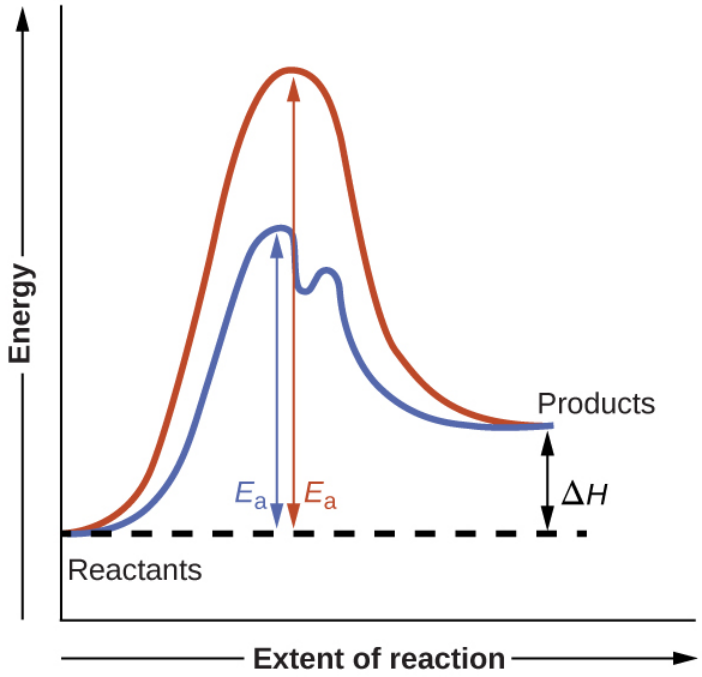
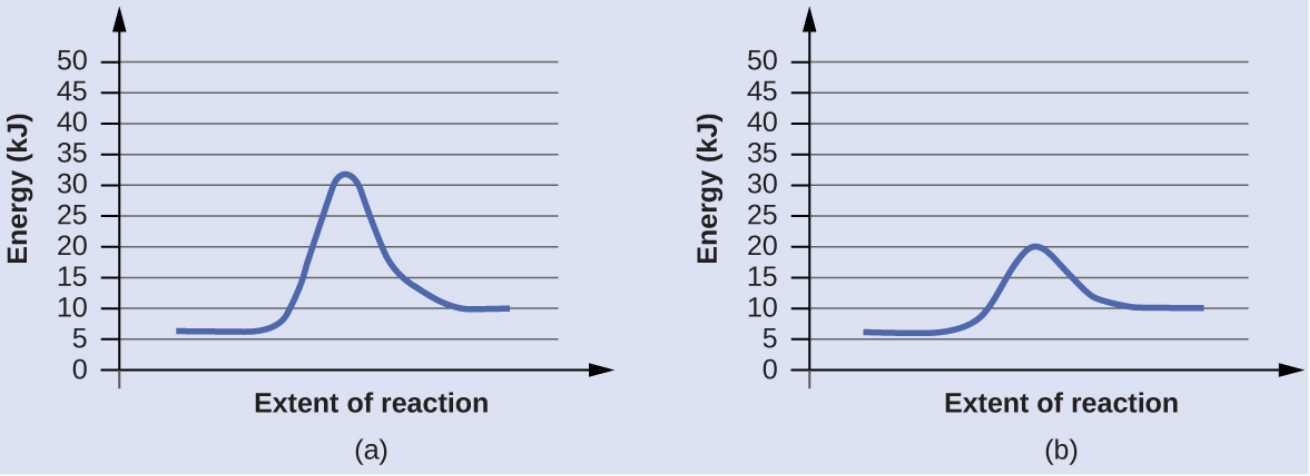


Figure 17.21 Reaction diagrams for an endothermic process in the absence (red curve) and presence (blue curve) of a catalyst. The catalyzed pathway involves a two-step mechanism (note the presence of two transition states) and an intermediate species (represented by the valley between the two transitions states).

Example 17.15

Reaction Diagrams for Catalyzed Reactions

The two reaction diagrams here represent the same reaction: one without a catalyst and one with a catalyst. Estimate the activation energy for each process, and identify which one involves a catalyst.



Solution

Activation energies are calculated by subtracting the reactant energy from the transition state energy.

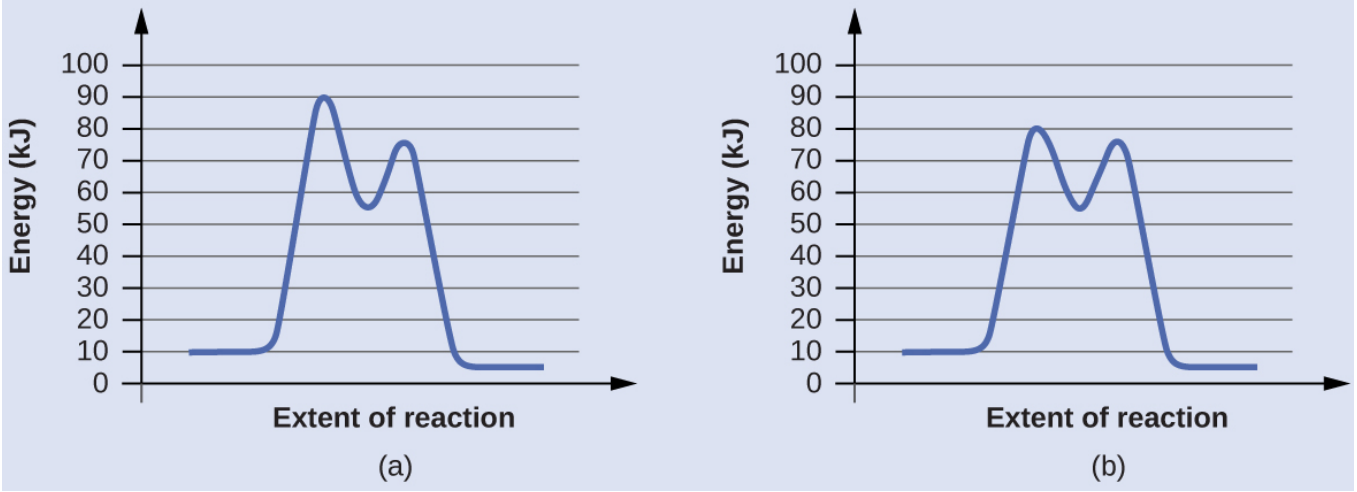
$$\text{diagram (a): } E_a = 32\ \text{kJ} - 6\ \text{kJ} = 26\ \text{kJ}$$

$$\text{diagram (b): } E_a = 20\ \text{kJ} - 6\ \text{kJ} = 14\ \text{kJ}$$

The catalyzed reaction is the one with lesser activation energy, in this case represented by diagram (b).

Check Your Learning

Reaction diagrams for a chemical process with and without a catalyst are shown below. Both reactions involve a two-step mechanism with a rate-determining first step. Compute activation energies for the first step of each mechanism, and identify which corresponds to the catalyzed reaction. How do the second steps of these two mechanisms compare?



Answer:

For the first step, E_a = 80 kJ for (a) and 70 kJ for (b), so diagram (b) depicts the catalyzed reaction. Activation energies for the second steps of both mechanisms are the same, 20 kJ.

A homogeneous catalyst is present in the same phase as the reactants. It interacts with a reactant to form an intermediate substance, which then decomposes or reacts with another reactant in one or more steps to regenerate the original catalyst and form product.

As an important illustration of homogeneous catalysis, consider the earth's ozone layer. Ozone in the upper atmosphere, which protects the earth from ultraviolet radiation, is formed when oxygen molecules absorb ultraviolet light and undergo the reaction:

$$3O_2(g) \xrightarrow{h\nu} 2O_3(g)$$

Ozone is a relatively unstable molecule that decomposes to yield diatomic oxygen by the reverse of this equation. This decomposition reaction is consistent with the following two-step mechanism:

$$\begin{aligned} O_3 &\longrightarrow O_2 + O \\ O + O_3 &\longrightarrow 2O_2 \end{aligned}$$

A number of substances can catalyze the decomposition of ozone. For example, the nitric oxide -catalyzed decomposition of ozone is believed to occur via the following three-step mechanism:

$$\begin{aligned} NO(g) + O_3(g) &\longrightarrow NO_2(g) + O_2(g) \\ O_3(g) &\longrightarrow O_2(g) + O(g) \\ NO_2(g) + O(g) &\longrightarrow NO(g) + O_2(g) \end{aligned}$$

As required, the overall reaction is the same for both the two-step uncatalyzed mechanism and the three-step NO-catalyzed mechanism:

$$2O_3(g) \longrightarrow 3O_2(g)$$

Notice that NO is a reactant in the first step of the mechanism and a product in the last step. This is another characteristic trait of a catalyst: Though it participates in the chemical reaction, it is not consumed by the reaction.

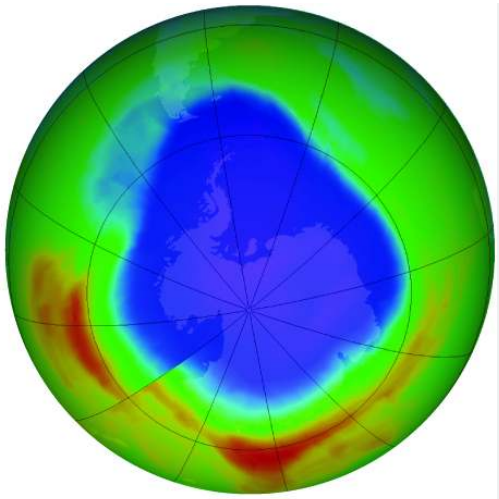
Portrait of a Chemist

Mario J. Molina

The 1995 Nobel Prize in Chemistry was shared by Paul J. Crutzen, Mario J. Molina (Figure 17.22), and F. Sherwood Rowland "for their work in atmospheric chemistry, particularly concerning the formation and decomposition of ozone."[2] Molina, a Mexican citizen, carried out the majority of his work at the Massachusetts Institute of Technology (MIT).



(a)



(b)

Figure 17.22 (a) Mexican chemist Mario Molina (1943 –) shared the Nobel Prize in Chemistry in 1995 for his research on (b) the Antarctic ozone hole. (credit a: courtesy of Mario Molina; credit b: modification of work by NASA)

In 1974, Molina and Rowland published a paper in the journal *Nature* detailing the threat of chlorofluorocarbon gases to the stability of the ozone layer in earth's upper atmosphere. The ozone layer protects earth from solar radiation by absorbing ultraviolet light. As chemical reactions deplete the amount of ozone in the upper atmosphere, a measurable "hole" forms above Antarctica, and an increase in the amount of solar ultraviolet radiation— strongly linked to the prevalence of skin cancers—reaches earth's surface. The work of Molina and Rowland was instrumental in the adoption of the Montreal Protocol, an international treaty signed in 1987 that successfully began phasing out production of chemicals linked to ozone destruction.

Molina and Rowland demonstrated that chlorine atoms from human-made chemicals can catalyze ozone destruction in a process similar to that by which NO accelerates the depletion of ozone. Chlorine atoms are generated when chlorocarbons or chlorofluorocarbons—once widely used as refrigerants and propellants—are photochemically decomposed by ultraviolet light or react with hydroxyl radicals. A sample mechanism is shown here using methyl chloride:

$$\mathrm{CH_3Cl + OH \longrightarrow Cl + \text{other products}}$$

Chlorine radicals break down ozone and are regenerated by the following catalytic cycle:

$$\mathrm{Cl + O_3 \longrightarrow ClO + O_2}$$
$$\mathrm{ClO + O \longrightarrow Cl + O_2}$$
$$\text{overall Reaction: } \mathrm{O_3 + O \longrightarrow 2O_2}$$

A single monatomic chlorine can break down thousands of ozone molecules. Luckily, the majority of atmospheric chlorine exists as the catalytically inactive forms Cl_2 and $ClONO_2$.

Since receiving his portion of the Nobel Prize, Molina has continued his work in atmospheric chemistry at MIT.

How Sciences Interconnect

Glucose-6-Phosphate Dehydrogenase Deficiency

Enzymes in the human body act as catalysts for important chemical reactions in cellular metabolism. As such, a deficiency of a particular enzyme can translate to a life-threatening disease. G6PD (glucose-6-phosphate dehydrogenase) deficiency, a genetic condition that results in a shortage of the enzyme glucose-6-phosphate dehydrogenase, is the most common enzyme deficiency in humans. This enzyme, shown in Figure 17.23, is the rate-limiting enzyme for the metabolic pathway that supplies NADPH to cells (Figure 17.24).

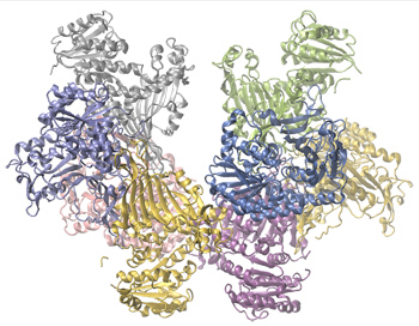


Figure 17.23 Glucose-6-phosphate dehydrogenase is a rate-limiting enzyme for the metabolic pathway that supplies NADPH to cells.

A disruption in this pathway can lead to reduced glutathione in red blood cells; once all glutathione is consumed, enzymes and other proteins such as hemoglobin are susceptible to damage. For example, hemoglobin can be metabolized to bilirubin, which leads to jaundice, a condition that can become severe. People who suffer from G6PD deficiency must avoid certain foods and medicines containing chemicals that can trigger damage their glutathione-deficient red blood cells.

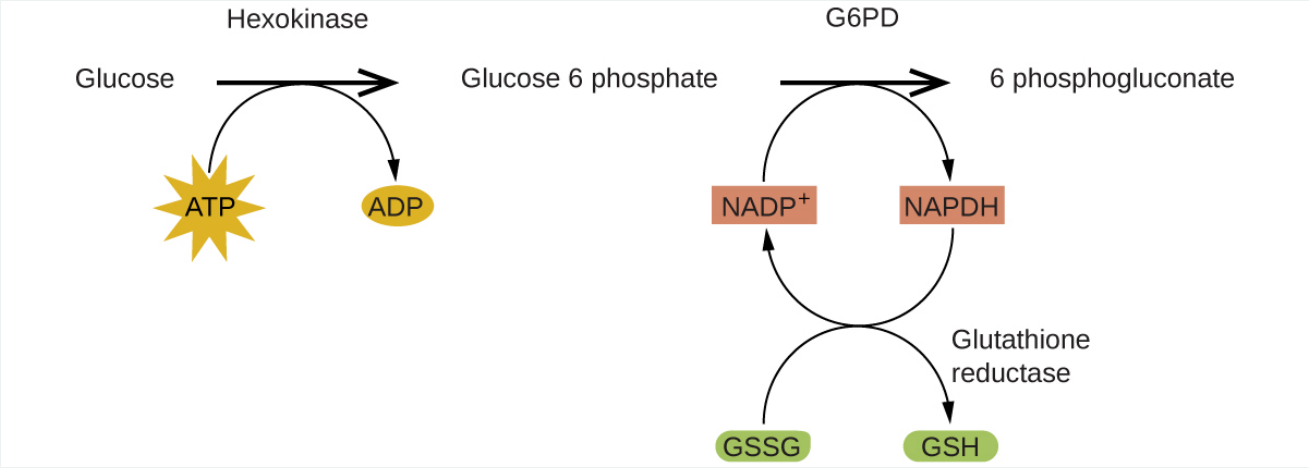


Figure 17.24 In the mechanism for the pentose phosphate pathway, G6PD catalyzes the reaction that regulates NADPH, a co-enzyme that regulates glutathione, an antioxidant that protects red blood cells and other cells from oxidative damage.

A heterogeneous catalyst is a catalyst that is present in a different phase (usually a solid) than the reactants. Such catalysts generally function by furnishing an active surface upon which a reaction can occur. Gas and liquid phase reactions catalyzed by heterogeneous catalysts occur on the surface of the catalyst rather than within the gas or liquid phase.

Heterogeneous catalysis typically involves the following processes:

1. Adsorption of the reactant(s) onto the surface of the catalyst
2. Activation of the adsorbed reactant(s)
3. Reaction of the adsorbed reactant(s)
4. Desorption of product(s) from the surface of the catalyst

Figure 17.25 illustrates the steps of a mechanism for the reaction of compounds containing a carbon–carbon double bond with hydrogen on a nickel catalyst. Nickel is the catalyst used in the hydrogenation of polyunsaturated fats and oils (which contain several carbon–carbon double bonds) to produce saturated fats and oils (which contain only carbon–carbon single bonds).

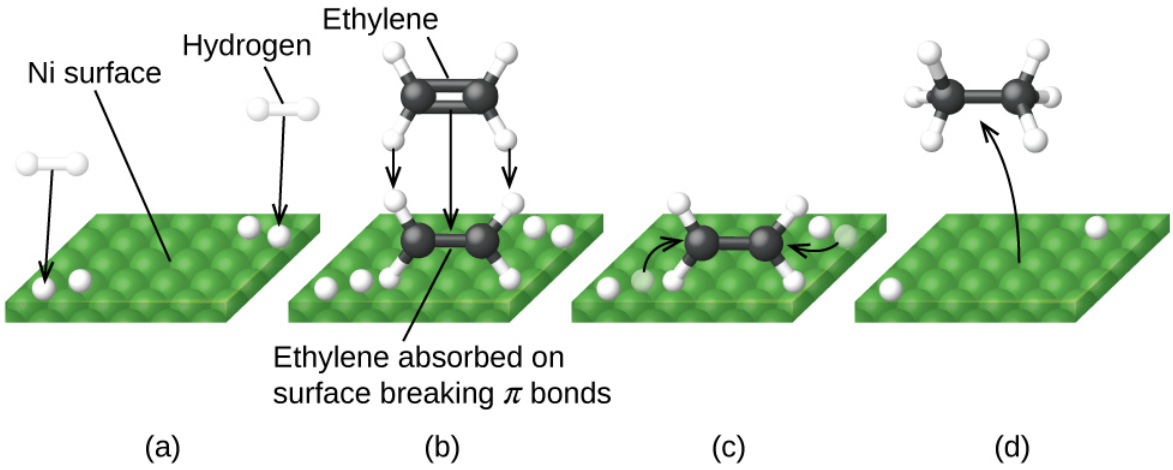


Figure 17.25 Mechanism for the Ni-catalyzed reaction $C_2H_4 + H_2 \longrightarrow C_2H_6$. (a) Hydrogen is adsorbed on the surface, breaking the H–H bonds and forming Ni–H bonds. (b) Ethylene is adsorbed on the surface, breaking the C-C π-bond and forming Ni–C bonds. (c) Atoms diffuse across the surface and form new C–H bonds when they collide. (d) C_2H_6 molecules desorb from the Ni surface.

Many important chemical products are prepared via industrial processes that use heterogeneous catalysts, including ammonia, nitric acid, sulfuric acid, and methanol. Heterogeneous catalysts are also used in the catalytic converters found on most gasoline-powered automobiles (Figure 17.26).

Chemistry in Everyday Life

Automobile Catalytic Converters

Scientists developed catalytic converters to reduce the amount of toxic emissions produced by burning gasoline in internal combustion engines. By utilizing a carefully selected blend of catalytically active metals, it is possible to effect complete combustion of all carbon-containing compounds to carbon dioxide while also reducing the output of nitrogen oxides. This is particularly impressive when we consider that one step involves adding more oxygen to the molecule and the other involves removing the oxygen (Figure 17.26).

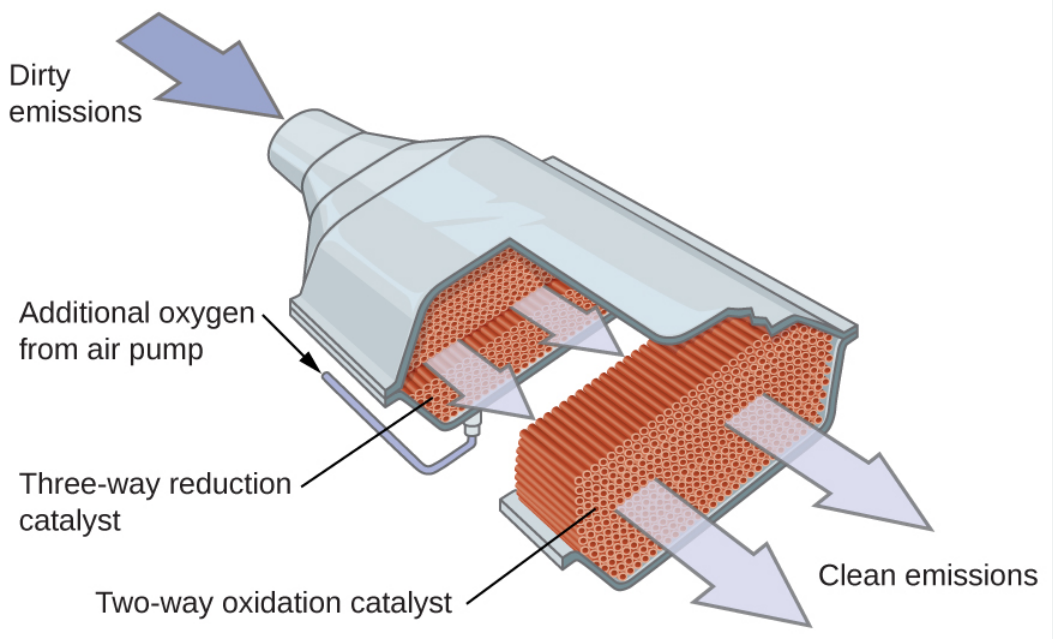


Figure 17.26 A catalytic converter allows for the combustion of all carbon-containing compounds to carbon dioxide, while at the same time reducing the output of nitrogen oxide and other pollutants in emissions from gasoline-burning engines.

Most modern, three-way catalytic converters possess a surface impregnated with a platinum-rhodium catalyst, which catalyzes the conversion of nitric oxide into dinitrogen and oxygen as well as the conversion of carbon monoxide and hydrocarbons such as octane into carbon dioxide and water vapor:

$$2NO_2(g) \longrightarrow N_2(g) + 2O_2(g)$$
$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$
$$2C_8H_{18}(g) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

In order to be as efficient as possible, most catalytic converters are preheated by an electric heater. This ensures that the metals in the catalyst are fully active even before the automobile exhaust is hot enough to maintain appropriate reaction temperatures.

Link to Learning

The University of California at Davis' "ChemWiki" provides a thorough explanation of how catalytic converters work.

Enzyme Structure and Function

The study of enzymes is an important interconnection between biology and chemistry. Enzymes are usually proteins (polypeptides) that help to control the rate of chemical reactions between biologically important compounds, particularly those that are involved in cellular metabolism. Different classes of enzymes perform a variety of functions, as shown in Table 17.3.

Classes of Enzymes and Their Functions

Class	Function
oxidoreductases	redox reactions
transferases	transfer of functional groups
hydrolases	hydrolysis reactions
lyases	group elimination to form double bonds
isomerases	isomerization
ligases	bond formation with ATP hydrolysis

Table 17.3

Enzyme molecules possess an active site, a part of the molecule with a shape that allows it to bond to a specific substrate (a reactant molecule), forming an enzyme-substrate complex as a reaction intermediate. There are two models that attempt to explain how this active site works. The most simplistic model is referred to as the lock-and-key hypothesis, which suggests that the molecular shapes of the active site and substrate are complementary, fitting together like a key in a lock. The induced fit hypothesis, on the other hand, suggests that the enzyme molecule is flexible and changes shape to accommodate a bond with the substrate. This is not to suggest that an enzyme's active site is completely malleable, however. Both the lock-and-key model and the induced fit model account for the fact that enzymes can only bind with specific substrates, since in general a particular enzyme only catalyzes a particular reaction (Figure 17.27).

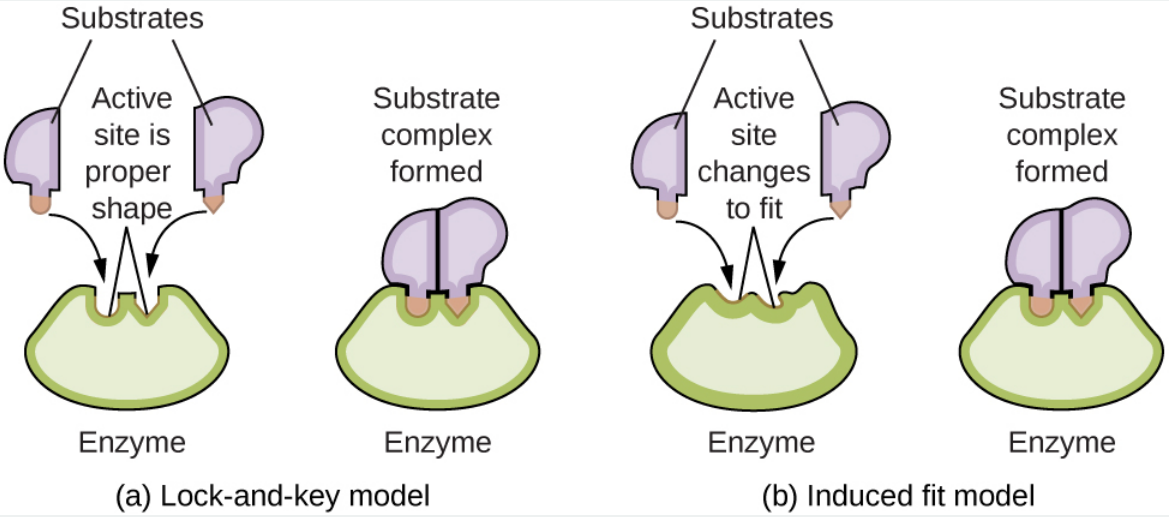


Figure 17.27 (a) According to the lock-and-key model, the shape of an enzyme's active site is a perfect fit for the substrate. (b) According to the induced fit model, the active site is somewhat flexible, and can change shape in order to bond with the substrate.

Link to Learning

The Royal Society of Chemistry provides an excellent introduction to enzymes for students and teachers.

The connection between the rate of a reaction and its equilibrium constant is one we can easily determine with just a bit of algebraic substitution. For a reaction where substance A forms B (and the reverse)

$$A \rightleftharpoons B$$

The rate of the forward reaction is

$$\text{Rate}(f) = k(f)[A]$$

And the rate of the reverse reaction is

$$\text{Rate}(r) = k(r)[B]$$

Once equilibrium is established, the rates of the forward and reverse reactions are equal:

$$\text{Rate}(f) = \text{Rate}(r) = k(f)[A] = k(r)[B]$$

Rearranging a bit, we get

$$\text{Rate}_f = \text{Rate}_r \text{ so } k(f)[A] = k(r)[B]$$

Also recall that the equilibrium constant is simply the ratio of product to reactant concentration at equilibrium:

$$\frac{k(f)}{k(r)} = \frac{[B]}{[A]}$$

$$K = \frac{[B]}{[A]}$$

So the equilibrium constant turns out to be the ratio of the forward to the reverse rate constants. This relationship also helps cement our understanding of the nature of a catalyst. That is, a catalyst does not change the fundamental equilibrium (or the underlying thermodynamics) of a reaction. Rather, what it does is alter the rate constant for the reaction – that is, both rate constants, forward and reverse, equally. In doing so, catalysts usually speed up the rate at which reactions attain equilibrium (though they can be used to slow down the rate of reaction as well!).

Footnotes

- 1Herrlich, P. “The Responsibility of the Scientist: What Can History Teach Us About How Scientists Should Handle Research That Has the Potential to Create Harm?” *EMBO Reports* 14 (2013): 759–764.
- 2“The Nobel Prize in Chemistry 1995,” Nobel Prize.org, accessed February 18, 2015, http://www.nobelprize.org/nobel_priz...aureates/1995/.

This page titled 12.8: Catalysis is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

12.9: Key Terms

activated complex | (also, transition state) unstable combination of reactant species formed during a chemical reaction

activation energy (E_a) | minimum energy necessary in order for a reaction to take place

Arrhenius equation | mathematical relationship between a reaction's rate constant, activation energy, and temperature

average rate | rate of a chemical reaction computed as the ratio of a measured change in amount or concentration of substance to the time interval over which the change occurred

bimolecular reaction | elementary reaction involving two reactant species

catalyst | substance that increases the rate of a reaction without itself being consumed by the reaction

collision theory | model that emphasizes the energy and orientation of molecular collisions to explain and predict reaction kinetics

elementary reaction | reaction that takes place in a single step, precisely as depicted in its chemical equation

frequency factor (A) | proportionality constant in the Arrhenius equation, related to the relative number of collisions having an orientation capable of leading to product formation

half-life of a reaction ($t_{1/2}$) | time required for half of a given amount of reactant to be consumed

heterogeneous catalyst | catalyst present in a different phase from the reactants, furnishing a surface at which a reaction can occur

homogeneous catalyst | catalyst present in the same phase as the reactants

initial rate | instantaneous rate of a chemical reaction at $t = 0$ s (immediately after the reaction has begun)

instantaneous rate | rate of a chemical reaction at any instant in time, determined by the slope of the line tangential to a graph of concentration as a function of time

integrated rate law | equation that relates the concentration of a reactant to elapsed time of reaction

intermediate | species produced in one step of a reaction mechanism and consumed in a subsequent step

method of initial rates | common experimental approach to determining rate laws that involves measuring reaction rates at varying initial reactant concentrations

molecularity | number of reactant species involved in an elementary reaction

overall reaction order | sum of the reaction orders for each substance represented in the rate law

rate constant (k) | proportionality constant in a rate law

rate expression | mathematical representation defining reaction rate as change in amount, concentration, or pressure of reactant or product species per unit time

rate law | (also, rate equation) (also, differential rate laws) mathematical equation showing the dependence of reaction rate on the rate constant and the concentration of one or more reactants

rate of reaction | measure of the speed at which a chemical reaction takes place

rate-determining step | (also, rate-limiting step) slowest elementary reaction in a reaction mechanism; determines the rate of the overall reaction

reaction diagram | used in chemical kinetics to illustrate various properties of a reaction

reaction mechanism | stepwise sequence of elementary reactions by which a chemical change takes place

reaction order | value of an exponent in a rate law (for example, zero order for 0, first order for 1, second order for 2, and so on)

termolecular reaction | elementary reaction involving three reactant species

unimolecular reaction | elementary reaction involving a single reactant species

---

This page titled 12.9: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **12.8: Key Terms** is licensed CC BY 4.0.

12.10: Key Equations

integrated rate law for zero-order reactions:
half-life for a zero-order reaction
integrated rate law for first-order reactions:
half-life for a first-order reaction
integrated rate law for second-order reactions:
half-life for a second-order reaction

This page titled 12.10: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **12.9: Key Equations** is licensed CC BY 4.0.

12.11: Summary

The rate of a reaction can be expressed either in terms of the decrease in the amount of a reactant or the increase in the amount of a product per unit time. Relations between different rate expressions for a given reaction are derived directly from the stoichiometric coefficients of the equation representing the reaction.

The rate of a chemical reaction is affected by several parameters. Reactions involving two phases proceed more rapidly when there is greater surface area contact. If temperature or reactant concentration is increased, the rate of a given reaction generally increases as well. A catalyst can increase the rate of a reaction by providing an alternative pathway with a lower activation energy.

Rate laws (*differential rate laws*) provide a mathematical description of how changes in the concentration of a substance affect the rate of a chemical reaction. Rate laws are determined experimentally and cannot be predicted by reaction stoichiometry. The order of reaction describes how much a change in the concentration of each substance affects the overall rate, and the overall order of a reaction is the sum of the orders for each substance present in the reaction. Reaction orders are typically first order, second order, or zero order, but fractional and even negative orders are possible.

Integrated rate laws are mathematically derived from differential rate laws, and they describe the time dependence of reactant and product concentrations.

The half-life of a reaction is the time required to decrease the amount of a given reactant by one-half. A reaction's half-life varies with rate constant and, for some reaction orders, reactant concentration. The half-life of a zero-order reaction decreases as the initial concentration of the reactant in the reaction decreases. The half-life of a first-order reaction is independent of concentration, and the half-life of a second-order reaction decreases as the concentration increases.

Chemical reactions typically require collisions between reactant species. These reactant collisions must be of proper orientation and sufficient energy in order to result in product formation. Collision theory provides a simple but effective explanation for the effect of many experimental parameters on reaction rates. The Arrhenius equation describes the relation between a reaction's rate constant, activation energy, temperature, and dependence on collision orientation.

The sequence of individual steps, or elementary reactions, by which reactants are converted into products during the course of a reaction is called the reaction mechanism. The molecularity of an elementary reaction is the number of reactant species involved, typically one (unimolecular), two (bimolecular), or, less commonly, three (termolecular). The overall rate of a reaction is determined by the rate of the slowest in its mechanism, called the rate-determining step. Unimolecular elementary reactions have first-order rate laws, while bimolecular elementary reactions have second-order rate laws. By comparing the rate laws derived from a reaction mechanism to that determined experimentally, the mechanism may be deemed either incorrect or plausible.

Catalysts affect the rate of a chemical reaction by altering its mechanism to provide a lower activation energy. Catalysts can be homogenous (in the same phase as the reactants) or heterogeneous (a different phase than the reactants).

This page titled 12.11: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **12.10: Summary** is licensed CC BY 4.0.

12.12: Exercises

12.1: Chemical Reaction Rates

Q12.1.1

What is the difference between average rate, initial rate, and instantaneous rate?

Solution

First, a general reaction rate must be defined to know what any variation of a rate is. The reaction rate is defined as the measure of the change in concentration of the reactants or products per unit time. The rate of a chemical reaction is not a constant and rather changes continuously, and can be influenced by temperature. Rate of a reaction can be defined as the disappearance of any reactant or appearance of any product. Thus, an average rate is the average reaction rate over a given period of time in the reaction, the instantaneous rate is the reaction rate at a specific given moment during the reaction, and the initial rate is the instantaneous rate at the very start of the reaction (when the product begins to form).

The instantaneous rate of a reaction can be denoted as

$$\lim_{\Delta t \to 0} \frac{\Delta [concentration]}{\Delta t}$$

Q12.1.2

Ozone decomposes to oxygen according to the equation $2\,O_3(g) \longrightarrow 3\,O_2(g)$. Write the equation that relates the rate expressions for this reaction in terms of the disappearance of O_3 and the formation of oxygen.

Solution

For the general reaction, aA ---> bB, the rate of the reaction can be expressed in terms of the disappearance of A or the appearance of B over a certain time period as follows.

$$-\frac{1}{a}\frac{\Delta [A]}{\Delta t} = -\frac{1}{b}\frac{\Delta [B]}{\Delta t} = \frac{1}{c}\frac{\Delta [C]}{\Delta t} = \frac{1}{d}\frac{\Delta [D]}{\Delta t} \tag{12.12.1}$$

We want the rate of a reaction to be positive, but the change in the concentration of a reactant, A, will be negative because it is being used up to be transformed into product, B. Therefore, when expressing the rate of the reaction in terms of the change in the concentration of A, it is important to add a negative sign in front to ensure the overall rate positive.

Lastly, the rate must be normalized according to the stoichiometry of the reaction. In the decomposition of ozone to oxygen, two moles of ozone form three moles of oxygen gas. This means that the increase in oxygen gas will be 1.5 times as great as the decrease in ozone. Because the rate of the reaction should be able to describe both species, we divide the change in concentration by its stoichiometric coefficient in the balanced reaction equation to deal with this issue.

Therefore, the rate of the reaction of the decomposition of ozone into oxygen gas can be described as follows:

$$Rate = -\frac{\Delta [O3]}{2\Delta T} = \frac{\Delta [O2]}{3\Delta T} \tag{12.12.2}$$

Answer

$$Rate=-\frac{\Δ[O3]}{2\ΔT}=\frac{\Δ[O2]}{3\ΔT}\]

Q12.1.3

In the nuclear industry, chlorine trifluoride is used to prepare uranium hexafluoride, a volatile compound of uranium used in the separation of uranium isotopes. Chlorine trifluoride is prepared by the reaction $Cl_2(g) + 3\,F_2(g) \longrightarrow 2\,ClF_3(g)$. Write the equation that relates the rate expressions for this reaction in terms of the disappearance of Cl_2 and F_2 and the formation of ClF_3.

Solution

In this problem we are asked to write the equation that relates rate expressions in terms of disappearance of the reactants of the equation and in terms of the formation of the product. A reaction rate gives insight to how rate is affected as a function of

concentration of the substances in the equation. Rates can often be expressed on graphs of concentration vs time expressed in change (Δ) of concentration and time and in a short enough time interval, the instantaneous rate can be approximated. If we were to analyze the reaction given, the graph would demonstrate that Cl_2 decreases, that F_2 decreases 3 times as quickly, and then ClF_3 increases at a rate doubles. The reactants are being used and converted to product so they decrease while products increase.

For this problem, we can apply the general formula of a rate to the specific aspects of a problem where the general form follows:

$$aA + bB \longrightarrow cC + dD$$

.

And the rate can then be written as $rate = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$. Here the negative signs are used to keep the convention of expressing rates as positive numbers.

In this specific case we use the stoichiometry to get the specific rates of disappearance and formation (back to what was said in the first paragraph). So, the problem just involves referring the to the equation and its balanced coefficients. Based upon the equation we see that Cl_2 is a reactant and has no coefficient, F_2 has a coefficient of 3 and is also used up, and then ClF_3 is a product that increases two-fold with a coefficient of 2. So, the rate here can be written as:

$$rate = -\frac{\Delta[Cl_2]}{\Delta t} = -\frac{1}{3}\frac{\Delta[F_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[ClF_3]}{\Delta t}$$

Answer

$$\text{rate} = +\frac{1}{2}\frac{\Delta[\text{ClF}_3]}{\Delta t} = -\frac{\Delta[\text{Cl}_2]}{\Delta t} = -\frac{1}{3}\frac{\Delta[\text{F}_2]}{\Delta t}$$

Q12.1.4

A study of the rate of dimerization of C_4H_6 gave the data shown in the table:

$$2\,\text{C}_4\text{H}_6 \longrightarrow \text{C}_8\text{H}_{12}$$

Time (s)	0	1600	3200	4800	6200
[C_4H_6] (*M*)	1.00×10^{-2}	5.04×10^{-3}	3.37×10^{-3}	2.53×10^{-3}	2.08×10^{-3}

a. Determine the average rate of dimerization between 0 s and 1600 s, and between 1600 s and 3200 s.
b. Estimate the instantaneous rate of dimerization at 3200 s from a graph of time versus [C_4H_6]. What are the units of this rate?
c. Determine the average rate of formation of C_8H_{12} at 1600 s and the instantaneous rate of formation at 3200 s from the rates found in parts (a) and (b).

Solution

1.) The average rate of dimerization is the change in concentration of a reactant per unit time. In this case it would be:

$$rate\ of\ dimerization = -\frac{\Delta[C_4H_6]}{\Delta t}$$

Rate of dimerization between 0 s and 1600 s:

$$rate\ of\ dimerization = -\frac{5.04\times10^{-3}M - 1.00\times10^{-2}M}{1600s - 0s}$$

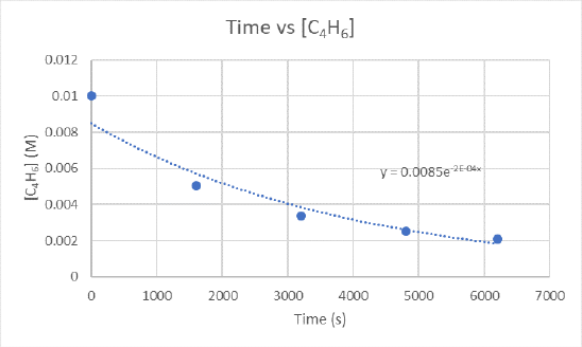
$$rate\ of\ dimerization = 3.10 \times 10^{-6}\,\frac{M}{s}$$

Rate of dimerization between 1600 s and 3200 s:

$$rate\ of\ dimerization = -\frac{3.37\times10^{-3}M - 5.04\times10^{-3}M}{3200s - 1600s}$$

$$rate\ of\ dimerization = 1.04 \times 10^{-6}\,\frac{M}{s}$$

2.) The instantaneous rate of dimerization at 3200 s can be found by graphing time versus [C_4H_6].



Because you want to find the rate of dimerization at 3200 s, you need to find the slope between 1600 s and 3200 s and also 3200 s and 4800 s.

For the slope between 1600 s and 3200 s use the points (1600 s, 5.04 x 10^{-3} M) and (3200 s, 3.37 x 10^{-3} M)

$$\frac{3.37\times10^{-3}M-5.04\times10^{-3}M}{3200s-1600s}$$

$$\frac{-0.00167M}{1600s}$$

$$-1.04\times10^{-6}\,\frac{M}{s}$$

For the slope between 3200 s and 4800 s use the points (3200s, 3.37 x 10^{-3} M) and (4800s, 2.53 x 10^{-3} M)

$$\frac{2.53\times10^{-3}M-3.37\times10^{-3}M}{4800s-3200s}$$

$$\frac{-8.4\times10^{-4}M}{1600s}$$

$$-5.25\times10^{-7}\,\frac{M}{s}$$

Take the two slopes you just found and find the average of them to get the instantaneous rate of dimerization.

$$\frac{-1.04\times10^{-6}\frac{M}{s}+-5.25\times x10^{-7}\frac{M}{s}}{2}$$

$$\frac{-1.565\times10^{-6}\frac{M}{s}}{2}$$

$$-7.83\times10^{-}7\frac{M}{s}$$

The instantaneous rate of dimerization is $-7.83\times10^{-}7\frac{M}{s}$ and the units of this rate is $\frac{M}{s}$.

3.) The average rate of formation of C_8H_{12} at 1600 s and the instantaneous rate of formation at 3200 s can be found by using our answers from part a and b. If you look back up at the original equation, you could see that C_4H_6 and C_8H_{12} are related in a two to one ratio. For every two moles of C_4H_6 used, there is one mole of C_8H_{12} produced.

For this reaction, the average rate of dimerization and the average rate of formation can be linked through this equation:

$$\frac{-1}{2}\frac{\Delta[C_4H_6]}{\Delta t}=\frac{\Delta[C_8H_{12}]}{\Delta t}$$

Notice that reactant side is negative because the reactants are being used up in the reaction.

So, for the average rate of formation of C_8H_{12} at 1600 s, use the rate of dimerization between 0 s and 1600 s we found earlier and plug into the equation:

$$\frac{-1}{2}\times3.10\times10^{-6}\,\frac{M}{s}=\frac{\Delta[C_8H_{12}]}{\Delta t}$$

$$\frac{\Delta[C_8H_{12}]}{\Delta t}=1.55\times10^{-6}\,\frac{M}{s}$$

The average rate of formation for C_8H_{12} at 1600 s is $1.55\times10^{-6}\,\frac{M}{s}$. The rate of formation will be positive because products are being formed.

The instantaneous rate of formation for C_8H_{12} can be linked to the instantaneous rate of dimerization by this equation:

$$\frac{-1}{2}\frac{d[C_4H_6]}{dt} = \frac{d[C_8H_{12}]}{dt}$$

So, for the instantaneous rate of formation for C_8H_{12} at 3200 s, use the value of instantaneous rate of dimerization at 3200 s found earlier and plug into the equation:

$$\frac{-1}{2} \times -7.83 \times 10^-7 \frac{M}{s} = \frac{d[C_8H_{12}]}{dt}$$

$$\frac{d[C_8H_{12}]}{dt} = -3.92 \times 10^{-7}\,\frac{M}{s}$$

The instantaneous rate of formation for C_8H_{12} at 3200 s is $-3.92 \times 10^-7 \frac{M}{s}$

Answer

a. $3.10 \times 10^{-6}\,\frac{M}{s}$ and $1.04 \times 10^{-6}\,\frac{M}{s}$
b. $-7.83 \times 10^-7 \frac{M}{s}$ and $\frac{M}{s}$
c. $-3.92 \times 10^-7 \frac{M}{s}$

Q12.1.5

A study of the rate of the reaction represented as $2A \longrightarrow B$ gave the following data:

Time (s)	0.0	5.0	10.0	15.0	20.0	25.0	35.0
[A] (M)	1.00	0.952	0.625	0.465	0.370	0.308	0.230

a. Determine the average rate of disappearance of *A* between 0.0 s and 10.0 s, and between 10.0 s and 20.0 s.
b. Estimate the instantaneous rate of disappearance of *A* at 15.0 s from a graph of time versus [*A*]. What are the units of this rate?
c. Use the rates found in parts (a) and (b) to determine the average rate of formation of *B* between 0.00 s and 10.0 s, and the instantaneous rate of formation of *B* at 15.0 s.

Solution

Equations: $\frac{-\triangle A}{\triangle time}$ and Rate=$\frac{-\triangle A}{2\triangle time} = \frac{\triangle B}{time}$

Solve: 1.)The change in A from 0s to 10s is .625-1=-.375 so $\frac{-\triangle A}{\triangle time}$ =.375/10= **0.0374 M/s**

Similarly, the change in A from 10 to 20 seconds is .370-.625=-.255 so $\frac{-\triangle A}{\triangle time}$ =.255/20-10= **0.0255M/s**

2.) We can estimate the rate law graphing the points against different order equations to determine the right order.

Zero Order:

$$\frac{d[A]}{dt} = -k$$

$$\int_{A_\circ}^{A} d[A] = -k\int_{0}^{t} dt$$

$$[A] = -kt + [A_\circ]$$

First Order:

$$\frac{d[A]}{dt} = -k[A]$$

$$\int_{A_\circ}^{A} \frac{d[A]}{[A]} = -kdt$$

$$Ln(A) = -kt + Ln(A_\circ)$$

Second Order:

$$\frac{d[A]}{dt} = -k[A]^2$$

$$\int_{A\circ}^{A} \frac{d[A]}{[A]^2} = -k\int_0^t dt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A_\circ]}$$

Now that we have found the linear from of each order we will plot the points vs an [A] y-axis, a Ln(A) y-axis, and a 1/[A] y-axis. whichever of the plots has the most linear points will give us a good idea of the order and the slope will be the k value.



Here we notice that the second order is most linear so we conclude the Rate to be..

$$\frac{-d[A]}{2dt} = k[A]^2$$

At 15 seconds [A]=.465 and from the slope of the graph we find k=.116.so if we plug this data in and multiply both sides by 2 to get rid of the 2 in the denominator on the left side of the equation we find that the rate of disappearance of A is .05 M/s where the units are equivalent to [mol*L^{-1}*s^{-1}]

3.) Using the equation $\frac{-\triangle A}{2\triangle time} = \frac{\triangle B}{time}$ we divide the rates in part a and b in half to get .0188 M/s from 0 to 10 seconds and .025 M/s for the estimated instantaneous rate at 15s.

Answer

(a) average rate, 0 − 10 s = 0.0375 mol L^{-1} s^{-1}; average rate, 12 − 18 s = 0.0225 mol L^{-1} s^{-1}; (b) instantaneous rate, 15 s = 0.0500 mol L^{-1} s^{-1}; (c) average rate for B formation = 0.0188 mol L^{-1} s^{-1}; instantaneous rate for B formation = 0.0250 mol L^{-1} s^{-1}

Q12.1.6

Consider the following reaction in aqueous solution:

$$5\,\mathrm{Br}^-(aq) + \mathrm{BrO}_3^-(aq) + 6\,\mathrm{H}^+(aq) \longrightarrow 3\,\mathrm{Br}_2(aq) + 3\,\mathrm{H}_2\mathrm{O}(l)$$

If the rate of disappearance of $\mathrm{Br}^-(aq)$ at a particular moment during the reaction is $3.5 \times 10^{-4}\ M\ s^{-1}$, what is the rate of appearance of $\mathrm{Br}_2(aq)$ at that moment?

Solution

Step 1. Define the rate of the reaction.

Recall:

For the general reaction: aA + bB → cC+ dD

$rate = -\frac{\Delta[A]}{a\Delta t} = -\frac{\Delta[B]}{b\Delta t} = \frac{\Delta[C]}{c\Delta t} = \frac{\Delta[D]}{d\Delta t}$

So, for the reaction: $5Br^-(aq) + BrO_3^-(aq) + 6H^+ \rightarrow 3Br_2(aq) + 3H_2O(l)$

The rate would be: $rate = -\frac{\Delta[Br^-]}{5\Delta t} = -\frac{\Delta[BrO_3^-]}{\Delta t} = -\frac{\Delta[H^+]}{6\Delta t} = \frac{\Delta[Br_2]}{3\Delta t} = \frac{H_2O}{3\Delta t}$

Step 2. Since we are given the rate for the disappearance of Br^-(aq) is $3.5x10^-4Ms^{-1}$, and we want to find the rate of appearance of Br_2(aq). Therefore we set the two rates equal to each other.

$$rate = -\frac{\Delta[Br^-]}{5\Delta t} = \frac{\Delta[Br_2]}{3\Delta t}$$

$$\text{And,} -\frac{\Delta[Br^-]}{\Delta t} = -3.5x10^{-4}Ms^{-1}$$

$$\text{So, } 3.5x10^{-4}Ms^{-1} = \frac{5}{3}\frac{\Delta[Br_2]}{\Delta t}$$

Step 3. Now solve the equation.

$$\frac{(3.5x10^{-4})(3)}{5} = \frac{\Delta[Br_2]}{\Delta t}$$

$$\frac{\Delta[Br_2]}{\Delta t} = 2.1x10^{-4}Ms^{-1}$$

Answer

$$\frac{\Delta[Br_2]}{\Delta t} = 2.1x10^{-4}Ms^{-1}$$

12.2: Factors Affecting Reaction Rates

Q12.2.1

Describe the effect of each of the following on the rate of the reaction of magnesium metal with a solution of hydrochloric acid: the molarity of the hydrochloric acid, the temperature of the solution, and the size of the pieces of magnesium.

Solution

Molarity of Hydrochloric Acid

- Reaction rates are affected by the frequency at which molecules collide. High Molarity=High Concentration which means more molecules are available to collide thus a faster reaction that one with a low molarity of HCl at a fixed volume.

Temperature of Solution

- Higher temperatures increase the rate of reaction because molecules move faster thus colliding more frequently
- increasing temperatures allows for more particles to move past activation energy barrier to start the reaction

Size of pieces of Magnesium

- reaction rate is dependent on solid reactant size; smaller pieces increases the chance of collision because they enable a greater surface area thus faster reaction rate

Q12.2.2

Go to the PhET Reactions & Rates interactive. Use the Single Collision tab to represent how the collision between monatomic oxygen (O) and carbon monoxide (CO) results in the breaking of one bond and the formation of another. Pull back on the red plunger to release the atom and observe the results. Then, click on “Reload Launcher” and change to “Angled shot” to see the difference.

a. What happens when the angle of the collision is changed?
b. Explain how this is relevant to rate of reaction.

Solution

According to the collision theory, there are many factors that cause a reaction to happen, with three of the factors being how often the molecules or atoms collide, the molecules' or atoms' orientations, and if there is sufficient energy for the reaction to happen. So, if the angle of the plunger is changed, the atom that is shot (a lone Oxygen atom in this case) will hit the other molecule (CO in this case) at a different spot and at a different angle, therefore changing the orientation and the number of proper collisions will most likely not cause for a reaction to happen. Thanks to the simulation, we can see that this is true: depending on the angle selected, the atom may take a long time to collide with the molecule and, when a collision does occur, it may not result in the breaking of the bond and the forming of the other (no reaction happens).

In this particular case, the rate of the reaction will decrease because, by changing the angle, the molecules or atoms won't collide with the correct orientation or as often with the correct orientation.

Q12.2.3

In the PhET Reactions & Rates interactive, use the “Many Collisions” tab to observe how multiple atoms and molecules interact under varying conditions. Select a molecule to pump into the chamber. Set the initial temperature and select the current amounts of each reactant. Select “Show bonds” under Options. How is the rate of the reaction affected by concentration and temperature?

S12.2.3

Based on the Collision Theory, a reaction will only occur if the molecules collide with proper orientation and with sufficient energy required for the reaction to occur. The minimum energy the molecules must collide with is called the activation energy (energy of transition state).

Increasing the concentration of reactants increases the probability that reactants will collide in the correct orientation since there are more reactants in the same volume of space. Therefore, increasing the concentration of reactants would increase the rate of the reaction. Decreasing the concentration of reactants would decrease the rate of reaction because the overall number of possible collisions would decrease.

Temperature is directly related the the kinetic energy of molecules and activation energy E_a is the minimum energy required for a reaction to occur and doesn't change for a reaction. Increasing the temperature increases the kinetic energy of the reactants meaning the reactants will move faster and collide with each other more frequently. Therefore, increasing the temperature increase the rate of the reaction. Decreasing the temperature decreases the rate of reaction since the molecules will have less kinetic energy, move slower, and therefore collide with each other less frequently.

Q12.2.4

In the PhET Reactions & Rates interactive, on the Many Collisions tab, set up a simulation with 15 molecules of A and 10 molecules of BC. Select “Show Bonds” under Options.

a. Leave the Initial Temperature at the default setting. Observe the reaction. Is the rate of reaction fast or slow?
b. Click “Pause” and then “Reset All,” and then enter 15 molecules of A and 10 molecules of BC once again. Select “Show Bonds” under Options. This time, increase the initial temperature until, on the graph, the total average energy line is completely above the potential energy curve. Describe what happens to the reaction.

Solution

a. On the simulation, we select the default setting and the reaction A+BC. In the default setting, we see frequent collisions, a low initial temperature, and a total average energy lower than the energy of activation. The collision theory states that the rate of a reaction is directly proportional to (the fraction of molecules with required orientation), (fractions of collisions with required energy), and (collision frequency). Although we see moving and frequently colliding reactants, the rate of the forward reaction is actually slow because it takes a long time for the products, AB and C, to start appearing. This is mainly because the fractions of collisions with required energy is low, coming from the average energy of the molecules being lower than the energy of activation.

b. The reaction proceeds at an even faster rate. Again, the collision theory states that the rate of a reaction is directly proportional to (the fraction of molecules with required orientation), (fractions of collisions with required energy), and (collision frequency). Because molecules have a higher amount of energy, they have more kinetic energy. With an increased kinetic energy, the molecules not only collide more but also increase in the fraction of collision. However, the forward reaction and the backward reaction both proceed at a fast rate, so both happen almost simultaneously. It takes a shorter time for both reactions to happen. With both of the reactions adding up together overall, there is eventually a state of equilibrium. The process at which equilibrium is reached, however, is faster. Therefore, the amount of products of A+BC stays the same after a while.

12.3: Rate Laws

Q12.3.1

How do the rate of a reaction and its rate constant differ?

S12.3.1

The rate of a reaction or reaction rate is the change in the concentration of either the reactant or the product over a period of time. If the concentrations change, the rate also changes.

Rate for A → B:

$$rate = \frac{\Delta[B]}{\Delta t} = -\frac{\Delta[A]}{\Delta t}$$

The rate constant (k) is a proportionality constant that relates the reaction rates to reactants. If the concentrations change, the rate constant does not change.

For a reaction with the general equation: $aA + bB \rightarrow cC + dD$

the experimentally determined rate law usually has the following form:

$$rate = k[A]^m[B]^n$$

Q12.3.2

Doubling the concentration of a reactant increases the rate of a reaction four times. With this knowledge, answer the following questions:

a. What is the order of the reaction with respect to that reactant?
b. Tripling the concentration of a different reactant increases the rate of a reaction three times. What is the order of the reaction with respect to that reactant?

Solution

(a) 2; (b) 1

Q12.3.3

Tripling the concentration of a reactant increases the rate of a reaction nine times. With this knowledge, answer the following questions:

a. What is the order of the reaction with respect to that reactant?
b. Increasing the concentration of a reactant by a factor of four increases the rate of a reaction four times. What is the order of the reaction with respect to that reactant?

Q12.3.4

How much and in what direction will each of the following affect the rate of the reaction: $\mathrm{CO}(g) + \mathrm{NO_2}(g) \longrightarrow \mathrm{CO_2}(g) + \mathrm{NO}(g)$ if the rate law for the reaction is $\text{rate} = k[\mathrm{NO_2}]^2$?

a. Decreasing the pressure of NO_2 from 0.50 atm to 0.250 atm.
b. Increasing the concentration of CO from 0.01 *M* to 0.03 *M*.

Solution

(a) The process reduces the rate by a factor of 4. (b) Since CO does not appear in the rate law, the rate is not affected.

Q12.3.5

How will each of the following affect the rate of the reaction: $\mathrm{CO}(g) + \mathrm{NO_2}(g) \longrightarrow \mathrm{CO_2}(g) + \mathrm{NO}(g)$ if the rate law for the reaction is $\text{rate} = k[\mathrm{NO_2}][\mathrm{CO}]$?

a. Increasing the pressure of NO_2 from 0.1 atm to 0.3 atm
b. Increasing the concentration of CO from 0.02 *M* to 0.06 *M*.

Q12.3.6

Regular flights of supersonic aircraft in the stratosphere are of concern because such aircraft produce nitric oxide, NO, as a byproduct in the exhaust of their engines. Nitric oxide reacts with ozone, and it has been suggested that this could contribute to depletion of the ozone layer. The reaction $\mathrm{NO} + \mathrm{O_3} \longrightarrow \mathrm{NO_2} + \mathrm{O_2}$ is first order with respect to both NO and O_3 with a rate

constant of 2.20 × 10^7 L/mol/s. What is the instantaneous rate of disappearance of NO when [NO] = 3.3 × 10^{-6} *M* and [O_3] = 5.9 × 10^{-7} *M*?

Solution

4.3 × 10^{-5} mol/L/s

Q12.3.7

Radioactive phosphorus is used in the study of biochemical reaction mechanisms because phosphorus atoms are components of many biochemical molecules. The location of the phosphorus (and the location of the molecule it is bound in) can be detected from the electrons (beta particles) it produces:

$$^{32}_{15}\mathrm{P} \longrightarrow ^{32}_{16}\mathrm{S} + \mathrm{e}^-$$

Rate = 4.85 × 10^{-2} day^{-1} $[^{32}\mathrm{P}]$

What is the instantaneous rate of production of electrons in a sample with a phosphorus concentration of 0.0033 *M*?

Q12.3.8

The rate constant for the radioactive decay of ^{14}C is 1.21 × 10^{-4} year^{-1}. The products of the decay are nitrogen atoms and electrons (beta particles):

$$^{6}_{14}\mathrm{C} \longrightarrow ^{6}_{14}\mathrm{N} + \mathrm{e}^-$$

$$\text{rate} = k[^{6}_{14}\mathrm{C}]$$

What is the instantaneous rate of production of N atoms in a sample with a carbon-14 content of 6.5 × 10^{-9} *M*?

Solution

7.9 × 10^{-13} mol/L/year

Q12.3.9

What is the instantaneous rate of production of N atoms Q12.3.8 in a sample with a carbon-14 content of 1.5 × 10^{-9} *M*?

Q12.3.10

The decomposition of acetaldehyde is a second order reaction with a rate constant of 4.71 × 10^{-8} L/mol/s. What is the instantaneous rate of decomposition of acetaldehyde in a solution with a concentration of 5.55 × 10^{-4} *M*?

Q12.3.11

Alcohol is removed from the bloodstream by a series of metabolic reactions. The first reaction produces acetaldehyde; then other products are formed. The following data have been determined for the rate at which alcohol is removed from the blood of an average male, although individual rates can vary by 25–30%. Women metabolize alcohol a little more slowly than men:

[C_2H_5OH] (*M*)	4.4 × 10^{-2}	3.3 × 10^{-2}	2.2 × 10^{-2}
Rate (mol/L/h)	2.0 × 10^{-2}	2.0 × 10^{-2}	2.0 × 10^{-2}

Determine the rate equation, the rate constant, and the overall order for this reaction.

Solution

rate = *k*; *k* = 2.0 × 10^{-2} mol/L/h (about 0.9 g/L/h for the average male); The reaction is zero order.

Q12.3.12

Under certain conditions the decomposition of ammonia on a metal surface gives the following data:

[NH_3] (*M*)	1.0 × 10^{-3}	2.0 × 10^{-3}	3.0 × 10^{-3}
Rate (mol/L/h[1])	1.5 × 10^{-6}	1.5 × 10^{-6}	1.5 × 10^{-6}

Determine the rate equation, the rate constant, and the overall order for this reaction.

Q12.3.13

Nitrosyl chloride, NOCl, decomposes to NO and Cl_2.

$$2\,NOCl(g) \longrightarrow 2\,NO(g) + Cl_2(g)$$

Determine the rate equation, the rate constant, and the overall order for this reaction from the following data:

[NOCl] (*M*)	0.10	0.20	0.30
Rate (mol/L/h)	8.0×10^{-10}	3.2×10^{-9}	7.2×10^{-9}

Solution

Before we can figure out the rate constant first we must first determine the basic rate equation and rate order. The basic rate equation for this reaction, where n is the rate order of NOCl and k is the rate constant, is

$$rate = k[NOCl]^n$$

since NOCl is the reactant in the reaction.

In order to figure out the order of the reaction we must find the order of [NOCl] as it is the only reactant in the reaction. To do this we must examine how the rate of the reaction changes as the concentration of NOCl changes.

As [NOCl] doubles in concentration from 0.10 M to 0.20 M the rate goes from 8.0 x 10^{-10} to 3.2 x 10^{-9}

(3.2 x 10^{-9}(mol/L/h))/(8.0 x 10^{-10}(mol/L/h)) = 4

so we conclude that as [NOCl] doubles, the rate goes up by 4. Since 2^2= 4 we can say that the order of [NOCl] is 2 so our updated rate law is

$$rate = k[NOCl]^2$$

Now that we have the order, we can substitute the first experimental values from the given table to find the rate constant, k

(8.0 x 10^{-10}(mol/L/h)) = k(0.10 M)^2 so

$$k = \frac{8.0 \times 10^{-10}}{(0.10\,M)^2} = 8 \times 10^{-8}\,M^{-1}\,sec^{-1}$$

We were able to find the units of k using rate order, when the rate order is 2 units of k are M^{-1} x sec^{-1}

So the rate equation is: rate = k[NOCl]^2, it is second order, and k = 8 x 10^{-8} M^{-1} x sec^{-1}

Overall rate law :

$$rate = \underbrace{(8 \times 10^{-8})}_{1/(\text{M x sec})}[NOCl]^2$$

Answer

rate = *k*[NOCl]^2; $k = 8.0 \times 10^{-8}$ L/mol/s; second order

Q12.3.14

From the following data, determine the rate equation, the rate constant, and the order with respect to *A* for the reaction $A \longrightarrow 2C$.

[*A*] (*M*)	1.33×10^{-2}	2.66×10^{-2}	3.99×10^{-2}
Rate (mol/L/h)	3.80×10^{-7}	1.52×10^{-6}	3.42×10^{-6}

Solution

A. Using the experimental data, we can compare the effects of changing [A] on the rate of reaction by relating ratios of [A] to ratios of rates

$$\frac{2.66 \times 10^{-2}}{1.33 \times 10^{-2}} = 2$$

and

$$\frac{1.52 \times 10^{-6}}{3.8 \times 10^{-7}} = 4$$

B. From this we know that doubling the concentration of A will result in quadrupling the rate of reaction. The order of this reaction is 2.

C. We can now write the rate equation since we know the order:

$$rate = k[A]^2$$

D. By plugging in one set of experimental data into our rate equation we can solve for the rate constant, k:

$$3.8 \times 10^{-7} = k \times (1.33 \times 10^{-2})^2$$

$$k = \frac{3.8 \times 10^{-7}}{1.769 \times 10^{-4}}$$

$$k = .00215 M^{-1} s^{-1}$$

Answer

$k = .00215 M^{-1} s^{-1}$

2nd Order

Q12.3.15

Nitrogen(II) oxide reacts with chlorine according to the equation:

$$2\,\mathrm{NO}(g) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{NOCl}(g)$$

The following initial rates of reaction have been observed for certain reactant concentrations:

[NO] (mol/L^1)	[Cl_2] (mol/L)	Rate (mol/L/h)
0.50	0.50	1.14
1.00	0.50	4.56
1.00	1.00	9.12

What is the rate equation that describes the rate's dependence on the concentrations of NO and Cl_2? What is the rate constant? What are the orders with respect to each reactant?

Solution

For the general equation,

$aA + bB \rightarrow cC + dD$

The rate can be written as

$rate = k[A]^m[B]^n$ where k is the rate constant, and m and n are the reaction orders.

For our equation

$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$

the $rate = k[NO]^m[Cl_2]^n$

Now, we need to find the reaction orders. Reaction orders can only be found through experimental values. We can compare two reactions where one of the reactants has the same concentration for both trials, and solve for the reaction order.

$$\frac{rate_1}{rate_2} = \frac{[NO]_1^m[Cl_2]_1^n}{[NO]_2^m[Cl_2]_2^n}$$

We can use the data in the table provided. If we plug in the values for rows 1 and 2, we see that the values for the concentration of Cl will cancel, leaving just the rates and the concentrations of NO.

$$\frac{1.14}{4.56} = \frac{[0.5]^m}{[1.0]^m}$$

We can now solve for m, and we find that m =2. This means that the reaction order for [NO] is 2.

Now we must find the value of n. To do so, we can use the same equation but with the values from rows 2 and 3. This time, the concentration of NO will cancel out.

$$\frac{4.56}{9.12} = \frac{[0.5]^n}{[1.0]^n}$$

When we solve for n, we find that n = 1. This means that the reaction order for [Cl_2] is 1.

We are one step closer to finishing our rate equation.

$$rate = k[NO]^2[Cl_2]$$

Finally, we can solve for the rate constant. To do this, we can use one of the trials of the experiment, and plug in the values for the rate, and concentrations of reactants, then solve for k.

$$1.14 mol/L/h = k[0.5 mol/L]^2[0.5 mol/L]$$

$$k = 9.12 L^2 mol^{-2} h^{-1}$$

So, our final rate equation is:

$$rate = (9.12 L^2 mol^{-2} h^{-1})[NO]^2[Cl_2]$$

*A common mistake is forgetting units. Make sure to track your units throughout the process of determining your rate constant. Be careful because the units will change relative to the reaction order.

Answer

rate = k[NO]^2[Cl]$_2$; k = 9.12 L^2 mol^{-2} h^{-1}; second order in NO; first order in Cl_2

Q12.3.17

Hydrogen reacts with nitrogen monoxide to form dinitrogen monoxide (laughing gas) according to the equation:

$$H_2(g) + 2\,NO(g) \longrightarrow N_2O(g) + H_2O(g)$$

Determine the rate equation, the rate constant, and the orders with respect to each reactant from the following data:

[NO] (*M*)	0.30	0.60	0.60
[H_2] (*M*)	0.35	0.35	0.70
Rate (mol/L/s)	2.835 × 10^{-3}	1.134 × 10^{-2}	2.268 × 10^{-2}

Solution

Determine the rate equation, the rate constant, and the orders with respect to each reactant.

The rate constant and the orders can be determined through the differential rate law. The general form of the differential rate law is given below:

aA + bB + cC ==> products

rate = - d[A]/dt = k[A]^n[B]^m[C]^p

where A, B, and C are the concentrations of the reactants, k is the rate constant, and n,m, and p refer to the order of each reactant.

To find the orders of each reactant, we see that when [NO] doubles but [H_2] doesn't change, the rate quadruples, meaning that [NO] is a second order reaction ([NO]^2). When [H_2] doubles but [NO] doesn't change, the rate doubles, meaning that [H_2] is a first order reaction. So the rate law would look something like this:

$$\text{Rate} = k[NO]^2[H_2]$$

We can use this rate law to determine the value of the rate constant. Plug in the data for reactant concentration and rate from one of the trials to solve for k the rate constant. In this case, we chose to use the data from trial 1 from the second column of the data table.

$$2.835 \times 10^{-3} = k[0.3]^2[0.35]$$

$$k = .09\ M^{-2}/s^{-1}$$

Q12.3.18

For the reaction $A \longrightarrow B + C$, the following data were obtained at 30 °C:

[*A*] (*M*)	0.230	0.356	0.557
Rate (mol/L/s)	4.17×10^{-4}	9.99×10^{-4}	2.44×10^{-3}

a. What is the order of the reaction with respect to [*A*], and what is the rate equation?
b. What is the rate constant?

Solution

1. The rate equation for an n order reaction is given as $\frac{dr}{dt} = k[A]^n$. Where $[A]$ is the concentration in M, and $\frac{dr}{dt}$ is the rate in M/s.

We can then use each set of data points, plug its values into the rate equation and solve for n. Note you can use any of the data points as long as the concentration corresponds to its rate.

Rate equation 1: $4.17 \times 10^{-4} = k[0.230]^n$

Rate equation 2: $9.99 \times 10^{-4} = k[0.356]^n$

We divide Rate equation 1 by Rate equation 2 in order to cancel out k, the rate constant.

$\frac{4.17 \times 10^{-4}}{9.99 \times 10^{-4}} = \frac{k[0.230]^n}{k[0.356]^n}$

$0.417 = 0.646^n$

Now the only unknown we have is n. Using logarithm rules one can solve for it.

$ln\ 0.417 = n \cdot ln\ 0.646$

$\frac{ln\ 0.417}{ln\ 0.646} = n = 2$

The rate equation is second order with respect to A and is written as $\frac{dr}{dt} = k[A]^2$.

2. We can solve for k by plugging in any data point into our rate equation $\frac{dr}{dt} = k[A]^2$.

Using the first data points for instance $[A] = 0.230\ \frac{mol}{L}$ and $\frac{dr}{dt} = 4.17 \times 10^{-4}\ \frac{mol}{L \cdot s}$] we get the equation $4.17 \times 10^{-4}\ \frac{mol}{L \cdot s} = k[0.230\ \frac{mol}{L}]^2$

Which solves for $k = 7.88 \times 10^{-3}\ \frac{L}{mol \cdot s}$

Since we know this is a second order reaction the appropriate units for k can also be written as $\frac{1}{M \cdot s}$

Answer

(a) The rate equation is second order in A and is written as rate = $k[A]^2$. (b) k = 7.88×10^{-13} L mol^{-1} s^{-1}

Q12.3.19

For the reaction $Q \longrightarrow W + X$, the following data were obtained at 30 °C:

$[Q]_{initial}$ (M)	0.170	0.212	0.357
Rate (mol/L/s)	6.68×10^{-3}	1.04×10^{-2}	2.94×10^{-2}

a. What is the order of the reaction with respect to [Q], and what is the rate equation?
b. What is the rate constant?

Solution

What is the order of the reaction with respect to [Q], and what is the rate equation?

- Order reaction: 2 because when you use the ratio trial 3:2, it will look like this:
- $(\frac{2.94 * 10^{-2}}{1.04 * 10^{-2}}) = (\frac{0.357^{x}}{0.212^{x}})$
- $2.82 = 1.7^{x}$
- x = 2 so the order of reaction is 2
- Rate reaction equation: Rate=k[Q]^2

What is the rate constant?

- To find the rate constant (k) simply plug and calculate one of the trials into the rate equation
- 1.04 x 10^{-2}=k[0.212]^2
- k=0.231 $M^{-1}s^{-1}$

Answer

Order: 2

k=0.231 $M^{-1}s^{-1}$

Q12.3.20

The rate constant for the first-order decomposition at 45 °C of dinitrogen pentoxide, N_2O_5, dissolved in chloroform, $CHCl_3$, is 6.2 × 10^{-4} min^{-1}.

$$2\,N_2O_5 \longrightarrow 4\,NO_2 + O_2$$

What is the rate of the reaction when [N_2O_5] = 0.40 M?

Solution

Step 1: The first step is to write the rate law. We know the general formula for for a first-order rate law. It is as follows: Rate=k[A]

Step 2: We now plug in [N_2O_5] in for [A] in our general rate law. We also plug in our rate constant (k), which was given to us. Now our equation looks as follows:

Rate=(6.2x10^{-4} min^{-1})[N_2O_5]

Step 3: We now plug in our given molarity. [N_2O_5]=0.4 M. Now our equation looks as follows:

Rate=(6.2x10^{-4} min^{-1})(0.4 M)

Step 4: We now solve our equation. Rate=(6.2x10^{-4} min^{-1})(0.4 M)= 2.48x10^{-4} M/min.
Step 5: Use significant figures and unit conversion to round 2.48x10^{-4} M/min to 2.5 × 10^{-4} (moles)$L^{-1}min^{-1}$

Answer

(a) 2.5 × 10^{-4} mol/L/min

Q12.3.21

The annual production of HNO_3 in 2013 was 60 million metric tons Most of that was prepared by the following sequence of reactions, each run in a separate reaction vessel.

a. $4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(g)$
b. $2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g)$
c. $3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(aq) + NO(g)$

The first reaction is run by burning ammonia in air over a platinum catalyst. This reaction is fast. The reaction in equation (c) is also fast. The second reaction limits the rate at which nitric acid can be prepared from ammonia. If equation (b) is second order in NO and first order in O_2, what is the rate of formation of NO_2 when the oxygen concentration is 0.50 *M* and the nitric oxide concentration is 0.75 *M*? The rate constant for the reaction is 5.8×10^{-6} $L^2/mol^2/s$.

Solution

To determine the rate law for an equation we need to look at its slow step. Since both equation a and c are fast, equation b can be considered the slow step of the reaction. The slow step is also considered the rate determining step of the system.

Hence, The rate determining step is the second step because it's the slow step.

rate of production of $NO_2 = k[A]^m[B]^n$

$rate = k[NO]^2[O_2]^1\ M/s$

$rate = (5.8 * 10^{-6})[0.75]^2[0.5]^1\ M/s$

$rate = 1.6 * 10^{-6}\ M/s$

Answer

$rate = 1.6 * 10^{-6}\ M/s$

Q12.3.22

The following data have been determined for the reaction:

$$I^- + OCl^- \longrightarrow IO^- + Cl^-$$

	1	2	3
$[I^-]_{initial}$ (*M*)	0.10	0.20	0.30
$[OCl^-]_{initial}$ (*M*)	0.050	0.050	0.010
Rate (mol/L/s)	3.05×10^{-4}	6.20×10^{-4}	1.83×10^{-4}

Determine the rate equation and the rate constant for this reaction.

Solution

Using the reactants, we can form the rate law of the reaction: $$ r=k[OCl^-]^n[I^-]^m \]

From there, we need to use the data to determine the order of both $[OCl^-]$ and $[I^-]$. In doing so, we need to compare r_1 to r_2 such that:

$$\frac{r_1}{r_2} = \frac{(0.10^m)(0.050^n)}{(0.20^m)(0.050^n)} = \frac{3.05 \times 10^{-4}}{6.20 \times 10^{-4}} \tag{12.12.3}$$

$$0.5^m = 0.5 \tag{12.12.4}$$

$$m = 1 \tag{12.12.5}$$

We can "cross out" the concentration of $[OCl^-]$ because it has the same concentration in both of the trials used.

Now that we know m ($[I^-]$) has a first order of 1.

We cannot "cross out" $[I^-]$ to find $[OCl^-]$ because no two trials have the same concentration. In order to solve for n we will plug in 1 for m.

$$\frac{r_1}{r_3} = \frac{(0.10^1)(0.050^n)}{(0.30^1)(0.010^n)} = \frac{3.05 \times 10^{-4}}{1.83 \times 10^{-4}} \tag{12.12.6}$$

$$\frac{1}{3}(5^n) = 1.6666667 \tag{12.12.7}$$

$$5^n = 5 \tag{12.12.8}$$

$$n = 1 \tag{12.12.9}$$

Since we know that orders of both n and m are equal to one, we can not substitute them into the rate law equation along with the respective concentrations (from either the first, second, or third reaction) and solve for the rate constant, k.

$$r = k[OCl^-]^n[I^-]^m \tag{12.12.10}$$

$$3.05 * 10^{-4} = k[0.05]^1[0.10]^1 \tag{12.12.11}$$

$$k = 6.1 * 10^{-2} \frac{L}{mol \times s} \tag{12.12.12}$$

Thus the overall rate law is: $$ r = (6.1 * 10^{-2} \frac {L}{mol \times s})[OCl^-][I^-] \]

The units for K depend on the overall order of the reaction. To find the overall order we add m and n together. By doing this we find an overall order of 2. This is why the units for K are $$ \frac {L}{mol \times s} \]

Answer

rate = k[I$^-$][OCl^{-1}]; k = 6.1 × 10^{-2} L mol $^{-1}$ s^{-1}

Q12.3.23

In the reaction

$$2NO + Cl_2 \rightarrow 2NOCl$$

the reactants and products are gases at the temperature of the reaction. The following rate data were measured for three experiments:

Initial p{NO}$	Initial p{Cl$_2$}$	Initial rate
(atm)	(atm)	(moles of A consumed atm sec^{-1})
0.50	0.50	5.1 x 10^{-3}
1.0	1.0	4.0 x 10^{-2}
0.50	1.0	1.0 x 10^{-2}

a. From these data, write the rate equation for this gas reaction. What order is the reaction in NO, Cl_2, and overall?
b. Calculate the specific rate constant for this reaction.

Solution

a. The rate equation can be determined by designing experiments that measure the concentration(s) of one or more reactants or products as a function of time. For the reaction $A + B \rightarrow products$, for example, we need to determine k and the exponents m and n in the following equation:

$$rate = k[A]^m[B]^n$$

To do this, the initial concentration of B can be kept constant while varying the initial concentration of A and calculating the initial reaction rate. This information would deduce the reaction order with respect to A. The same process can be done to find

the reaction order with respect to B.
In this particular example,

$$\frac{rate_2}{rate_3} = \frac{k[A_2]^m[B_2]^n}{k[A_3]^m[B_3]^n}$$

So taking the values from the table,

$$\frac{4.0*10^{-2}}{1.0*10^{-2}} = \frac{k[1.0]^m[1.0]^n}{k[0.5]^m[1.0]^n}$$

and by canceling like terms, you are left with

$$\frac{4.0*10^{-2}}{1.0*10^{-2}} = \frac{[1.0]^m}{[0.5]^m}$$

Now, solve for m

$4 = 2^m \Longrightarrow m = 2$ Because m=2, the reaction with respect to NO is 2. NO **is second order.**

You can repeat the same process to find n.

$$\frac{rate_3}{rate_1} = \frac{k[A_3]^m[B_3]^n}{k[A_1]^m[B_1]^n}$$

Taking the values from the table,

$$\frac{1.0*10^{-2}}{5.1*10^{-3}} = \frac{k[0.5]^m[1.0]^n}{k[0.5]^m[0.5]^n}$$

and by canceling like terms, you are left with

$$\frac{1.0*10^{-2}}{5.1*10^{-3}} = \frac{[1.0]^n}{[0.5]^n}$$

Now this time, solve for n

$2 = 2^n \Longrightarrow n = 1$

Because n=1, the reaction with respect to Cl_2 is 1. Cl_2 **is first order.**

So the rate equation is

$$rate = k[NO]^2[Cl_2]^1$$

To find the overall rate order, you simply add the orders together. Second order + first order makes the **overall reaction third order.**

b. The rate constant is calculated by inserting the data from any row of the table into the experimentally determined rate law and solving for k. For a third order reaction, the units of k are $frac1atm^2 * sec$. Using Experiment 1,

$$rate = k[NO]^2[Cl_2]^1 \Longrightarrow 5.1 * 10^{-3} \frac{atm}{sec} = k[0.5matm]^2[0.5atm]^1$$

$$k = 0.0408 \frac{1}{atm^2 * sec}$$

Answer

NO is second order.

Cl_2 is first order.

Overall reaction order is three.

b)

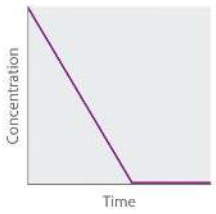
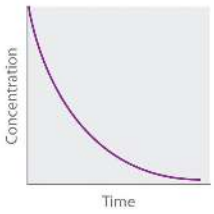
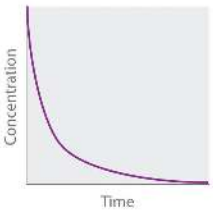
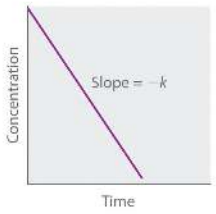
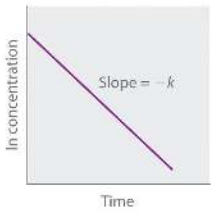
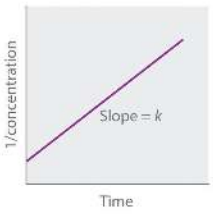
$k = 0.0408\ atm^{-2} * sec^{-1}$

12.4: Integrated Rate Laws

Q12.4.1

Describe how graphical methods can be used to determine the order of a reaction and its rate constant from a series of data that includes the concentration of *A* at varying times.

Solution

	Zeroth Order	First Order	Second Order
Differential rate law	$\text{Rate} = -\frac{\Delta[A]}{\Delta t} = k$	$\text{Rate} = -\frac{\Delta[A]}{\Delta t} = k[A]$	$\text{Rate} = -\frac{\Delta[A]}{\Delta t} = k[A]^2$
Concentration vs. time			
Integrated rate law	$[A] = [A]_0 - kt$	$[A] = [A]_0 e^{-kt}$ or $\ln[A] = \ln[A]_0 - kt$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$
Straight-line plot to determine rate constant			
Relative rate vs. concentration	[A], *M* / Rate, M/s 1 / 1 2 / 1 3 / 1	[A], *M* / Rate, M/s 1 / 1 2 / 2 3 / 3	[A], *M* / Rate, M/s 1 / 1 2 / 4 3 / 9
Half-life	$t_{1/2} = \frac{[A]_0}{2k}$	$t_{1/2} = \frac{0.693}{k}$	$t_{1/2} = \frac{1}{k[A]_0}$
Units of *k*, rate constant	M/s	1/s	$M^{-1}\cdot s^{-1}$

To determine the order of a reaction when given the data series, one must graph the data how it is, graph it as natural log of [A], and graph it as 1/[A]. Whichever method yields a straight line will determine the order. Respective of the methods of graphing above, if a straight line is yielded by the first graphing method its a 0 order, if by the second method it's a 1st order, and the third graphing method its a 2nd order. When the order of the graph is known, a series of equations, given in the above image, can be used with the various points on the graph to determine the value of k. We can see that we need an initial value of A and a final value of A, and both of these would be given by the data.

Zeroth order when plotting initial concentrating versus final concentration you have a negative linear slope.

$$[A] = [A]_0 - kt$$

First order when plotting ln[initial concentration] versus ln[final concentration] you have a negative linear slope.

$$\ln[A] = \ln[A]_0 - kt$$

Second order when plotting the 1/[initial concentration] versus 1/[final concentration] you have a positive linear slope.

$$\frac{1}{[\text{A}]} = \frac{1}{[\text{A}]_0} + kt$$

Q12.4.2

Use the data provided to graphically determine the order and rate constant of the following reaction: $\text{SO}_2\text{Cl}_2 \longrightarrow \text{SO}_2 + \text{Cl}_2$

Time (s)	0	5.00×10^3	1.00×10^4	1.50×10^4	2.50×10^4	3.00×10^4	4.00×10^4

$[SO_2Cl_2]$ (*M*)	0.100	0.0896	0.0802	0.0719	0.0577	0.0517	0.0415

Solution

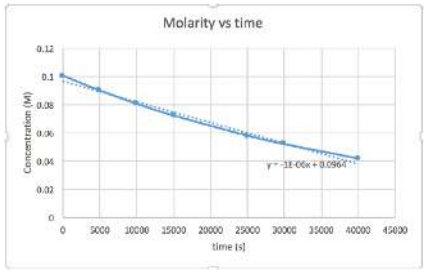
Use the data to graphically determine the order and rate constant of the following reaction.

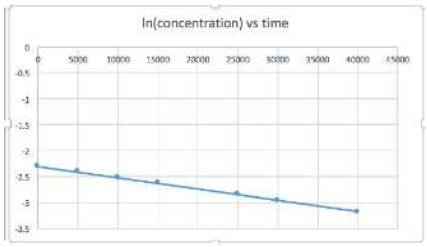
In order to determine the rate law for a reaction from a set of data consisting of concentration (or the values of some function of concentration) versus time, make three graphs of the data based on the integrated rate laws of each order reaction.

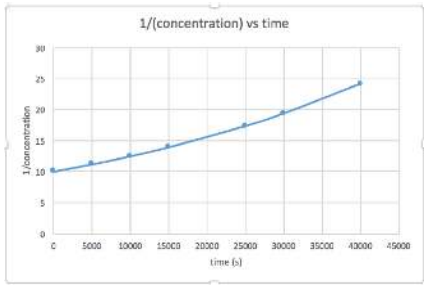
[concentration] versus time (linear for a zero order reaction) ln [concentration] versus time (linear for a 1st order reaction) 1 / [concentration] versus time (linear for a 2nd order reaction)

slope= -2.0×10^{-5}

k = 2.0×10^{-5}





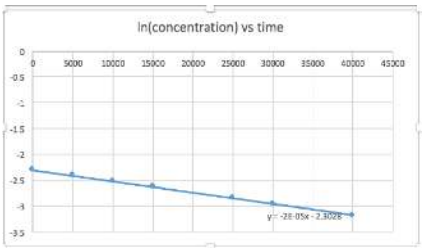


The graph that is linear indicates the order of the reaction. Then, you can find the correct rate equation:

zero order reaction	rate = k	(k = - slope of line)
1st order reaction	rate = k[A]	(k = - slope of line)
2nd order reaction	rate = $k[A]^2$	(k = slope of line)

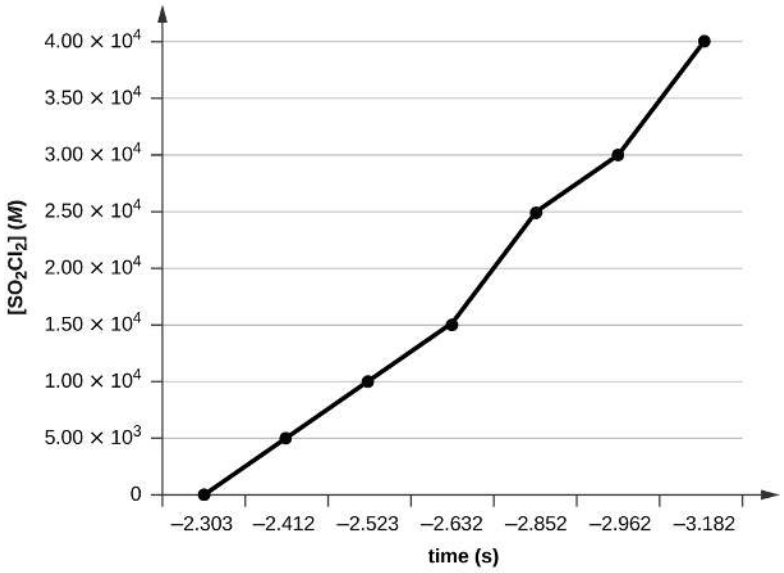
In this graph, ln(concentration) vs time is linear, indicating that the **reaction is first order**.

k=-slope of line



Answer

Plotting a graph of $\ln[SO_2Cl_2]$ versus t reveals a linear trend; therefore we know this is a first-order reaction:



$k = -2.20 \times 10^5\ s^{-1}$

Q12.4.3

Use the data provided in a graphical method to determine the order and rate constant of the following reaction:

$$2P \longrightarrow Q + W$$

Time (s)	9.0	13.0	18.0	22.0	25.0
[P] (*M*)	1.077×10^{-3}	1.068×10^{-3}	1.055×10^{-3}	1.046×10^{-3}	1.039×10^{-3}

Solution

Add texts here. Do not delete this text first.

Q12.4.4

Pure ozone decomposes slowly to oxygen, $2\,O_3(g) \longrightarrow 3\,O_2(g)$. Use the data provided in a graphical method and determine the order and rate constant of the reaction.

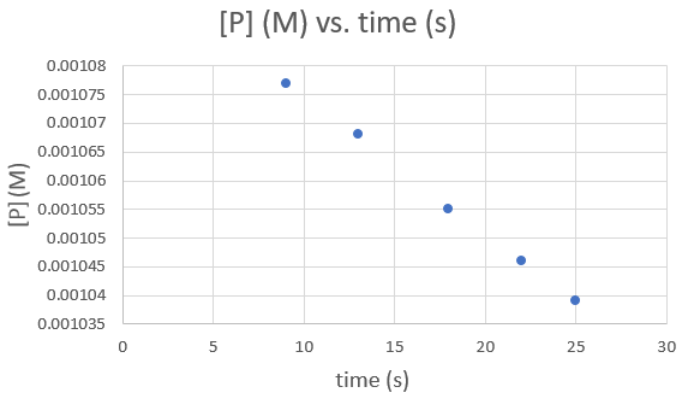
Time (h)	0	2.0×10^3	7.6×10^3	1.23×10^4	1.70×10^4	1.70×10^4
$[O_3]$ (*M*)	1.00×10^{-5}	4.98×10^{-6}	2.07×10^{-6}	1.39×10^{-6}	1.22×10^{-6}	1.05×10^{-6}

Solution

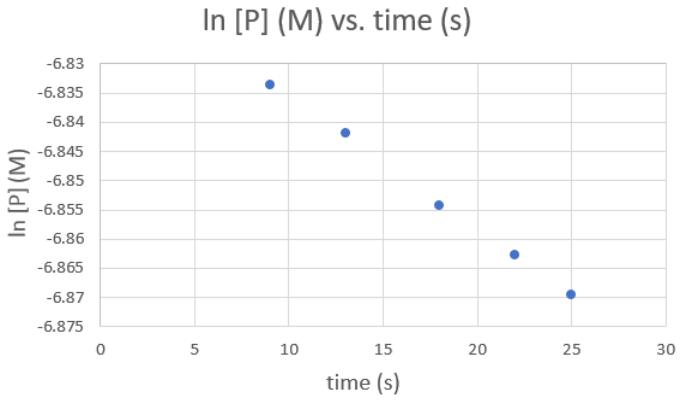
To determine the order and rate constant, you need to graph the data for zero order, first order, and second order by plotting concentration versus time- [A] vs. time, natural logarithm (ln) of [A] vs. time, and 1/[A] vs. time respectively. The order of the

reaction is determined by identifying which of these three graphs produces a straight line. The rate constant *k* is represented by the slope of the graph. The graphs with their respective data values would be

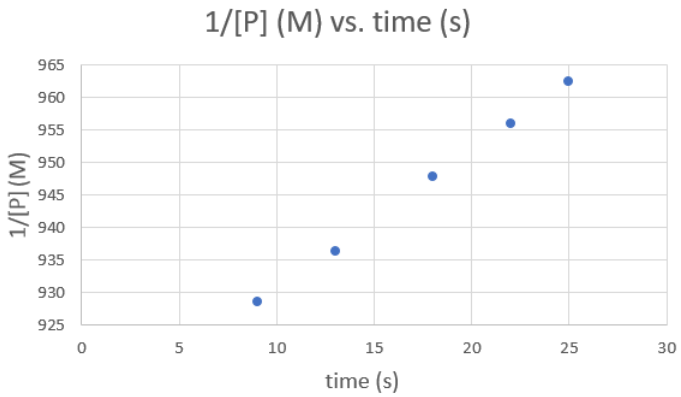
Time (s)	9.0	13.0	18.0	22.0	25.0
[P] (*M*)	1.077 × 10−3	1.068 × 10−3	1.055 × 10−3	1.046 × 10−3	1.039 × 10−3



Time (s)	9.0	13.0	18.0	22.0	25.0
ln [P] (*M*)	-6.83358	-6.84197	-6.85421	-6.86278	-6.8695

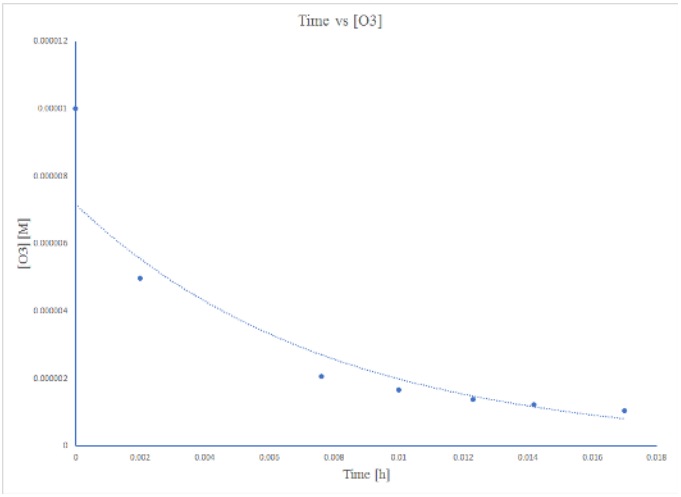


Time (s)	9.0	13.0	18.0	22.0	25.0
1/[P] (*M*)	928.5051	936.3296	947.8673	956.0229	962.4639

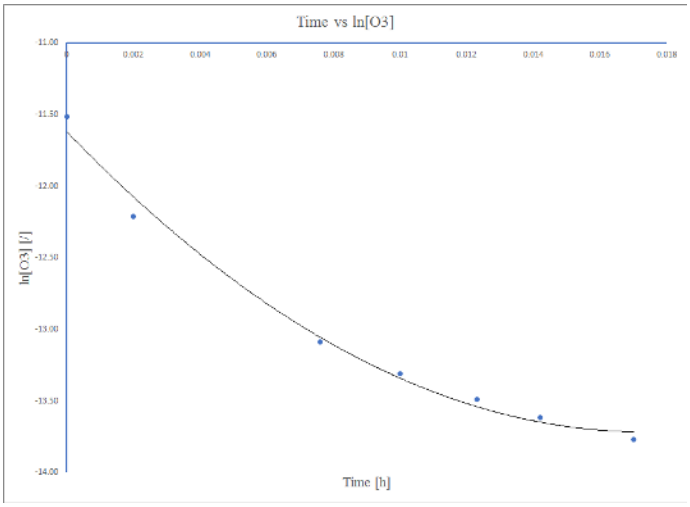


Since each graph yields a straight line **the order and rate constant of the reaction cannot be determined**.

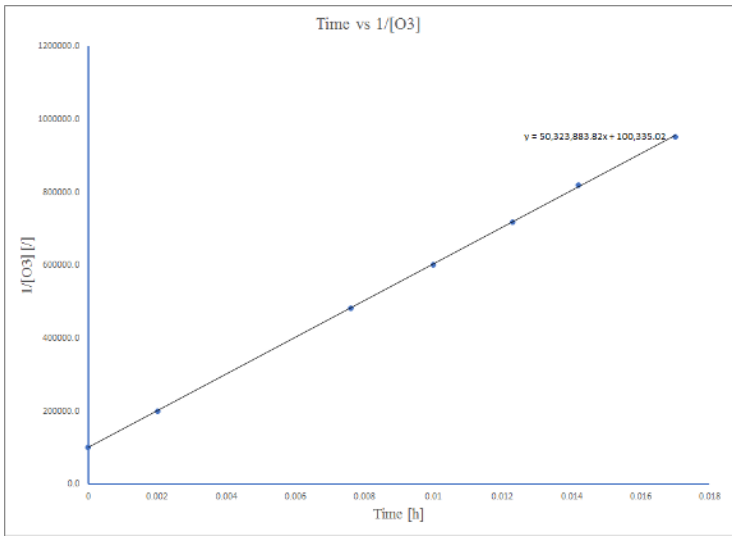
To identify how the concentrations changes a function of time, requires solving the appropriate differential equation (i.e., the differential rate law).
The zero-order rate law predicts in a linear decay of concentration with time
The 1st-order rate law predicts in an exponential decay of concentration with time
The 2nd-order rate law predicts in an reciprocal decay of concentration with time



The plot is not linear, so the reaction is not zero order.



The plot is not linear, so the reaction is not first order.



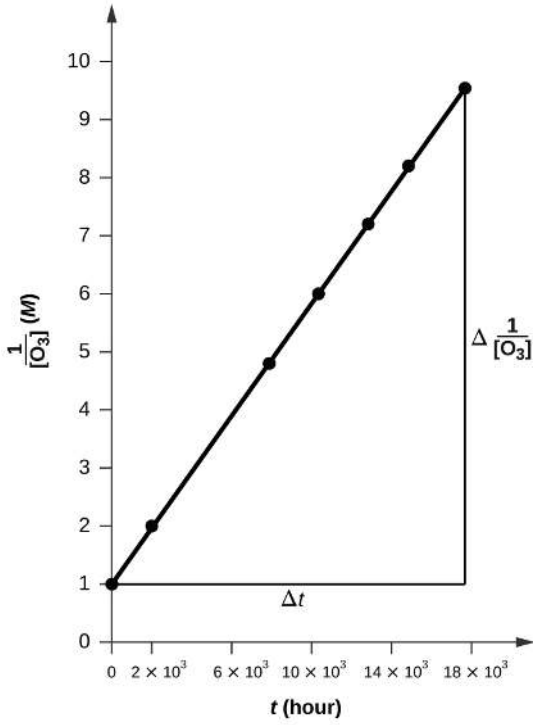
The plot is nicely linear, so the reaction is second order.

To a second order equation, $1/[A] = k*t+1/[A_0]$

Thus, the value of K is the slope of the graph Time vs $\frac{1}{O_3}$,

k = 50.3*10^6 L mol^{-1} h^{-1}

Answer



The plot is nicely linear, so the reaction is second order.

k = 50.1 L mol^{-1} h^{-1}

Q12.4.5

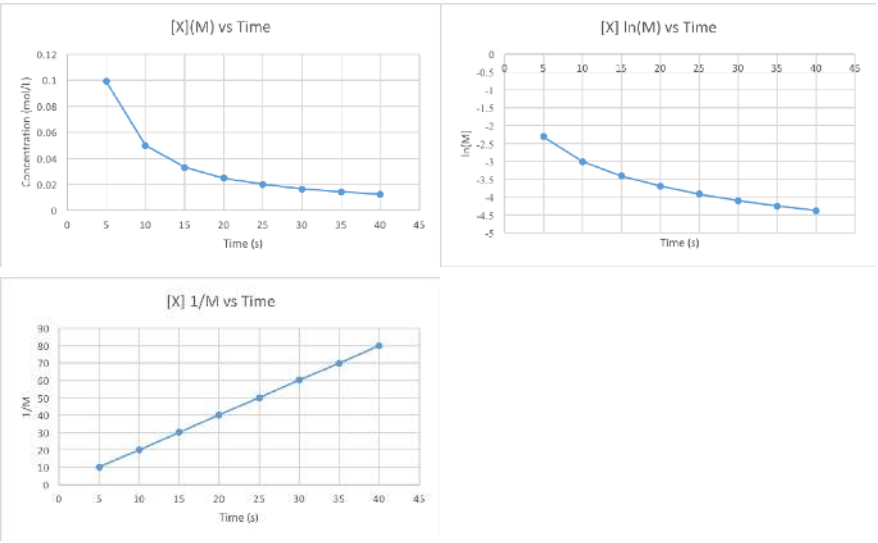
From the given data, use a graphical method to determine the order and rate constant of the following reaction:

$$2X \longrightarrow Y + Z \tag{12.12.13}$$

Time (s)	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0
[*X*] (*M*)	0.0990	0.0497	0.0332	0.0249	0.0200	0.0166	0.0143	0.0125

Solution

In order to determine the order of the reaction we need to plot the data using three different graphs. All three graphs will have time in seconds as the x-axis, but the y-axis is what will differ. One graph will plot concentration versus time, the second will plot natural log of concentration versus time, and the other will plot 1/concentration versus times. Whichever graph results in a line, we know that must be the order of the reaction. If we get a line using the first graph, it will be zero order, if it is a line for the second graph it will be first order, and if it is a line for the third graph it will be a second order reaction. Now lets plot the data to determine the order.



We can clearly see that the third graph, which plots 1/M versus time, is a straight line while the other two are slightly curved. Therefore, we can determine that the rate of this reaction is second order. This also tells us that the units of the rate constant which should be $M^{-2}s^{-1}$ for a second order reaction.

To determine the rate constant, called k, we simple need to figure out the slope of the third graph since that is the order of this reaction. To find the slope of the line, we take two points and subtract the y values and then divide them by the difference of the x values. This is how to do it:

Use the points (5, 10.101) and (40, 80).

Now use these to get the slop, aka the rate constant: (80-10.101)/(40-5) = 1.997 = k

So the rate constant for this second order reaction is 1.997 $M^{-1}s^{-1}$.

Q12.4.6

What is the half-life for the first-order decay of phosphorus-32? $\left({}^{32}_{15}\mathrm{P} \longrightarrow {}^{32}_{16}\mathrm{S} + \mathrm{e}^{-}\right)$ The rate constant for the decay is 4.85×10^{-2} day^{-1}.

Solution

This is a first order reaction, so we can use our half life equation below:

$$t_{1/2} = \frac{0.693}{k}$$

The rate constant is given to us in units per day. All we have to do, is to plug it into the equation.

$$t_{1/2} = \frac{0.693}{4.85 * 10^{-2}}$$

$$= 14.3\ days$$

A12.4.6

14.3 d

Q12.4.7

What is the half-life for the first-order decay of carbon-14? $({}^{6}_{14}\mathrm{C} \longrightarrow {}^{7}_{14}\mathrm{N} + \mathrm{e}^{-})$ The rate constant for the decay is 1.21 × 10^{-4} $year^{-1}$.

Solution

To find the half life, we need to use the first-order half-life equation. All half life reactions undergo first order reactions.

The half-life equation for first order is

$$t_{1/2} = ln2/k$$

with k being the rate constant. The rate constant for carbon-14 was given as $1.21 \times 10^{-4} year^{-1}$.

Plug it in the equation.

$$t_{1/2} = ln2/(1.21 \times 10^{-4} year^{-1})$$

and solve for $t_{1/2}$.

When you calculate it, the half life for carbon-14 is 5.73*10^{3}

Answer

The half-life for carbon-14 is calculated to be 5.73*10^{3}

Q12.4.8

What is the half-life for the decomposition of NOCl when the concentration of NOCl is 0.15 *M*? The rate constant for this second-order reaction is 8.0 × 10^{-8} L/mol/s.

Solution

The half-life of a reaction, $t_{1/2}$, is the amount of time that is required for a reactant concentration to decrease by half compared to its initial concentration. When solving for the half-life of a reaction, we should first consider the order of reaction to determine it's rate law. In this case, we are told that this reaction is second-order, so we know that the integrated rate law is given as:

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

Isolating for time, we find that:

$$t_{1/2} = \frac{1}{k[A]_0}$$

Now it is just a matter of substituting the information we have been given to calculate $t_{1/2}$, where the rate constant, k, is equal to 8.0 × 10^{-8} L/mol/s and initial concentration, $[A]_0$, is equal to 0.15*M*:

$$t_{1/2} = \frac{1}{(8.0 \times 10^{-8})(0.15)} = 8.33 \times 10^{7}\ seconds$$

Answer

8.33 × 10^{7} s

Q12.4.9

What is the half-life for the decomposition of O_3 when the concentration of O_3 is $2.35 \times 10^{-6}\ M$? The rate constant for this second-order reaction is 50.4 L/mol/h.

Solution

Add texts here. Do not delete this text first.

Since the reaction is second order, its half-life is

$$t_{\frac{1}{2}} = \frac{1}{k[\mathrm{A}]_0}$$

$$t_{1/2} = \frac{1}{(50.4 M^{-1}/h)[2.35 \times 10^{-6} M]}$$

So, half-life is 8443 hours.

Q12.4.10

The reaction of compound *A* to give compounds *C* and *D* was found to be second-order in *A*. The rate constant for the reaction was determined to be 2.42 L/mol/s. If the initial concentration is 0.500 mol/L, what is the value of $t_{1/2}$?

Solution

As mentioned in the question the reaction of compound *A* will result in the formation of compounds C and D. This reaction was found to be second-order in *A*. Therefore, we should use the second order equation for half-life which relates the rate constant and initial concentrations to the half-life:

$$t_{\frac{1}{2}} = \frac{1}{k[A]_0}$$

Since we were given *k* (rate constant) and Initial concentration of A, we have everything needed to calculate the half life of A.

$$k = 0.5 \frac{\frac{L}{mol}}{s}$$

$$[A]_0 = 2.42 \frac{mol}{L}$$

When we plug in the given information notice that the units cancel out to seconds.

$$t_{\frac{1}{2}} = \frac{1}{\frac{2.42 L mol^{-}}{s}[0.500 \frac{mol}{L}]} = 0.826 s$$

Answer

0.826 s

Q12.4.11

The half-life of a reaction of compound *A* to give compounds *D* and *E* is 8.50 minutes when the initial concentration of *A* is 0.150 mol/L. How long will it take for the concentration to drop to 0.0300 mol/L if the reaction is (a) first order with respect to *A* or (b) second order with respect to *A*?

Solution

Organize the given variables:

(half-life of *A*) $t_{1/2} = 8.50 min$
(initial concentration of *A*) $[A]_0 = 0.150 mol/L$
(target concentration of *A*) $[A] = 0.0300 mol/L$

Find the the rate constant k, using the half-life formulas for each respective order. After finding k, use the integrated rate law respective to each order and the initial and target concentrations of *A* to find the time it took for the concentration to drop.

(a) first order with respect to *A*

(half-life) $t_{1/2} = \frac{ln(2)}{k} = \frac{0.693}{k}$
(rearranged for k) $k = \frac{0.693}{t_{1/2}}$
(plug in $t_{1/2}$ = 8.50 min) $k = \frac{0.693}{8.50min} = 0.0815min^{-1}$
(integrated rate law) $ln[A] = -kt + ln[A]_0$
(rearranged for t) $ln(\frac{[A]}{[A]_0}) = -kt$
$-ln(\frac{[A]}{[A]_0}) = kt$
$ln(\frac{[A]}{[A]_0})^{-1} = kt$
$ln(\frac{[A]_0}{[A]}) = kt$
$t = \frac{ln(\frac{[A]_0}{[A]})}{k}$
(plug in variables) $t = \frac{ln(\frac{0.150mol/L}{0.0300mol/L})}{0.0815min^{-1}} = \frac{ln(5.00)}{0.0815min^{-1}} = 19.7min$

(b) second order with respect to *A*

(half-life) $t_{1/2} = \frac{1}{k[A]_0}$
(rearranged for k) $k = \frac{1}{t_{1/2}[A]_0}$
(plug in variables) $k = \frac{1}{(8.50min)(0.150mol/L)} = \frac{1}{1.275min \cdot mol/L} = 0.784L/mol \cdot min$
(integrated rate law) $\frac{1}{[A]} = kt + \frac{1}{[A]_0}$
(rearranged for t) $\frac{1}{[A]} - \frac{1}{[A]_0} = kt$
$t = \frac{1}{k}(\frac{1}{[A]} - \frac{1}{[A]_0})$
(plug in variables) $t = \frac{1}{0.784L/mol \cdot min}(\frac{1}{0.0300mol/L} - \frac{1}{0.150mol/L}) = \frac{1}{0.784L/mol \cdot min}(\frac{80}{3}L/mol) = 34.0min$

Answer

a) 19.7 min

b) 34.0 min

Q12.4.12

Some bacteria are resistant to the antibiotic penicillin because they produce penicillinase, an enzyme with a molecular weight of 3 × 10^4 g/mol that converts penicillin into inactive molecules. Although the kinetics of enzyme-catalyzed reactions can be complex, at low concentrations this reaction can be described by a rate equation that is first order in the catalyst (penicillinase) and that also involves the concentration of penicillin. From the following data: 1.0 L of a solution containing 0.15 µg (0.15 × 10^{-6} g) of penicillinase, determine the order of the reaction with respect to penicillin and the value of the rate constant.

[Penicillin] (*M*)	**Rate (mol/L/min)**
2.0 × 10^{-6}	1.0 × 10^{-10}
3.0 × 10^{-6}	1.5 × 10^{-10}
4.0 × 10^{-6}	2.0 × 10^{-10}

Solution

The first step is to solve for the order or the reaction. This can be done by setting up two expressions which equate the rate to the rate constant times the molar concentration of penicillin raised to the power of it's order. Once we have both expressions set

up, we can divide them to cancel out k (rate constant) and use a basic logarithm to solve for the exponent, which is the order. It will look like this.

$$\text{rate(mol/L/min)}=k[M]^x$$

$$(1.0 \times 10^{-10})=k[2.0 \times 10^{-6}]^x$$

$$(1.5 \times 10^{-10})=k[3.0 \times 10^{-6}]^x$$

Dividing the two equations results in the expression:

$$(2/3)=(2/3)^x$$

*A single ratio equation can also be set up to solve for the reaction order:

*

$$\frac{rate_1}{rate_2}=\frac{k[Penicillin]_1^x}{k[Penicillin]_2^x}$$

*We then solve for x in a similar fashion.

*

$$\frac{1.0x10^{-10}}{1.5x10^{-10}}=\frac{[2.0x10^{-6}]^x}{[3.0x10^{-6}]^x}$$

We can now use the natural logarithm to solve for x, or simply and intuitively see that in order for the equation to work, **x must be equal to one. Thus, the reaction is of the first order.**

Now that we have the order of the reaction, we can proceed to solve for the value of the rate constant. Substituting x=1 into our first equation yields the expression:

$$(1 \times 10^{-10})=k[2.0 \times 10^{-6}]^1$$

$$k=(1 \times 10^{-10})/(2 \times 10^{-6})$$

$$k= (5 \times 10^{-5}) \text{ min}^{-1}$$

We have a unit of min^{-1} because we divided (mol/L/min) by molarity, which is in (mol/L), yielding a unit of min^{-1}.

We were given two important pieces of information to finish the problem. It is stated that the enzyme has a molecular weight of 3×10^4 g/mol, and that we have a one liter solution that contains $(0.15 \times 10^{-6}$ g) of penicillinase. Dividing the amount of grams by the molecular weight yields **5×10^{-12} moles**.

$$(0.15 \times 10^{-6}) \text{ g} / (3 \times 10^4) \text{ g/mol} = (5 \times 10^{-12}) \text{ mol}$$

Now that we have the amount of moles, we can divide our rate constant by this value.

$$(5 \times 10^{-5}) \text{ min}^{-1} / (5 \times 10^{-12}) \text{ mol} = \mathbf{(1 \times 10^7) \text{ mol}^{-1} \text{ min}^{-1}}$$

Answer

The reaction is first order with $k = 1.0 \times 10^7$ mol^{-1} min^{-1}

Q12.4.13

Both technetium-99 and thallium-201 are used to image heart muscle in patients with suspected heart problems. The half-lives are 6 h and 73 h, respectively. What percent of the radioactivity would remain for each of the isotopes after 2 days (48 h)?

Solution

This problem is asking us for the percentage of radioactivity remaining after a certain time for both isotopes after 48 hours. We must identify an equation that will help us solve this and we can determine that we can determine this information using the first order equation.

This equation $Ln(N/N_o)$= -kt tells that the Natural log of the fraction remaining is equal to the rate constant times time. To determine the rate constant, we can also compute .693 over the half-life given in the information.

For Technetium-99 we can determine the rate constant by plugging into the second equation: .693/6 hrs= .1155 h^{-1}

Now that we have the rate constant we can plug in : $Ln(N/N_o)$=-(.1155h^{-1})(48h) so $Ln(N/N_o)$=-5.544 and if we take the inverse of the natural log, we get (N/N_o)=3.9x10 $^{-3}$ and if we multiply this by 100, we get .39% remaining.

We can do this same process for Thallium-201 and plugin: .693/73 hrs= .009493151 h^{-1} and when we plug this into the first order equation we get:

$Ln(N/N_o)$=-(.009493h^{-1})(48h) so $Ln(N/N_o)$=-.45567248 and when we take the inverse of the natural log, we get (N/N_o)=.6340 and when multiplied by 100, we get 63.40% remaining which makes sense since its half-life is 73 hours and only 48 hours have passed, half of the amount has yet to be consumed.

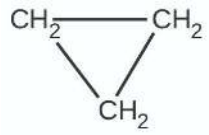
Answer

Technetium-99: 0.39%

Thallium-201: 63.40%

Q12.4.14

There are two molecules with the formula C_3H_6. Propene, $CH_3CH{=}CH_2$, is the monomer of the polymer polypropylene, which is used for indoor-outdoor carpets. Cyclopropane is used as an anesthetic:



When heated to 499 °C, cyclopropane rearranges (isomerizes) and forms propene with a rate constant of 5.95 × 10^{-4} s^{-1}. What is the half-life of this reaction? What fraction of the cyclopropane remains after 0.75 h at 499 °C?

Solution

Use the equation

$$t_{1/2} = \frac{ln2}{k}$$

since this is a first-order reaction. You can tell that this is a **first order reaction** due to the units of measurement of the rate constant, which is s^{-1}. Different orders of reactions lead to different rate constants, and a rate constant of s^{-1} will always be first order.

Plug into the equation, and you get half life = 1164.95 seconds. To convert this to hours, we would divide this number by 3600 seconds/hour, to get **0.324 hours**.

Use the integrated first order rate law

$$ln\frac{[A]}{[A]_0} = -kt$$

. In this equation, $[A]_0$ represents the initial amount of compound present at time 0, while [A] represents the amount of compound that is left after the reaction has occurred. Therefore, the fraction

$$\frac{[A]}{[A]_0}$$

is equal to the fraction of cyclopropane that remains after a certain amount of time, in this case, 0.75 hours.

Substitute x for the fraction of

$$\frac{[A]}{[A]_0}$$

into the integrated rate law:

$$ln\frac{[A]}{[A]_0} = -kt$$

$$ln(x) = -5.95x10^{-4}(0.75)$$

$$x = e^{(-0.000595)(0.75)}$$

= 0.20058 = 20%.

So, the half life is 0.324 hours, and 20% of the cyclopropane will remain as 80% will have formed propene.

Answer

0.324 hours. ; 20% remains

Q12.4.16

Fluorine-18 is a radioactive isotope that decays by positron emission to form oxygen-18 with a half-life of 109.7 min. (A positron is a particle with the mass of an electron and a single unit of positive charge; the nuclear equation is $^{18}_{9}\text{F} \longrightarrow {}^{18}_{8}\text{O} + {}^{0}_{1}\text{e}^{+}$.) Physicians use ^{18}F to study the brain by injecting a quantity of fluoro-substituted glucose into the blood of a patient. The glucose accumulates in the regions where the brain is active and needs nourishment.

a. What is the rate constant for the decomposition of fluorine-18?
b. If a sample of glucose containing radioactive fluorine-18 is injected into the blood, what percent of the radioactivity will remain after 5.59 h?
c. How long does it takFe for 99.99% of the ^{18}F to decay?

Solution

a) The nuclear decay of an isotope of an element is represented by the first order equation:

$$\ln(N/N0) = -kt$$

Where t is time, N0 is the initial amount of the substance, N is the amount of the substance after time t, and k is the rate constant. We can rearrange the equation and isolate k so that we could solve for the rate constant:

$$k = [-\ln(N/N0)] / t$$

We are given that fluorine-18 has a half-life of 109.7 minutes. Since we have the half-life, we can choose an arbitrary value for N_0 and use half of that value for N. In this case, we choose 100 for N_0 and 50 for N. Now we can plug in those values into the equation above and solve for k.

$$k = [-\ln(50/100)] / 109.7$$

$$k = 0.6931 / 109.7 = 0.006319 \text{ min}^{-1}$$

The rate constant for this reaction is 0.006319 min^{-1}.

b) For this problem, we are able to use the same equation from part a:

$$\ln(N/N0) = -kt$$

However, this time we are given the amount of time elapsed instead of the half-life, and we are asked to determine the percent of fluorine-18 radioactivity remaining after that time. In this problem, we must plug in values for N0, k (determined from part a), and t.

But first, since we are given the elapsed time in hours, we must convert it into minutes:

$$5.59 \text{ hours} \times (60 \text{ minutes} / 1 \text{ hours}) = 335.4 \text{ minutes}$$

This gives us the value for t. We also have values for k (0.006319 min^{-1}) and N_0 (again an arbitrary number.) Now we can plug values into the original equation, giving us:

$$\ln(N/100) = -(0.006319)(335.4)$$

We solve this equation by taking the exponential of both sides:

$$e^{\ln(N/100)} = e^{-(0.006319)(335.4)}$$

where $e^{\ln}$ equals 1 and now we can just solve for N:

$$N/100 = e^{-(0.006319)(335.4)}$$

$$N = [e^{-(0.006319)(335.4)}] \times 100 = 12.0$$

Since 100 was used as the initial amount and 12.0 was determined as the remaining amount, 12.0 can be used as the percentage of remaining amount of radioactivity of fluorine-18. **Thus the percent of fluorine-18 radioactivity remaining after 5.59 hours is 12.0%.**

c) This part of the question is much like the previous two parts, but this time we are given the initial amount of radioactivity, the final amount of radioactivity and we are asked do determine how long it took for that amount of radioactivity to decay. We are able to use the same equation:

$$\ln(N/N0) = -kt$$

However, now we are given N and N_0 and we have already determined k from before. We are told that 99.99% of the radioactivity has decayed, so we can use 100 and 0.01 for N_0 and N respectively. We plug these values in to the equation, solve for t, and get

$$\ln(0.01/1000) = -0.006319t$$

$$-9.21 = -0.006319t$$

$$t = 1458 \text{ minutes}$$

Therefore, its takes 1458 minutes for 99.99% of the radioactivity to decay.

Answer

a) 0.006319 min^{-1}

b) 12.0%

c) 1458 minutes

Q12.4.17

Suppose that the half-life of steroids taken by an athlete is 42 days. Assuming that the steroids biodegrade by a first-order process, how long would it take for $\frac{1}{64}$ of the initial dose to remain in the athlete's body?

Solution

252 days

for first order reaction: $t_{1/2} = 0.693 / k$

$k = 0.693 / 42$

$k = 0.0165$

for first order reaction: $[A] = [A]_0 e^{-kt}$

1/64 initial means that: $[A] = 1/64 [A]_0$

therefore: $1/64 [A]_0 = [A]_0 e^{-0.0165t}$

t = 252 days

Q12.4.18

Recently, the skeleton of King Richard III was found under a parking lot in England. If tissue samples from the skeleton contain about 93.79% of the carbon-14 expected in living tissue, what year did King Richard III die? The half-life for carbon-14 is 5730 years.

Solution

In order to find out what year King Richard III died, set [A]/[A_0] (the percent of carbon-14 still contained) equal to $0.5^{\text{time(t)/half life } (t_{1/2})}$ or use the equation $N(t) = N_0e^{-rt}$.

Using the first equation:

$A/A_0 = 0.5^{t/t_{1/2}}$ plug in the given numbers $.9379 = 0.5^{t/5730}$ and solve for t.

$ln.9379 = (t/5730)(ln0.5)$ (using the rule of logs)

$-.0641 = (t/5730)(-.693)$

$-367.36 = -.693t$

$t = 530.1 years$

Using $N(t) = N_0e^{-rt}$ this problem is solved by the following:

$1/2 = e^{-5730r}$

$r = 0.000121$

Now that we know what r is, we can use this value in our original formula and solve for t, the amount of years that have passed.

This time, we use 93.78, the percent of the carbon-14 remaining as N(t) and 100 as the original, N_0.

$93.78 = 100e^{-0.000121t}$

$t = 530.7$ years

Another way of doing this is by using these two equations:

$\lambda = \frac{0.693}{t_{1/2}}$ and $\frac{n_t}{n_0} = -\lambda t$

n_t = concentration at time t (93.79)

n_0 = initial concentration (100)

First solve for lambda or the decay constant by plugging in the half life.

Then plug in lambda and the other numbers into the second equation, and solve for t- which should equal to 530.1 years as well.

If we want to find out what year King Richard III died, we take the current year, 2017, and subtract 530 years. Doing this, we find that King Richard III died in the year 1487.

Answer

King Richard III died in the year 1487

Q12.4.19

Nitroglycerine is an extremely sensitive explosive. In a series of carefully controlled experiments, samples of the explosive were heated to 160 °C and their first-order decomposition studied. Determine the average rate constants for each experiment using the following data:

Initial [$C_3H_5N_3O_9$] (*M*)	4.88	3.52	2.29	1.81	5.33	4.05	2.95	1.72
***t* (s)**	300	300	300	300	180	180	180	180
% Decomposed	52.0	52.9	53.2	53.9	34.6	35.9	36.0	35.4

Solution

First we need to understand what the question is asking for: the average rate constant. The average rate constant is the variable "k" when discussing kinetics and it can be defined as the proportionality constant in the equation that expresses the relationship

between the rate of a chemical reaction and the concentrations of the reacting substances. Knowing that we need to find K in this first order reaction, we can look to formulas that include "k," initial and final concentrations $[A]_o and [A]_t$, and half life time "t." Since this is a first order reaction, we can look to the first order equations, and doing that we find one that includes the variables given in the question:

$$\ln[A]_t = -kt + \ln[A]_o$$

For the first reaction, we have an initial concentration of 4.88 M, and a percentage decomposed. To find the final concentration, we must multiply the initial concentration by the percentage decomposed to know how much decomposed, and subtract that from the original to find out how much is left: 4.88M x 0.52= 2.54 M and 4.88M-2.54M=2.34M

Now, we have the variables we need, and we plug it into the equation above:

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[2.34M] = -k(300s) + \ln[4.88M]$

$k = \frac{-(\ln[2.34M] - \ln[4.88M])}{300}$

$k = 2.45x10^{-3}$

Since it asks for the rate constant of each experiment, we now must do the same procedure for each data set to find the rate constant:

Second experiment

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[1.66M] = -k(300s) + \ln[3.52M]$

$k = \frac{-(\ln[1.66M] - \ln[3.52M])}{300}$

$k = 2.51x10^{-3}$

Third experiment

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[1.07M] = -k(300s) + \ln[2.29M]$

$k = \frac{-(\ln[1.07M] - \ln[2.29M])}{300}$

$k = 2.54x10^{-3}$

Fourth experiment

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[0.834M] = -k(300s) + \ln[1.81M]$

$k = \frac{-(\ln[0.834M] - \ln[1.81M])}{300}$

$k = 2.58x10^{-3}$

Fifth Experiment

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[3.49M] = -k(180s) + \ln[5.33M]$

$k = \frac{-(\ln[3.49M] - \ln[5.33M])}{180}$

$k = 2.35x10^{-3}$

Sixth Experiment

$\ln[A]_t = -kt + \ln[A]_o$

$\ln[2.60M] = -k(180s) + \ln[4.05M]$

$k=\frac{-(\ln[2.60M]-\ln[4.05M])}{180}$

$k=2.46x10^{-3}$

Seventh Experiment

$\ln[A]_t=-kt+\ln[A]_o$

$\ln[1.89M]=-k(180s)+\ln[2.95M]$

$k=\frac{-(\ln[1.89M]-\ln[2.95M])}{180}$

$k=2.47x10^{-3}$

Eighth experiment

$\ln[A]_t=-kt+\ln[A]_o$

$\ln[1.11M]=-k(180s)+\ln[1.72M]$

$k=\frac{-(\ln[1.11M]-\ln[1.72M])}{180}$

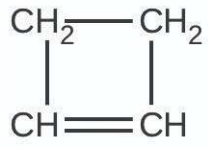
$k=2.43x10^{-3}$

Answer

$[A]_0$ (M)	$k \times 10^3$ (s^{-1})
4.88	2.45
3.52	2.51
2.29	2.54
1.81	2.58
5.33	2.35
4.05	2.44
2.95	2.47
1.72	2.43

Q12.4.20

For the past 10 years, the unsaturated hydrocarbon 1,3-butadiene ($CH_2=CH-CH=CH_2$) has ranked 38th among the top 50 industrial chemicals. It is used primarily for the manufacture of synthetic rubber. An isomer exists also as cyclobutene:



The isomerization of cyclobutene to butadiene is first-order and the rate constant has been measured as 2.0×10^{-4} s^{-1} at 150 °C in a 0.53-L flask. Determine the partial pressure of cyclobutene and its concentration after 30.0 minutes if an isomerization reaction is carried out at 150 °C with an initial pressure of 55 torr.

Solution

Since this is a first order reaction, the integrated rate law is: $[A_t]=[A_0]e^{-kt}$

Partial Pressure: Use the integrated rate law to find the partial pressure at 30 minutes:

Use A_0 = 55 torr, t = 30 min, and k = $2.0*10^{-4}s^{-1}$ to solve the integrated rate law equation:

$[A_{30}]=(55torr)*e^{-(2.0x10^{-4}\frac{1}{sec})(30min\cdot\frac{60sec}{1min})}$

Solve this equation to get:

$[A_{30}] = (55torr) * e^{-0.36}$

$A_{30}]$ = 38.37 torr.

Initial Concentration: Find the initial concentration using the ideal gas law.

The ideal gas law is given by $PV = nRT \rightarrow n = \frac{PV}{RT}$. Use this form of the gas law to solve for the initial concentration n.

Use V = 0.53L, R = 0.08206 $\frac{L*atm}{mol*L}$, T = 423.15 K, and P = $\frac{1atm}{760}$ = 0.07237 atm .

Solve the ideal gas equation using these values:

$n = \frac{(55torr)(0.53L)}{(0.08206\frac{L*atm}{mol*K})(423.15K)} = 0.00110$ moles cyclobutene.

Now find the initial concentration of cyclobutene A_0 using the equation $[A_0] = \frac{n}{V}$:

$A_0 = \frac{n}{V} = \frac{0.00110moles}{0.53L} = 0.00208M$

Concentration at 30 minutes: Find the concentration of cyclobutene at 30 minutes by using the integrated rate law given above, using time t = 30 minutes, or 1800 seconds.

$[A_{30}] = (0.00208M)e^{-0.36} = 0.00145M$

So at 30 minutes, the cyclobutene concentration is 0.00145 M, and the partial pressure is 38.37 torr.

Answer

Partial Pressure: 38.37 torr.

Concentration: 0.00145 M

12.5: Collision Theory

Q12.5.1

Chemical reactions occur when reactants collide. What are two factors that may prevent a collision from producing a chemical reaction?

Solution

The two factors that may prevent a collision from producing a chemical reaction are:

1. **Kinetic energy of the molecule**

In order for chemical reactions to occur, molecules require enough velocity to overcome the minimum activation energy needed to break the old bonds and form new bonds with other molecules. At higher temperatures, the molecules possess the minimum amount of kinetic energy needed which ensures the collisions will be energetic enough to lead to a reaction.

2. **The orientation of molecules during the collision**

Two molecules have to collide in the right orientation in order for the reaction to occur. Molecules have to orient properly for another molecule to collide at the right activation state.

Q12.5.2

When every collision between reactants leads to a reaction, what determines the rate at which the reaction occurs?

Solution

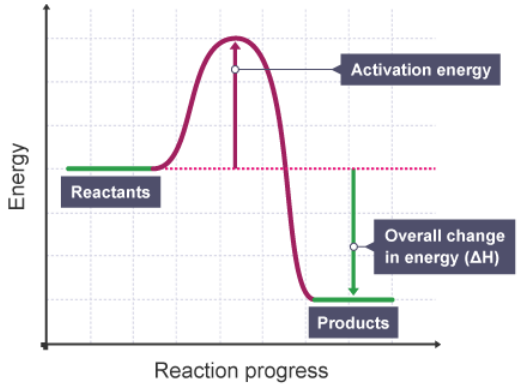
There has to be contact between reactants for a reaction to occur. The more the reactants collide, the more often reactions can occur. Factors that determine reaction rates include concentration of reactants, temperature, physical states of reactants, surface area, and the use of a catalyst. The reaction rate usually increases as the concentration of a reactant increases. Increasing the temperature increases the average kinetic energy of molecules, causing them to collide more frequently, which increases the reaction rate. When two reactants are in the same fluid phase, their particles collide more frequently, which increases the reaction rate. If the surface area of a reactant is increased, more particles are exposed to the other reactant therefore more

collisions occur and the rate of reaction increases. A catalyst participates in a chemical reaction and increases the reaction rate without changing itself.

Q12.5.3

What is the activation energy of a reaction, and how is this energy related to the activated complex of the reaction?

Solution



Activation energy is the energy barrier that must be overcome in order for a reaction to occur. To get the molecules into a state that allows them to break and form bonds, the molecules must be contorted (deformed, or bent) into an unstable state called the transition state. The transition state is a high-energy state, and some amount of energy – the activation energy – must be added in order for the molecule reach it. Because the transition state is unstable, reactant molecules don't stay there long, but quickly proceed to the next step of the chemical reaction.The activated complex is the highest energy of the transition state of the reaction.

Q12.5.5

Describe how graphical methods can be used to determine the activation energy of a reaction from a series of data that includes the rate of reaction at varying temperatures.

Solution

This method is based on the **Arrhenius equation** which can be used to show the effect of a change of temperature on the **rate constant**, and therefore on the rate of reaction.

The rate constant is different from reaction rat in that the reaction rate is the measure of how fast or slow a chemical reaction takes place while a rate constant is a constant that shows the relationship between the reaction rate and the concentrations of the reactants or products.

For example, for the reaction $A + B \rightarrow C$, the rate law would be:

$$rate = k[A]^a[B]^b$$

k = rate constant

[A] = concentration of reactant A

a = order of reaction with respect to A

[B] = concentration of reactant B

b = order of reaction with respect to B

However, the rate constant remains constant only if you are changing the concentration of the reactants. If you change the temperature or the catalyst of the reaction, the rate constant will change and this is demonstrated by the Arrhenius equation:

$$k = Ae^{\frac{-E_a}{RT}}$$

$$ln\left(\frac{k_1}{k_2}\right) = \left(\frac{-E_a}{R}\right)\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

k = rate constant

A = frequency factor

E_a = activation energy

e = exponential function, e^x

R = gas constant

T = temperature (K)

In other words, the activation energy of a reaction, E_a, from a series of data that includes the rate of reaction, k, at varying temperatures can be determined by graphing it on a plot of $\ln k$ versus $\frac{1}{T}$. You can then use the slope of the graph you have plotted to solve for E_a by setting the slope equal to $\frac{-E_a}{R}$.

Q12.5.6

How does an increase in temperature affect rate of reaction? Explain this effect in terms of the collision theory of the reaction rate.

S12.5.6

Collision theory states that the rates of chemical reactions depend on the fraction of molecules with the correct orientation, fraction of collisions with required energy, and the collision frequency. Because the fraction of collisions with required energy is a function of temperature, as temperature increases, the fraction of collisions with required energy also increases. The kinetic energy of reactants also increases with temperature which means molecules will collide more often increasing collisions frequency. With increased fraction of collisions with required energy and collisions frequency, the rate of chemical reaction increases. We see mathematically, from the Arrhenius equation, that temperature and the rate constant are related.

$$k = Ae^{\frac{E_a}{RT}} \tag{12.12.14}$$

where k is the rate constant, A is a specific constant, R is 8.3145 J/K, Ea is the reaction-specific activation energy in J, and T is temperature in K. We see from the equation that k is very sensitive to changes in the temperature.

Q12.5.7

The rate of a certain reaction doubles for every 10 °C rise in temperature.

a. How much faster does the reaction proceed at 45 °C than at 25 °C?
b. How much faster does the reaction proceed at 95 °C than at 25 °C?

Solution

By finding the difference in temperature, 45 °C - 25 °C, we get 20 °C. Since the rate of the reaction doubles every 10 °C increase in temperature and the rate of the reaction experienced a 20 °C increase in temperature, we see that the reaction rate doubled twice ($2^2 = 4$). As a result, the **reaction proceeds four times faster**.

Following the same process as in part a, we get the difference in temperature to be 70 °C. Since the rate of the reaction doubles every 10 °C increase in temperature and the system experienced a 70 °C change, we see that the reaction doubled seven times ($2^7 = 128$). We can see the **reaction proceeds 128 times faster**.

(a) 4-times faster (b) 128-times faster

Q12.5.8

In an experiment, a sample of $NaClO_3$ was 90% decomposed in 48 min. Approximately how long would this decomposition have taken if the sample had been heated 20 °C higher?

S12.5.8

First off, it is important to recognize that this decomposition reaction is a **first-order reaction**, which can be written as follows: $2NaClO_3 \rightarrow 2NaCl + 3O_2$

Understanding this, it is important to be able to then be able to recognize which equation would be most useful given the initial conditions presented by the question. Since we are dealing with time, percentage of material left, and temperature, the only viable equation that could relate all of this would be the Arrhenius Equation, which is written as follows: $\ln(\frac{k_2}{k_1}) = \frac{Ea}{R}(\frac{1}{t_1} - \frac{1}{t_2})$

However, this problem does not give us enough information such as what the activation energy is or the initial temperature in order to mathematically solve this problem. Additionally, the problem tells us to approximate how long the decomposition would take, which means we are asked to answer this question conceptually based on our knowledge of thermodynamics and reaction rates. As a general rule of thumb, we know that for every 10˚C rise in temperature the rate of reaction doubles. Since the question tells us that there is a 20˚C rise in temperature we can deduce that the reaction rate doubles twice, as per the general rule mentioned before. This means **the overall reaction rate for this decomposition would quadruple**, or would be 4 times faster than the reaction rate at the initial temperature.

We can gut check this answer by recalling how an increase in the average kinetic energy (temperature) decreases the time it takes for the reaction to take place and increase the reaction rate. Thus, if we increase the temperature we should have a faster reaction rate.

Q12.5.9

The rate constant at 325 °C for the decomposition reaction $\mathrm{C_4H_8} \longrightarrow 2\,\mathrm{C_2H_4}$ is 6.1 × 10^{-8} s^{-1}, and the activation energy is 261 kJ per mole of C_4H_8. Determine the frequency factor for the reaction.

Solution

S12.5.9

Using the Arrhenius equation allows me to find the frequency factor, A.

$k=Ae^{-Ea/RT}$

k, Ea, R, and T are all known values. k, Ea, and T are given in the problem as $6.1x10^{-8}$, 261 kJ, and 598 K, respectively.

So, plugging them into the equation gives:

$6.1x10^{-8}\ s^{-1}=Ae^{(-261000\ J)/(8.3145\ J/mol)(598\ K)}$

Take $e^{(-261000\ J)/(8.3145\ J/mol)(598)}$ and get 1.59×10^{-23}. Divide k, 6.1×10^{-8}, by 1.59×10^{-23} and get $A=3.9 \times 10^{15} s^{-1}$

A12.5.9

$3.9 \times 10^{15}\ s^{-1}$

Q12.5.10

The rate constant for the decomposition of acetaldehyde (CH_3CHO), to methane (CH_4), and carbon monoxide (CO), in the gas phase is 1.1 × 10^{-2} L/mol/s at 703 K and 4.95 L/mol/s at 865 K. Determine the activation energy for this decomposition.

S12.5.10

The equation for relating the rate constant and activation energy of a reaction is the Arrhenius equation:

$$k = Ae^{-\frac{E_a}{RT}} \tag{12.12.15}$$

When given two rate constants at two different temperatures but for the same reaction, the Arrhenius equation can be rewritten as:

$$ln(\frac{k_2}{k_1}) = \frac{E_a}{R}(\frac{1}{T_1} - \frac{1}{T_2}) \tag{12.12.16}$$

In this problem, all the variables are given except for the E_a (activation energy).

k_1 = 1.1 × 10^{-2} L/mol/s

T_1 = 703 K

k_2 = 4.95 L/mol/s

T_2 = 865 K

R = 8.314 J/(mol K) (Ideal Gas Constant)

Now plug in all these values into the equation, and solve for E_a.

$$ln(\frac{4.95\frac{L}{mol\times s}}{1.1\times10^{-2}\frac{L}{mol\times s}}) = \frac{E_a}{8.314\times10^{-3}\frac{kJ}{mol\times K}}(\frac{1}{703}-\frac{1}{865}) \quad (12.12.17)$$

E_a = 190 kJ (2 sig figs)

Q12.5.11

An elevated level of the enzyme alkaline phosphatase (ALP) in the serum is an indication of possible liver or bone disorder. The level of serum ALP is so low that it is very difficult to measure directly. However, ALP catalyzes a number of reactions, and its relative concentration can be determined by measuring the rate of one of these reactions under controlled conditions. One such reaction is the conversion of p-nitrophenyl phosphate (PNPP) to p-nitrophenoxide ion (PNP) and phosphate ion. Control of temperature during the test is very important; the rate of the reaction increases 1.47 times if the temperature changes from 30 °C to 37 °C. What is the activation energy for the ALP–catalyzed conversion of PNPP to PNP and phosphate?

Solution

43.0 kJ/mol

Q12.5.12

In terms of collision theory, to which of the following is the rate of a chemical reaction proportional?

a. the change in free energy per second
b. the change in temperature per second
c. the number of collisions per second
d. the number of product molecules

Q12.5.13

Hydrogen iodide, HI, decomposes in the gas phase to produce hydrogen, H_2, and iodine, I_2. The value of the rate constant, *k*, for the reaction was measured at several different temperatures and the data are shown here:

Temperature (K)	k (M^{-1} s^{-1})
555	6.23×10^{-7}
575	2.42×10^{-6}
645	1.44×10^{-4}
700	2.01×10^{-3}

What is the value of the activation energy (in kJ/mol) for this reaction?

Solution

177 kJ/mol

Q12.5.14

The element Co exists in two oxidation states, Co(II) and Co(III), and the ions form many complexes. The rate at which one of the complexes of Co(III) was reduced by Fe(II) in water was measured. Determine the activation energy of the reaction from the following data:

T (K)	k (s^{-1})

T (K)	k (s^{-1})
293	0.054
298	0.100

Q12.5.15

The hydrolysis of the sugar sucrose to the sugars glucose and fructose,

$$C_{12}H_{22}O_{11} + H_2O \longrightarrow C_6H_{12}O_6 + C_6H_{12}O_6$$

follows a first-order rate equation for the disappearance of sucrose: Rate = $k[C_{12}H_{22}O_{11}]$ (The products of the reaction, glucose and fructose, have the same molecular formulas but differ in the arrangement of the atoms in their molecules.)

a. In neutral solution, $k = 2.1 \times 10^{-11}$ s^{-1} at 27 °C and 8.5×10^{-11} s^{-1} at 37 °C. Determine the activation energy, the frequency factor, and the rate constant for this equation at 47 °C (assuming the kinetics remain consistent with the Arrhenius equation at this temperature).
b. When a solution of sucrose with an initial concentration of 0.150 *M* reaches equilibrium, the concentration of sucrose is 1.65×10^{-7} *M*. How long will it take the solution to reach equilibrium at 27 °C in the absence of a catalyst? Because the concentration of sucrose at equilibrium is so low, assume that the reaction is irreversible.
c. Why does assuming that the reaction is irreversible simplify the calculation in part (b)?

Solution

E_a = 108 kJ

$A = 2.0 \times 10^8$ s^{-1}

$k = 3.2 \times 10^{-10}$ s^{-1}

(b) 1.81×10^8 h or 7.6×10^6 day. (c) Assuming that the reaction is irreversible simplifies the calculation because we do not have to account for any reactant that, having been converted to product, returns to the original state.

Q12.5.16

Use the PhET Reactions & Rates interactive simulation to simulate a system. On the “Single collision” tab of the simulation applet, enable the “Energy view” by clicking the “+” icon. Select the first $A + BC \longrightarrow AB + C$ reaction (A is yellow, B is purple, and C is navy blue). Using the “straight shot” default option, try launching the *A* atom with varying amounts of energy. What changes when the Total Energy line at launch is below the transition state of the Potential Energy line? Why? What happens when it is above the transition state? Why?

Q12.5.17

Use the PhET Reactions & Rates interactive simulation to simulate a system. On the “Single collision” tab of the simulation applet, enable the “Energy view” by clicking the “+” icon. Select the first $A + BC \longrightarrow AB + C$ reaction (A is yellow, B is purple, and C is navy blue). Using the “angled shot” option, try launching the *A* atom with varying angles, but with more Total energy than the transition state. What happens when the *A* atom hits the *BC* molecule from different directions? Why?

Solution

The *A* atom has enough energy to react with *BC*; however, the different angles at which it bounces off of *BC* without reacting indicate that the orientation of the molecule is an important part of the reaction kinetics and determines whether a reaction will occur.

12.6: Reaction Mechanisms

Q12.6.1

Why are elementary reactions involving three or more reactants very uncommon?

Q12.6.2

In general, can we predict the effect of doubling the concentration of *A* on the rate of the overall reaction $A + B \longrightarrow C$? Can we predict the effect if the reaction is known to be an elementary reaction?

Solution

Add texts here. Do not delete this text first.

No. In general, for the overall reaction, we cannot predict the effect of changing the concentration without knowing the rate equation. Yes. If the reaction is an elementary reaction, then doubling the concentration of *A* doubles the rate.

Q12.6.3

Phosgene, $COCl_2$, one of the poison gases used during World War I, is formed from chlorine and carbon monoxide. The mechanism is thought to proceed by:

step 1:	Cl + CO → COCl
step 2:	COCl + Cl_2 → $COCl_2$ + Cl

a. Write the overall reaction equation.
b. Identify any reaction intermediates.
c. Identify any intermediates.

Q12.6.4

Define these terms:

a. unimolecular reaction
b. bimolecular reaction
c. elementary reaction
d. overall reaction

Q12.6.5

What is the rate equation for the elementary termolecular reaction $A + 2B \longrightarrow$ products ? For $3A \longrightarrow$ products?

Solution

Add texts here. Do not delete this text first.

We are given that both of these reactions are elementary termolecular. The molecularity of a reaction refers to the number of reactant particles that react together with the proper and energy and orientation. Termolecular reactions have three atoms to collide simultaneously. As it is termolecular, and there are no additional reactants aside from the three given in each reaction, there are no intermediate reactions. The rate law for elementary reactions is determined by the stoichiometry of the reaction without needed experimental data.

The basic rate form for the elementary step is what follows:

$rate = k \cdot reactant\ 1^{i} \cdot reactant\ 2^{ii} \cdot \ldots$ Where i and ii are the stochiometric coefficient from reactant 1 and 2 respectively.

For: $3A \rightarrow products$

$k \cdot A^3 = rate$

For: $A + 2B \rightarrow products$

$k \cdot [A] \cdot [B]^2 = rate$

Note that the order of these reactions are both three.

Answer

Add texts here. Do not delete this text first.

Rate = $k[A][B]^2$; Rate = $k[A]^3$

Q12.6.6

Given the following reactions and the corresponding rate laws, in which of the reactions might the elementary reaction and the overall reaction be the same?

(a) $\mathrm{Cl_2 + CO \longrightarrow Cl_2CO}$

$\text{rate} = k[\mathrm{Cl_2}]^{\frac{3}{2}}[\mathrm{CO}]$

(b) $\mathrm{PCl_3 + Cl_2 \longrightarrow PCl_5}$

$\text{rate} = k[\mathrm{PCl_3}][\mathrm{Cl_2}]$

(c) $\mathrm{2\,NO + H_2 \longrightarrow N_2 + H_2O}$

$\text{rate} = k[\mathrm{NO}][\mathrm{H_2}]$

(d) $\mathrm{2\,NO + O_2 \longrightarrow 2\,NO_2}$

$\text{rate} = k[\mathrm{NO}]^2[\mathrm{O_2}]$

(e) $\mathrm{NO + O_3 \longrightarrow NO_2 + O_2}$

$\text{rate} = k[\mathrm{NO}][\mathrm{O_3}]$

Solution

Add texts here. Do not delete this text first.

An elementary reaction is a chemical reaction in which the reactants directly form products in a single step. In another words, the rate law for the overall reaction is same as experimentally found rate law. Out of 5 options, option (b),(d), and (e) are such reactions

Q12.6.7

Write the rate equation for each of the following elementary reactions:

a. $\mathrm{O_3 \xrightarrow{sunlight} O_2 + O}$
b. $\mathrm{O_3 + Cl \longrightarrow O_2 + ClO}$
c. $\mathrm{ClO + O \longrightarrow Cl + O_2}$
d. $\mathrm{O_3 + NO \longrightarrow NO_2 + O_2}$
e. $\mathrm{NO_2 + O \longrightarrow NO + O_2}$

Solution

Add texts here. Do not delete this text first.

Rate equations are **dependent on the reactants and not the products.**

The rate law of a reaction can be found using a rate constant (which is found experimentally), and the initial concentrations of reactants.

A general solution for the equation

$aA + bB \rightarrow cC + dD$

is $rate = k[A]^m[B]^n$ where m and n are reaction orders.

However, reaction orders are found experimentally, and since we do not have experimental data for these reactions, we can disregard that part of the equation.

To find the rate laws, all we have to do is plug the reactants into the rate formula. This is only due to the case that these are elementary reactions. Further reading on elementary reactions can be found on Libre Texts.

a. $\mathrm{O_3 \longrightarrow O_2 + O}$

To write this reaction's rate equation, only focus on the reactant(s) and its/their concentration and multiplying that by a rate constant, "k".

$$\text{Rate} = k[O_3]$$

b. $O_3 + Cl \longrightarrow O_2 + ClO$

To write this reaction's rate equation, only focus on the reactant(s) and its/their concentration and multiplying that by a rate constant, "k".

$$\text{Rate} = k[O_3][Cl]$$

c. $ClO + O \longrightarrow Cl + O_2$

To write this reaction's rate equation, only focus on the reactant(s) and its/their concentration and multiplying that by a rate constant, "k".

$$\text{Rate} = k[ClO][O]$$

d. $O_3 + NO \longrightarrow NO_2 + O_2$

To write this reaction's rate equation, only focus on the reactant(s) and its/their concentration and multiplying that by a rate constant, "k".

$$\text{Rate} = k[O_3][NO]$$

e. $NO_2 + O \longrightarrow NO + O_2$

To write this reaction's rate equation, only focus on the reactant(s) and its/their concentration and multiplying that by a rate constant, "k".

$$\text{Rate} = k[NO_2][O]$$

Answer

Add texts here. Do not delete this text first.

(a) $\text{Rate}_1 = k[O_3]$; (b) $\text{Rate}_2 = k[O_3][Cl]$; (c) $\text{Rate}_3 = k[ClO][O]$; (d) $\text{Rate}_2 = k[O_3][NO]$; (e) $\text{Rate}_3 = k[NO_2][O]$

Q12.6.8

Nitrogen(II) oxide, NO, reacts with hydrogen, H_2, according to the following equation:

$$2\,NO + 2\,H_2 \longrightarrow N_2 + 2\,H_2O$$

What would the rate law be if the mechanism for this reaction were:

$$2\,NO + H_2 \longrightarrow N_2 + H_2O_2 \text{ (slow)}$$

$$H_2O_2 + H_2 \longrightarrow 2\,H_2O \text{ (fast)}$$

The rate law of the mechanism is determined by the slow step of the reaction. Since the slow step is an elementary step, the rate law can be drawn from the coefficients of the chemical equation. So therefore, the rate law is as follows: $\text{rate}=k[NO]^2[H_2]$. Since both NO and H_2 are reactants in the overall reaction (therefore are not intermediates in the reaction), no further steps have to be done to determine the rate law.

Q12.6.9

Consider the reaction

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ (occurs under light)

The mechanism is a chain reaction involving Cl atoms and CH_3 radicals. Which of the following steps does not terminate this chain reaction?

a. $CH_3 + Cl \rightarrow CH_3CI$
b. $CH_3 + HCl \rightarrow CH_4 + Cl$
c. $CH_3 + CH_3 \rightarrow C_2H_2$

d. Cl + Cl → Cl_2

Solution

Add texts here. Do not delete this text first.

Chain reactions involve reactions that create products necessary for more reactions to occur. In this case, a reaction step will continue the chain reaction if a radical is generated. Radicals are highly reactive particles, so more reactions in the chain will take place as long as they are present. The chlorine is considered a free radical as it has an unpaired electron; for this reason it is very reactive and propagates a chain reaction. It does so by taking an electron from a stable molecule and making that molecule reactive, and that molecule goes on to react with stable species, and in that manner a long series of "chain" reactions are initiated. A chlorine radical will continue the chain by completing the following reaction:

$$Cl\cdot + CH_4 \rightarrow CH_3\cdot + HCl$$

The CH_3 generated by this reaction can then react with other species, continuing to propagate the chain reaction.

Option 1 is incorrect because the only species it produces is CH_3Cl, a product in the overall reaction that is unreactive. This terminates the chain reaction because it fails to produce any Cl or CH_3 radicals that are necessary for further propagating the overall reaction.

Option 2 is the correct answer because it produces a Cl radical. This Cl radical can continue the chain by colliding with CH_4 molecules.

Option 3 is incorrect because it fails to produce a radical capable of continuing the chain.

Option 4 is incorrect because it produces Cl_2, a molecule that does not react unless additional light is supplied. Therefore, this step breaks the chain.

Answer

Add texts here. Do not delete this text first.

Answer: Option 2: $CH_3 + HCl \rightarrow CH_4 + Cl$

Q12.6.10

Experiments were conducted to study the rate of the reaction represented by this equation.

$$2\,\mathrm{NO}(g) + 2\,\mathrm{H_2}(g) \longrightarrow \mathrm{N_2}(g) + 2\,\mathrm{H_2O}(g)$$

Initial concentrations and rates of reaction are given here.

Experiment	Initial Concentration [NO] (mol/L)	Initial Concentration, [H_2] (mol/L)	Initial Rate of Formation of N_2 (mol/L min)
1	0.0060	0.0010	1.8×10^{-4}
2	0.0060	0.0020	3.6×10^{-4}
3	0.0010	0.0060	0.30×10^{-4}
4	0.0020	0.0060	1.2×10^{-4}

Consider the following questions:

a. Determine the order for each of the reactants, NO and H_2, from the data given and show your reasoning.
b. Write the overall rate law for the reaction.
c. Calculate the value of the rate constant, *k*, for the reaction. Include units.
d. For experiment 2, calculate the concentration of NO remaining when exactly one-half of the original amount of H_2 had been consumed.
e. The following sequence of elementary steps is a proposed mechanism for the reaction.

Step 1: $\mathrm{NO} + \mathrm{NO} \rightleftharpoons \mathrm{N_2O_2}$

Step 2: $N_2O_2 + H_2 \rightleftharpoons H_2O + N_2O$

Step 3: $N_2O + H_2 \rightleftharpoons N_2 + H_2O$

Based on the data presented, which of these is the rate determining step? Show that the mechanism is consistent with the observed rate law for the reaction and the overall stoichiometry of the reaction.

S12.6.10

1. i) Find the order for [NO] by using experiment 3 and 4 where [H_2] is constant

Notice that [NO] doubles from experiment 3 to 4 and the rate quadruples. So the order for [NO] is **2**

ii) Find the order for [H_2] by using experiment 1 and 2 where [NO] is constant

Notice that [H_2] doubles from experiment 1 to 2 and the rate doubles as well. So the order for [H_2] is **1**

2. Put in the order for each product as the exponents for the corresponding reactant.

$rate = k[NO]^2[H_2]$

3. Put in the concentrations and the rate from one of the experiments into the rate law and solve for k. (Here, experiement 1 is used but any of them will work)

$rate = k[NO]^2[H_2]\ \ .00018 = k[.006]^2[.001] k = 5000 M^{-2}s^{-1}$

4. Plug in values for experiment 2 into the rate law equation and solve for the concentration of NO

$.00036 = 5000[NO]^2[.001] [NO]^2 = 7.2x10^{-5}\ [NO] = .0085M$

5. Write the rate laws for each step and then see which matches the rate law we found in question 2. The rate determining step (the slow step) is the one that gives the rate for the overall reaction. Because of this, only those concentrations will influence the overall reaction, contrary to what we would believe if we just looked at the overall reaction.

Step 1: $NO + NO \rightleftharpoons N_2O_2$

$rate = k_1[NO]^2$ This rate law is not the same as the one we calculate in question 2 so this **can not** be the rate determining step

Step 2: $N_2O_2 + H_2 \rightleftharpoons N_2O + N_2O$

$rate = k_2[N_2O_2][H_2]$

Since N_2O_2 is an intermediate you must replace it in the rate law equation. Intermediates can not be in the rate law because they do not appear in the overall reaction. Here you can take the reverse of equation 1 (k_{-1}) and substitute the other side (the reactants of equation 1) for the intermediate in the rate law equation.

$$rate_1 = rate_{-1}$$

$$k_1[NO]^2 = k_{-1}[N_2O_2]$$

$$[N_2O_2] = \frac{k_1[NO]^2}{k_{-1}}$$

$rate = \frac{k_2 k_1 [NO]^2 [H_2]}{k_{-1}}$

Overall: $rate = k[NO]^2[H_2]$ This **is the same** so it is the rate determining step.

So $N_2O_2 + H_2 \rightleftharpoons N_2O + N_2O$ is the rate determining step step.

Answer

Add texts here. Do not delete this text first.

(a) NO: 2

H_2 : 1

(b) Rate = k [NO]^2[H_2];

https://chem.libretexts.org/@go/page/479000

(c) $k = 5.0 \times 10^3\ \text{mol}^{-2}\ \text{L}^{-2}\ \text{min}^{-1}$;

(d) 0.0050 mol/L;

(e) Step II is the rate-determining step.

Q12.6.11

The reaction of CO with Cl_2 gives phosgene ($COCl_2$), a nerve gas that was used in World War I. Use the mechanism shown here to complete the following exercises:

- $\text{Cl}_2(g) \rightleftharpoons 2\,\text{Cl}(g)$ (fast, k_1 represents the forward rate constant, k_{-1} the reverse rate constant)
- $\text{CO}(g) + \text{Cl}(g) \longrightarrow \text{COCl}(g)$ (slow, k_2 the rate constant)
- $\text{COCl}(g) + \text{Cl}(g) \longrightarrow \text{COCl}_2(g)$ (fast, k_3 the rate constant)

a. Write the overall reaction.
b. Identify all intermediates.
c. Write the rate law for each elementary reaction.
d. Write the overall rate law expression.

Solution

Add texts here. Do not delete this text first.

1. To write the overall reaction you have to identify the intermediates and leave them out. The easiest way to do this is to write out all the products and reactants and cross out anything that is on both sides.

$$Cl_2(g) + CO(g) + \cancel{2Cl(g)} + \cancel{COCl(g)} \Rightarrow \cancel{2Cl(g)} + \cancel{COCl(g)} + COCl_2(g)$$

In this you will cross out the 2Cl(g) molecules and the COCl(g). What is left after that is the overall reaction.

$Cl_2(g) + CO(g) \Rightarrow + COCl_2(g)$

2. For part two you will just list the intermediates that you crossed out.

Cl and COCl are intermediates

3. Each rate law will be the rate equal to the rate constant times the concentrations of the reactants

reaction 1

(forward) rate=$k_1[Cl_2]$ (reverse) rate=$k_{-1}[Cl]$

reaction 2

rate=$k_2[CO][Cl]$

Reaction 3

rate=$k_3[COCl][Cl]$

4. The overall rate law is based off the slowest step (step #2), since it is the rate determining step, but Cl is present in that rate law so we have to replace it with an equivalent that does not contain an intermediate. To do this you use the equilibrium since the rates are the same you can set up the rate laws of the forward and reverse equal to each other.

$k_1[Cl_2] = k_{-1}[Cl]$

$[Cl] = k_1[Cl_2]/k_{-1}$

rate=$k_2[CO]k_1[Cl_2]/k_{-1}$

rate=$k^{\circ}[CO][Cl_2]$

Steps to replacing and intermediate

1. Set the forward and reverse reaction equal to each other using separate constants
2. Solve for the intermediate using algebra
3. Plug into the rate determining formula

4. All the k's will be condensed into a K prime constant

12.7: Catalysis

Q12.7.1

Account for the increase in reaction rate brought about by a catalyst.

Q12.7.2

Compare the functions of homogeneous and heterogeneous catalysts.

Q12.7.3

Consider this scenario and answer the following questions: Chlorine atoms resulting from decomposition of chlorofluoromethanes, such as CCl_2F_2, catalyze the decomposition of ozone in the atmosphere. One simplified mechanism for the decomposition is:

$$\begin{aligned} O_3 &\xrightarrow{\text{sunlight}} O_2 + O \\ O_3 + Cl &\longrightarrow O_2 + ClO \\ ClO + O &\longrightarrow Cl + O_2 \end{aligned}$$

a. Explain why chlorine atoms are catalysts in the gas-phase transformation:

$$2\,O_3 \longrightarrow 3\,O_2$$

b. Nitric oxide is also involved in the decomposition of ozone by the mechanism:

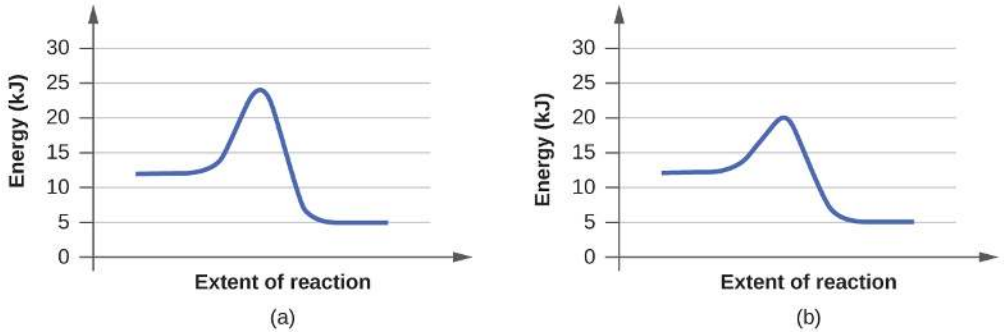
$$\begin{aligned} O_3 &\xrightarrow{\text{sunlight}} O_2 + O \\ O_3 + NO &\longrightarrow NO_2 + O_2 \\ NO_2 + O &\longrightarrow NO + O_2 \end{aligned}$$

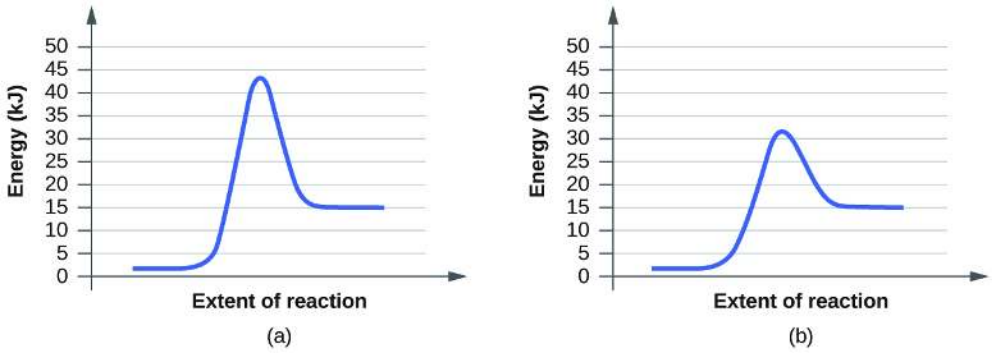
Is NO a catalyst for the decomposition? Explain your answer.

Q12.7.4

For each of the following pairs of reaction diagrams, identify which of the pair is catalyzed:

(a)

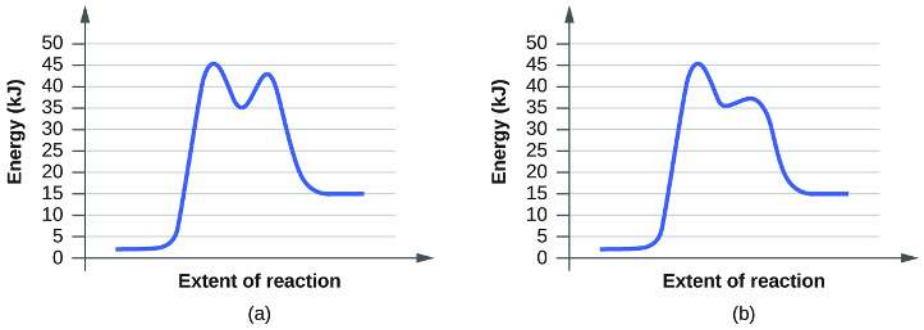




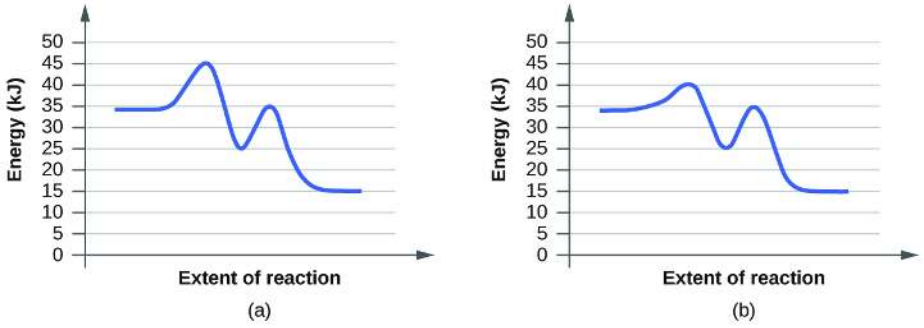
Q12.7.5

For each of the following pairs of reaction diagrams, identify which of the pairs is catalyzed:

(a)



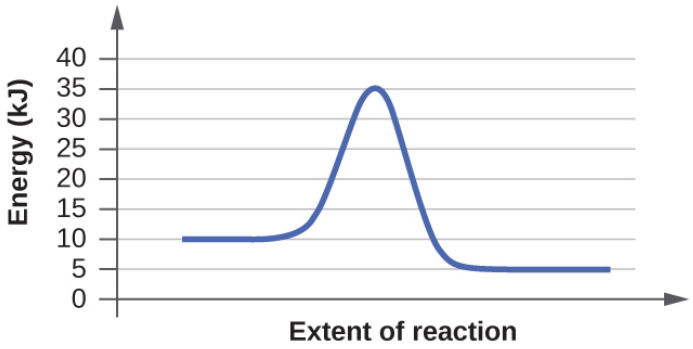
(b)



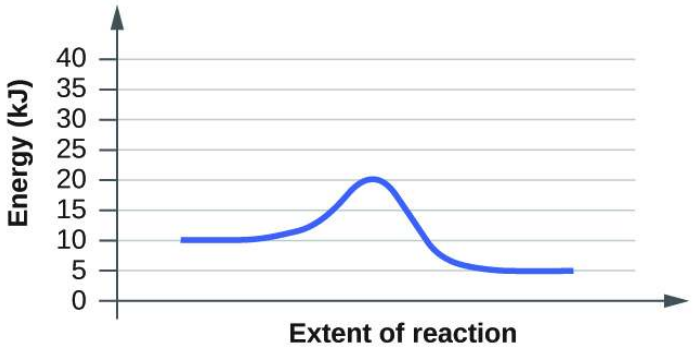
Q12.7.6

For each of the following reaction diagrams, estimate the activation energy (E_a) of the reaction:

(a)



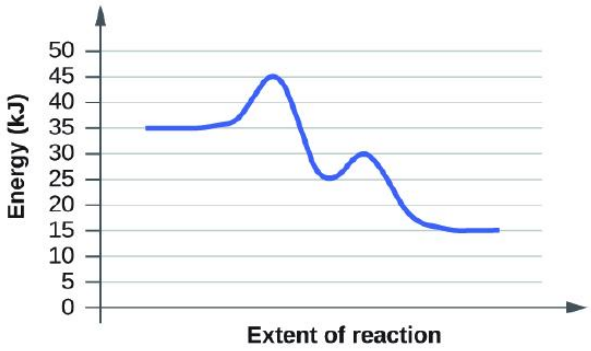
(b)



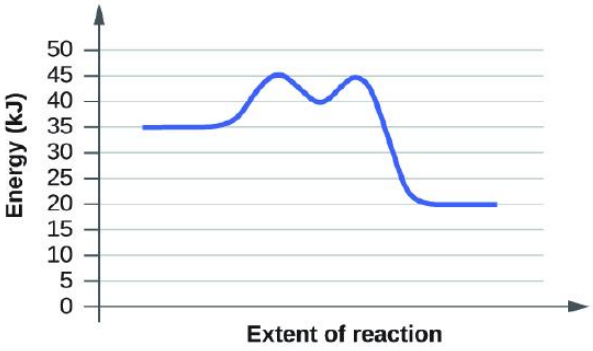
Q12.7.7

For each of the following reaction diagrams, estimate the activation energy (E_a) of the reaction:

(a)



(b)



Q12.7.8

1. Based on the diagrams in Question Q12.7.6, which of the reactions has the fastest rate? Which has the slowest rate?
2. Based on the diagrams in Question Q12.7.7, which of the reactions has the fastest rate? Which has the slowest rate?

This page titled 12.12: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

13: Solutions and their Physical Properties

13.1: Types of Solutions - Some Terminology
13.2: Solution Concentration
13.3: Intermolecular Forces and the Solution Process
13.4: Solution Formation and Equilibrium
13.5: Solubilities of Gases
13.6: Vapor Pressures of Solutions
13.7: Osmotic Pressure
13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions
13.9: Solutions of Electrolytes
13.10: Colloidal Mixtures
13.11: Exercises

This page titled 13: Solutions and their Physical Properties is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

13.1: Types of Solutions - Some Terminology

Learning Objectives

- To understand how enthalpy and entropy changes affect solution formation.
- To use the magnitude of the changes in both enthalpy and entropy to predict whether a given solute–solvent combination will spontaneously form a solution.

In all solutions, whether gaseous, liquid, or solid, the substance present in the greatest amount is the solvent, and the substance or substances present in lesser amounts are the solute(s). The solute does not have to be in the same physical state as the solvent, but the physical state of the solvent usually determines the state of the solution. As long as the solute and solvent combine to give a homogeneous solution, the solute is said to be soluble in the solvent. Table $13.1.1$ lists some common examples of gaseous, liquid, and solid solutions and identifies the physical states of the solute and solvent in each.

Table 13.1.1: Types of Solutions

Solution	Solute	Solvent	Examples
gas	gas	gas	air, natural gas
liquid	gas	liquid	seltzer water (CO_2 gas in water)
liquid	liquid	liquid	alcoholic beverage (ethanol in water), gasoline
liquid	solid	liquid	tea, salt water
solid	gas	solid	H_2 in Pd (used for H_2 storage)
solid	solid	liquid	mercury in silver or gold (amalgam often used in dentistry)

The formation of a solution from a solute and a solvent is a physical process, not a chemical one. That is, both solute and solvent can be recovered in chemically unchanged forms using appropriate separation methods. For example, solid zinc nitrate dissolves in water to form an aqueous solution of zinc nitrate:

$$Zn(NO_3)_2(s) + H_2O(l) \rightarrow Zn^{2+}(aq) + 2\,NO_3^-(aq)$$

Because $Zn(NO_3)_2$ can be recovered easily by evaporating the water, this is a physical process. In contrast, metallic zinc appears to dissolve in aqueous hydrochloric acid. In fact, the two substances undergo a chemical reaction to form an aqueous solution of zinc chloride with evolution of hydrogen gas:

$$Zn(s) + 2\,H^+(aq) + 2\,Cl^-(aq) \rightarrow Zn^{2+}(aq) + 2\,Cl^-(aq) + H_2(g)$$

> *Dissolution of a solute in a solvent to form a solution typically does not involve a chemical transformation and largely consist of breaking intermolecular forces rather than covalent bonds.*

When the solution evaporates, we do not recover metallic zinc, so we cannot say that metallic zinc is soluble in aqueous hydrochloric acid because it is chemically transformed when it dissolves. The dissolution of a solute in a solvent to form a solution does not involve a chemical transformation.

Mercury Amalgams: Examples of solid Solutions

An amalgam is an alloy of mercury with another metal. It may be a liquid, a soft paste or a solid, depending upon the proportion of mercury. These alloys are formed through metallic bonding, with the electrostatic attractive force of the conduction electrons working to bind all the positively charged metal ions together into a crystal lattice structure. Almost all metals can form amalgams with mercury including aluminum (Video $13.1.1$), the notable exceptions being iron, platinum,

tungsten, and tantalum. Silver-mercury amalgams are important in dentistry, and gold-mercury amalgam is used in the extraction of gold from ore.



Video 13.1.1: Aluminum dissolving into Mercury.

Substances that form a single homogeneous phase in all proportions are said to be completely **miscible** in one another. Ethanol and water are miscible, just as mixtures of gases are miscible. If two substances are essentially insoluble in each other, such as oil and water, they are immiscible. Examples of gaseous solutions that we have already discussed include Earth's atmosphere.

This page titled 13.1: Types of Solutions - Some Terminology is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.1: Types of Solutions - Some Terminology** is licensed CC BY-NC-SA 4.0.

13.2: Solution Concentration

Learning Objectives

- To describe the concentration of a solution in the way that is most appropriate for a particular problem or application.
- To be familiar with the different units used to express the concentrations of a solution.

There are several different ways to quantitatively describe the concentration of a solution. For example, molarity is a useful way to describe solution concentrations for reactions that are carried out in solution. Mole fractions are used not only to describe gas concentrations but also to determine the vapor pressures of mixtures of similar liquids. Example 13.2.1 reviews the methods for calculating the molarity and mole fraction of a solution when the masses of its components are known.

Example 13.2.1: Molarity and Mole Fraction

Commercial vinegar is essentially a solution of acetic acid in water. A bottle of vinegar has 3.78 g of acetic acid per 100.0 g of solution. Assume that the density of the solution is 1.00 g/mL.

a. What is its molarity?
b. What is its mole fraction?

Given: mass of substance and mass and density of solution

Asked for: molarity and mole fraction

Strategy:

A. Calculate the number of moles of acetic acid in the sample. Then calculate the number of liters of solution from its mass and density. Use these results to determine the molarity of the solution.
B. Determine the mass of the water in the sample and calculate the number of moles of water. Then determine the mole fraction of acetic acid by dividing the number of moles of acetic acid by the total number of moles of substances in the sample.

Solution:

A: The molarity is the number of moles of acetic acid per liter of solution. We can calculate the number of moles of acetic acid as its mass divided by its molar mass.

$$\text{moles } CH_3CO_2H = \frac{3.78 \text{ g } CH_3CO_2H}{60.05 \text{ g/mol}}$$

$$= 0.0629 \text{ mol}$$

The volume of the solution equals its mass divided by its density.

$$\text{volume} = \frac{\text{mass}}{\text{density}}$$

$$= \frac{100.0 \text{ g solution}}{1.00 \text{ g/mL}} = 100 \text{ } mL$$

Then calculate the molarity directly.

$$\text{molarity of } CH_3CO_2H = \frac{\text{moles } CH_3CO_2H}{\text{liter solution}}$$

$$= \frac{0.0629 \text{ } mol \text{ } CH_3CO_2H}{(100 \text{ mL})(1 \text{ } L/1000 \text{ mL})} = 0.629 \text{ } M \text{ } CH_3CO_2H$$

This result makes intuitive sense. If 100.0 g of aqueous solution (equal to 100 mL) contains 3.78 g of acetic acid, then 1 L of solution will contain 37.8 g of acetic acid, which is a little more than $\frac{1}{2}$ mole. Keep in mind, though, that the mass and volume of a solution are related by its density; concentrated aqueous solutions often have densities greater than 1.00 g/mL.

B: To calculate the mole fraction of acetic acid in the solution, we need to know the number of moles of both acetic acid and water. The number of moles of acetic acid is 0.0629 mol, as calculated in part (a). We know that 100.0 g of vinegar contains 3.78 g of acetic acid; hence the solution also contains (100.0 g − 3.78 g) = 96.2 g of water. We have

$$moles\ \mathrm{H_2O} = \frac{96.2\ \cancel{g}\ \mathrm{H_2O}}{18.02\ \cancel{g}/mol} = 5.34\ mol\ \mathrm{H_2O}$$

The mole fraction χ of acetic acid is the ratio of the number of moles of acetic acid to the total number of moles of substances present:

$$\begin{aligned}\chi_{\mathrm{CH_3CO_2H}} &= \frac{moles\ \mathrm{CH_3CO_2H}}{moles\ \mathrm{CH_3CO_2H} + moles\ \mathrm{H_2O}} \\ &= \frac{0.0629\ mol}{0.0629\ mol + 5.34\ mol} \\ &= 0.0116 = 1.16 \times 10^{-2}\end{aligned}$$

This answer makes sense, too. There are approximately 100 times as many moles of water as moles of acetic acid, so the ratio should be approximately 0.01.

? Exercise 13.2.1: Molarity and Mole Fraction

A solution of HCl gas dissolved in water (sold commercially as “muriatic acid,” a solution used to clean masonry surfaces) has 20.22 g of HCl per 100.0 g of solution, and its density is 1.10 g/mL.

a. What is its molarity?
b. What is its mole fraction?

Answer a

6.10 M HCl

Answer b

$\chi_{HCl} = 0.111$

The concentration of a solution can also be described by its molality (m), the number of moles of solute per kilogram of solvent:

$$\text{molality (m)} = \frac{\text{moles solute}}{\text{kilogram solvent}} \tag{13.2.1}$$

Molality, therefore, has the same numerator as molarity (the number of moles of solute) but a different denominator (kilogram of solvent rather than liter of solution). For dilute aqueous solutions, the molality and molarity are nearly the same because dilute solutions are mostly solvent. Thus because the density of water under standard conditions is very close to 1.0 g/mL, the volume of 1.0 kg of H_2O under these conditions is very close to 1.0 L, and a 0.50 M solution of KBr in water, for example, has approximately the same concentration as a 0.50 m solution.

Another common way of describing concentration is as the ratio of the mass of the solute to the total mass of the solution. The result can be expressed as mass percentage, parts per million (ppm), or parts per billion (ppb):

$$\text{mass percentage} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100 \tag{13.2.2}$$

$$\text{parts per million (ppm)} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 \tag{13.2.3}$$

$$\text{parts per billion (ppb)} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9 \tag{13.2.4}$$

In the health sciences, the concentration of a solution is often expressed as parts per thousand (ppt), indicated as a proportion. For example, adrenalin, the hormone produced in high-stress situations, is available in a 1:1000 solution, or one gram of adrenalin per

1000 g of solution.

The labels on bottles of commercial reagents often describe the contents in terms of mass percentage. Sulfuric acid, for example, is sold as a 95% aqueous solution, or 95 g of H_2SO_4 per 100 g of solution. Parts per million and parts per billion are used to describe concentrations of highly dilute solutions. These measurements correspond to milligrams and micrograms of solute per kilogram of solution, respectively. For dilute aqueous solutions, this is equal to milligrams and micrograms of solute per liter of solution (assuming a density of 1.0 g/mL).

✔ Example 13.2.2: Molarity and Mass

Several years ago, millions of bottles of mineral water were contaminated with benzene at ppm levels. This incident received a great deal of attention because the lethal concentration of benzene in rats is 3.8 ppm. A 250 mL sample of mineral water has 12.7 ppm of benzene. Because the contaminated mineral water is a very dilute aqueous solution, we can assume that its density is approximately 1.00 g/mL.

a. What is the molarity of the solution?
b. What is the mass of benzene in the sample?

Given: volume of sample, solute concentration, and density of solution

Asked for: molarity of solute and mass of solute in 250 mL

Strategy:

A. Use the concentration of the solute in parts per million to calculate the molarity.
B. Use the concentration of the solute in parts per million to calculate the mass of the solute in the specified volume of solution.

Solution:

a. A To calculate the molarity of benzene, we need to determine the number of moles of benzene in 1 L of solution. We know that the solution contains 12.7 ppm of benzene. Because 12.7 ppm is equivalent to 12.7 mg/1000 g of solution and the density of the solution is 1.00 g/mL, the solution contains 12.7 mg of benzene per liter (1000 mL). The molarity is therefore

$$\begin{aligned}\text{molarity} &= \frac{\text{moles}}{\text{liter solution}} \\ &= \frac{(12.7\ \cancel{mg})\left(\frac{1\ \cancel{g}}{1000\ \cancel{mg}}\right)\left(\frac{1\ mol}{78.114\ \cancel{g}}\right)}{1.00\ L} \\ &= 1.63 \times 10^{-4} M\end{aligned}$$

b. B We are given that there are 12.7 mg of benzene per 1000 g of solution, which is equal to 12.7 mg/L of solution. Hence the mass of benzene in 250 mL (250 g) of solution is

$$\begin{aligned}\text{mass of benzene} &= \frac{(12.7\ mg\ \text{benzene})(250\ \cancel{mL})}{1000\ \cancel{mL}} \\ &= 3.18\ mg \\ &= 3.18 \times 10^{-3}\ g\ \text{benzene}\end{aligned}$$

? Exercise 13.2.2: Molarity of Lead Solution

The maximum allowable concentration of lead in drinking water is 9.0 ppb.

a. What is the molarity of Pb^{2+} in a 9.0 ppb aqueous solution?
b. Use your calculated concentration to determine how many grams of Pb^{2+} are in an 8 oz glass of water.

Answer a

4.3×10^{-8} M

Answer b

2×10^{-6} g

How do chemists decide which units of concentration to use for a particular application? Although molarity is commonly used to express concentrations for reactions in solution or for titrations, it does have one drawback—molarity is the number of moles of solute divided by the volume of the solution, and the volume of a solution depends on its density, which is a function of temperature. Because volumetric glassware is calibrated at a particular temperature, typically 20°C, the molarity may differ from the original value by several percent if a solution is prepared or used at a significantly different temperature, such as 40°C or 0°C. For many applications this may not be a problem, but for precise work these errors can become important. In contrast, mole fraction, molality, and mass percentage depend on only the masses of the solute and solvent, which are independent of temperature.

Mole fraction is not very useful for experiments that involve quantitative reactions, but it is convenient for calculating the partial pressure of gases in mixtures, as discussed previously. Mole fractions are also useful for calculating the vapor pressures of certain types of solutions. Molality is particularly useful for determining how properties such as the freezing or boiling point of a solution vary with solute concentration. Because mass percentage and parts per million or billion are simply different ways of expressing the ratio of the mass of a solute to the mass of the solution, they enable us to express the concentration of a substance even when the molecular mass of the substance is unknown. Units of ppb or ppm are also used to express very low concentrations, such as those of residual impurities in foods or of pollutants in environmental studies.

Table 13.2.1 summarizes the different units of concentration and typical applications for each. When the molar mass of the solute and the density of the solution are known, it becomes relatively easy with practice to convert among the units of concentration we have discussed, as illustrated in Example 13.2.3

Table 13.2.1: Different Units for Expressing the Concentrations of Solutions*

Unit	Definition	Application
molarity (M)	moles of solute/liter of solution (mol/L)	Used for quantitative reactions in solution and titrations; mass and molecular mass of solute and volume of solution are known.
mole fraction (χ)	moles of solute/total moles present (mol/mol)	Used for partial pressures of gases and vapor pressures of some solutions; mass and molecular mass of each component are known.
molality (m)	moles of solute/kg of solvent (mol/kg)	Used in determining how colligative properties vary with solute concentration; masses and molecular mass of solute are known.
mass percentage (%)	[mass of solute (g)/mass of solution (g)] × 100	Useful when masses are known but molecular masses are unknown.
parts per thousand (ppt)	[mass of solute/mass of solution] × 10^3 (g solute/kg solution)	Used in the health sciences, ratio solutions are typically expressed as a proportion, such as 1:1000.
parts per million (ppm)	[mass of solute/mass of solution] × 10^6 (mg solute/kg solution)	Used for trace quantities; masses are known but molecular masses may be unknown.
parts per billion (ppb)	[mass of solute/mass of solution] × 10^9 (µg solute/kg solution)	Used for trace quantities; masses are known but molecular masses may be unknown.

*The molarity of a solution is temperature dependent, but the other units shown in this table are independent of temperature.

Example 13.2.3: Vodka

Vodka is essentially a solution of ethanol in water. Typical vodka is sold as "80 proof," which means that it contains 40.0% ethanol by volume. The density of pure ethanol is 0.789 g/mL at 20°C. If we assume that the volume of the solution is the sum of the volumes of the components (which is not strictly correct), calculate the following for the ethanol in 80-proof vodka.

a. the mass percentage
b. the mole fraction
c. the molarity and the molality

Given: volume percent and density

Asked for: mass percentage, mole fraction, molarity, and molality

Strategy:

1. Use the density of the solute to calculate the mass of the solute in 100.0 mL of solution. Calculate the mass of water in 100.0 mL of solution.
2. Determine the mass percentage of solute by dividing the mass of ethanol by the mass of the solution and multiplying by 100.
3. Convert grams of solute and solvent to moles of solute and solvent. Calculate the mole fraction of solute by dividing the moles of solute by the total number of moles of substances present in solution.
4. Calculate the molarity of the solution: moles of solute per liter of solution. Determine the molality of the solution by dividing the number of moles of solute by the kilograms of solvent.

Solution:

The key to this problem is to use the density of pure ethanol to determine the mass of ethanol (CH_3CH_2OH), abbreviated as EtOH, in a given volume of solution. We can then calculate the number of moles of ethanol and the concentration of ethanol in any of the required units. A Because we are given a percentage by volume, we assume that we have 100.0 mL of solution. The volume of ethanol will thus be 40.0% of 100.0 mL, or 40.0 mL of ethanol, and the volume of water will be 60.0% of 100.0 mL, or 60.0 mL of water. The mass of ethanol is obtained from its density:

$$mass\ of\ EtOH = (40.0\ \cancel{mL}) \left(\frac{0.789\ g}{\cancel{mL}} \right) = 31.6\ g\ EtOH$$

If we assume the density of water is 1.00 g/mL, the mass of water is 60.0 g. We now have all the information we need to calculate the concentration of ethanol in the solution.

A The mass percentage of ethanol is the ratio of the mass of ethanol to the total mass of the solution, expressed as a percentage:

$$\begin{aligned} \%EtOH &= \left(\frac{mass\ of\ EtOH}{mass\ of\ solution} \right) (100) \\ &= \left(\frac{31.6\ \cancel{g}\ EtOH}{31.6\ \cancel{g}\ EtOH + 60.0\ \cancel{g}\ H_2O} \right) (100) \\ &= 34.5\% \end{aligned}$$

B The mole fraction of ethanol is the ratio of the number of moles of ethanol to the total number of moles of substances in the solution. Because 40.0 mL of ethanol has a mass of 31.6 g, we can use the molar mass of ethanol (46.07 g/mol) to determine the number of moles of ethanol in 40.0 mL:

$$\begin{aligned} moles\ CH_3CH_2OH &= (31.6\ \ g\ \cancel{CH_3CH_2OH}) \left(\frac{1\ mol}{46.07\ \ g\ \cancel{CH_3CH_2OH}} \right) \\ &= 0.686\ mol\ CH_3CH_2OH \end{aligned}$$

Similarly, the number of moles of water is

$$moles\ H_2O = (60.0\ \ g\ H_2O)\left(\frac{1\ mol\ H_2O}{18.02\ \ g\ H_2O}\right) = 3.33\ mol\ H_2O$$

The mole fraction of ethanol is thus

$$\chi_{CH_3CH_2OH} = \frac{0.686\ \ mol}{0.686\ \ mol + 3.33\ \ mol} = 0.171$$

C and D The molarity of the solution is the number of moles of ethanol per liter of solution. We already know the number of moles of ethanol per 100.0 mL of solution, so the molarity is

$$M_{CH_3CH_2OH} = \left(\frac{0.686\ mol}{100\ \ mL}\right)\left(\frac{1000\ \ mL}{L}\right) = 6.86\ M$$

The molality of the solution is the number of moles of ethanol per kilogram of solvent. Because we know the number of moles of ethanol in 60.0 g of water, the calculation is again straightforward:

$$m_{CH_3CH_2OH} = \left(\frac{0.686\ mol\ EtOH}{60.0\ \ g\ H_2O}\right)\left(\frac{1000\ \ g}{kg}\right) = \frac{11.4\ mol\ EtOH}{kg\ H_2O} = 11.4\ m$$

? Exercise 13.2.3: Toluene/Benzene Solution

A solution is prepared by mixing 100.0 mL of toluene with 300.0 mL of benzene. The densities of toluene and benzene are 0.867 g/mL and 0.874 g/mL, respectively. Assume that the volume of the solution is the sum of the volumes of the components. Calculate the following for toluene.

a. mass percentage
b. mole fraction
c. molarity
d. molality

Answer a

mass percentage toluene = 24.8%

Answer b

$\chi_{toluene} = 0.219$

Answer c

2.35 M toluene

Answer d

3.59 m toluene



A Video Discussing Different Measures of Concentration. Video Link: Measures of Concentration, YouTube (opens in new window) [youtu.be]



A Video Discussing how to Convert Measures of Concentration. Video Link: Converting Units of Concentration, YouTube(opens in new window) [youtu.be]

Summary

Different units are used to express the concentrations of a solution depending on the application. The concentration of a solution is the quantity of solute in a given quantity of solution. It can be expressed in several ways: molarity (moles of solute per liter of solution); mole fraction, the ratio of the number of moles of solute to the total number of moles of substances present; mass percentage, the ratio of the mass of the solute to the mass of the solution times 100; parts per thousand (ppt), grams of solute per kilogram of solution; parts per million (ppm), milligrams of solute per kilogram of solution; parts per billion (ppb), micrograms of solute per kilogram of solution; and molality (m), the number of moles of solute per kilogram of solvent.

This page titled 13.2: Solution Concentration is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.4: Ways of Expressing Concentration** by Anonymous is licensed CC BY-NC-SA 3.0.
- **13.2: Solution Concentration** is licensed CC BY-NC-SA 4.0.

13.3: Intermolecular Forces and the Solution Process

Learning Objectives

- To understand the relationship between solubility and molecular structure.
- To demonstrate how the strength of intramolecular bonding determines the solubility of a solute in a given solvent.

The interactions that determine the solubility of a substance in a liquid depend largely on the chemical nature of the solute (such as whether it is ionic or molecular) rather than on its physical state (solid, liquid, or gas). We will first describe the general case of forming a solution of a molecular species in a liquid solvent and then describe the formation of a solution of an ionic compound.

The Role of Enthalpy in Solution Formation

Energy is required to overcome the intermolecular interactions in a solute, which can be supplied only by the new interactions that occur in the solution, when each solute particle is surrounded by particles of the solvent in a process called **solvation**, or hydration when the solvent is water. Thus all of the solute–solute interactions and many of the solvent–solvent interactions must be disrupted for a solution to form. In this section, we describe the role of enthalpy in this process.

Because enthalpy is a state function, we can use a thermochemical cycle to analyze the energetics of solution formation. The process occurs in three discrete steps, indicated by ΔH_1, ΔH_2, and ΔH_3 in Figure 13.3.1. The overall enthalpy change in the formation of the solution (ΔH_{soln}) is the sum of the enthalpy changes in the three steps:

$$\Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3 \tag{13.3.1}$$

When a solvent is added to a solution, steps 1 and 2 are both endothermic because energy is required to overcome the intermolecular interactions in the solvent (ΔH_1) and the solute (ΔH_2). Because ΔH is positive for both steps 1 and 2, the solute–solvent interactions (ΔH_3) must be stronger than the solute–solute and solvent–solvent interactions they replace in order for the dissolution process to be exothermic ($\Delta H_{soln} < 0$). When the solute is an ionic solid, ΔH_2 corresponds to the lattice energy that must be overcome to form a solution. The higher the charge of the ions in an ionic solid, the higher the lattice energy. Consequently, solids that have very high lattice energies, such as MgO (−3791 kJ/mol), are generally insoluble in all solvents.

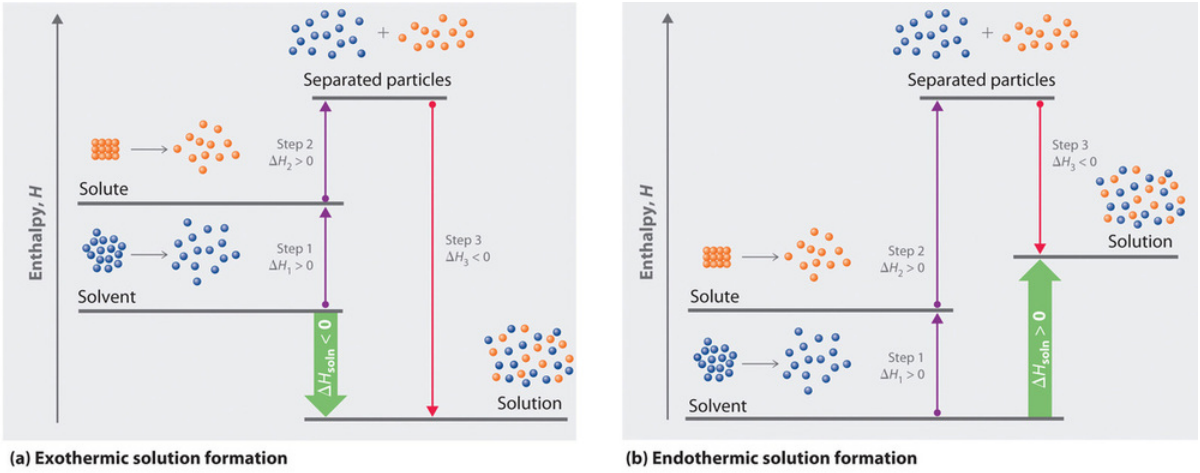


Figure 13.3.1: Enthalpy Changes That Accompany the Formation of a Solution. Solvation can be an exothermic or endothermic process depending on the nature of the solute and solvent. In both cases, step 1, separation of the solvent particles, is energetically uphill ($\Delta H_1 > 0$), as is step 2, separation of the solute particles ($\Delta H_2 > 0$). In contrast, energy is released in step 3 ($\Delta H_3 < 0$) because of interactions between the solute and solvent. (a) When ΔH_3 is larger in magnitude than the sum of ΔH_1 and ΔH_2, the overall process is exothermic ($\Delta H_{soln} < 0$), as shown in the thermochemical cycle. (b) When ΔH_3 is smaller in magnitude than the sum of ΔH_1 and ΔH_2, the overall process is endothermic ($\Delta H_{soln} > 0$).

A positive value for ΔH_{soln} does not mean that a solution will not form. Whether a given process, including formation of a solution, occurs spontaneously depends on whether the total energy of the system is lowered as a result. Enthalpy is only one of the contributing factors. A high ΔH_{soln} is usually an indication that the substance is not very soluble. Instant cold packs used to treat athletic injuries, for example, take advantage of the large positive ΔH_{soln} of ammonium nitrate during dissolution (+25.7 kJ/mol), which produces temperatures less than 0°C (Figure 13.3.2).



Figure 13.3.2: Commercial Cold Packs for Treating Injuries. These packs contain solid NH_4NO_3 and water in separate compartments. When the seal between the compartments is broken, the NH_4NO_3 dissolves in the water. Because ΔH_{soln} for NH_4NO_3 is much greater than zero, heat is absorbed by the cold pack during the dissolution process, producing local temperatures less than 0°C.

Solutions of Molecular Substances in Liquids

The London dispersion forces, dipole–dipole interactions, and hydrogen bonds that hold molecules to other molecules are generally weak. Even so, energy is required to disrupt these interactions. Unless some of that energy is recovered in the formation of new, favorable solute–solvent interactions, the increase in entropy on solution formation is not enough for a solution to form.

For solutions of gases in liquids, we can safely ignore the energy required to separate the solute molecules ($\Delta H_2 = 0$) because the molecules are already separated. Thus we need to consider only the energy required to separate the solvent molecules (ΔH_1) and the energy released by new solute–solvent interactions (ΔH_3). Nonpolar gases such as N_2, O_2, and Ar have no dipole moment and cannot engage in dipole–dipole interactions or hydrogen bonding. Consequently, the only way they can interact with a solvent is by means of London dispersion forces, which may be weaker than the solvent–solvent interactions in a polar solvent. It is not surprising, then, that nonpolar gases are most soluble in nonpolar solvents. In this case, ΔH_1 and ΔH_3 are both small and of similar magnitude. In contrast, for a solution of a nonpolar gas in a polar solvent, ΔH_1 is far greater than ΔH_3. As a result, nonpolar gases are less soluble in polar solvents than in nonpolar solvents. For example, the concentration of N_2 in a saturated solution of N_2 in water, a polar solvent, is only $7.07 \times 10^{-4}\ M$ compared with $4.5 \times 10^{-3}\ M$ for a saturated solution of N_2 in benzene, a nonpolar solvent.

The solubilities of nonpolar gases in water generally increase as the molecular mass of the gas increases, as shown in Table 13.3.1. This is precisely the trend expected: as the gas molecules become larger, the strength of the solvent–solute interactions due to London dispersion forces increases, approaching the strength of the solvent–solvent interactions.

Table 13.3.1: Solubilities of Selected Gases in Water at 20°C and 1 atm Pressure

Gas	Solubility (M) × 10^{-4}
He	3.90
Ne	4.65
Ar	15.2
Kr	27.9
Xe	50.2

Gas	Solubility (M) × 10^{-4}
H_2	8.06
N_2	7.07
CO	10.6
O_2	13.9
N_2O	281
CH_4	15.5

Virtually all common organic liquids, whether polar or not, are miscible. The strengths of the intermolecular attractions are comparable; thus the enthalpy of solution is expected to be small ($\Delta H_{soln} \approx 0$), and the increase in entropy drives the formation of a solution. If the predominant intermolecular interactions in two liquids are very different from one another, however, they may be immiscible. For example, organic liquids such as benzene, hexane, CCl_4, and CS_2 (S=C=S) are nonpolar and have no ability to act as hydrogen bond donors or acceptors with hydrogen-bonding solvents such as H_2O, HF, and NH_3; hence they are immiscible in these solvents. When shaken with water, they form separate phases or layers separated by an interface (Figure 13.3.3), the region between the two layers. Just because two liquids are immiscible, however, does not mean that they are completely insoluble in each other. For example, 188 mg of benzene dissolves in 100 mL of water at 23.5°C. Adding more benzene results in the separation of an upper layer consisting of benzene with a small amount of dissolved water (the solubility of water in benzene is only 178 mg/100 mL of benzene).



Figure 13.3.3: Immiscible Liquids. Separatory funnel demonstrating the separation of oil and colored water. from Wikipedia

The solubilities of simple alcohols in water are given in Table 13.3.2

Table 13.3.2: Solubilities of Straight-Chain Organic Alcohols in Water at 20°C

Alcohol	Solubility (mol/100 g of H_2O)
methanol	completely miscible
ethanol	completely miscible
n-propanol	completely miscible
n-butanol	0.11
n-pentanol	0.030
n-hexanol	0.0058
n-heptanol	0.0008

Only the three lightest alcohols (methanol, ethanol, and n-propanol) are completely miscible with water. As the molecular mass of the alcohol increases, so does the proportion of hydrocarbon in the molecule. Correspondingly, the importance of hydrogen bonding and dipole–dipole interactions in the pure alcohol decreases, while the importance of London dispersion forces increases, which leads to progressively fewer favorable electrostatic interactions with water. Organic liquids such as acetone, ethanol, and

tetrahydrofuran are sufficiently polar to be completely miscible with water yet sufficiently nonpolar to be completely miscible with all organic solvents.



Tetrahydrofuran (THF)

The same principles govern the solubilities of molecular solids in liquids. For example, elemental sulfur is a solid consisting of cyclic S_8 molecules that have no dipole moment. Because the S_8 rings in solid sulfur are held to other rings by London dispersion forces, elemental sulfur is insoluble in water. It is, however, soluble in nonpolar solvents that have comparable London dispersion forces, such as CS_2 (23 g/100 mL). In contrast, glucose contains five –OH groups that can form hydrogen bonds. Consequently, glucose is very soluble in water (91 g/120 mL of water) but essentially insoluble in nonpolar solvents such as benzene. The structure of one isomer of glucose is shown here.



D-Glucose

Low-molecular-mass hydrocarbons with highly electronegative and polarizable halogen atoms, such as chloroform ($CHCl_3$) and methylene chloride (CH_2Cl_2), have both significant dipole moments and relatively strong London dispersion forces. These hydrocarbons are therefore powerful solvents for a wide range of polar and nonpolar compounds. Naphthalene, which is nonpolar, and phenol (C_6H_5OH), which is polar, are very soluble in chloroform. In contrast, the solubility of ionic compounds is largely determined not by the polarity of the solvent but rather by its dielectric constant, a measure of its ability to separate ions in solution, as you will soon see.

Example 13.3.1

Identify the most important solute–solvent interactions in each solution.

a. iodine in benzene
b. aniline ($C_6H_5NH_2$) in dichloromethane (CH_2Cl_2)



Aniline

c. iodine in water

Given: components of solutions

Asked for: predominant solute–solvent interactions

Strategy:

Identify all possible intermolecular interactions for both the solute and the solvent: London dispersion forces, dipole–dipole interactions, or hydrogen bonding. Determine which is likely to be the most important factor in solution formation.

Solution

a. Benzene and $I2$ are both nonpolar molecules. The only possible attractive forces are London dispersion forces.
b. Aniline is a polar molecule with an $-\mathrm{NH_2}$ group, which can act as a hydrogen bond donor. Dichloromethane is also polar, but it has no obvious hydrogen bond acceptor. Therefore, the most important interactions between aniline and CH_2Cl_2 are likely to be London interactions.
c. Water is a highly polar molecule that engages in extensive hydrogen bonding, whereas I_2 is a nonpolar molecule that cannot act as a hydrogen bond donor or acceptor. The slight solubility of I_2 in water ($1.3 \times 10^{-3}\ mol/L$ at 25°C) is due to London dispersion forces.

? Exercise 13.3.1

Identify the most important interactions in each solution:

a. ethylene glycol ($HOCH_2CH_2OH$) in acetone
b. acetonitrile ($\mathrm{CH_3C \equiv N}$) in acetone
c. n-hexane in benzene

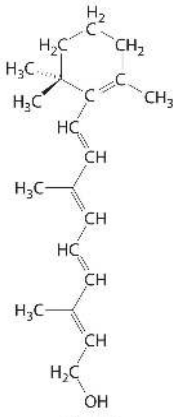
Answer:

a. hydrogen bonding
b. London interactions
c. London dispersion forces

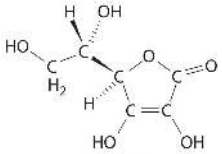
Hydrophilic and Hydrophobic Solutes

A solute can be classified as hydrophilic (literally, “water loving”), meaning that it has an electrostatic attraction to water, or hydrophobic (“water fearing”), meaning that it repels water. A hydrophilic substance is polar and often contains O–H or N–H groups that can form hydrogen bonds to water. For example, glucose with its five O–H groups is hydrophilic. In contrast, a hydrophobic substance may be polar but usually contains C–H bonds that do not interact favorably with water, as is the case with naphthalene and n-octane. Hydrophilic substances tend to be very soluble in water and other strongly polar solvents, whereas hydrophobic substances are essentially insoluble in water and soluble in nonpolar solvents such as benzene and cyclohexane.

The difference between hydrophilic and hydrophobic substances has substantial consequences in biological systems. For example, vitamins can be classified as either fat soluble or water soluble. Fat-soluble vitamins, such as vitamin A, are mostly nonpolar, hydrophobic molecules. As a result, they tend to be absorbed into fatty tissues and stored there. In contrast, water-soluble vitamins, such as vitamin C, are polar, hydrophilic molecules that circulate in the blood and intracellular fluids, which are primarily aqueous. Water-soluble vitamins are therefore excreted much more rapidly from the body and must be replenished in our daily diet. A comparison of the chemical structures of vitamin A and vitamin C quickly reveals why one is hydrophobic and the other hydrophilic.



Vitamin A
(*trans*-Retinol)

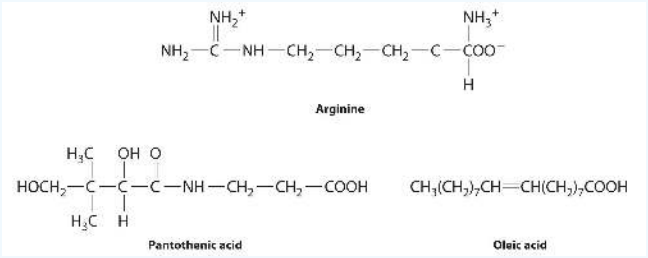


Vitamin C
(L-Ascorbic acid)

Because water-soluble vitamins are rapidly excreted, the risk of consuming them in excess is relatively small. Eating a dozen oranges a day is likely to make you tired of oranges long before you suffer any ill effects due to their high vitamin C content. In contrast, fat-soluble vitamins constitute a significant health hazard when consumed in large amounts. For example, the livers of polar bears and other large animals that live in cold climates contain large amounts of vitamin A, which have occasionally proven fatal to humans who have eaten them.

Example 13.3.2

The following substances are essential components of the human diet:



Arginine

Pantothenic acid

Oleic acid

Using what you know of hydrophilic and hydrophobic solutes, classify each as water soluble or fat soluble and predict which are likely to be required in the diet on a daily basis.

a. arginine
b. pantothenic acid
c. oleic acid

Given: chemical structures

Asked for: classification as water soluble or fat soluble; dietary requirement

Strategy:

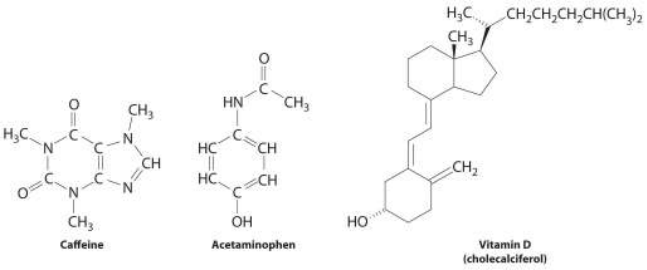
Based on the structure of each compound, decide whether it is hydrophilic or hydrophobic. If it is hydrophilic, it is likely to be required on a daily basis.

Solution:

a. Arginine is a highly polar molecule with two positively charged groups and one negatively charged group, all of which can form hydrogen bonds with water. As a result, it is hydrophilic and required in our daily diet.
b. Although pantothenic acid contains a hydrophobic hydrocarbon portion, it also contains several polar functional groups ($-OH$ and $-CO_2H$) that should interact strongly with water. It is therefore likely to be water soluble and required in the diet. (In fact, pantothenic acid is almost always a component of multiple-vitamin tablets.)
c. Oleic acid is a hydrophobic molecule with a single polar group at one end. It should be fat soluble and not required daily.

? Exercise 13.3.2

These compounds are consumed by humans: caffeine, acetaminophen, and vitamin D. Identify each as primarily hydrophilic (water soluble) or hydrophobic (fat soluble), and predict whether each is likely to be excreted from the body rapidly or slowly.



Answer: Caffeine and acetaminophen are water soluble and rapidly excreted, whereas vitamin D is fat soluble and slowly excreted

Solid Solutions

Solutions are not limited to gases and liquids; solid solutions also exist. For example, amalgams, which are usually solids, are solutions of metals in liquid mercury. Because most metals are soluble in mercury, amalgams are used in gold mining, dentistry, and many other applications. A major difficulty when mining gold is separating very small particles of pure gold from tons of crushed rock. One way to accomplish this is to agitate a suspension of the crushed rock with liquid mercury, which dissolves the gold (as well as any metallic silver that might be present). The very dense liquid gold–mercury amalgam is then isolated and the mercury distilled away.

An alloy is a solid or liquid solution that consists of one or more elements in a metallic matrix. A solid alloy has a single homogeneous phase in which the crystal structure of the solvent remains unchanged by the presence of the solute. Thus the microstructure of the alloy is uniform throughout the sample. Examples are substitutional and interstitial alloys such as brass or solder. Liquid alloys include sodium/potassium and gold/mercury. In contrast, a partial alloy solution has two or more phases that can be homogeneous in the distribution of the components, but the microstructures of the two phases are not the same. As a liquid solution of lead and tin is cooled, for example, different crystalline phases form at different cooling temperatures. Alloys usually have properties that differ from those of the component elements.

Network solids such as diamond, graphite, and $ceSiO_2$ are insoluble in all solvents with which they do not react chemically. The covalent bonds that hold the network or lattice together are simply too strong to be broken under normal conditions. They are certainly much stronger than any conceivable combination of intermolecular interactions that might occur in solution. Most metals are insoluble in virtually all solvents for the same reason: the delocalized metallic bonding is much stronger than any favorable metal atom–solvent interactions. Many metals react with solutions such as aqueous acids or bases to produce a solution. However, in these instances the metal undergoes a chemical transformation that cannot be reversed by simply removing the solvent.

Solids with very strong intermolecular bonding tend to be insoluble.

Solubilities of Ionic Substances in Liquids

Ionic substances are generally most soluble in polar solvents; the higher the lattice energy, the more polar the solvent must be to overcome the lattice energy and dissolve the substance. Because of its high polarity, water is the most common solvent for ionic compounds. Many ionic compounds are soluble in other polar solvents, however, such as liquid ammonia, liquid hydrogen fluoride,

and methanol. Because all these solvents consist of molecules that have relatively large dipole moments, they can interact favorably with the dissolved ions.

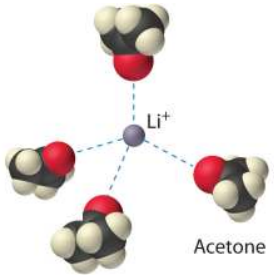


Figure 13.3.4: Ion–Dipole Interactions in the Solvation of Li+ Ions by Acetone, a Polar Solvent

The interaction of water with Na+ and Cl− ions in an aqueous solution of NaCl was illustrated in Figure 4.3. The ion–dipole interactions between Li+ ions and acetone molecules in a solution of LiCl in acetone are shown in Figure 13.3.4 The energetically favorable Li+–acetone interactions make ΔH_3 in Figure 13.1 sufficiently negative to overcome the positive ΔH_1 and ΔH_2. Because the dipole moment of acetone (2.88 D), and thus its polarity, is actually larger than that of water (1.85 D), one might even expect that LiCl would be more soluble in acetone than in water. In fact, the opposite is true: 83 g of LiCl dissolve in 100 mL of water at 20°C, but only about 4.1 g of LiCl dissolve in 100 mL of acetone. This apparent contradiction arises from the fact that the dipole moment is a property of a single molecule in the gas phase. A more useful measure of the ability of a solvent to dissolve ionic compounds is its dielectric constant (ε), which is the ability of a bulk substance to decrease the electrostatic forces between two charged particles. By definition, the dielectric constant of a vacuum is 1. In essence, a solvent with a high dielectric constant causes the charged particles to behave as if they have been moved farther apart. At 25°C, the dielectric constant of water is 80.1, one of the highest known, and that of acetone is only 21.0. Hence water is better able to decrease the electrostatic attraction between Li+ and Cl− ions, so LiCl is more soluble in water than in acetone. This behavior is in contrast to that of molecular substances, for which polarity is the dominant factor governing solubility.

> A solvent's dielectric constant is the most useful measure of its ability to dissolve ionic compounds. A solvent's polarity is the dominant factor in dissolving molecular substances.

Summary

The solubility of a substance is the maximum amount of a solute that can dissolve in a given quantity of solvent; it depends on the chemical nature of both the solute and the solvent and on the temperature and pressure. The solubility of a substance in a liquid is determined by intermolecular interactions, which also determine whether two liquids are miscible. Solutes can be classified as hydrophilic (water loving) or hydrophobic (water fearing). Vitamins with hydrophilic structures are water soluble, whereas those with hydrophobic structures are fat soluble. Many metals dissolve in liquid mercury to form amalgams. Covalent network solids and most metals are insoluble in nearly all solvents. The solubility of ionic compounds is largely determined by the dielectric constant (ε) of the solvent, a measure of its ability to decrease the electrostatic forces between charged particles. Solutions of many ionic compounds in organic solvents can be dissolved using crown ethers, cyclic polyethers large enough to accommodate a metal ion in the center, or cryptands, compounds that completely surround a cation.

This page titled 13.3: Intermolecular Forces and the Solution Process is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.3: Intermolecular Forces and the Solution Process** is licensed CC BY-NC-SA 4.0.

13.4: Solution Formation and Equilibrium

Learning Objectives

- To understand the relationship among temperature, pressure, and solubility.
- The understand that the solubility of a solid may increase or decrease with increasing temperature,

When a solute dissolves, its individual atoms, molecules, or ions interact with the solvent, become solvated, and are able to diffuse independently throughout the solution (part (a) in Figure 13.4.1). This is not, however, a unidirectional process. If the molecule or ion happens to collide with the surface of a particle of the undissolved solute, it may adhere to the particle in a process called crystallization. Dissolution and crystallization continue as long as excess solid is present, resulting in a dynamic equilibrium analogous to the equilibrium that maintains the vapor pressure of a liquid. We can represent these opposing processes as follows:

$$\text{solute} + \text{solvent} \underset{\text{dissolution}}{\overset{\text{crystallization}}{\rightleftharpoons}} \text{solution} \tag{13.4.1}$$

Although the terms precipitation and crystallization are both used to describe the separation of solid solute from a solution, crystallization refers to the formation of a solid with a well-defined crystalline structure, whereas precipitation refers to the formation of any solid phase, often one with very small particles.

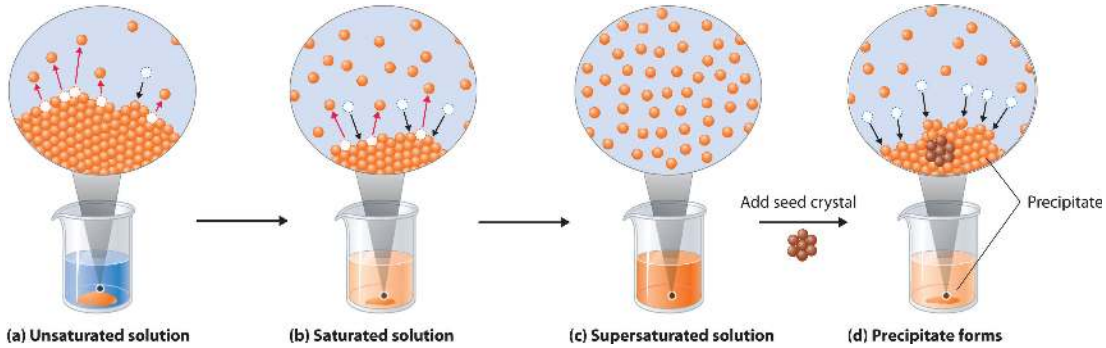


Figure 13.4.1 Dissolution and Precipitation. (a) When a solid is added to a solvent in which it is soluble, solute particles leave the surface of the solid and become solvated by the solvent, initially forming an unsaturated solution. (b) When the maximum possible amount of solute has dissolved, the solution becomes saturated. If excess solute is present, the rate at which solute particles leave the surface of the solid equals the rate at which they return to the surface of the solid. (c) A supersaturated solution can usually be formed from a saturated solution by filtering off the excess solute and lowering the temperature. (d) When a seed crystal of the solute is added to a supersaturated solution, solute particles leave the solution and form a crystalline precipitate. (CC BY-SA-NC 3.0; anonymous)

Solubility as a Function of Temperature

Figure 13.4.2 shows plots of the solubilities of several organic and inorganic compounds in water as a function of temperature. Although the solubility of a solid generally increases with increasing temperature, there is no simple relationship between the structure of a substance and the temperature dependence of its solubility. Many compounds (such as glucose and CH_3CO_2Na) exhibit a dramatic increase in solubility with increasing temperature. Others (such as $NaCl$ and K_2SO_4) exhibit little variation, and still others (such as Li_2SO_4) become less soluble with increasing temperature.

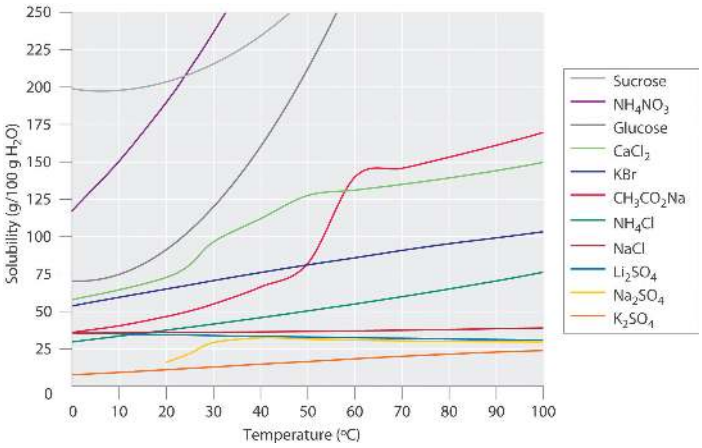


Figure 13.4.2: Solubilities of Several Inorganic and Organic Solids in Water as a Function of Temperature. Solubility may increase or decrease with temperature; the magnitude of this temperature dependence varies widely among compounds. (CC BY-SA-NC 3.0; anonymous)

Notice in particular the curves for NH_4NO_3 and $CaCl_2$. The dissolution of ammonium nitrate in water is endothermic ($\Delta H_{soln} = +25.7\ kJ/mol$), whereas the dissolution of calcium chloride is exothermic ($\Delta H_{soln} = -68.2\ kJ/mol$), yet Figure 13.4.2 shows that the solubility of both compounds increases sharply with increasing temperature. In fact, the magnitudes of the changes in both enthalpy and entropy for dissolution are temperature dependent. Because the solubility of a compound is ultimately determined by relatively small differences between large numbers, there is generally no good way to predict how the solubility will vary with temperature.

Fractional Crystallization

The variation of solubility with temperature has been measured for a wide range of compounds, and the results are published in many standard reference books. Chemists are often able to use this information to separate the components of a mixture by fractional crystallization, the separation of compounds on the basis of their solubilities in a given solvent (Figure 13.4.3). For example, if we have a mixture of 150 g of sodium acetate (CH_3CO_2Na) and 50 g of KBr, we can separate the two compounds by dissolving the mixture in 100 g of water at 80°C and then cooling the solution slowly to 0°C. According to the temperature curves in Figure 13.4.2, both compounds dissolve in water at 80°C, and all 50 g of KBr remains in solution at 0°C. Only about 36 g of CH_3CO_2Na are soluble in 100 g of water at 0°C, however, so approximately 114 g (150 g − 36 g) of CH_3CO_2Na crystallizes out on cooling. The crystals can then be separated by filtration. Thus fractional crystallization allows us to recover about 75% of the original CH_3CO_2Na in essentially pure form in only one step.

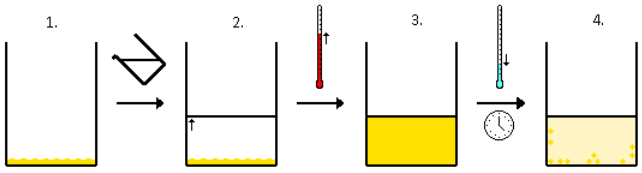


Figure 13.4.3: Steps in a Fractional crystallization technique: (1 to 2) Solvent added (clear) to compound (orange) (2 to 3) Solvent heated to give saturated compound solution (orange) (3 to 4) Saturated compound solution (orange) allowed to cool over time to give crystals (orange) and a saturated solution (pale-orange). (Public Domain; Quantockgoblin via Wikipedia)

Fractional crystallization is a common technique for purifying compounds as diverse as those shown in Figure 13.4.2 and from antibiotics to enzymes to drugs (Figure 13.4.4). For the technique to work properly, the compound of interest must be more soluble at high temperature than at low temperature, so that lowering the temperature causes it to crystallize out of solution. In addition, the impurities must be more soluble than the compound of interest (as was KBr in this example) and preferably present in relatively small amounts.



Figure 13.4.4: Image of a single solvent, slowly cooling ibuprofen salt crystallizing. (CC BY-SA 2.5 Generic; Jacopo Werther via Wikipedia).

Summary

When a solution contains the maximum amount of solute that can dissolve under a given set of conditions, it is a saturated solution. Otherwise, it is unsaturated. Supersaturated solutions, which contain more dissolved solute than allowed under particular conditions, are not stable; the addition of a seed crystal, a small particle of solute, will usually cause the excess solute to crystallize. A system in which crystallization and dissolution occur at the same rate is in dynamic equilibrium. The solubility of most substances depends strongly on the temperature. The solubility of most solid or liquid solutes increases with increasing temperature. The components of a mixture can often be separated using fractional crystallization, which separates compounds according to their solubilities. The solubility of a gas decreases with increasing temperature.

This page titled 13.4: Solution Formation and Equilibrium is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.4: Solution Formation and Equilibrium** is licensed CC BY-NC-SA 4.0.

13.5: Solubilities of Gases

The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility. For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium, but 100 times less than the solubility of chloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the liquid hydrocarbon hexane, C_6H_{14}, is approximately 20 times greater than it is in water.

Other factors also affect the solubility of a given substance in a given solvent. Temperature is one such factor, with gas solubility typically decreasing as temperature increases (Figure 13.5.1). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

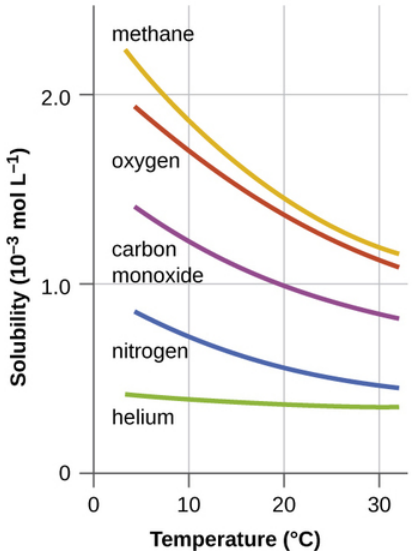


Figure 13.5.1: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (Figure 13.5.2).



Figure 13.5.2: (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service)

The solubility of a gaseous solute is also affected by the partial pressure of solute in the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at this pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution

in the form of small bubbles (Figure 13.5.3). At this point, the beverage is *supersaturated* with carbon dioxide and, with time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become "flat."

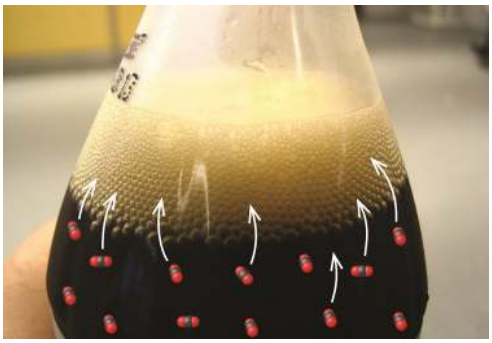


Figure 13.5.3: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

For many gaseous solutes, the relation between solubility, C_g, and partial pressure, P_g, is a proportional one:

$$C_g = kP_g \tag{13.5.1}$$

where k is a proportionality constant that depends on the identities of the gaseous solute and solvent, and on the solution temperature. This is a mathematical statement of Henry's law: *The quantity of an ideal gas that dissolves in a definite volume of liquid is directly proportional to the pressure of the gas.*

Example 13.5.1: Application of Henry's Law

At 20 °C, the concentration of dissolved oxygen in water exposed to gaseous oxygen at a partial pressure of 101.3 kPa (760 torr) is 1.38×10^{-3} mol L^{-1}. Use Henry's law to determine the solubility of oxygen when its partial pressure is 20.7 kPa (155 torr), the approximate pressure of oxygen in earth's atmosphere.

Solution

According to Henry's law, for an ideal solution the solubility, C_g, of a gas (1.38×10^{-3} mol L^{-1}, in this case) is directly proportional to the pressure, P_g, of the undissolved gas above the solution (101.3 kPa, or 760 torr, in this case). Because we know both C_g and P_g, we can rearrange this expression to solve for k.

$$\begin{aligned} C_g &= kP_g \\ k &= \frac{C_g}{P_g} \\ &= \frac{1.38 \times 10^{-3}\ \text{mol L}^{-1}}{101.3\ \text{kPa}} \\ &= 1.36 \times 10^{-5}\ \text{mol L}^{-1}\ \text{kPa}^{-1} \\ &\quad (1.82 \times 10^{-6}\ \text{mol L}^{-1}\ \text{torr}^{-1}) \end{aligned}$$

Now we can use k to find the solubility at the lower pressure.

$$C_g = kP_g \tag{13.5.2}$$

$$1.36 \times 10^{-5}\ \text{mol L}^{-1}\ \text{kPa}^{-1} \times 20.7\ \text{kPa} \tag{13.5.3}$$

$$(\text{or } 1.82 \times 10^{-6}\ \text{mol L}^{-1}\ \text{torr}^{-1} \times 155\ \text{torr})$$

$$= 2.82 \times 10^{-4}\ \text{mol L}^{-1}$$

Note that various units may be used to express the quantities involved in these sorts of computations. Any combination of units that yield to the constraints of dimensional analysis are acceptable.

Exercise 13.5.1

Exposing a 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 20.26 kPa (152 torr) resulted in the dissolution of 1.45×10^{-3} g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 101.3 kPa (760 torr).

Answer

7.25×10^{-3} g in 100.0 mL or 0.0725 g/L

Case Study: Decompression Sickness ("The Bends")

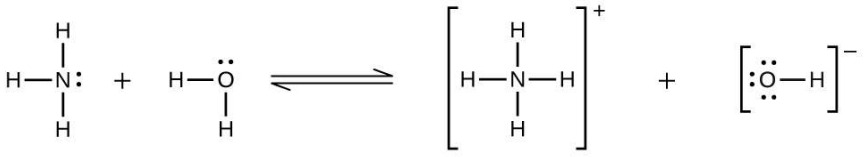
Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing an increase of approximately 1 atm for each 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher ambient pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.



Figure 13.5.4: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

As the diver ascends to the surface of the water, the ambient pressure decreases and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that can cause a variety of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompression stops, pausing for several minutes at given depths during the ascent. When these preventive measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (Figure 13.5.4).

Deviations from Henry's law are observed when a chemical reaction takes place between the gaseous solute and the solvent. Thus, for example, the solubility of ammonia in water does not increase as rapidly with increasing pressure as predicted by the law because ammonia, being a base, reacts to some extent with water to form ammonium ions and hydroxide ions.



Gases can form supersaturated solutions. If a solution of a gas in a liquid is prepared either at low temperature or under pressure (or both), then as the solution warms or as the gas pressure is reduced, the solution may become supersaturated. In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (Figure 13.5.5), a deep lake in a volcanic crater. The water at the bottom of Lake Nyos is saturated with carbon dioxide by volcanic activity beneath the lake. It is believed that the lake underwent a turnover due to gradual heating from below the lake, and the warmer, less-dense water saturated with carbon dioxide reached the surface. Consequently, tremendous quantities of dissolved CO_2 were released, and the colorless gas, which is denser than air, flowed down the valley below the lake and suffocated humans and animals living in the valley.



(a) (b)

Figure 13.5.5: (a) It is believed that the 1986 disaster that killed more than 1700 people near Lake Nyos in Cameroon resulted when a large volume of carbon dioxide gas was released from the lake. (b) A CO_2 vent has since been installed to help outgas the lake in a slow, controlled fashion and prevent a similar catastrophe from happening in the future. (credit a: modification of work by Jack Lockwood; credit b: modification of work by Bill Evans)



Henry's Law (The Solubility of Gases in Solvents): https://youtu.be/fiJZCGpArJI

This page titled 13.5: Solubilities of Gases is shared under a CC BY license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **11.3: Solubility** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

13.6: Vapor Pressures of Solutions

Learning Objectives

- To describe the relationship between solute concentration and the physical properties of a solution.
- To understand that the total number of nonvolatile solute particles determines the decrease in vapor pressure, increase in boiling point, and decrease in freezing point of a solution versus the pure solvent.

Raoult's Law

Adding a nonvolatile solute, one whose vapor pressure is too low to measure readily, to a volatile solvent decreases the vapor pressure of the solvent. We can understand this phenomenon qualitatively by examining Figure 13.6.1, which is a schematic diagram of the surface of a solution of glucose in water. In an aqueous solution of glucose, a portion of the surface area is occupied by nonvolatile glucose molecules rather than by volatile water molecules. As a result, fewer water molecules can enter the vapor phase per unit time, even though the surface water molecules have the same kinetic energy distribution as they would in pure water. At the same time, the rate at which water molecules in the vapor phase collide with the surface and reenter the solution is unaffected. The net effect is to shift the dynamic equilibrium between water in the vapor and the liquid phases, decreasing the vapor pressure of the solution compared with the vapor pressure of the pure solvent.

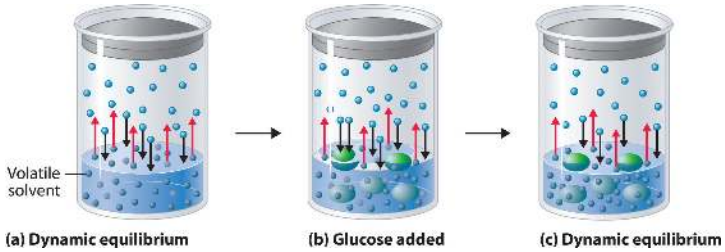


Figure 13.6.1: A Model Depicting Why the Vapor Pressure of a Solution of Glucose Is Less Than the Vapor Pressure of Pure Water. (a) When water or any volatile solvent is in a closed container, water molecules move into and out of the liquid phase at the same rate in a dynamic equilibrium. (b) If a nonvolatile solute such as glucose is added, some fraction of the surface area is occupied by solvated solute molecules. As a result, the rate at which water molecules evaporate is decreased, although initially their rate of condensation is unchanged. (c) When the glucose solution reaches equilibrium, the concentration of water molecules in the vapor phase, and hence the vapor pressure, is less than that of pure water.

Figure 13.6.2 shows two beakers, one containing pure water and one containing an aqueous glucose solution, in a sealed chamber. We can view the system as having two competing equilibria: water vapor will condense in both beakers at the same rate, but water molecules will evaporate more slowly from the glucose solution because fewer water molecules are at the surface. Eventually all of the water will evaporate from the beaker containing the liquid with the higher vapor pressure (pure water) and condense in the beaker containing the liquid with the lower vapor pressure (the glucose solution). If the system consisted of only a beaker of water inside a sealed container, equilibrium between the liquid and vapor would be achieved rather rapidly, and the amount of liquid water in the beaker would remain constant.

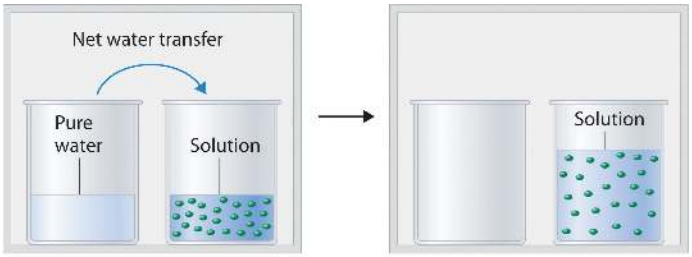


Figure 13.6.2: Transfer of Water to a Beaker Containing a Glucose Solution. (top) One beaker contains an aqueous solution of glucose, and the other contains pure water. If they are placed in a sealed chamber, the lower vapor pressure of water in the glucose solution results in a net transfer of water from the beaker containing pure water to the beaker containing the glucose solution. (bottom) Eventually, all of the water is transferred to the beaker that has the glucose solution.

If the particles of a solute are essentially the same size as those of the solvent and both solute and solvent have roughly equal probabilities of being at the surface of the solution, then the effect of a solute on the vapor pressure of the solvent is proportional to the number of sites occupied by solute particles at the surface of the solution. Doubling the concentration of a given solute causes

twice as many surface sites to be occupied by solute molecules, resulting in twice the decrease in vapor pressure. The relationship between solution composition and vapor pressure is therefore

$$P_A = X_A P_A^0 \tag{13.6.1}$$

where P_A is the vapor pressure of component A of the solution (in this case the solvent), XA is the mole fraction of A in solution, and P_A^0 is the vapor pressure of pure A. Equation 13.6.1 is known as **Raoult's law**, after the French chemist who developed it. If the solution contains only a single nonvolatile solute (B), then $X_A + X_B = 1$, and we can substitute $X_A = 1 - X_B$ to obtain

$$P_A = (1 - X_B)P_A^0 \tag{13.6.2}$$

$$= P_A^0 - X_B P_A^0 \tag{13.6.3}$$

Rearranging and defining $\Delta P_A = P_A^0 - P_A$, we obtain a relationship between the decrease in vapor pressure and the mole fraction of nonvolatile solute:

$$P_A^0 - P_A = \Delta P_A = X_B P_A^0 \tag{13.6.4}$$

We can solve vapor pressure problems in either of two ways: by using Equation 13.6.1 to calculate the actual vapor pressure above a solution of a nonvolatile solute, or by using Equation 13.6.4 to calculate the decrease in vapor pressure caused by a specified amount of a nonvolatile solute.

Example 13.6.1

Ethylene glycol ($HOCH_2CH_2OH$), the major ingredient in commercial automotive antifreeze, increases the boiling point of radiator fluid by lowering its vapor pressure. At 100°C, the vapor pressure of pure water is 760 mmHg. Calculate the vapor pressure of an aqueous solution containing 30.2% ethylene glycol by mass, a concentration commonly used in climates that do not get extremely cold in winter.

Given: identity of solute, percentage by mass, and vapor pressure of pure solvent

Asked for: vapor pressure of solution

Strategy:

A. Calculate the number of moles of ethylene glycol in an arbitrary quantity of water, and then calculate the mole fraction of water.
B. Use Raoult's law to calculate the vapor pressure of the solution.

Solution:

A A 30.2% solution of ethylene glycol contains 302 g of ethylene glycol per kilogram of solution; the remainder (698 g) is water. To use Raoult's law to calculate the vapor pressure of the solution, we must know the mole fraction of water. Thus we must first calculate the number of moles of both ethylene glycol (EG) and water present:

$$moles\ EG = (302\ \cancel{g}) \left(\frac{1\ mol}{62.07\ \cancel{g}} \right) = 4.87\ mol\ EG$$

$$moles\ \mathrm{H_2O} = (698\ \cancel{g}) \left(\frac{1\ mol}{18.02\ \cancel{g}} \right) = 38.7\ mol\ H_2O$$

The mole fraction of water is thus

$$X_{\mathrm{H_2O}} = \frac{38.7\ \cancel{mol}\ H_2O}{38.7\ \cancel{mol}\ H_2O + 4.87\ \cancel{mol}\ EG} = 0.888$$

B From Raoult's law (Equation 13.6.1), the vapor pressure of the solution is

$$P_{\mathrm{H_2O}} = (X_{H2_O})(P^0_{H2_O)} = (0.888)(760\ mmHg) = 675\ mmHg$$

Alternatively, we could solve this problem by calculating the mole fraction of ethylene glycol and then using Equation 13.6.4 to calculate the resulting decrease in vapor pressure:

$$X_{EG} = \frac{4.87\ mol\ EG}{4.87\ mol\ EG + 38.7\ mol\ H_2O} = 0.112$$

$$\Delta P_{\mathrm{H_2O}} = (X_{EG})(P^0_{H_2O}) = (0.112)(760\ mmHg) = 85.1\ mmHg$$

$$P_{\mathrm{H_2O}} = P^0 H_2O - \Delta P_{H_2O} = 760\ mmHg - 85.1\ mmHg = 675\ mmHg$$

The same result is obtained using either method.

? Exercise 13.6.1

Seawater is an approximately 3.0% aqueous solution of $NaCl$ by mass with about 0.5% of other salts by mass. Calculate the decrease in the vapor pressure of water at 25°C caused by this concentration of $NaCl$, remembering that 1 mol of $NaCl$ produces 2 mol of solute particles. The vapor pressure of pure water at 25°C is 23.8 mmHg.

Answer

0.45 mmHg. This may seem like a small amount, but it constitutes about a 2% decrease in the vapor pressure of water and accounts in part for the higher humidity in the north-central United States near the Great Lakes, which are freshwater lakes. The decrease therefore has important implications for climate modeling.



Finding the Vapor Pressure of a Solution (Ionic-Nonvolatile Solute):

https://youtu.be/sRBaRXsql9s

Volatile Solutes

Even when a solute is volatile, meaning that it has a measurable vapor pressure, we can still use Raoult's law. In this case, we calculate the vapor pressure of each component separately. The total vapor pressure of the solution (P_T) is the sum of the vapor pressures of the components:

$$P_T = P_A + P_B \tag{13.6.5}$$

$$= X_A P^0_A + X_B P^0_B \tag{13.6.6}$$

Because $X_B = 1 - X_A$ for a two-component system,

$$P_T = X_A P^0_A + (1 - X_A) P^0_B \tag{13.6.7}$$

Thus we need to specify the mole fraction of only one of the components in a two-component system. Consider, for example, the vapor pressure of solutions of benzene and toluene of various compositions. At 20°C, the vapor pressures of pure benzene and toluene are 74.7 and 22.3 mmHg, respectively. The vapor pressure of benzene in a benzene–toluene solution is

$$P_{C_6H_6} = X_{C_6H_6} P^0_{C_6H_6} \tag{13.6.8}$$

and the vapor pressure of toluene in the solution is

$$P_{C_6H_5CH_3} = X_{C_6H_5CH_3} P^0_{C_6H_5CH3} \tag{13.6.9}$$

Equations 13.6.8 and 13.6.9 are both in the form of the equation for a straight line: $y = mx + b$, where $b = 0$. Plots of the vapor pressures of both components versus the mole fractions are therefore straight lines that pass through the origin, as shown in Figure Figure 13.6.3. Furthermore, a plot of the total vapor pressure of the solution versus the mole fraction is a straight line that represents the sum of the vapor pressures of the pure components. Thus the vapor pressure of the solution is always greater than the vapor pressure of either component.

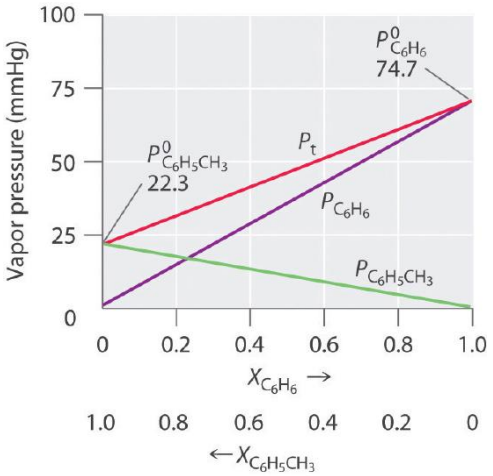


Figure 13.6.3: Vapor Pressures of Benzene–Toluene Solutions. Plots of the vapor pressures of benzene (C_6H_6) and toluene ($C_6H_5CH_3$) versus the mole fractions at 20°C are straight lines. For a solution like this, which approximates an ideal solution, the total vapor pressure of the solution (Pt) is the sum of the vapor pressures of the components.

A solution of two volatile components that behaves like the solution in Figure 13.6.3, which is defined as a solution that obeys Raoult's law. Like an ideal gas, an ideal solution is a hypothetical system whose properties can be described in terms of a simple model. Mixtures of benzene and toluene approximate an ideal solution because the intermolecular forces in the two pure liquids are almost identical in both kind and magnitude. Consequently, the change in enthalpy on solution formation is essentially zero ($\Delta H_{soln} \approx 0$), which is one of the defining characteristics of an ideal solution.

Ideal solutions and ideal gases are both simple models that ignore intermolecular interactions.

Most real solutions, however, do not obey Raoult's law precisely, just as most real gases do not obey the ideal gas law exactly. Real solutions generally deviate from Raoult's law because the intermolecular interactions between the two components A and B differ. We can distinguish between two general kinds of behavior, depending on whether the intermolecular interactions between molecules A and B are stronger or weaker than the A–A and B–B interactions in the pure components. If the A–B interactions are stronger than the A–A and B–B interactions, each component of the solution exhibits a lower vapor pressure than expected for an ideal solution, as does the solution as a whole. The favorable A–B interactions effectively stabilize the solution compared with the vapor. This kind of behavior is called a negative deviation from Raoult's law. Systems stabilized by hydrogen bonding between two molecules, such as acetone and ethanol, exhibit negative deviations from Raoult's law. Conversely, if the A–B interactions are weaker than the A–A and B–B interactions yet the entropy increase is enough to allow the solution to form, both A and B have an increased tendency to escape from the solution into the vapor phase. The result is a higher vapor pressure than expected for an ideal solution, producing a positive deviation from Raoult's law. In a solution of CCl_4 and methanol, for example, the nonpolar CCl_4 molecules interrupt the extensive hydrogen bonding network in methanol, and the lighter methanol molecules have weaker London dispersion forces than the heavier CCl_4 molecules. Consequently, solutions of CCl_4 and methanol exhibit positive deviations from Raoult's law.

✔ Example 13.6.2

For each system, compare the intermolecular interactions in the pure liquids and in the solution to decide whether the vapor pressure will be greater than that predicted by Raoult's law (positive deviation), approximately equal to that predicted by Raoult's law (an ideal solution), or less than the pressure predicted by Raoult's law (negative deviation).

a. cyclohexane and ethanol
b. methanol and acetone
c. n-hexane and isooctane (2,2,4-trimethylpentane)

Given: identity of pure liquids

Asked for: predicted deviation from Raoult's law

Strategy:

Identify whether each liquid is polar or nonpolar, and then predict the type of intermolecular interactions that occur in solution.

Solution:

a. Liquid ethanol contains an extensive hydrogen bonding network, and cyclohexane is nonpolar. Because the cyclohexane molecules cannot interact favorably with the polar ethanol molecules, they will disrupt the hydrogen bonding. Hence the A–B interactions will be weaker than the A–A and B–B interactions, leading to a higher vapor pressure than predicted by Raoult's law (a positive deviation).
b. Methanol contains an extensive hydrogen bonding network, but in this case the polar acetone molecules create A–B interactions that are stronger than the A–A or B–B interactions, leading to a negative enthalpy of solution and a lower vapor pressure than predicted by Raoult's law (a negative deviation).
c. Hexane and isooctane are both nonpolar molecules (isooctane actually has a very small dipole moment, but it is so small that it can be ignored). Hence the predominant intermolecular forces in both liquids are London dispersion forces. We expect the A–B interactions to be comparable in strength to the A–A and B–B interactions, leading to a vapor pressure in good agreement with that predicted by Raoult's law (an ideal solution).

? Exercise 13.6.2

For each system, compare the intermolecular interactions in the pure liquids with those in the solution to decide whether the vapor pressure will be greater than that predicted by Raoult's law (positive deviation), approximately equal to that predicted by Raoult's law (an ideal solution), or less than the pressure predicted by Raoult's law (negative deviation):

a. benzene and n-hexane
b. ethylene glycol and CCl_4
c. acetic acid and n-propanol

Answer a

approximately equal

Answer b

positive deviation (vapor pressure greater than predicted)

Answer c

negative deviation (vapor pressure less than predicted)



Finding Vapor Pressure of a Solution (Nonionic-Volatile Solute):

https://youtu.be/s06fzZZtLl0

Summary

- **Henry's law:**

$$C = kP$$

- **Raoult's law**:

$$P_A = X_A P_A^0$$

- **vapor pressure lowering**:

$$P_A^0 - P_A = \Delta P_A = X_B P_A^0$$

- **vapor pressure of a system containing two volatile components**:

$$P_T = X_A P_A^0 + (1 - X_A) P_B^0$$

The vapor pressure of the solution is proportional to the mole fraction of solvent in the solution, a relationship known as **Raoult's law**. Solutions that obey Raoult's law are called ideal solutions. Most real solutions exhibit positive or negative deviations from Raoult's law.

Contributors and Attributions

- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

---

This page titled 13.6: Vapor Pressures of Solutions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.6: Vapor Pressures of Solutions** is licensed CC BY-NC-SA 4.0.

13.7: Osmotic Pressure

Learning Objectives

- To describe the relationship between solute concentration and the physical properties of a solution.
- To understand that the total number of nonvolatile solute particles determines the decrease in vapor pressure, increase in boiling point, and decrease in freezing point of a solution versus the pure solvent.

Osmotic pressure is a colligative property of solutions that is observed using a semipermeable membrane, a barrier with pores small enough to allow solvent molecules to pass through but not solute molecules or ions. The net flow of solvent through a semipermeable membrane is called osmosis (from the Greek osmós, meaning “push”). The direction of net solvent flow is always from the side with the lower concentration of solute to the side with the higher concentration.

Osmosis can be demonstrated using a U-tube like the one shown in Figure 13.7.1, which contains pure water in the left arm and a dilute aqueous solution of glucose in the right arm. A net flow of water through the membrane occurs until the levels in the arms eventually stop changing, which indicates that equilibrium has been reached. The osmotic pressure (Π) of the glucose solution is the difference in the pressure between the two sides, in this case the heights of the two columns. Although the semipermeable membrane allows water molecules to flow through in either direction, the rate of flow is not the same in both directions because the concentration of water is not the same in the two arms. The net flow of water through the membrane can be prevented by applying a pressure to the right arm that is equal to the osmotic pressure of the glucose solution.

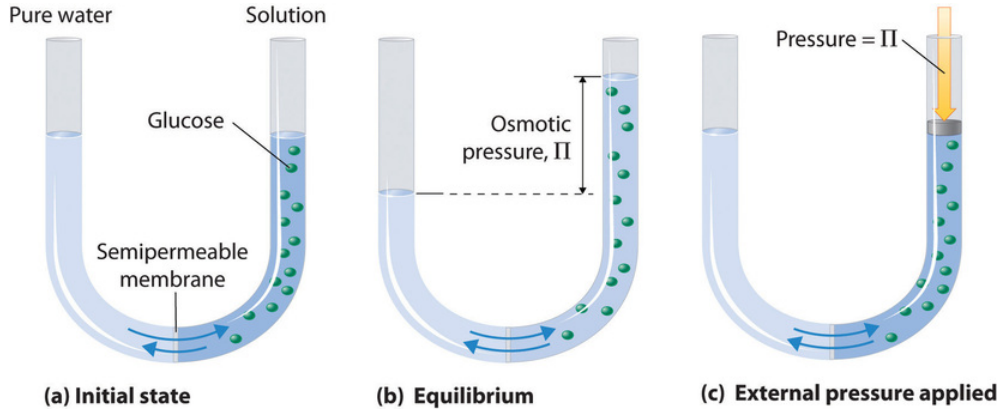


Figure 13.7.1: Osmotic Pressure. (a) A dilute solution of glucose in water is placed in the right arm of a U-tube, and the left arm is filled to the same height with pure water; a semipermeable membrane separates the two arms. Because the flow of pure solvent through the membrane from left to right (from pure water to the solution) is greater than the flow of solvent in the reverse direction, the level of liquid in the right tube rises. (b) At equilibrium, the pressure differential, equal to the osmotic pressure of the solution (Π_{soln}), *equalizes the flow rate of solvent in both directions. (c) Applying an external pressure equal to the osmotic pressure of the original glucose solution to the liquid in the right arm reverses the flow of solvent and restores the original situation.*

The osmotic pressure of a solution depends on the concentration of dissolved solute particles. Osmotic pressure obeys a law that resembles the ideal gas equation:

$$\Pi = \frac{nRT}{V} = MRT \tag{13.7.1}$$

where

- M is the number of moles of solute per unit volume of solution (i.e., the molarity of the solution),
- R is the ideal gas constant, and
- T is the absolute temperature.

As shown in Example 13.7.1, osmotic pressures tend to be quite high, even for rather dilute solutions.

Example 13.7.1: Yeast Cells

When placed in a concentrated salt solution, certain yeasts are able to produce high internal concentrations of glycerol to counteract the osmotic pressure of the surrounding medium. Suppose that the yeast cells are placed in an aqueous solution containing 4.0% $NaCl$ by mass; the solution density is 1.02 g/mL at 25°C.

Calculate the osmotic pressure of a 4.0% aqueous $NaCl$ solution at 25°C.

If the normal osmotic pressure inside a yeast cell is 7.3 atm, corresponding to a total concentration of dissolved particles of 0.30 M, what concentration of glycerol must the cells synthesize to exactly balance the external osmotic pressure at 25°C?

Given: concentration, density, and temperature of $NaCl$ solution; internal osmotic pressure of cell

Asked for: osmotic pressure of $NaCl$ solution and concentration of glycerol needed

Strategy:

A. Calculate the molarity of the $NaCl$ solution using the formula mass of the solute and the density of the solution. Then calculate the total concentration of dissolved particles.
B. Use Equation 13.7.1 to calculate the osmotic pressure of the solution.
C. Subtract the normal osmotic pressure of the cells from the osmotic pressure of the salt solution to obtain the additional pressure needed to balance the two. Use Equation 13.7.1 to calculate the molarity of glycerol needed to create this osmotic pressure.

Solution:

A The solution contains 4.0 g of $NaCl$ per 100 g of solution. Using the formula mass of $NaCl$ (58.44 g/mol) and the density of the solution (1.02 g/mL), we can calculate the molarity:

$$\begin{aligned} M_{NaCl} &= \frac{moles\ NaCl}{\text{liter solution}} \\ &= \left(\frac{4.0\ \cancel{g}\ NaCl}{58.44\ \cancel{g}/mol\ NaCl}\right)\left(\frac{1}{100\ \cancel{g\ solution}}\right)\left(\frac{1.02\ \cancel{g\ solution}}{1.00\ \cancel{mL\ solution}}\right)\left(\frac{1000\ \cancel{mL}}{1\ L}\right) \\ &= 0.70\ M\ NaCl \end{aligned}$$

Because 1 mol of $NaCl$ produces 2 mol of particles in solution, the total concentration of dissolved particles in the solution is (2)(0.70 M) = 1.4 M.

B Now we can use Equation 13.7.1 to calculate the osmotic pressure of the solution:

$$\begin{aligned} \Pi &= MRT \\ &= (1.4\ mol/L)\,[0.0821\ (L\cdot atm)/(K\cdot mol)]\,(298\ K) \\ &= 34\ atm \end{aligned}$$

C If the yeast cells are to exactly balance the external osmotic pressure, they must produce enough glycerol to give an additional internal pressure of (34 atm − 7.3 atm) = 27 atm. Glycerol is a nonelectrolyte, so we can solve Equation 13.7.1 for the molarity corresponding to this osmotic pressure:

$$\begin{aligned} M &= \frac{\Pi}{RT} \\ &= \frac{27\ \cancel{atm}}{[0.0821(L\cdot\ \cancel{atm})/(\ \cancel{K}\cdot mol)](298\ \cancel{K})} \\ &= 1.1\ M\ glycerol \end{aligned}$$

In solving this problem, we could also have recognized that the only way the osmotic pressures can be the same inside the cells and in the solution is if the concentrations of dissolved particles are the same. We are given that the normal concentration of dissolved particles in the cells is 0.3 M, and we have calculated that the $NaCl$ solution is effectively 1.4 M in dissolved

particles. The yeast cells must therefore synthesize enough glycerol to increase the internal concentration of dissolved particles from 0.3 M to 1.4 M—that is, an additional 1.1 M concentration of glycerol.

? Exercise 13.7.1

Assume that the fluids inside a sausage are approximately 0.80 M in dissolved particles due to the salt and sodium nitrite used to prepare them. Calculate the osmotic pressure inside the sausage at 100°C to learn why experienced cooks pierce the semipermeable skin of sausages before boiling them.

Answer: 24 atm

Because of the large magnitude of osmotic pressures, osmosis is extraordinarily important in biochemistry, biology, and medicine. Virtually every barrier that separates an organism or cell from its environment acts like a semipermeable membrane, permitting the flow of water but not solutes. The same is true of the compartments inside an organism or cell. Some specialized barriers, such as those in your kidneys, are slightly more permeable and use a related process called dialysis, which permits both water and small molecules to pass through but not large molecules such as proteins.

The same principle has long been used to preserve fruits and their essential vitamins over the long winter. High concentrations of sugar are used in jams and jellies not for sweetness alone but because they greatly increase the osmotic pressure. Thus any bacteria not killed in the cooking process are dehydrated, which keeps them from multiplying in an otherwise rich medium for bacterial growth. A similar process using salt prevents bacteria from growing in ham, bacon, salt pork, salt cod, and other preserved meats. The effect of osmotic pressure is dramatically illustrated in Figure 13.7.2, which shows what happens when red blood cells are placed in a solution whose osmotic pressure is much lower or much higher than the internal pressure of the cells.

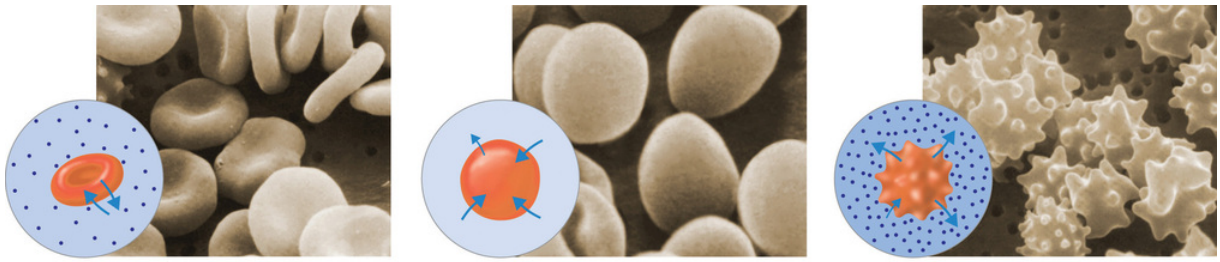


Figure 13.7.2: Effect on Red Blood Cells of the Surrounding Solution's Osmotic Pressure. (a) When red blood cells are placed in a dilute salt solution having the same osmotic pressure as the intracellular fluid, the rate of flow of water into and out of the cells is the same and their shape does not change. (b) When cells are placed in distilled water whose osmotic pressure is less than that of the intracellular fluid, the rate of flow of water into the cells is greater than the rate of flow out of the cells. The cells swell and eventually burst. (c) When cells are placed in a concentrated salt solution with an osmotic pressure greater than that of the intracellular fluid, the rate of flow of water out of the cells is greater than the rate of flow into the cells. The cells shrivel and become so deformed that they cannot function.

In addition to capillary action, trees use osmotic pressure to transport water and other nutrients from the roots to the upper branches. Evaporation of water from the leaves results in a local increase in the salt concentration, which generates an osmotic pressure that pulls water up the trunk of the tree to the leaves.

Finally, a process called reverse osmosis can be used to produce pure water from seawater. As shown schematically in Figure 13.7.3, applying high pressure to seawater forces water molecules to flow through a semipermeable membrane that separates pure water from the solution, leaving the dissolved salt behind. Large-scale desalinization plants that can produce hundreds of thousands of gallons of freshwater per day are common in the desert lands of the Middle East, where they supply a large proportion of the freshwater needed by the population. Similar facilities are now being used to supply freshwater in southern California. Small, hand-operated reverse osmosis units can produce approximately 5 L of freshwater per hour, enough to keep 25 people alive, and are now standard equipment on US Navy lifeboats.

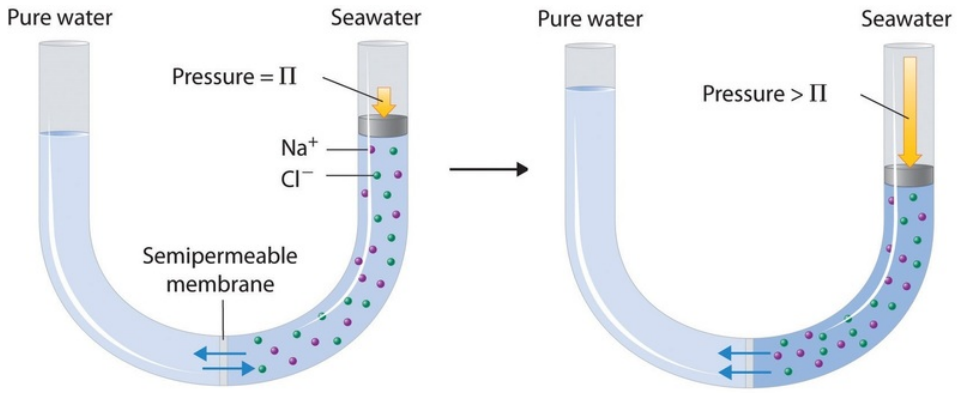


Figure 13.7.3: Desalinization of Seawater by Reverse Osmosis. (top) When the pressure applied to seawater equals its osmotic pressure (Π_{soln}), *there is no net flow of water across the semipermeable membrane. (bottom) The application of pressure greater than the osmotic pressure of seawater forces water molecules to flow through the membrane, leaving behind a concentrated salt solution. In desalinization plants, seawater is continuously introduced under pressure and pure water is collected, so the process continues indefinitely.*



Osmotic Pressure: https://youtu.be/uYQxI4mi3DA

Summary

When a solution and a pure solvent are separated by a semipermeable membrane, a barrier that allows solvent molecules but not solute molecules to pass through, the flow of solvent in opposing directions is unequal and produces an osmotic pressure, which is the difference in pressure between the two sides of the membrane. Osmosis is the net flow of solvent through such a membrane due to different solute concentrations. Dialysis uses a semipermeable membrane with pores that allow only small solute molecules and solvent molecules to pass through.

This page titled 13.7: Osmotic Pressure is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.7: Osmotic Pressure** is licensed CC BY-NC-SA 4.0.

13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions

Learning Objectives

- To describe the relationship between solute concentration and the physical properties of a solution.
- To understand that the total number of nonvolatile solute particles determines the decrease in vapor pressure, increase in boiling point, and decrease in freezing point of a solution versus the pure solvent.

Many of the physical properties of solutions differ significantly from those of the pure substances discussed in earlier chapters, and these differences have important consequences. For example, the limited temperature range of liquid water (0°C–100°C) severely limits its use. Aqueous solutions have both a lower freezing point and a higher boiling point than pure water. Probably one of the most familiar applications of this phenomenon is the addition of ethylene glycol (“antifreeze”) to the water in an automobile radiator. This solute lowers the freezing point of the water, preventing the engine from cracking in very cold weather from the expansion of pure water on freezing. Antifreeze also enables the cooling system to operate at temperatures greater than 100°C without generating enough pressure to explode.

Changes in the freezing point and boiling point of a solution depend primarily on the number of solute particles present rather than the kind of particles. Such properties of solutions are called colligative properties (from the Latin colligatus, meaning “bound together” as in a quantity). As we will see, the vapor pressure and osmotic pressure of solutions are also colligative properties.

When we determine the number of particles in a solution, it is important to remember that not all solutions with the same molarity contain the same concentration of solute particles. Consider, for example, 0.01 M aqueous solutions of sucrose, $NaCl$, and $\mathrm{CaCl_2}$. Because sucrose dissolves to give a solution of neutral molecules, the concentration of solute particles in a 0.01 M sucrose solution is 0.01 M. In contrast, both NaCl and $\mathrm{CaCl_2}$ are ionic compounds that dissociate in water to yield solvated ions. As a result, a 0.01 M aqueous solution of NaCl contains 0.01 M Na+ ions and 0.01 M Cl^- ions, for a total particle concentration of 0.02 M. Similarly, the $\mathrm{CaCl_2}$ solution contains 0.01 M Ca^{2+} ions and 0.02 M Cl^- ions, for a total particle concentration of 0.03 M.These values are correct for dilute solutions, where the dissociation of the compounds to form separately solvated ions is complete. At higher concentrations (typically >1 M), especially with salts of small, highly charged ions (such as Mg^{2+} or Al^{3+}), or in solutions with less polar solvents, dissociation to give separate ions is often incomplete. The sum of the concentrations of the dissolved solute particles dictates the physical properties of a solution. In the following discussion, we must therefore keep the chemical nature of the solute firmly in mind.

Boiling Point Elevation

Recall that the normal boiling point of a substance is the temperature at which the vapor pressure equals 1 atm. If a nonvolatile solute lowers the vapor pressure of a solvent, it must also affect the boiling point. Because the vapor pressure of the solution at a given temperature is less than the vapor pressure of the pure solvent, achieving a vapor pressure of 1 atm for the solution requires a higher temperature than the normal boiling point of the solvent. Thus the boiling point of a solution is always greater than that of the pure solvent. We can see why this must be true by comparing the phase diagram for an aqueous solution with the phase diagram for pure water (Figure $13.8.1$). The vapor pressure of the solution is less than that of pure water at all temperatures. Consequently, the liquid–vapor curve for the solution crosses the horizontal line corresponding to P = 1 atm at a higher temperature than does the curve for pure water.

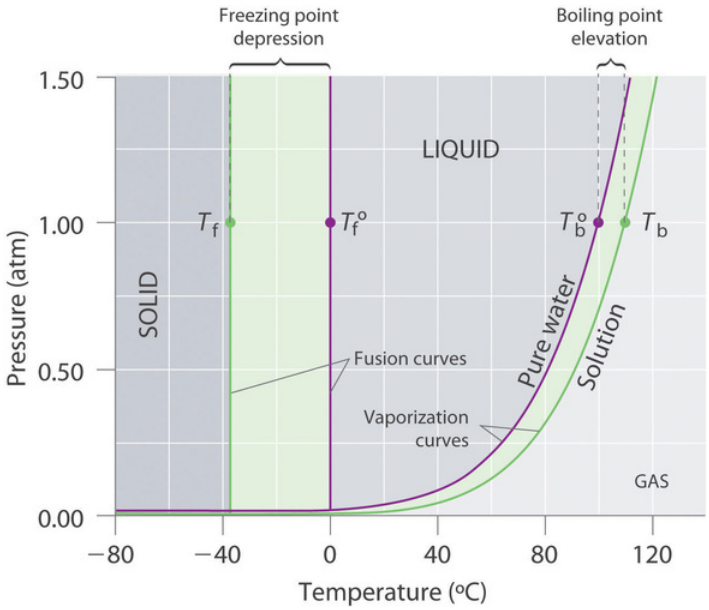


Figure 13.8.1: Phase Diagrams of Pure Water and an Aqueous Solution of a Nonvolatile Solute. The vaporization curve for the solution lies below the curve for pure water at all temperatures, which results in an increase in the boiling point and a decrease in the freezing point of the solution.

The boiling point of a solution with a nonvolatile solute is always greater than the boiling point of the pure solvent.

The magnitude of the increase in the boiling point is related to the magnitude of the decrease in the vapor pressure. As we have just discussed, the decrease in the vapor pressure is proportional to the concentration of the solute in the solution. Hence the magnitude of the increase in the boiling point must also be proportional to the concentration of the solute (Figure 13.8.2). We can define the boiling point elevation (ΔT_b) as the difference between the boiling points of the solution and the pure solvent:

$$\Delta T_b = T_b - T_b^0 \tag{13.8.1}$$

where T_b is the boiling point of the solution and T_b^0 is the boiling point of the pure solvent. We can express the relationship between ΔT_b and concentration as follows

$$\Delta T_b = mK_b \tag{13.8.2}$$

where m is the concentration of the solute expressed in molality, and K_b is the molal boiling point elevation constant of the solvent, which has units of °C/m. Table 13.8.1 lists characteristic Kb values for several commonly used solvents.

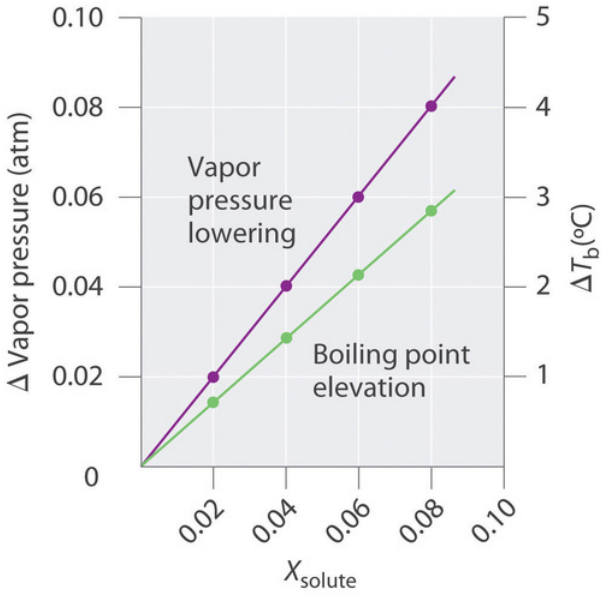


Figure 13.8.2: Vapor Pressure Decrease and Boiling Point Increase as Functions of the Mole Fraction of a Nonvolatile Solute

For relatively dilute solutions, the magnitude of both properties is proportional to the solute concentration.

Table 13.8.1: Boiling Point Elevation Constants (K_b) and Freezing Point Depression Constants (K_f) for Some Solvents

Solvent	Boiling Point (°C)	K_b (°C/m)	Freezing Point (°C)	K_f (°C/m)
acetic acid	117.90	3.22	16.64	3.63
benzene	80.09	2.64	5.49	5.07
d-(+)-camphor	207.4	4.91	178.8	37.8
carbon disulfide	46.2	2.42	−112.1	3.74
carbon tetrachloride	76.8	5.26	−22.62	31.4
chloroform	61.17	3.80	−63.41	4.60
nitrobenzene	210.8	5.24	5.70	6.87
water	100.00	0.51	0.00	1.86

The concentration of the solute is typically expressed as molality rather than mole fraction or molarity for two reasons. First, because the density of a solution changes with temperature, the value of molarity also varies with temperature. If the boiling point depends on the solute concentration, then by definition the system is not maintained at a constant temperature. Second, molality and mole fraction are proportional for relatively dilute solutions, but molality has a larger numerical value (a mole fraction can be only between zero and one). Using molality allows us to eliminate nonsignificant zeros.

According to Table 13.8.1, the molal boiling point elevation constant for water is 0.51°C/m. Thus a 1.00 m aqueous solution of a nonvolatile molecular solute such as glucose or sucrose will have an increase in boiling point of 0.51°C, to give a boiling point of 100.51°C at 1.00 atm. The increase in the boiling point of a 1.00 m aqueous NaCl solution will be approximately twice as large as that of the glucose or sucrose solution because 1 mol of NaCl produces 2 mol of dissolved ions. Hence a 1.00 m NaCl solution will have a boiling point of about 101.02°C.

✔ Example 13.8.3

In Example 13.8.1, we calculated that the vapor pressure of a 30.2% aqueous solution of ethylene glycol at 100°C is 85.1 mmHg less than the vapor pressure of pure water. We stated (without offering proof) that this should result in a higher boiling point for the solution compared with pure water. Now that we have seen why this assertion is correct, calculate the boiling point of the aqueous ethylene glycol solution.

Given: composition of solution

Asked for: boiling point

Strategy:

Calculate the molality of ethylene glycol in the 30.2% solution. Then use Equation 13.8.2 to calculate the increase in boiling point.

Solution:

From Example 13.8.1, we know that a 30.2% solution of ethylene glycol in water contains 302 g of ethylene glycol (4.87 mol) per 698 g of water. The molality of the solution is thus

$$\text{molality of ethylene glycol} = \left(\frac{4.87\ mol}{698\ \cancel{g}\ H_2O}\right)\left(\frac{1000\ \cancel{g}}{1\ kg}\right) = 6.98m \tag{13.8.3}$$

From Equation 13.8.2, the increase in boiling point is therefore

$$\Delta T_b = mK_b = (6.98\ \cancel{m})(0.51^\circ C/\ \cancel{m}) = 3.6^\circ C \tag{13.8.4}$$

The boiling point of the solution is thus predicted to be 104°C. With a solute concentration of almost 7 m, however, the assumption of a dilute solution used to obtain Equation 13.8.2 may not be valid.

? Exercise 13.8.3

Assume that a tablespoon (5.00 g) of NaCl is added to 2.00 L of water at 20.0°C, which is then brought to a boil to cook spaghetti. At what temperature will the water boil?

Answer

100.04°C, or 100°C to three significant figures. (Recall that 1 mol of NaCl produces 2 mol of dissolved particles. The small increase in temperature means that adding salt to the water used to cook pasta has essentially no effect on the cooking time.)

Freezing Point Depression

The phase diagram in Figure 13.8.1 shows that dissolving a nonvolatile solute in water not only raises the boiling point of the water but also lowers its freezing point. The solid–liquid curve for the solution crosses the line corresponding to P = 1 atm at a lower temperature than the curve for pure water.

This phenomenon is exploited in “de-icing” schemes that use salt (Figure 13.8.3), calcium chloride, or urea to melt ice on roads and sidewalks, and in the use of ethylene glycol as an “antifreeze” in automobile radiators. Seawater freezes at a lower temperature than fresh water, and so the Arctic and Antarctic oceans remain unfrozen even at temperatures below 0 °C (as do the body fluids of fish and other cold-blooded sea animals that live in these oceans).



Figure 13.8.3: Rock salt (NaCl), calcium chloride ($CaCl_2$), or a mixture of the two are used to melt ice. (credit: modification of work by Eddie Welker)

We can understand this result by imagining that we have a sample of water at the normal freezing point temperature, where there is a dynamic equilibrium between solid and liquid. Water molecules are continuously colliding with the ice surface and entering the solid phase at the same rate that water molecules are leaving the surface of the ice and entering the liquid phase. If we dissolve a nonvolatile solute such as glucose in the liquid, the dissolved glucose molecules will reduce the number of collisions per unit time

between water molecules and the ice surface because some of the molecules colliding with the ice will be glucose. Glucose, though, has a very different structure than water, and it cannot fit into the ice lattice. Consequently, the presence of glucose molecules in the solution can only decrease the rate at which water molecules in the liquid collide with the ice surface and solidify. Meanwhile, the rate at which the water molecules leave the surface of the ice and enter the liquid phase is unchanged. The net effect is to cause the ice to melt. The only way to reestablish a dynamic equilibrium between solid and liquid water is to lower the temperature of the system, which decreases the rate at which water molecules leave the surface of the ice crystals until it equals the rate at which water molecules in the solution collide with the ice.

By analogy to our treatment of boiling point elevation,the freezing point depression (ΔT_f) is defined as the difference between the freezing point of the pure solvent and the freezing point of the solution:

$$\Delta T_f = T_f^0 - T_f \tag{13.8.5}$$

where

- T_f^0 is the freezing point of the pure solvent and
- T_f is the freezing point of the solution.

The order of the terms is reversed compared with Equation 13.8.1 to express the freezing point depression as a positive number. The relationship between ΔT_f and the solute concentration is given by an equation analogous to Equation 13.8.2:

$$\Delta T_f = mK_f \tag{13.8.6}$$

where

- m is the molality of the solution and
- K_f is the molal freezing point depression constant for the solvent (in units of °C/m).

Like K_b, each solvent has a characteristic value of K_f (Table 13.8.1). Freezing point depression depends on the total number of dissolved nonvolatile solute particles, just as with boiling point elevation. Thus an aqueous NaCl solution has twice as large a freezing point depression as a glucose solution of the same molality.

People who live in cold climates use freezing point depression to their advantage in many ways. For example, ethylene glycol is added to engine coolant water to prevent an automobile engine from being destroyed, and methanol is added to windshield washer fluid to prevent the fluid from freezing. Heated glycols are often sprayed onto the surface of airplanes prior to takeoff in inclement weather in the winter to remove ice that has already formed and prevent the formation of more ice, which would be particularly dangerous if formed on the control surfaces of the aircraft (Video 13.8.1).



Video* 13.8.1*: *Freezing point depression is exploited to remove ice from the control surfaces of aircraft.*

The decrease in vapor pressure, increase in boiling point, and decrease in freezing point of a solution versus a pure liquid all depend on the total number of dissolved nonvolatile solute particles.

Example 13.8.4

In colder regions of the United States, $NaCl$ or $CaCl_2$ is often sprinkled on icy roads in winter to melt the ice and make driving safer. Use the data in Figure 13.9 to estimate the concentrations of two saturated solutions at 0°C, one of $NaCl$ and one of $CaCl_2$, and calculate the freezing points of both solutions to see which salt is likely to be more effective at melting ice.

Given: solubilities of two compounds

Asked for: concentrations and freezing points

Strategy:

A. Estimate the solubility of each salt in 100 g of water from Figure 13.9. Determine the number of moles of each in 100 g and calculate the molalities.
B. Determine the concentrations of the dissolved salts in the solutions. Substitute these values into Equation 13.8.4to calculate the freezing point depressions of the solutions.

Solution:

A From Figure 13.9, we can estimate the solubilities of $NaCl$ and $CaCl_2$ to be about 36 g and 60 g, respectively, per 100 g of water at 0°C. The corresponding concentrations in molality are

$$m_{NaCl} = \left(\frac{36\ \cancel{g\ NaCl}}{100\ \cancel{g}\ H_2O}\right)\left(\frac{1\ mol\ NaCl}{58.44\ \cancel{g\ NaCl}}\right)\left(\frac{1000\ \cancel{g}}{1\ kg}\right) = 6.2\ m \tag{13.8.7}$$

$$m_{CaCl_2} = \left(\frac{60\ \cancel{g\ CaCl_2}}{100\ \cancel{g}\ H_2O}\right)\left(\frac{1\ mol\ CaCl_2}{110.98\ \cancel{g\ CaCl_2}}\right)\left(\frac{1000\ \cancel{g}}{1kg}\right) = 5.4\ m \tag{13.8.8}$$

The lower formula mass of $NaCl$ more than compensates for its lower solubility, resulting in a saturated solution that has a slightly higher concentration than $CaCl_2$.

B Because these salts are ionic compounds that dissociate in water to yield two and three ions per formula unit of $NaCl$ and $CaCl_2$, respectively, the actual concentrations of the dissolved species in the two saturated solutions are 2 × 6.2 m = 12 m for $NaCl$ and 3 × 5.4 m = 16 m for $CaCl_2$. The resulting freezing point depressions can be calculated using Equation 13.8.4:

$$NaCl: \Delta T_f = mK_f = (12\ \cancel{m})(1.86^\circ C/\ \cancel{m}) = 22^\circ C \tag{13.8.9}$$

$$CaCl_2: \Delta T_f = mK_f = (16\ \cancel{m})(1.86^\circ C/\ \cancel{m}) = 30^\circ C \tag{13.8.10}$$

Because the freezing point of pure water is 0°C, the actual freezing points of the solutions are −22°C and −30°C, respectively. Note that $CaCl_2$ is substantially more effective at lowering the freezing point of water because its solutions contain three ions per formula unit. In fact, $CaCl_2$ is the salt usually sold for home use, and it is also often used on highways.

Because the solubilities of both salts decrease with decreasing temperature, the freezing point can be depressed by only a certain amount, regardless of how much salt is spread on an icy road. If the temperature is significantly below the minimum temperature at which one of these salts will cause ice to melt (say −35°C), there is no point in using salt until it gets warmer

Exercise 13.8.4

Calculate the freezing point of the 30.2% solution of ethylene glycol in water whose vapor pressure and boiling point we calculated in Example 13.8.68 and Example 13.8.610.

Answer

−13.0°C

Example 13.8.5

Arrange these aqueous solutions in order of decreasing freezing points: 0.1 m KCl, 0.1 m glucose, 0.1 m SrCl2, 0.1 m ethylene glycol, 0.1 m benzoic acid, and 0.1 m HCl.

Given: molalities of six solutions

Asked for: relative freezing points

Strategy:

A. Identify each solute as a strong, weak, or nonelectrolyte, and use this information to determine the number of solute particles produced.
B. Multiply this number by the concentration of the solution to obtain the effective concentration of solute particles. The solution with the highest effective concentration of solute particles has the largest freezing point depression.

Solution:

A Because the molal concentrations of all six solutions are the same, we must focus on which of the substances are strong electrolytes, which are weak electrolytes, and which are nonelectrolytes to determine the actual numbers of particles in solution. KCl, $SrCl_2$, and HCl are **strong electrolytes**, producing two, three, and two ions per formula unit, respectively. Benzoic acid is a weak electrolyte (approximately one particle per molecule), and glucose and ethylene glycol are both nonelectrolytes (one particle per molecule).

B The molalities of the solutions in terms of the total particles of solute are: KCl and HCl, 0.2 m; $SrCl_2$, 0.3 m; glucose and ethylene glycol, 0.1 m; and benzoic acid, 0.1–0.2 m. Because the magnitude of the decrease in freezing point is proportional to the concentration of dissolved particles, the order of freezing points of the solutions is: glucose and ethylene glycol (highest freezing point, smallest freezing point depression) > benzoic acid > HCl = KCl > $SrCl_2$.

Exercise 13.8.5

Arrange these aqueous solutions in order of increasing freezing points: 0.2 m $NaCl$, 0.3 m acetic acid, 0.1 m $\mathrm{CaCl_2}$, and 0.2 m sucrose.

Answer

0.2 m NaCl (lowest freezing point) < 0.3 m acetic acid ≈ 0.1 m $\mathrm{CaCl_2}$ < 0.2 m sucrose (highest freezing point)



Boiling Point Elevation and Freezing Point Depression: https://youtu.be/0MZm1Ay6LhU

Determination of Molar Masses

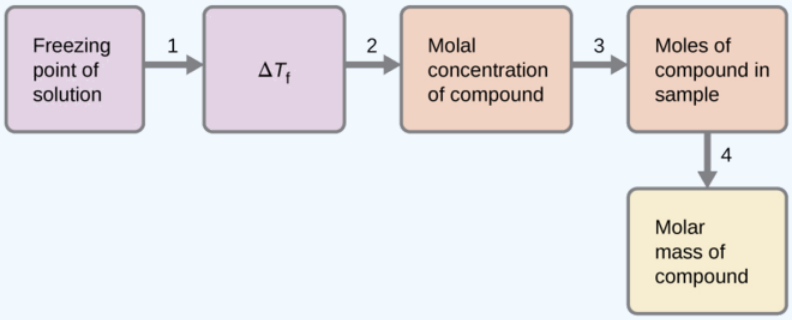
Osmotic pressure and changes in freezing point, boiling point, and vapor pressure are directly proportional to the concentration of solute present. Consequently, we can use a measurement of one of these properties to determine the molar mass of the solute from the measurements.

Determining Molar Mass from Freezing Point Depression

A solution of 4.00 g of a nonelectrolyte dissolved in 55.0 g of benzene is found to freeze at 2.32 °C. What is the molar mass of this compound?

Solution

We can solve this problem using the following steps.



1. *Determine the change in freezing point from the observed freezing point and the freezing point of pure benzene* (Table 13.8.1).

$\Delta T_f = 5.5\ ^\circ\text{C} - 2.32\ ^\circ\text{C} = 3.2\ ^\circ\text{C}$

- *Determine the molal concentration from K_f, the freezing point depression constant for benzene* (Table 13.8.1), *and* ΔT_f.

  $\Delta T_f = K_f m$

  $$m = \frac{\Delta T_f}{K_f} = \frac{3.2\ ^\circ\text{C}}{5.12\ ^\circ\text{C} m^{-1}} = 0.63\ m$$

- *Determine the number of moles of compound in the solution from the molal concentration and the mass of solvent used to make the solution.*

  $$\text{Moles of solute} = \frac{0.62\ \text{mol solute}}{1.00\ \text{kg solvent}} \times 0.0550\ \text{kg solvent} = 0.035\ \text{mol}$$

- *Determine the molar mass from the mass of the solute and the number of moles in that mass.*

  $$\text{Molar mass} = \frac{4.00\ \text{g}}{0.034\ \text{mol}} = 1.2 \times 10^2\ \text{g/mol}$$

Exercise 13.8.6

A solution of 35.7 g of a nonelectrolyte in 220.0 g of chloroform has a boiling point of 64.5 °C. What is the molar mass of this compound?

Answer

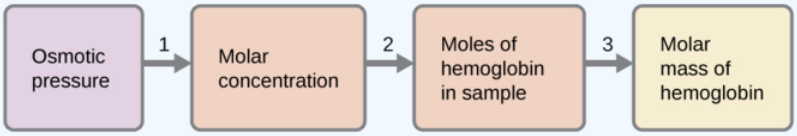
1.8×10^2 g/mol

Determination of a Molar Mass from Osmotic Pressure

A 0.500 L sample of an aqueous solution containing 10.0 g of hemoglobin has an osmotic pressure of 5.9 torr at 22 °C. What is the molar mass of hemoglobin?

Solution

Here is one set of steps that can be used to solve the problem:



1. *Convert the osmotic pressure to atmospheres, then determine the molar concentration from the osmotic pressure.*

$$\Pi = \frac{5.9 \text{ torr} \times 1 \text{ atm}}{760 \text{ torr}} = 7.8 \times 10^{-3} \text{ atm} \tag{13.8.11}$$

$$\Pi = MRT \tag{13.8.12}$$

$$M = \frac{\Pi}{RT} = \frac{7.8 \times 10^{-3} \text{ atm}}{(0.08206 \text{ L atm/mol K})(295 \text{ K})} = 3.2 \times 10^{-4} \text{ M}$$

- *Determine the number of moles of hemoglobin in the solution from the concentration and the volume of the solution.*

$$\text{moles of hemoglobin} = \frac{3.2 \times 10^{-4} \text{ mol}}{1 \text{ L solution}} \times 0.500 \text{ L solution} = 1.6 \times 10^{-4} \text{ mol}$$

- *Determine the molar mass from the mass of hemoglobin and the number of moles in that mass.*

$$\text{molar mass} = \frac{10.0 \text{ g}}{1.6 \times 10^{-4} \text{ mol}} = 6.2 \times 10^{4} \text{ g/mol}$$

? Exercise 13.8.7

What is the molar mass of a protein if a solution of 0.02 g of the protein in 25.0 mL of solution has an osmotic pressure of 0.56 torr at 25 °C?

Answer

2.7×10^4 g/mol



Finding the Molecular Weight of an Unknown using Colligative Properties:

https://youtu.be/faSk2REYy74

Summary

- **vapor pressure lowering**:

$$P_A^0 - P_A = \Delta P_A = X_B P_A^0 \tag{13.8.13}$$

- **vapor pressure of a system containing two volatile components**:

$$P_T = X_A P_A^0 + (1 - X_A) P_B^0 \tag{13.8.14}$$

- **boiling point elevation**:

$$\Delta T_b = m K_b \tag{13.8.15}$$

- **freezing point depression:**

$$\Delta T_f = m K_f \tag{13.8.16}$$

The colligative properties of a solution depend on only the total number of dissolved particles in solution, not on their chemical identity. Colligative properties include vapor pressure, boiling point, freezing point, and osmotic pressure. The addition of a nonvolatile solute (one without a measurable vapor pressure) decreases the vapor pressure of the solvent. The vapor pressure of the solution is proportional to the mole fraction of solvent in the solution, a relationship known as **Raoult's law**. Solutions that obey Raoult's law are called ideal solutions. Most real solutions exhibit positive or negative deviations from Raoult's law. The boiling point elevation (ΔT_b) and freezing point depression (ΔT_f) of a solution are defined as the differences between the boiling and freezing points, respectively, of the solution and the pure solvent. Both are proportional to the molality of the solute.

This page titled 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions** is licensed CC BY-NC-SA 4.0.

13.9: Solutions of Electrolytes

Thus far we have assumed that we could simply multiply the molar concentration of a solute by the number of ions per formula unit to obtain the actual concentration of dissolved particles in an electrolyte solution. We have used this simple model to predict such properties as freezing points, melting points, vapor pressure, and osmotic pressure. If this model were perfectly correct, we would expect the freezing point depression of a 0.10 m solution of sodium chloride, with 2 mol of ions per mole of $NaCl$ in solution, to be exactly twice that of a 0.10 m solution of glucose, with only 1 mol of molecules per mole of glucose in solution. In reality, this is not always the case. Instead, the observed change in freezing points for 0.10 m aqueous solutions of $NaCl$ and KCl are significantly less than expected (−0.348°C and −0.344°C, respectively, rather than −0.372°C), which suggests that fewer particles than we expected are present in solution.

The relationship between the actual number of moles of solute added to form a solution and the apparent number as determined by colligative properties is called the van't Hoff factor (i) and is defined as follows:Named for Jacobus Hendricus van't Hoff (1852–1911), a Dutch chemistry professor at the University of Amsterdam who won the first Nobel Prize in Chemistry (1901) for his work on thermodynamics and solutions.

$$i = \frac{\text{apparent number of particles in solution}}{\text{number of moles of solute dissolved}} \tag{13.9.1}$$

> As the solute concentration increases the van't Hoff factor decreases.

The van't Hoff factor is therefore a measure of a deviation from ideal behavior. The lower the van 't Hoff factor, the greater the deviation. As the data in Table 13.9.1 show, the van't Hoff factors for ionic compounds are somewhat lower than expected; that is, their solutions apparently contain fewer particles than predicted by the number of ions per formula unit. As the concentration of the solute increases, the van't Hoff factor decreases because ionic compounds generally do not totally dissociate in aqueous solution.

Table 13.9.1: van 't Hoff Factors for 0.0500 M Aqueous Solutions of Selected Compounds at 25°C

Compound	i (measured)	i (ideal)
glucose	1.0	1.0
sucrose	1.0	1.0
$NaCl$	1.9	2.0
HCl	1.9	2.0
$MgCl_2$	2.7	3.0
$FeCl_3$	3.4	4.0
$Ca(NO_3)_2$	2.5	3.0
$AlCl_3$	3.2	4.0
$MgSO_4$	1.4	2.0

Instead, some of the ions exist as ion pairs, a cation and an anion that for a brief time are associated with each other without an intervening shell of water molecules (Figure 13.9.1). Each of these temporary units behaves like a single dissolved particle until it dissociates. Highly charged ions such as Mg^{2+}, Al^{3+}, SO_4^{2-}, and PO_4^{3-} have a greater tendency to form ion pairs because of their strong electrostatic interactions. The actual number of solvated ions present in a solution can be determined by measuring a colligative property at several solute concentrations.

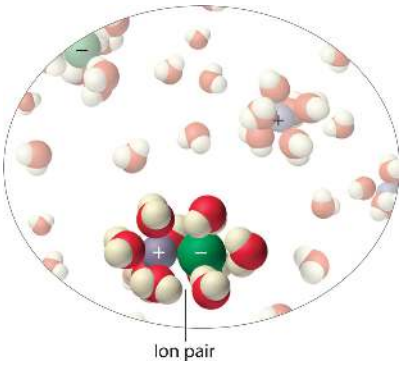


Figure 13.9.1: Ion Pairs. In concentrated solutions of electrolytes like $NaCl$, some of the ions form neutral ion pairs that are not separated by solvent and diffuse as single particles.

Example 13.9.1: Iron Chloride in Water

A 0.0500 M aqueous solution of $FeCl_3$ has an osmotic pressure of 4.15 atm at 25°C. Calculate the van't Hoff factor i for the solution.

Given: solute concentration, osmotic pressure, and temperature

Asked for: van't Hoff factor

Strategy:

A. Use Equation 13.9.12 to calculate the expected osmotic pressure of the solution based on the effective concentration of dissolved particles in the solvent.
B. Calculate the ratio of the observed osmotic pressure to the expected value. Multiply this number by the number of ions of solute per formula unit, and then use Equation 13.9.1 to calculate the van't Hoff factor.

Solution:

A If $FeCl_3$ dissociated completely in aqueous solution, it would produce four ions per formula unit [Fe3+(aq) plus 3Cl−(aq)] for an effective concentration of dissolved particles of 4 × 0.0500 M = 0.200 M. The osmotic pressure would be

$$\Pi = MRT = (0.200\ mol/L)\,[0.0821\ (L\cdot atm)/(K\cdot mol)]\,(298\ K) = 4.89\ atm \tag{13.9.2}$$

B The observed osmotic pressure is only 4.15 atm, presumably due to ion pair formation. The ratio of the observed osmotic pressure to the calculated value is 4.15 atm/4.89 atm = 0.849, which indicates that the solution contains (0.849)(4) = 3.40 particles per mole of $FeCl_3$ dissolved. Alternatively, we can calculate the observed particle concentration from the osmotic pressure of 4.15 atm:

$$4.15\ atm = M\,[0.0821\ (L\cdot atm)/(K\cdot mol)]\,(298\ K) \tag{13.9.3}$$

$$0.170 mol/L = M \tag{13.9.4}$$

The ratio of this value to the expected value of 0.200 M is 0.170 M/0.200 M = 0.850, which again gives us (0.850)(4) = 3.40 particles per mole of $FeCl_3$ dissolved. From Equation 13.9.1, the van't Hoff factor for the solution is

$$i = \frac{3.40\ \text{particles observed}}{1\ \text{formula unit}\ FeCl_3} = 3.40 \tag{13.9.5}$$

Exercise 13.9.1: Magnesium Chloride in Water

Calculate the van't Hoff factor for a 0.050 m aqueous solution of $MgCl_2$ that has a measured freezing point of −0.25°C.

Answer:

2.7 (versus an ideal value of 3

Key Concepts and Summary

Ionic compounds may not completely dissociate in solution due to **activity** effects, in which case observed colligative effects may be less than predicted.

This page titled 13.9: Solutions of Electrolytes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.9: Solutions of Electrolytes** is licensed CC BY-NC-SA 4.0.

13.10: Colloidal Mixtures

A suspension is a heterogeneous mixture of particles of one substance distributed throughout a second phase; the dispersed particles separate from the dispersing phase on standing. In contrast, the particles in a colloid are smaller and do not separate on standing. A colloid can be classified as a sol, a dispersion of solid particles in a liquid or solid; a gel, a semisolid sol in which all of the liquid phase has been absorbed by the solid particles; an aerosol, a dispersion of solid or liquid particles in a gas; or an emulsion, a dispersion of one liquid phase in another. A colloid can be distinguished from a true solution by its ability to scatter a beam of light, known as the Tyndall effect. Hydrophilic colloids contain an outer shell of groups that interact favorably with water, whereas hydrophobic colloids have an outer surface with little affinity for water. Emulsions are prepared by dispersing a hydrophobic liquid in water. In the absence of a dispersed hydrophobic liquid phase, solutions of detergents in water form organized spherical aggregates called micelles. Phospholipids are a class of detergent-like molecules that have two hydrophobic tails attached to a hydrophilic head. A bilayer is a two-dimensional sheet consisting of a double layer of phospholipid molecules arranged tail to tail with a hydrophobic interior and a hydrophilic exterior. Cells are collections of molecules that are surrounded by a phospholipid bilayer called a cell membrane and are able to reproduce themselves.

This page titled 13.10: Colloidal Mixtures is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.10: Colloidal Mixtures** is licensed CC BY-NC-SA 4.0.

13.11: Exercises

11.2: The Dissolution Process

Q11.2.1

How do solutions differ from compounds? From other mixtures?

S11.2.1

A solution can vary in composition, while a compound cannot vary in composition. Solutions are homogeneous at the molecular level, while other mixtures are heterogeneous.

Q11.2.2

Which of the principal characteristics of solutions can we see in the solutions of $K_2Cr_2O_7$ shown in



Figure: When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott)

S11.2.2

The solutions are the same throughout (the color is constant throughout), and the composition of a solution of $K_2Cr_2O_7$ in water can vary.

Q11.2.3

When KNO_3 is dissolved in water, the resulting solution is significantly colder than the water was originally.

a. Is the dissolution of KNO_3 an endothermic or an exothermic process?
b. What conclusions can you draw about the intermolecular attractions involved in the process?
c. Is the resulting solution an ideal solution?

S11.2.3

(a) The process is endothermic as the solution is consuming heat. (b) Attraction between the K^+ and NO_3^- ions is stronger than between the ions and water molecules (the ion-ion interactions have a lower, more negative energy). Therefore, the dissolution process increases the energy of the molecular interactions, and it consumes the thermal energy of the solution to make up for the difference. (c) No, an ideal solution is formed with no appreciable heat release or consumption.

Q11.2.4

Give an example of each of the following types of solutions:

a. a gas in a liquid
b. a gas in a gas
c. a solid in a solid

S11.2.4

(a) CO_2 in water; (b) O_2 in N_2 (air); (c) bronze (solution of tin or other metals in copper)

Q11.2.5

Indicate the most important types of intermolecular attractions in each of the following solutions:

a. The solution in Figure.

b. NO(l) in CO(l)
c. $Cl_2(g)$ in $Br_2(l)$
d. HCl(aq) in benzene $C_6H_6(l)$
e. Methanol $CH_3OH(l)$ in $H_2O(l)$

S11.2.5

(a) ion-dipole forces; (b) dipole-dipole forces; (c) dispersion forces; (d) dispersion forces; (e) hydrogen bonding

Q11.2.5

Predict whether each of the following substances would be more soluble in water (polar solvent) or in a hydrocarbon such as heptane (C_7H_{16}, nonpolar solvent):

a. vegetable oil (nonpolar)
b. isopropyl alcohol (polar)
c. potassium bromide (ionic)

S11.2.5

(a) heptane; (b) water; (c) water

Q11.2.6

Heat is released when some solutions form; heat is absorbed when other solutions form. Provide a molecular explanation for the difference between these two types of spontaneous processes.

S11.2.6

Heat is released when the total intermolecular forces (IMFs) between the solute and solvent molecules are stronger than the total IMFs in the pure solute and in the pure solvent: Breaking weaker IMFs and forming stronger IMFs releases heat. Heat is absorbed when the total IMFs in the solution are weaker than the total of those in the pure solute and in the pure solvent: Breaking stronger IMFs and forming weaker IMFs absorbs heat.

Q11.2.7

Solutions of hydrogen in palladium may be formed by exposing Pd metal to H_2 gas. The concentration of hydrogen in the palladium depends on the pressure of H_2 gas applied, but in a more complex fashion than can be described by Henry's law. Under certain conditions, 0.94 g of hydrogen gas is dissolved in 215 g of palladium metal.

a. Determine the molarity of this solution (solution density = 1.8 g/cm^3).
b. Determine the molality of this solution (solution density = 1.8 g/cm^3).
c. Determine the percent by mass of hydrogen atoms in this solution (solution density = 1.8 g/cm^3).

S11.2.7

http://cnx.org/contents/mH6aqegx@2/The-Dissolution-Process

11.3: Electrolytes

Q11.3.1

Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules.

S11.3.2

Crystals of NaCl dissolve in water, a polar liquid with a very large dipole moment, and the individual ions become strongly solvated. Hexane is a nonpolar liquid with a dipole moment of zero and, therefore, does not significantly interact with the ions of the NaCl crystals.

Q11.3.2

Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.

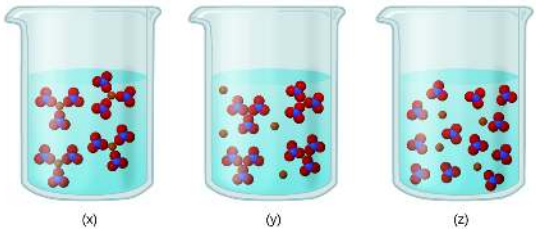
S11.3.2

HBr is an acid and so its molecules react with water molecules to form H_3O^+ and Br^- ions that provide conductivity to the solution. Though HBr is soluble in benzene, it does not react chemically but remains dissolved as neutral HBr molecules. With no ions present in the benzene solution, it is electrically nonconductive.

Q11.3.3

Consider the solutions presented:

(a) Which of the following sketches best represents the ions in a solution of $Fe(NO_3)_3(aq)$?



(b) Write a balanced chemical equation showing the products of the dissolution of $Fe(NO_3)_3$.

S11.3.3

(a) $Fe(NO_3)_3$ is a strong electrolyte, thus it should completely dissociate into Fe^{3+} and (NO_3^-) ions. Therefore, (z) best represents the solution. (b) $\mathrm{Fe(NO_3)_3}(s) \longrightarrow \mathrm{Fe^{3+}}(aq) + 3\,\mathrm{NO_3^-}(aq)$

Q11.3.4

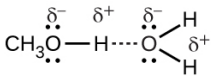
Compare the processes that occur when methanol (CH_3OH), hydrogen chloride (HCl), and sodium hydroxide (NaOH) dissolve in water. Write equations and prepare sketches showing the form in which each of these compounds is present in its respective solution.

S11.3.3

Methanol, CH_3OH, dissolves in water in all proportions, interacting via hydrogen bonding.

Methanol:

$$CH_3OH_{(l)} + H_2O_{(l)} \longrightarrow CH_3OH_{(aq)} \tag{13.11.1}$$



hydrogen bonding

Hydrogen chloride, HCl, dissolves in and reacts with water to yield hydronium cations and chloride anions that are solvated by strong ion-dipole interactions.

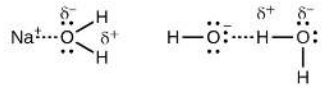
Hydrogen chloride:

$$HCl(g) + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)} \tag{13.11.2}$$

Sodium hydroxide, NaOH, dissolves in water and dissociates to yield sodium cations and hydroxide anions that are strongly solvated by ion-dipole interactions and hydrogen bonding, respectively.

Sodium hydroxide:

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)} \tag{13.11.3}$$



Ion-dipole attraction hydrogen bonding

https://chem.libretexts.org/@go/page/494813

Q11.3.5

What is the expected electrical conductivity of the following solutions?

a. NaOH(aq)
b. HCl(aq)
c. $C_6H_{12}O_6$(aq) (glucose)
d. NH_3(l)

S11.3.5

(a) high conductivity (solute is an ionic compound that will dissociate when dissolved); (b) high conductivity (solute is a strong acid and will ionize completely when dissolved); (c) nonconductive (solute is a covalent compound, neither acid nor base, unreactive towards water); (d) low conductivity (solute is a weak base and will partially ionize when dissolved)

Q11.3.6

Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.

S11.3.6

A medium must contain freely mobile, charged entities to be electrically conductive. The ions present in a typical ionic solid are immobilized in a crystalline lattice and so the solid is not able to support an electrical current. When the ions are mobilized, either by melting the solid or dissolving it in water to dissociate the ions, current may flow and these forms of the ionic compound are conductive.

Q11.3.7

Indicate the most important type of intermolecular attraction responsible for solvation in each of the following solutions:

a. the solutions in Figure
b. methanol, CH_3OH, dissolved in ethanol, C_2H_5OH
c. methane, CH_4, dissolved in benzene, C_6H_6
d. the polar halocarbon CF_2Cl_2 dissolved in the polar halocarbon $CF_2ClCFCl_2$
e. O_2(l) in N_2(l)

S11.3.7

(a) ion-dipole; (b) hydrogen bonds; (c) dispersion forces; (d) dipole-dipole attractions; (e) dispersion forces

11.4: Solubility

Q11.4.1

Suppose you are presented with a clear solution of sodium thiosulfate, $Na_2S_2O_3$. How could you determine whether the solution is unsaturated, saturated, or supersaturated?

S11.4.1

Add a small crystal of $Na_2S_2O_3$. It will dissolve in an unsaturated solution, remain apparently unchanged in a saturated solution, or initiate precipitation in a supersaturated solution.

Q11.4.2

Supersaturated solutions of most solids in water are prepared by cooling saturated solutions. Supersaturated solutions of most gases in water are prepared by heating saturated solutions. Explain the reasons for the difference in the two procedures.

S11.4.2

The solubility of solids usually decreases upon cooling a solution, while the solubility of gases usually decreases upon heating.

Q11.4.3

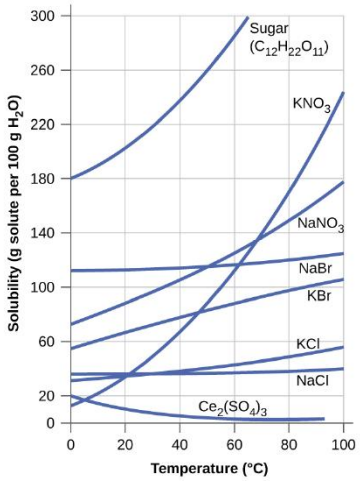
Suggest an explanation for the observations that ethanol, C_2H_5OH, is completely miscible with water and that ethanethiol, C_2H_5SH, is soluble only to the extent of 1.5 g per 100 mL of water.

S11.4.3

The hydrogen bonds between water and C_2H_5OH are much stronger than the intermolecular attractions between water and C_2H_5SH.

Q11.4.4

Calculate the percent by mass of KBr in a saturated solution of KBr in water at 10 °C using the following figure for useful data, and report the computed percentage to one significant digit.



This graph shows how the solubility of several solids changes with temperature.

S11.4.4

At 10 °C, the solubility of KBr in water is approximately 60 g per 100 g of water.

$$\%\ KBr = \frac{60\ g\ KBr}{(60+100)\ g\ solution} = 40\% \tag{13.11.4}$$

Q11.4.5

Which of the following gases is expected to be most soluble in water? Explain your reasoning.

a. CH_4
b. CCl_4
c. $CHCl_3$

S11.4.5

(c) $CHCl_3$ is expected to be most soluble in water. Of the three gases, only this one is polar and thus capable of experiencing relatively strong dipole-dipole attraction to water molecules.

Q11.4.6

At 0 °C and 1.00 atm, as much as 0.70 g of O_2 can dissolve in 1 L of water. At 0 °C and 4.00 atm, how many grams of O_2 dissolve in 1 L of water?

S11.4.6

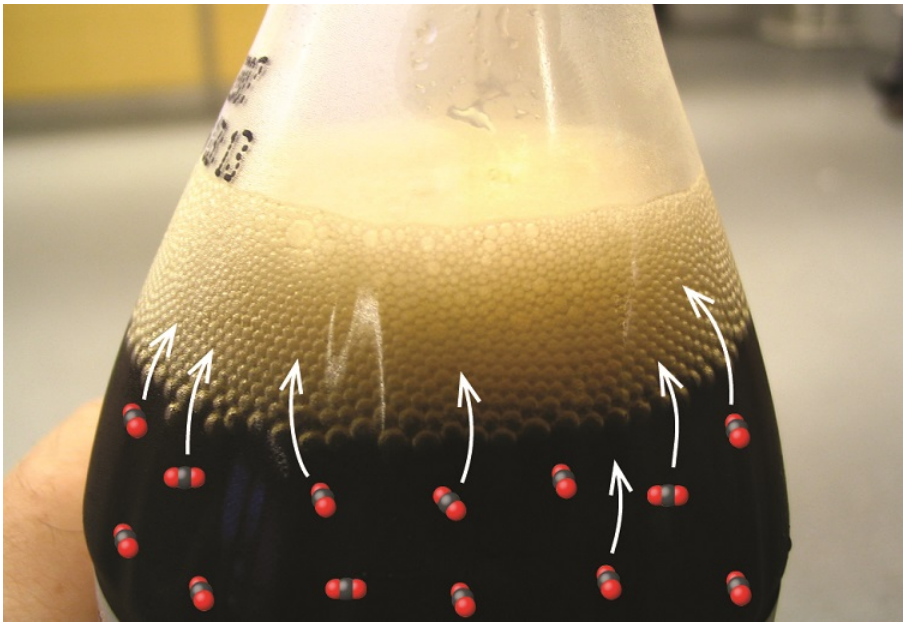
This problem requires the application of Henry's law. The governing equation is $C_g = kP_g$.

$$k = \frac{C_g}{P_g} = \frac{0.70\ g}{1.00\ atm} = 0.70\ g\ atm^{-1} \tag{13.11.5}$$

Under the new conditions, $C_g = 0.70\ g\ atm^{-1} \times 4.00\ atm = 2.80\ g$.

Q11.4.7

Refer to following figure for the following three questions:



a. How did the concentration of dissolved CO_2 in the beverage change when the bottle was opened?
b. What caused this change?
c. Is the beverage unsaturated, saturated, or supersaturated with CO_2?

S11.4.7

(a) It decreased as some of the CO_2 gas left the solution (evidenced by effervescence). (b) Opening the bottle released the high-pressure CO_2 gas above the beverage. The reduced CO_2 gas pressure, per Henry's law, lowers the solubility for CO_2. (c) The dissolved CO_2 concentration will continue to slowly decrease until equilibrium is reestablished between the beverage and the very low CO_2 gas pressure in the opened bottle. Immediately after opening, the beverage, therefore, contains dissolved CO_2 at a concentration greater than its solubility, a nonequilibrium condition, and is said to be supersaturated.

Q11.4.8

The Henry's law constant for CO_2 is $3.4 \times 10^{-2}\ M/atm$ at 25 °C. What pressure of carbon dioxide is needed to maintain a CO_2 concentration of 0.10 *M* in a can of lemon-lime soda?

S11.4.8

$$P_g = \frac{C_g}{k} = \frac{0.10\ M}{3.4 \times 10^{-2}\ M/atm} = 2.9\ atm \tag{13.11.6}$$

Q11.4.9

The Henry's law constant for O_2 is $1.3 \times 10^{-3}\ M/atm$ at 25 °C. What mass of oxygen would be dissolved in a 40-L aquarium at 25 °C, assuming an atmospheric pressure of 1.00 atm, and that the partial pressure of O_2 is 0.21 atm?

S11.4.9

Start with Henry's law

$$C_g = kP_g \tag{13.11.7}$$

and apply it to O_2

$$C(O_2) = (1.3 \times 10^{-3}\ M/atm)(0.21\ atm) = 2.7 \times 10^{-4}\ mol/L \tag{13.11.8}$$

The total amount is \((2.7 \times 10^{-4}\; mol/L)(40\;L=1.08 \times 10^{-2} \;mol\]

The mass of oxygen is $(1.08 \times 10^{-2}\ mol)(32.0\ g/mol) = 0.346\ g$

or, using two significant figures, $0.35\ g$.

Q11.4.10

How many liters of HCl gas, measured at 30.0 °C and 745 torr, are required to prepare 1.25 L of a 3.20-*M* solution of hydrochloric acid?

S11.4.10

First, calculate the moles of HCl needed. Then use the ideal gas law to find the volume required.

M = mol L−1

3.20M=xmol1.25L

x = 4.00 mol HCl

Before using the ideal gas law, change pressure to atmospheres and convert temperature from °C to kelvin.

\[1\;atmx=760torr745torr

x = 0.9803 atm

V=nRTP=(4.000molHCl)(0.08206LatmK−1mol−1)(303.15K)0.9803atm=102 L HCl

102 L HCl

more http://cnx.org/contents/2488fW6W@2/Solubility

11.5: Colligative Properties

Q11.5.1

Which is/are part of the macroscopic domain of solutions and which is/are part of the microscopic domain: boiling point elevation, Henry's law, hydrogen bond, ion-dipole attraction, molarity, nonelectrolyte, nonstoichiometric compound, osmosis, solvated ion?

Q11.5.2

What is the microscopic explanation for the macroscopic behavior illustrated in [link]?

S11.5.2

The strength of the bonds between like molecules is stronger than the strength between unlike molecules. Therefore, some regions will exist in which the water molecules will exclude oil molecules and other regions will exist in which oil molecules will exclude water molecules, forming a heterogeneous region.

Q11.5.3

Sketch a qualitative graph of the pressure versus time for water vapor above a sample of pure water and a sugar solution, as the liquids evaporate to half their original volume.

Q11.5.4

A solution of potassium nitrate, an electrolyte, and a solution of glycerin ($C_3H_5(OH)_3$), a nonelectrolyte, both boil at 100.3 °C. What other physical properties of the two solutions are identical?

S11.5.4

Both form homogeneous solutions; their boiling point elevations are the same, as are their lowering of vapor pressures. Osmotic pressure and the lowering of the freezing point are also the same for both solutions.

Q11.5.5

What are the mole fractions of H_3PO_4 and water in a solution of 14.5 g of H_3PO_4 in 125 g of water?

Q11.5.6

What are the mole fractions of HNO_3 and water in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?

S11.5.6

a. Find number of moles of HNO_3 and H_2O in 100 g of the solution. Find the mole fractions for the components.
b. The mole fraction of HNO_3 is 0.378. The mole fraction of H_2O is 0.622.

Q11.5.7

Calculate the mole fraction of each solute and solvent:

a. 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
b. 0.86 g of NaCl in 1.00×10^2 g of water—a solution of sodium chloride for intravenous injection
c. 46.85 g of codeine, $C_{18}H_{21}NO_3$, in 125.5 g of ethanol, C_2H_5OH
d. 25 g of I_2 in 125 g of ethanol, C_2H_5OH

S11.5.7

a.
$$583\ g\,H_2SO_4 \times \frac{1\ mole\,H_2SO_4}{98.08\ g\,H_2SO_4} = 5.94\ mole\,H_2SO_4$$

$$1.50\ kg\,H_2O \times \frac{1000\ g}{1\ kg} \times \frac{1\ mole\,H_2O}{18.02\ g\,H_2O} = 83.2\ moles\,H_2O$$

Q11.5.8

Calculate the mole fraction of each solute and solvent:

a. 0.710 kg of sodium carbonate (washing soda), Na_2CO_3, in 10.0 kg of water—a saturated solution at 0 °C
b. 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack
c. 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2
d. 0.372 g of histamine, C_5H_9N, in 125 g of chloroform, $CHCl_3$

S11.5.8

a. $X_{Na_2CO_3} = 0.0119$; $X_{H_2O} = 0.988$;
b. $X_{NH_4NO_3} = 0.9927$; $X_{H_2O} = 0.907$;
c. $X_{Cl_2} = 0.192$; $X_{CH_2CI_2} = 0.808$;
d. $X_{C_5H_9N} = 0.00426$; $X_{CHCl_3} = 0.997$

Q11.5.9

Calculate the mole fractions of methanol, CH_3OH; ethanol, C_2H_5OH; and water in a solution that is 40% methanol, 40% ethanol, and 20% water by mass. (Assume the data are good to two significant figures.)

Q11.5.10

What is the difference between a 1 *M* solution and a 1 *m* solution?

S11.5.10

In a 1 *M* solution, the mole is contained in exactly 1 L of solution. In a 1 *m* solution, the mole is contained in exactly 1 kg of solvent.

Q11.5.11

What is the molality of phosphoric acid, H_3PO_4, in a solution of 14.5 g of H_3PO_4 in 125 g of water?

Q11.5.12

What is the molality of nitric acid in a concentrated solution of nitric acid (68.0% HNO_3 by mass)?

S11.5.12

(a) Determine the molar mass of HNO_3. Determine the number of moles of acid in the solution. From the number of moles and the mass of solvent, determine the molality. (b) 33.7 *m*

Q11.5.13

Calculate the molality of each of the following solutions:

a. 583 g of H_2SO_4 in 1.50 kg of water—the acid solution used in an automobile battery
b. 0.86 g of NaCl in 1.00×10^2 g of water—a solution of sodium chloride for intravenous injection
c. 46.85 g of codeine, $C_{18}H_{21}NO_3$, in 125.5 g of ethanol, C_2H_5OH
d. 25 g of I_2 in 125 g of ethanol, C_2H_5OH

Q11.5.14

Calculate the molality of each of the following solutions:

a. 0.710 kg of sodium carbonate (washing soda), Na_2CO_3, in 10.0 kg of water—a saturated solution at 0°C
b. 125 g of NH_4NO_3 in 275 g of water—a mixture used to make an instant ice pack
c. 25 g of Cl_2 in 125 g of dichloromethane, CH_2Cl_2
d. 0.372 g of histamine, C_5H_9N, in 125 g of chloroform, $CHCl_3$

S11.5.14

(a) 6.70×10^{-1} *m*; (b) 5.67 *m*; (c) 2.8 *m*; (d) 0.0358 *m*

Q11.5.15

The concentration of glucose, $C_6H_{12}O_6$, in normal spinal fluid is $\frac{75\text{ mg}}{100\text{ g}}$. What is the molality of the solution?

Q11.5.16

A 13.0% solution of K_2CO_3 by mass has a density of 1.09 g/cm^3. Calculate the molality of the solution.

S11.5.16

1.08 *m*

Q11.5.17

a. Why does 1 mol of sodium chloride depress the freezing point of 1 kg of water almost twice as much as 1 mol of glycerin?
b. What is the boiling point of a solution of 115.0 g of sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water?

S11.5.17

a. Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the boiling point of water and the boiling point of the solution; determine the new boiling point.
b. 100.5 °C

Q11.5.18

What is the boiling point of a solution of 9.04 g of I_2 in 75.5 g of benzene, assuming the I_2 is nonvolatile?

Q11.5.19

What is the freezing temperature of a solution of 115.0 g of sucrose, $C_{12}H_{22}O_{11}$, in 350.0 g of water, which freezes at 0.0 °C when pure?

S11.5.19

(a) Determine the molar mass of sucrose; determine the number of moles of sucrose in the solution; convert the mass of solvent to units of kilograms; from the number of moles and the mass of solvent, determine the molality; determine the difference between the freezing temperature of water and the freezing temperature of the solution; determine the new freezing temperature. (b) −1.8 °C

Q11.5.20

What is the freezing point of a solution of 9.04 g of I_2 in 75.5 g of benzene?

Q11.5.21

What is the osmotic pressure of an aqueous solution of 1.64 g of $Ca(NO_3)_2$ in water at 25 °C? The volume of the solution is 275 mL.

S11.5.21

(a) Determine the molar mass of $Ca(NO_3)_2$; determine the number of moles of $Ca(NO_3)_2$ in the solution; determine the number of moles of ions in the solution; determine the molarity of ions, then the osmotic pressure. (b) 2.67 atm

Q11.5.22

What is osmotic pressure of a solution of bovine insulin (molar mass, 5700 g mol^{-1}) at 18 °C if 100.0 mL of the solution contains 0.103 g of the insulin?

Q11.5.23

What is the molar mass of a solution of 5.00 g of a compound in 25.00 g of carbon tetrachloride (bp 76.8 °C; K_b = 5.02 °C/*m*) that boils at 81.5 °C at 1 atm?

S11.5.24

(a) Determine the molal concentration from the change in boiling point and K_b; determine the moles of solute in the solution from the molal concentration and mass of solvent; determine the molar mass from the number of moles and the mass of solute. (b) 2.1×10^2 g mol^{-1}

Q11.5.25

A sample of an organic compound (a nonelectrolyte) weighing 1.35 g lowered the freezing point of 10.0 g of benzene by 3.66 °C. Calculate the molar mass of the compound.

Q11.5.26

A 1.0 *m* solution of HCl in benzene has a freezing point of 0.4 °C. Is HCl an electrolyte in benzene? Explain.

S11.5.26

No. Pure benzene freezes at 5.5 °C, and so the observed freezing point of this solution is depressed by ΔT_f = 5.5 − 0.4 = 5.1 °C. The value computed, assuming no ionization of HCl, is ΔT_f = (1.0 m)(5.14 °C/*m*) = 5.1 °C. Agreement of these values supports the assumption that HCl is not ionized.

Q11.5.27

A solution contains 5.00 g of urea, $CO(NH_2)_2$, a nonvolatile compound, dissolved in 0.100 kg of water. If the vapor pressure of pure water at 25 °C is 23.7 torr, what is the vapor pressure of the solution?

Q11.5.28

A 12.0-g sample of a nonelectrolyte is dissolved in 80.0 g of water. The solution freezes at −1.94 °C. Calculate the molar mass of the substance.

S11.5.28

144 g mol^{-1}

Q11.5.29

Arrange the following solutions in order by their decreasing freezing points: 0.1 *m* Na_3PO_4, 0.1 *m* C_2H_5OH, 0.01 *m* CO_2, 0.15 *m* NaCl, and 0.2 *m* $CaCl_2$.

Q11.5.30

Calculate the boiling point elevation of 0.100 kg of water containing 0.010 mol of NaCl, 0.020 mol of Na_2SO_4, and 0.030 mol of $MgCl_2$, assuming complete dissociation of these electrolytes.

S11.5.30

0.870 °C

Q11.5.31

How could you prepare a 3.08 *m* aqueous solution of glycerin, $C_3H_8O_3$? What is the freezing point of this solution?

Q11.5.32

A sample of sulfur weighing 0.210 g was dissolved in 17.8 g of carbon disulfide, CS_2 (K_b = 2.43 °C/*m*). If the boiling point elevation was 0.107 °C, what is the formula of a sulfur molecule in carbon disulfide?

S11.5.32

S_8

Q11.5.33

In a significant experiment performed many years ago, 5.6977 g of cadmium iodide in 44.69 g of water raised the boiling point 0.181 °C. What does this suggest about the nature of a solution of CdI_2?

Q11.5.34

Lysozyme is an enzyme that cleaves cell walls. A 0.100-L sample of a solution of lysozyme that contains 0.0750 g of the enzyme exhibits an osmotic pressure of 1.32×10^{-3} atm at 25 °C. What is the molar mass of lysozyme?

S11.5.34

1.39×10^4 g mol^{-1}

Q11.5.35

The osmotic pressure of a solution containing 7.0 g of insulin per liter is 23 torr at 25 °C. What is the molar mass of insulin?

Q11.5.36

The osmotic pressure of human blood is 7.6 atm at 37 °C. What mass of glucose, $C_6H_{12}O_6$, is required to make 1.00 L of aqueous solution for intravenous feeding if the solution must have the same osmotic pressure as blood at body temperature, 37 °C?

S11.5.36

54 g

Q11.5.37

What is the freezing point of a solution of dibromobenzene, $C_6H_4Br_2$, in 0.250 kg of benzene, if the solution boils at 83.5 °C?

Q11.5.38

What is the boiling point of a solution of NaCl in water if the solution freezes at −0.93 °C?

S11.5.38

100.26 °C

Q11.5.39

The sugar fructose contains 40.0% C, 6.7% H, and 53.3% O by mass. A solution of 11.7 g of fructose in 325 g of ethanol has a boiling point of 78.59 °C. The boiling point of ethanol is 78.35 °C, and K_b for ethanol is 1.20 °C/*m*. What is the molecular formula of fructose?

Q11.5.40

The vapor pressure of methanol, CH_3OH, is 94 torr at 20 °C. The vapor pressure of ethanol, C_2H_5OH, is 44 torr at the same temperature.

a. Calculate the mole fraction of methanol and of ethanol in a solution of 50.0 g of methanol and 50.0 g of ethanol.
b. Ethanol and methanol form a solution that behaves like an ideal solution. Calculate the vapor pressure of methanol and of ethanol above the solution at 20 °C.

c. Calculate the mole fraction of methanol and of ethanol in the vapor above the solution.

S11.5.40

(a) $X_{CH_3OH} = 0.590$; $X_{C_2H_5OH} = 0.410$; (b) Vapor pressures are: CH_3OH: 55 torr; C_2H_5OH: 18 torr; (c) CH_3OH: 0.75; C_2H_5OH: 0.25

Q11.5.41

The triple point of air-free water is defined as 273.15 K. Why is it important that the water be free of air?

Q11.5.42

Meat can be classified as fresh (not frozen) even though it is stored at −1 °C. Why wouldn't meat freeze at this temperature?

S11.5.42

The ions and compounds present in the water in the beef lower the freezing point of the beef below −1 °C.

Q11.5.43

An organic compound has a composition of 93.46% C and 6.54% H by mass. A solution of 0.090 g of this compound in 1.10 g of camphor melts at 158.4 °C. The melting point of pure camphor is 178.4 °C. K_f for camphor is 37.7 °C/*m*. What is the molecular formula of the solute? Show your calculations.

Q11.5.44

A sample of $HgCl_2$ weighing 9.41 g is dissolved in 32.75 g of ethanol, C_2H_5OH (K_b = 1.20 °C/*m*). The boiling point elevation of the solution is 1.27 °C. Is $HgCl_2$ an electrolyte in ethanol? Show your calculations.

S11.5.44

$$\Delta \text{bp} = K_\text{b} m = (1.20\ ^\circ\text{C}/m)\left(\frac{9.41\ \text{g} \times \frac{1\ \text{mol HgCl}_2}{271.496\ \text{g}}}{0.03275\ \text{kg}}\right) = 1.27\ ^\circ\text{C}$$

The observed change equals the theoretical change; therefore, no dissociation occurs.

Q11.5.45

A salt is known to be an alkali metal fluoride. A quick approximate determination of freezing point indicates that 4 g of the salt dissolved in 100 g of water produces a solution that freezes at about −1.4 °C. What is the formula of the salt? Show your calculations.

11.6: Colloids

Q11.6.1

Identify the dispersed phase and the dispersion medium in each of the following colloidal systems: starch dispersion, smoke, fog, pearl, whipped cream, floating soap, jelly, milk, and ruby.

S11.6.1

Colloidal System	Dispersed Phase	Dispersion Medium
starch dispersion	starch	water
smoke	solid particles	air
fog	water	air
pearl	water	calcium carbonate ($CaCO_3$)
whipped cream	air	cream
floating soap	air	soap

Colloidal System	Dispersed Phase	Dispersion Medium
jelly	fruit juice	pectin gel
milk	butterfat	water
ruby	chromium(III) oxide (Cr_2O_3)	aluminum oxide (Al_2O_3)

Q11.6.2

Distinguish between dispersion methods and condensation methods for preparing colloidal systems.

S11.6.2

Dispersion methods use a grinding device or some other means to bring about the subdivision of larger particles. Condensation methods bring smaller units together to form a larger unit. For example, water molecules in the vapor state come together to form very small droplets that we see as clouds.

Q11.6.3

How do colloids differ from solutions with regard to dispersed particle size and homogeneity?

S11.6.3

Colloidal dispersions consist of particles that are much bigger than the solutes of typical solutions. Colloidal particles are either very large molecules or aggregates of smaller species that usually are big enough to scatter light. Colloids are homogeneous on a macroscopic (visual) scale, while solutions are homogeneous on a microscopic (molecular) scale.

Q11.6.4

Explain the cleansing action of soap.

S11.6.4

Soap molecules have both a hydrophobic and a hydrophilic end. The charged (hydrophilic) end, which is usually associated with an alkali metal ion, ensures water solubility The hydrophobic end permits attraction to oil, grease, and other similar nonpolar substances that normally do not dissolve in water but are pulled into the solution by the soap molecules.

Q11.6.5

How can it be demonstrated that colloidal particles are electrically charged?

S11.6.5

If they are placed in an electrolytic cell, dispersed particles will move toward the electrode that carries a charge opposite to their own charge. At this electrode, the charged particles will be neutralized and will coagulate as a precipitate.

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 13.11: Exercises is shared under a CC BY license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

14: Fundamental Equilibrium Concepts

In this chapter, you will learn how to predict the position of the balance and the yield of a product of a reaction under specific conditions, how to change a reaction's conditions to increase or reduce yield, and how to evaluate an equilibrium system's reaction to disturbances.

14.1: Introduction
14.2: Chemical Equilibria
14.3: Equilibrium Constants
14.4: Shifting Equilibria - Le Chatelier’s Principle
14.5: Equilibrium Calculations
14.6: Key Terms
14.7: Key Equations
14.8: Summary
14.9: Exercises

This page titled 14: Fundamental Equilibrium Concepts is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

14.1: Introduction

Imagine a beach populated with sunbathers and swimmers. As those basking in the sun get too hot, they enter the surf to swim and cool off. As the swimmers tire, they return to the beach to rest. If the rate at which sunbathers enter the surf were to equal the rate at which swimmers return to the sand, then the numbers (though not the identities) of sunbathers and swimmers would remain constant. This scenario illustrates a dynamic phenomenon known as *equilibrium*, in which opposing processes occur at equal rates. Chemical and physical processes are subject to this phenomenon; these processes are at equilibrium when the forward and reverse reaction rates are equal. Equilibrium systems are pervasive in nature; the various reactions involving carbon dioxide dissolved in blood are examples (see Figure 14.1.1). This chapter provides a thorough introduction to the essential aspects of chemical equilibria.

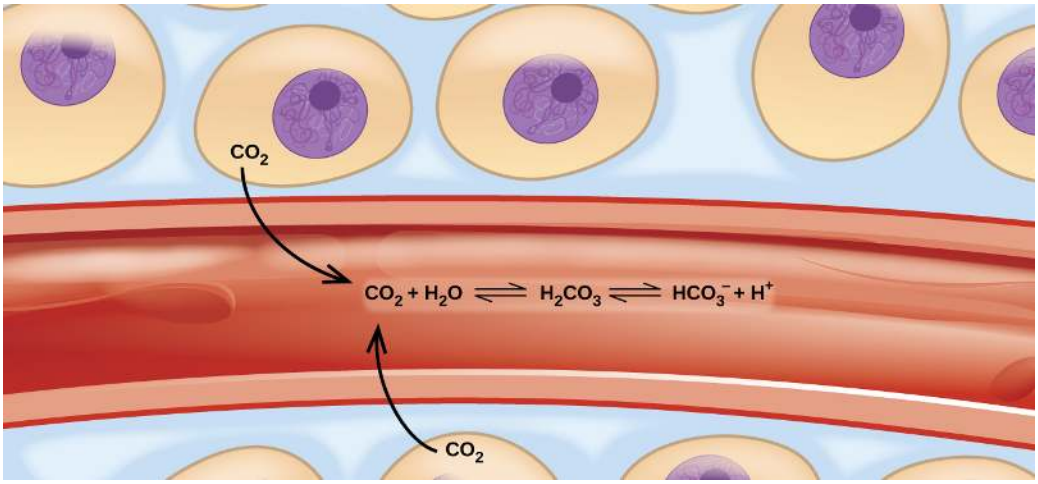


Figure 14.1.1: Transport of carbon dioxide in the body involves several reversible chemical reactions, including hydrolysis and acid ionization (among others).

This page titled 14.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.0: Introduction** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

14.2: Chemical Equilibria

Learning Objectives

By the end of this section, you will be able to:

- Describe the nature of equilibrium systems
- Explain the dynamic nature of a chemical equilibrium
- Relate the magnitude of an equilibrium constant to properties of the chemical system

The convention for writing chemical equations involves placing reactant formulas on the left side of a reaction arrow and product formulas on the right side. By this convention, and the definitions of “reactant” and “product,” a chemical equation represents the reaction in question as proceeding from left to right. **Reversible reactions**, however, may proceed in both forward (left to right) and reverse (right to left) directions. When the rates of the forward and reverse reactions are equal, the concentrations of the reactant and product species remain constant over time and the system is at **equilibrium**. The relative concentrations of reactants and products in equilibrium systems vary greatly; some systems contain mostly products at equilibrium, some contain mostly reactants, and some contain appreciable amounts of both.

Figure $14.2.1$ illustrates fundamental equilibrium concepts using the reversible decomposition of colorless dinitrogen tetroxide to yield brown nitrogen dioxide, an elementary reaction described by the equation:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Note that a special double arrow is used to emphasize the reversible nature of the reaction.

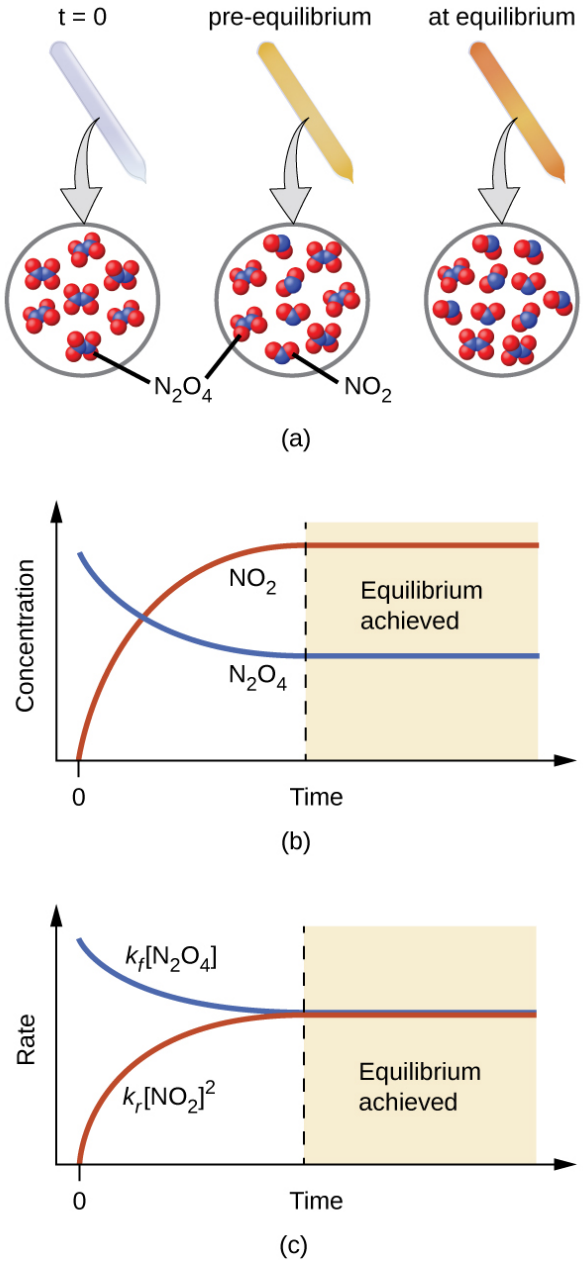


Figure 14.2.1: (a) A sealed tube containing colorless N_2O_4 darkens as it decomposes to yield brown NO_2. (b) Changes in concentration over time as the decomposition reaction achieves equilibrium. (c) At equilibrium, the forward and reverse reaction rates are equal.

For this elementary process, rate laws for the forward and reverse reactions may be derived directly from the reaction stoichiometry:

$$\text{rate}_f = k_f\,[N_2O_4]$$

$$\text{rate}_{\,r} = k_r[NO_2]^2$$

As the reaction begins (t = 0), the concentration of the N_2O_4 reactant is finite and that of the NO_2 product is zero, so the forward reaction proceeds at a finite rate while the reverse reaction rate is zero. As time passes, N_2O_4 is consumed and its concentration falls, while NO_2 is produced and its concentration increases (Figure 14.2.1*b*). The decreasing concentration of the reactant slows the forward reaction rate, and the increasing product concentration speeds the reverse reaction rate (Figure 14.2.1*c*). This process continues until *the forward and reverse reaction rates become equal*, at which time the reaction has reached equilibrium, as characterized by constant concentrations of its reactants and products (shaded areas of Figure 14.2.1*b* and Figure 14.2.1*c*). It's important to emphasize that chemical equilibria are dynamic; a reaction at equilibrium has not "stopped," but is proceeding in the

forward and reverse directions at the same rate. This dynamic nature is essential to understanding equilibrium behavior as discussed in this and subsequent chapters of the text.

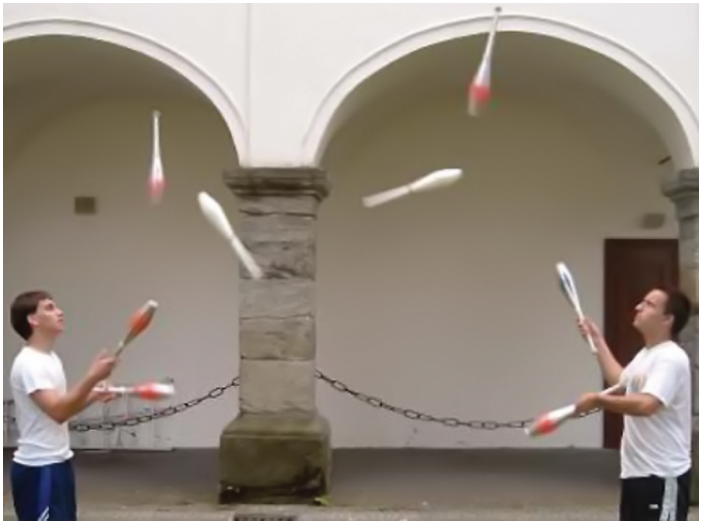


Figure 14.2.2: A two-person juggling act illustrates the dynamic aspect of chemical equilibria. Each person is throwing and catching clubs at the same rate, and each holds a (approximately) constant number of clubs.

Physical changes, such as phase transitions, are also reversible and may establish equilibria. This concept was introduced in another chapter of this text through discussion of the vapor pressure of a condensed phase (liquid or solid). As one example, consider the vaporization of bromine:

$$\mathrm{Br_2(l) \rightleftharpoons Br_2(g)}$$

When liquid bromine is added to an otherwise empty container and the container is sealed, the forward process depicted above (vaporization) will commence and continue at a roughly constant rate as long as the exposed surface area of the liquid and its temperature remain constant. As increasing amounts of gaseous bromine are produced, the rate of the reverse process (condensation) will increase until it equals the rate of vaporization and equilibrium is established. A photograph showing this phase transition equilibrium is provided in Figure 14.2.3



Figure 14.2.3: A sealed tube containing an equilibrium mixture of liquid and gaseous bromine. (credit: http://images-of-elements.com/bromine.php)

This page titled 14.2: Chemical Equilibria is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.1: Chemical Equilibria** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

14.3: Equilibrium Constants

Learning Objectives

By the end of this section, you will be able to:

- Derive reaction quotients from chemical equations representing homogeneous and heterogeneous reactions
- Calculate values of reaction quotients and equilibrium constants, using concentrations and pressures
- Relate the magnitude of an equilibrium constant to properties of the chemical system

The status of a reversible reaction is conveniently assessed by evaluating its reaction quotient (Q). For a reversible reaction described by

$$mA + nB + \rightleftharpoons xC + yD$$

the reaction quotient is derived directly from the stoichiometry of the balanced equation as

$$Q_c = \frac{[C]^x[D]^y}{[A]^m[B]^n}$$

where the subscript c denotes the use of molar concentrations in the expression. If the reactants and products are gaseous, a reaction quotient may be similarly derived using partial pressures:

$$Q_p = \frac{P_C^x P_D^y}{{P_A}^m {P_B}^n}$$

Note that the reaction quotient equations above are a simplification of more rigorous expressions that use *relative* values for concentrations and pressures rather than *absolute* values. These relative concentration and pressure values are dimensionless (they have no units); consequently, so are the reaction quotients. For purposes of this introductory text, it will suffice to use the simplified equations and to disregard units when computing Q. In most cases, this will introduce only modest errors in calculations involving reaction quotients.

Example 14.3.1: Writing Reaction Quotient Expressions

Write the concentration-based reaction quotient expression for each of the following reactions:

a. $3\,O_2(g) \rightleftharpoons 2\,O_3(g)$
b. $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$
c. $4\,NH_3(g) + 7\,O_2(g) \rightleftharpoons 4\,NO_2(g) + 6\,H_2O(g)$

Solution

a. $$Q_c = \frac{[O_3]^2}{[O_2]^3}$$

b. $$Q_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$$

c. $$Q_c = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^7}$$

Exercise 14.3.1

Write the concentration-based reaction quotient expression for each of the following reactions:

a. $2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$
b. $C_4H_8(g) \rightleftharpoons 2\,C_2H_4(g)$
c. $2\,C_4H_{10}(g) + 13\,O_2(g) \rightleftharpoons 8\,CO_2(g) + 10\,H_2O(g)$

Answer

a. $$Q_c = \frac{[SO_3]^2}{[SO_2]^2\,[O_2]}$$

b. $$Q_c = \frac{[C_2H_4]^2}{[C_4H_8]}$$

c. $$Q_c = \frac{[CO_2]^8[H_2O]^{10}}{[C_4H_{10}]^2[O_2]^{13}}$$

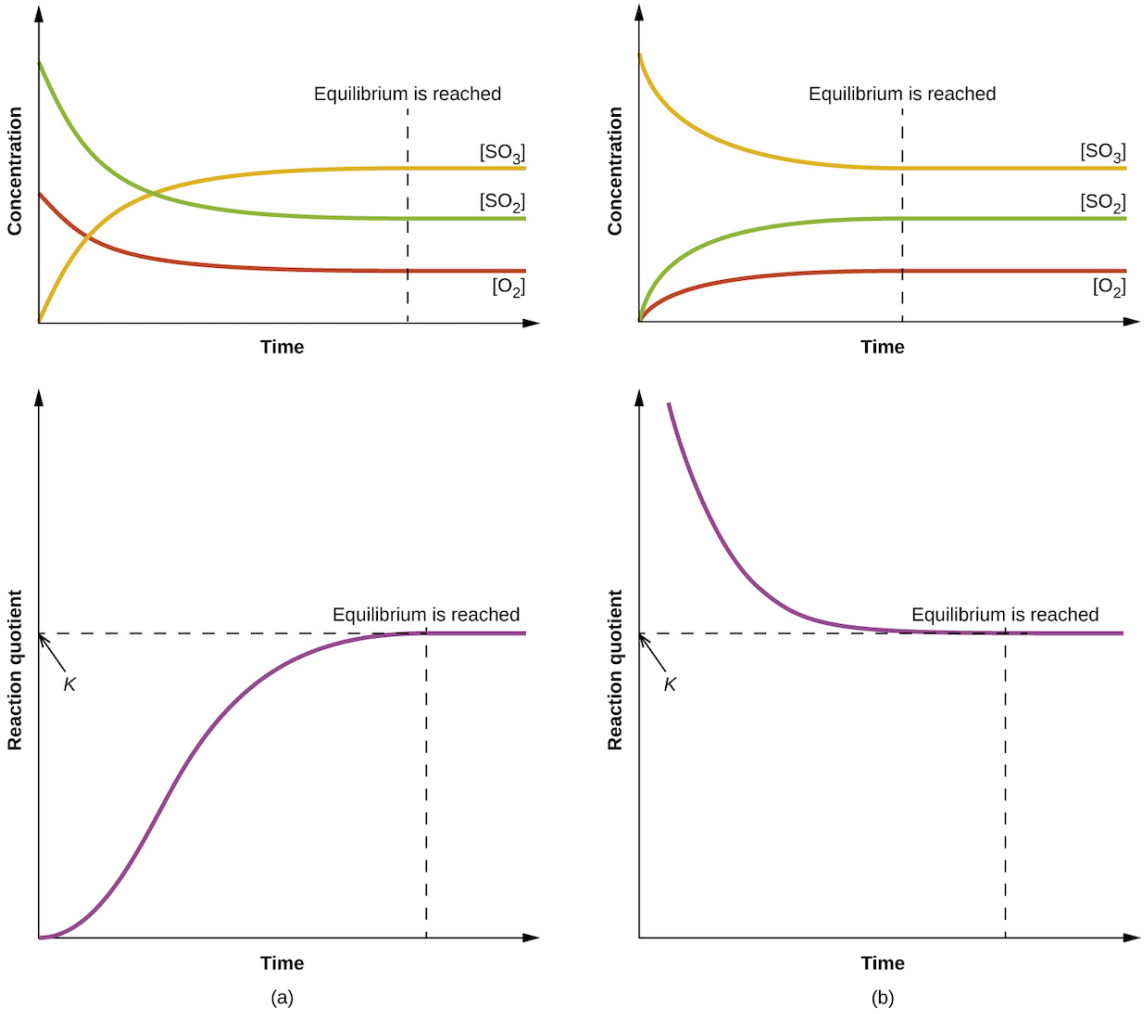


Figure 14.3.1 Changes in concentrations and Q_c for a chemical equilibrium achieved beginning with (a) a mixture of reactants only and (b) products only.

The numerical value of Q varies as a reaction proceeds towards equilibrium; therefore, it can serve as a useful indicator of the reaction's status. To illustrate this point, consider the oxidation of sulfur dioxide:

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$

Two different experimental scenarios are depicted in Figure 14.3.1, one in which this reaction is initiated with a mixture of reactants only, SO_2 and O_2, and another that begins with only product, SO_3. For the reaction that begins with a mixture of reactants only, Q is initially equal to zero:

$$Q_c = \frac{\left[\mathrm{SO_3}\right]^2}{\left[\mathrm{SO_2}\right]^2\left[\mathrm{O_2}\right]} = \frac{0^2}{\left[\mathrm{SO_2}\right]^2\left[\mathrm{O_2}\right]} = 0$$

As the reaction proceeds toward equilibrium in the forward direction, reactant concentrations decrease (as does the denominator of Q_c), product concentration increases (as does the numerator of Q_c), and the reaction quotient consequently increases. When equilibrium is achieved, the concentrations of reactants and product remain constant, as does the value of Q_c.

If the reaction begins with only product present, the value of Q_c is initially undefined (immeasurably large, or infinite):

$$Q_c = \frac{\left[\mathrm{SO_3}\right]^2}{\left[\mathrm{SO_2}\right]^2\left[\mathrm{O_2}\right]} = \frac{\left[\mathrm{SO_3}\right]^2}{0} \to \infty$$

In this case, the reaction proceeds toward equilibrium in the reverse direction. The product concentration and the numerator of Q_c decrease with time, the reactant concentrations and the denominator of Q_c increase, and the reaction quotient consequently decreases until it becomes constant at equilibrium.

The reaction quotient Q is not temperature dependent as it represents the ratio of concentrations (or partial pressures for gases) at a given time during the reaction.

The constant value of Q exhibited by a system at equilibrium is called the equilibrium constant, K:

$$K \equiv Q \text{ at equilibrium}$$

Comparison of the data plots in Figure $14.3.1$shows that both experimental scenarios resulted in the same value for the equilibrium constant. This is a general observation for all equilibrium systems, known as the **law of mass action**.

Definition: Law of Mass Action

At a given temperature, the reaction quotient for a system at equilibrium is constant.

Example 14.3.2: Evaluating a Reaction Quotient

Gaseous nitrogen dioxide forms dinitrogen tetroxide according to this equation:

$$2\,\mathrm{NO_2(g)} \rightleftharpoons \mathrm{N_2O_4(g)}$$

When 0.10 mol NO_2 is added to a 1.0-L flask at 25 °C, the concentration changes so that at equilibrium, [NO_2] = 0.016 *M* and [N_2O_4] = 0.042 *M*.

a. What is the value of the reaction quotient before any reaction occurs?
b. What is the value of the equilibrium constant for the reaction?

Solution

As for all equilibrium calculations in this text, use the simplified equations for Q and K and disregard any concentration or pressure units, as noted previously in this section.

(a) Before any product is formed

$$\left[\mathrm{NO_2}\right] = \frac{0.10\text{ mol}}{1.0\text{ L}} = 0.10\text{ M}$$

$$\left[\mathrm{N_2O_4}\right] = 0\text{ M}$$

Thus

$$Q_c = \frac{\left[\mathrm{N_2O_4}\right]}{\left[\mathrm{NO_2}\right]^2} = \frac{0}{0.10^2} = 0$$

(b) At equilibrium,

$$K_c = Q_c = \frac{[N_2O_4]}{[NO_2]^2} = \frac{0.042}{0.016^2} = 1.6 \times 10^2.$$

The equilibrium constant is 1.6×10^2.

Exercise 14.3.2

For the reaction

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$

the equilibrium concentrations are $[SO_2]$ = 0.90 *M*, $[O_2]$ = 0.35 *M*, and $[SO_3]$ = 1.1 *M*. What is the value of the equilibrium constant, K_c?

Answer

K_c = 4.3

By its definition, the magnitude of an equilibrium constant explicitly reflects the composition of a reaction mixture at equilibrium, and it may be interpreted with regard to the extent of the forward reaction. A reaction exhibiting a large *K* will reach equilibrium when most of the reactant has been converted to product, whereas a small *K* indicates the reaction achieves equilibrium after very little reactant has been converted. It's important to keep in mind that the magnitude of *K* does *not* indicate how rapidly or slowly equilibrium will be reached. Some equilibria are established so quickly as to be nearly instantaneous, and others so slowly that no perceptible change is observed over the course of days, years, or longer. The equilibrium constant is reported at a given temperature since the concentrations (or partial pressures) of the reaction components will change with temperature.

The equilibrium constant for a reaction can be used to predict the behavior of mixtures containing its reactants and/or products. As demonstrated by the sulfur dioxide oxidation process described above, a chemical reaction will proceed in whatever direction is necessary to achieve equilibrium. Comparing *Q* to *K* for an equilibrium system of interest allows prediction of what reaction (forward or reverse), if any, will occur.

To further illustrate this important point, consider the reversible reaction shown below:

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g) \quad K_c = 0.640 \quad T = 800^\circ C$$

The bar charts in Figure 14.3.2 represent changes in reactant and product concentrations for three different reaction mixtures. The reaction quotients for mixtures 1 and 3 are initially lesser than the reaction's equilibrium constant, so each of these mixtures will experience a net forward reaction to achieve equilibrium. The reaction quotient for mixture 2 is initially greater than the equilibrium constant, so this mixture will proceed in the reverse direction until equilibrium is established.

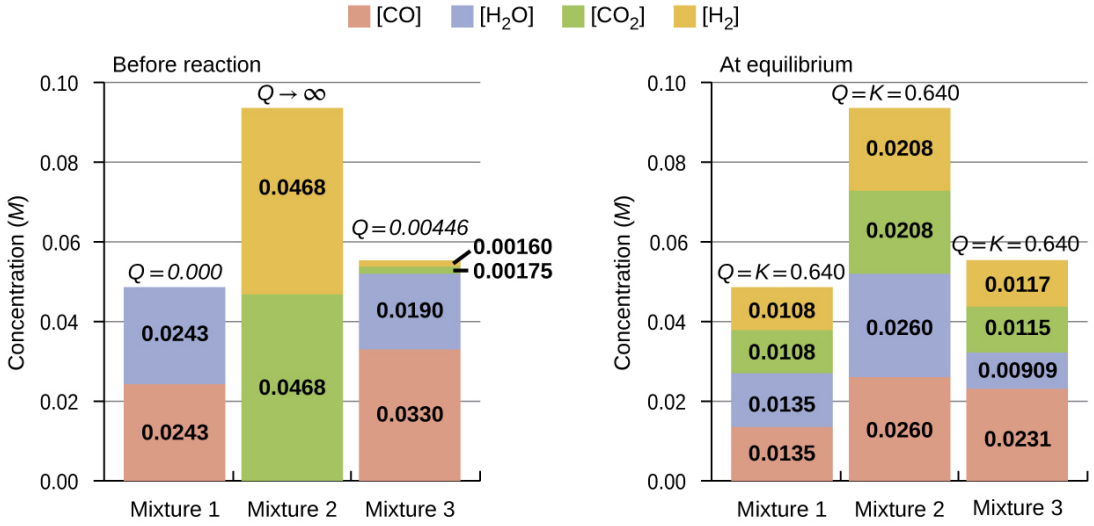


Figure 14.3.2: Compositions of three mixtures before ($Q_c \neq K_c$) and after ($Q_c = K_c$) equilibrium is established for the reaction

Example 14.3.3: Predicting the Direction of Reaction

Given here are the starting concentrations of reactants and products for three experiments involving this reaction:

$$\mathrm{CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)} \quad K_c = 0.64$$

Determine in which direction the reaction proceeds as it goes to equilibrium in each of the three experiments shown.

Reactants/Products	Experiment 1	Experiment 2	Experiment 3
$[CO]_i$	0.020 *M*	0.011 *M*	0.0094 *M*
$[H_2O]_i$	0.020 *M*	0.0011 *M*	0.0025 *M*
$[CO_2]_i$	0.0040 *M*	0.037 *M*	0.0015 *M*
$[H_2]_i$	0.0040 *M*	0.046 *M*	0.0076 *M*

Solution

Experiment 1:

$$Q_c = \frac{[\mathrm{CO_2}]\,[\mathrm{H_2}]}{[\mathrm{CO}]\,[\mathrm{H_2O}]} = \frac{(0.0040)(0.0040)}{(0.020)(0.020)} = 0.040$$

$Q_c < K_c$ (0.040 < 0.64)

The reaction will proceed in the forward direction.

Experiment 2:

$$Q_c = \frac{[\mathrm{CO_2}]\,[\mathrm{H_2}]}{[\mathrm{CO}]\,[\mathrm{H_2O}]} = \frac{(0.037)(0.046)}{(0.011)(0.0011)} = 1.4 \times 10^2$$

$Q_c > K_c$ (140 > 0.64)

The reaction will proceed in the reverse direction.

Experiment 3:

$$Q_c = \frac{[\mathrm{CO_2}]\,[\mathrm{H_2}]}{[\mathrm{CO}]\,[\mathrm{H_2O}]} = \frac{(0.0015)(0.0076)}{(0.0094)(0.0025)} = 0.48$$

$Q_c < K_c$ (0.48 < 0.64)

The reaction will proceed in the forward direction.

Exercise 14.3.3

Calculate the reaction quotient and determine the direction in which each of the following reactions will proceed to reach equilibrium.

a. A 1.00-L flask containing 0.0500 mol of NO(g), 0.0155 mol of Cl_2(g), and 0.500 mol of NOCl:

$$\mathrm{2\,NO(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g)} \quad K_c = 4.6 \times 10^4$$

b. A 5.0-L flask containing 17 g of NH_3, 14 g of N_2, and 12 g of H_2:

$$\mathrm{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)} \quad K_c = 0.060$$

c. A 2.00-L flask containing 230 g of SO_3(g):

$$\mathrm{2\,SO_3(g) \rightleftharpoons 2\,SO_2(g) + O_2(g)} \quad K_c = 0.230$$

Answer

(a) Q_c = 6.45 \times 10^{3} DELMAR, forward. (b) Q_c = 0.23, reverse. (c) Q_c = 0, forward.

Homogeneous Equilibria

A homogeneous equilibrium is one in which all reactants and products (and any catalysts, if applicable) are present in the same phase. By this definition, homogeneous equilibria take place in *solutions*. These solutions are most commonly either liquid or gaseous phases, as shown by the examples below:

$$\mathrm{C_2H_2(aq) + 2\,Br_2(aq) \rightleftharpoons C_2H_2Br_4(aq)} \qquad K_c = \frac{[C_2H_2Br_4]}{[C_2H_2]\,[Br_2]^2}$$

$$\mathrm{I_2(aq) + I^-(aq) \rightleftharpoons I_3^-(aq)} \qquad K_c = \frac{[I_3-}{[I_2]\,[I^-]}$$

$$\mathrm{HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)} \qquad K_c = \frac{[H_3O^+]\,[F^-]}{[HF]}$$

$$\mathrm{NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)} \qquad K_c = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}$$

These examples all involve aqueous solutions, those in which water functions as the solvent. In the last two examples, water also functions as a reactant, but its concentration is *not* included in the reaction quotient. The reason for this omission is related to the more rigorous form of the *Q* (or *K*) expression mentioned previously in this chapter, in which *relative concentrations for liquids and solids are equal to 1 and needn't be included*. Consequently, reaction quotients include concentration or pressure terms only for gaseous and solute species.

Note

It is a common error to claim that the molar concentration of the solvent is in some way involved in the equilibrium law. This error is a result of a misunderstanding of solution thermodynamics. For example, it is often claimed that $K_a = K_{eq}[H_2O]$ for aqueous solutions. This equation is incorrect because it is an erroneous interpretation of the correct equation $K_a = K_{eq}(a_{H_2O})$. Because a_{H_2O} = 1 for a dilute solution, $K_a = K_{eq}(1)$, or $K_a = K_{eq}$.

The equilibria below all involve gas-phase solutions:

$$\mathrm{C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)} \qquad K_c = \frac{[C_2H_4]\,[H_2]}{[C_2H_6]}$$

$$\mathrm{3\,O_2(g) \rightleftharpoons 2\,O_3(g)} \qquad K_c = \frac{[O_3]^2}{[O_2]^3}$$

$$\mathrm{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)} \qquad K_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$$

$$\mathrm{C_3H_8(g) + 5\,O_2(g) \rightleftharpoons 3\,CO_2(g) + 4\,H_2O(g)} \qquad K_c = \frac{[CO_2]^3[H_2O]^4}{[C_3H_8]\,[O_2]^5}$$

For gas-phase solutions, the equilibrium constant may be expressed in terms of either the molar concentrations (K_c) or partial pressures (K_p) of the reactants and products. A relation between these two *K* values may be simply derived from the ideal gas equation and the definition of molarity:

$$PV = nRT$$

$$P = \left(\frac{n}{V}\right)RT$$

$$= MRT$$

where P is partial pressure, V is volume, n is molar amount, R is the gas constant, T is temperature, and M is molar concentration.

For the gas-phase reaction

$$aA + bB \rightleftharpoons cC + dD$$

$$\begin{aligned} K_P &= \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b} \\ &= \frac{([C] \times RT)^c ([D] \times RT)^d}{([A] \times RT)^a ([B] \times RT)^b} \\ &= \frac{[C]^c [D]^d}{[A]^a [B]^b} \times \frac{(RT)^{c+d}}{(RT)^{a+b}} \\ &= K_c (RT)^{(c+d)-(a+b)} \\ &= K_c (RT)^{\Delta n} \end{aligned}$$

And so, the relationship between K_c and K_P is

$$K_P = K_c (RT)^{\Delta n}$$

where Δn is the difference in the molar amounts of product and reactant gases, in this case:

$$\Delta n = (c + d) - (a + b)$$

Example 14.3.4: Calculation of K_P

Write the equations relating K_c to K_P for each of the following reactions:

a. $\mathrm{C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)}$
b. $\mathrm{CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)}$
c. $\mathrm{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)}$
d. K_c is equal to 0.28 for the following reaction at 900 °C:

$$\mathrm{CS_2(g) + 4\,H_2(g) \rightleftharpoons CH_4(g) + 2\,H_2S(g)}$$

What is K_P at this temperature?

Solution

a. $\Delta n = (2) - (1) = 1$

$$K_P = K_c (RT)^{\Delta n} = K_c (RT)^1 = K_c (RT)$$

b. $\Delta n = (2) - (2) = 0$

$$K_P = K_c (RT)^{\Delta n} = K_c (RT)^0 = K_c$$

c. $\Delta n = (2) - (1 + 3) = -2$

$$K_P = K_c (RT)^{\Delta n} = K_c (RT)^{-2} = \frac{K_c}{(RT)^2}$$

d. $$K_P = K_c (RT)^{\Delta n} = (0.28)[(0.0821)(1173)]^{-2} = 3.0 \times 10^{-5}$$

Exercise 14.3.4

Write the equations relating K_c to K_P for each of the following reactions:

a. $\mathrm{2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)}$
b. $\mathrm{N_2O_4(g) \rightleftharpoons 2\,NO_2(g)}$

c. $C_3H_8(g) + 5\,O_2(g) \rightleftharpoons 3\,CO_2(g) + 4\,H_2O(g)$
d. At 227 °C, the following reaction has K_c = 0.0952:

$$CH_3OH(g) \rightleftharpoons CO(g) + 2\,H_2(g)$$

What would be the value of K_P at this temperature?

Answer

(a) $K_P = K_c\,(RT)^{-1}$; (b) $K_P = K_c\,(RT)$; (c) $K_P = K_c\,(RT)$; (d) 160 or 1.6 \times 10^{2} DELMAR

Heterogeneous Equilibria

A heterogeneous equilibrium involves reactants and products in two or more different phases, as illustrated by the following examples:

$$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Cl^-(aq) \qquad K_c = \left[Pb^{2+}\right]\left[Cl^-\right]^2$$

$$CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s) \qquad K_c = \frac{1}{[CO_2]}$$

$$C(s) + 2\,S(g) \rightleftharpoons CS_2(g) \qquad K_c = \frac{[CS_2]}{\left[S^2\right.}$$

$$Br_2(l) \rightleftharpoons Br_2(g) \qquad K_c = [Br_2(g)]$$

Again, note that concentration terms are only included for gaseous and solute species, as discussed previously.

Two of the above examples include terms for gaseous species only in their equilibrium constants, and so K_p expressions may also be written:

$$CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s) \qquad K_P = \frac{1}{P_{CO_2}}$$

$$C(s) + 2\,S(g) \rightleftharpoons CS_2(g) \qquad K_P = \frac{P_{CS_2}}{(P_S)^2}$$

Coupled Equilibria

The equilibrium systems discussed so far have all been relatively simple, involving just single reversible reactions. Many systems, however, involve two or more *coupled* equilibrium reactions, those which have in common one or more reactant or product species. Since the law of mass action allows for a straightforward derivation of equilibrium constant expressions from balanced chemical equations, the K value for a system involving coupled equilibria can be related to the K values of the individual reactions. Three basic manipulations are involved in this approach, as described below.

1. Changing the direction of a chemical equation essentially swaps the identities of “reactants” and “products,” and so the equilibrium constant for the reversed equation is simply the reciprocal of that for the forward equation.

$$A \rightleftharpoons B \quad K_c = \frac{[B]}{[A]}$$

$$B \rightleftharpoons A \quad K_{c'} = \frac{[A]}{[B]}$$

$$K_{c'} = \frac{1}{K_c}$$

2. Changing the stoichiometric coefficients in an equation by some factor x results in an exponential change in the equilibrium constant by that same factor:

$$A \rightleftharpoons B \qquad K_c = \frac{[B]}{[A]}$$

$$xA \rightleftharpoons xB \quad K_c = \frac{[B]^x}{[A]^x}$$

$$K_{c'} = {K_c}^x$$

3. Adding two or more equilibrium equations together yields an overall equation whose equilibrium constant is the mathematical product of the individual reaction's K values:

$$A \rightleftharpoons B \quad K_{c1} = \frac{[B]}{[A]}$$

$$B \rightleftharpoons C \quad K_{c2} = \frac{[C]}{[B]}$$

The net reaction for these coupled equilibria is obtained by summing the two equilibrium equations and canceling any redundancies:

$$A + B \rightleftharpoons B + C$$
$$A + B \rightleftharpoons B + C$$
$$A \rightleftharpoons C$$
$$K_{c'} = \frac{[C]}{[A]}$$

Comparing the equilibrium constant for the net reaction to those for the two coupled equilibrium reactions reveals the following relationship:

$$K_{c1}K_{c2} = \frac{[B]}{[A]} \times \frac{[C]}{[B]} = \frac{\cancel{[B]}[C]}{[A]\ \cancel{[B]}} = \frac{[C]}{[A]} = K_{c'}$$

$$K_{c'} = K_{c1}K_{c2}$$

Example $14.3.5$ demonstrates the use of this strategy in describing coupled equilibrium processes.

Example 14.3.5: Equilibrium Constants for Coupled Reactions

A mixture containing nitrogen, hydrogen, and iodine established the following equilibrium at 400 °C:

$$2\,\mathrm{NH_3(g)} + 3\,\mathrm{I_2(g)} \rightleftharpoons \mathrm{N_2(g)} + 6\,\mathrm{HI(g)}$$

Use the information below to calculate Kc for this reaction.

$$\mathrm{N_2(g)} + 3\,\mathrm{H_2(g)} \rightleftharpoons 2\,\mathrm{NH_3(g)}\ K_{c1} = 0.50 \text{ at } 400^\circ C$$
$$\mathrm{H_2(g)} + \mathrm{I_2(g)} \rightleftharpoons 2\,\mathrm{HI(g)}\ K_{c2} = 50 \text{ at } 400^\circ C$$

Solution

The equilibrium equation of interest and its K value may be derived from the equations for the two coupled reactions as follows.

Reverse the first coupled reaction equation:

$$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g) \quad {K_{c1}}' = \frac{1}{K_{c1}} = \frac{1}{0.50} = 2.0$$

Multiply the second coupled reaction by 3:

$$3H_2(g) + 3I_2(g) \rightleftharpoons 6HI(g) \quad {K_{c2}}' = K_{c2}^3 = 50^3 = 1.2 \times 10^5$$

Finally, add the two revised equations:

$$2\,\mathrm{NH_3(g)} + 3\,\mathrm{H_2(g)} + 3\,\mathrm{I_2(g)} \rightleftharpoons \mathrm{N_2(g)} + 3\,\mathrm{H_2(g)} + 6\,\mathrm{HI(g)}$$
$$2\,\mathrm{NH_3(g)} + 3\,\mathrm{I_2(g)} \rightleftharpoons \mathrm{N_2(g)} + 6\,\mathrm{HI(g)}$$
$$K_c = K_{c1}, K_{c2}, = (2.0)\left(1.2 \times 10^5\right) = 2.5 \times 10^5$$

Exercise 14.3.5

Use the provided information to calculate Kc for the following reaction at 550 °C:

$$\mathrm{H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g)} \qquad K_c = ?$$

$$\mathrm{CoO(s) + CO(g) \rightleftharpoons Co(s) + CO_2(g)} \qquad K_{c1} = 490$$

$$\mathrm{CoO(s) + H_2(g) \rightleftharpoons Co(s) + H_2O(g)} \qquad K_{c1} = 67$$

Answer

K_c = 0.14

This page titled 14.3: Equilibrium Constants is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.2: Equilibrium Constants** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

14.4: Shifting Equilibria - Le Chatelier's Principle

Learning Objectives

By the end of this section, you will be able to:

- Describe the ways in which an equilibrium system can be stressed
- Predict the response of a stressed equilibrium using Le Chatelier's principle

A system at equilibrium is in a state of dynamic balance, with forward and reverse reactions taking place at equal rates. If an equilibrium system is subjected to a change in conditions that affects these reaction rates differently (a *stress*), then the rates are no longer equal and the system is not at equilibrium. The system will subsequently experience a net reaction in the direction of greater rate (a *shift*) that will re-establish the equilibrium. This phenomenon is summarized by **Le Chatelier's principle**

Definition: Le Chatelier's Principle

if an equilibrium system is stressed, the system will experience a shift in response to the stress that re-establishes equilibrium.

Reaction rates are affected primarily by concentrations, as described by the reaction's rate law, and temperature, as described by the Arrhenius equation. Consequently, changes in concentration and temperature are the two stresses that can shift an equilibrium.

Effect of a Change in Concentration

If an equilibrium system is subjected to a change in the concentration of a reactant or product species, the rate of either the forward or the reverse reaction will change. As an example, consider the equilibrium reaction

$$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)} \quad K_c = 50.0 \text{ at } 400^\circ C$$

The rate laws for the forward and reverse reactions are

$$\begin{aligned} &\text{forward} \quad \mathrm{H_2(g) + I_2(g) \rightarrow 2\,HI(g)} && \text{rate}_f = k_f[\mathrm{H_2}]^m[\mathrm{I_2}]^n \\ &\text{reverse} \quad \mathrm{2\,HI(g) \rightarrow H_2(g) + I_2(g)} && \text{rate}_r = k_r[\mathrm{HI}]^x \end{aligned}$$

When this system is at equilibrium, the forward and reverse reaction rates are equal.

$$\text{rate}_f = \text{rate}_r$$

If the system is stressed by adding reactant, either $\mathrm{H_2}$ or $\mathrm{I_2}$, the resulting increase in concentration causes the rate of the forward reaction to increase, exceeding that of the reverse reaction:

$$\text{rate}_f > \text{rate}_r$$

The system will experience a temporary net reaction in the forward direction to re-establish equilibrium (the equilibrium will shift right). This same shift will result if some product HI is removed from the system, which decreases the rate of the reverse reaction, again resulting in the same imbalance in rates.

The same logic can be used to explain the left shift that results from either removing reactant or adding product to an equilibrium system. These stresses both result in an increased rate for the reverse reaction

$$\text{rate}_f < \text{rate}_r$$

and a temporary net reaction in the reverse direction to re-establish equilibrium.

As an alternative to this kinetic interpretation, the effect of changes in concentration on equilibria can be rationalized in terms of reaction quotients. When the system is at equilibrium,

$$Q_c = \frac{[\mathrm{HI}]^2}{[\mathrm{H_2}]\,[\mathrm{I_2}]} = K_c$$

If reactant is added (increasing the denominator of the reaction quotient) or product is removed (decreasing the numerator), then $Q_c < K_c$ and the equilibrium will shift right. Note that the three different ways of inducing this stress result in three different changes in the composition of the equilibrium mixture. If H_2 is added, the right shift will consume I_2 and produce HI as equilibrium is re-established, yielding a mixture with a greater concentrations of H_2 and HI and a lesser concentration of I_2 than was present before. If I_2 is added, the new equilibrium mixture will have greater concentrations of I_2 and HI and a lesser concentration of H_2. Finally, if HI is removed, the concentrations of all three species will be lower when equilibrium is reestablished. Despite these differences in composition, *the value of the equilibrium constant will be the same after the stress as it was before* (per the law of mass action). The same logic may be applied for stresses involving removing reactants or adding product, in which case $Q_c > K_c$ and the equilibrium will shift left.For gas-phase equilibria such as this one, some additional perspectives on changing the concentrations of reactants and products are worthy of mention. The partial pressure P of an ideal gas is proportional to its molar concentration M,

$$M = \frac{n}{V} = \frac{P}{RT}$$

and so changes in the partial pressures of any reactant or product are essentially changes in concentrations and thus yield the same effects on equilibria. Aside from adding or removing reactant or product, the pressures (concentrations) of species in a gas-phase equilibrium can also be changed by *changing the volume occupied by the system*. Since

all species of a gas-phase equilibrium occupy the same volume, a given change in volume will cause the same change in concentration for both reactants and products. In order to discern what shift, if any, this type of stress will induce the stoichiometry of the reaction must be considered.

At equilibrium, the reaction

$$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)}$$

is described by the reaction quotient

$$Q_P = \frac{(P_{\mathrm{HI}})^2}{P_{\mathrm{H_2}} P_{\mathrm{I_2}}} = K_p$$

If the volume occupied by an equilibrium mixture of these species is decreased by a factor of 3, the partial pressures of all three species will be increased by a factor of 3:

$$\begin{aligned} Q_{p'} &= \frac{(3P_{HI})^2}{3P_{H_2} 3P_{I_2}} = \frac{9P_{H^2}}{9P_{H_2} P_{I_2}} = \frac{P_{HI^2}}{P_{H_2} P_{I_2}} \\ &= Q_P = K_P \end{aligned}$$

And so, changing the volume of this gas-phase equilibrium mixture does not result in a shift of the equilibrium.

A similar treatment of a different system,

$$\mathrm{2\,NO_2(g) \rightleftharpoons 2\,NO(g) + O_2(g)}$$

however, yields a different result:

$$Q_P = \frac{(P_{NO})^2}{(P_{O_2})} (P_{NO_2})^2$$

$$\begin{aligned} Q'_P &= \frac{(3P_{\mathrm{NO}})^2 (3P_{\mathrm{O_2}})}{3(P_{\mathrm{NO_2}})^2} = \frac{9(P_{\mathrm{NO}})^2 (3P_{\mathrm{O_2}})}{9(P_{\mathrm{NO_2}})^2} = \frac{27(P_{\mathrm{NO}})^2 P_{\mathrm{O_2}}}{9(P_{\mathrm{NO_2}})^2} \\ &= 3Q_P > K_P \end{aligned}$$

In this case, the change in volume results in a reaction quotient greater than the equilibrium constant, and so the equilibrium will shift left.

These results illustrate the relationship between the stoichiometry of a gas-phase equilibrium and the effect of a volume-induced pressure (concentration) change. If the total molar amounts of reactants and products are equal, as in the first example, a change in volume does not shift the equilibrium. If the molar amounts of reactants and products are different, a change in volume will shift the equilibrium in a direction that better “accommodates” the volume change. In the second example, two moles of reactant (NO_2)

yield three moles of product (2NO + O_2), and so decreasing the system volume causes the equilibrium to shift left since the reverse reaction produces less gas (2 mol) than the forward reaction (3 mol). Conversely, increasing the volume of this equilibrium system would result in a shift towards products.

Link to Learning

Check out this link to see a dramatic visual demonstration of how equilibrium changes with pressure changes.

Chemistry in Everyday Life: Equilibrium and Soft Drinks

The connection between chemistry and carbonated soft drinks goes back to 1767, when Joseph Priestley (1733–1804) developed a method of infusing water with carbon dioxide to make carbonated water. Priestley's approach involved production of carbon dioxide by reacting oil of vitriol (sulfuric acid) with chalk (calcium carbonate).

The carbon dioxide was then dissolved in water, reacting to produce hydrogen carbonate, a weak acid that subsequently ionized to yield bicarbonate and hydrogen ions:

$$\text{dissolution} \qquad \mathrm{CO_2(g) \rightleftharpoons CO_2(aq)}$$

$$\text{hydrolysis} \qquad \mathrm{CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)}$$

$$\text{ionization} \qquad \mathrm{H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)}$$

These same equilibrium reactions are the basis of today's soft-drink carbonation process. Beverages are exposed to a high pressure of gaseous carbon dioxide during the process to shift the first equilibrium above to

the right, resulting in desirably high concentrations of dissolved carbon dioxide and, per similar shifts in the other two equilibria, its hydrolysis and ionization products. A bottle or can is then nearly filled with the carbonated beverage, leaving a relatively small volume of air in the container above the beverage surface (the *headspace*) before it is sealed. The pressure of carbon dioxide in the container headspace is very low immediately after sealing, but it rises as the dissolution equilibrium is re-established by shifting to the left. Since the volume of the beverage is significantly greater than the volume of the headspace, only a relatively small amount of dissolved carbon dioxide is lost to the headspace.

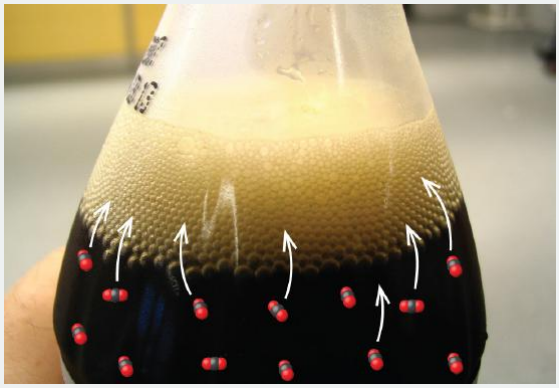


Figure 14.4.1: Opening a soft-drink bottle lowers the CO_2 pressure above the beverage, shifting the dissolution equilibrium and releasing dissolved CO_2 from the beverage. (credit: modification of work by "D Coetzee"/Flickr)

When a carbonated beverage container is opened, a hissing sound is heard as pressurized CO_2 escapes from the headspace. This causes the dissolution equilibrium to shift left, resulting in a decrease in the concentration of dissolved CO_2 and subsequent left-shifts of the hydrolysis and ionization equilibria. Fortunately for the consumer, the dissolution equilibrium is usually re-established slowly, and so the beverage may be enjoyed while its dissolved carbon dioxide concentration remains palatably high. Once the equilibria are re-established, the $CO_2(aq)$ concentration will be significantly lowered, and the beverage acquires a characteristic taste referred to as "flat."

Effect of a Change in Temperature

Consistent with the law of mass action, an equilibrium stressed by a change in concentration will shift to re-establish equilibrium without any change in the value of the

equilibrium constant, K. When an equilibrium shifts in response to a temperature change, however, it is re-established with a different relative composition that exhibits a different value for the equilibrium constant.

To understand this phenomenon, consider the elementary reaction

$$A \rightleftharpoons B$$

Since this is an elementary reaction, the rates laws for the forward and reverse may be derived directly from the balanced equation's stoichiometry:

$$\text{rate}_f = k_f[A]$$
$$\text{rate}_r = k_r[B]$$

When the system is at equilibrium,

$$\text{rate}_r = \text{rate}_f$$

Substituting the rate laws into this equality and rearranging gives

$$k_f[A] = k_r[B]$$
$$\frac{[B]}{[A]} = \frac{k_f}{k_r} = K_c$$

The equilibrium constant is seen to be a mathematical function of the rate constants for the forward and reverse reactions. Since the rate constants vary with temperature as

described by the Arrhenius equation, is stands to reason that the equilibrium constant will likewise vary with temperature (assuming the rate constants are affected to different extents by the temperature change). For more complex reactions involving multistep reaction mechanisms, a similar but more complex mathematical relation exists between the equilibrium constant and the rate constants of the steps in the mechanism. Regardless of how complex the reaction may be, the temperature-dependence of its equilibrium constant persists.

Predicting the shift an equilibrium will experience in response to a change in temperature is most conveniently accomplished by considering the enthalpy change of the reaction. For example, the decomposition of dinitrogen tetroxide is an endothermic (heat-consuming) process:

$$\mathrm{N_2O_4(g) \rightleftharpoons 2\,NO_2(g)} \quad \Delta H = +57.20\ \mathrm{kJ}$$

For purposes of applying Le Chatelier's principle, heat (q) may be viewed as a reactant:

$$\mathrm{heat + N_2O_4(g) \rightleftharpoons 2\,NO_2(g)}$$

Raising the temperature of the system is akin to increasing the amount of a reactant, and so the equilibrium will shift to the right. Lowering the system temperature will likewise cause the equilibrium to shift left. For exothermic processes, heat is viewed as a product of the reaction and so the opposite temperature dependence is observed.

Effect of a Catalyst

The kinetics chapter of this text identifies a *catalyst* as a substance that enables a reaction to proceed via a different mechanism with an accelerated rate. The catalyzed reaction mechanism involves a lower energy transition state than the uncatalyzed reaction, resulting in a lower activation energy, E_a, and a correspondingly greater rate constant.

To discern the effect of catalysis on an equilibrium system, consider the reaction diagram for a simple one-step (elementary) reaction shown in Figure 14.4.2. The lowered transition state energy of the catalyzed reaction results in lowered activation energies for both the forward and the reverse reactions. Consequently, both forward and reverse reactions are accelerated, and equilibrium is achieved more quickly *but without a change in the equilibrium constant*.

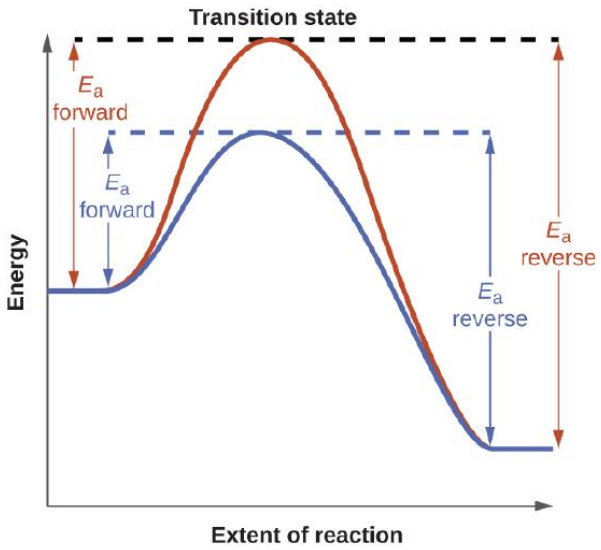


Figure 14.4.2 : Reaction diagrams for an elementary process in the absence (red) and presence (blue) of a catalyst. The presence of catalyst lowers the activation energies of both the forward and reverse reactions but does not affect the value of the equilibrium constant.

An interesting case study highlighting these equilibrium concepts is the industrial production of ammonia, NH_3. This substance is among the “top 10” industrial chemicals with regard to production, with roughly two billion pounds produced annually in the US. Ammonia is used as a chemical feedstock to synthesize a wide range of commercially useful compounds, including fertilizers, plastics, dyes, and explosives.

Most industrial production of ammonia uses the Haber-Bosch process based on the following equilibrium reaction:

$$\mathrm{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)} \quad \Delta H = -92.2\ \mathrm{kJ}$$

The traits of this reaction present challenges to its use in an efficient industrial process. The equilibrium constant is relatively small (Kp on the order of 10^{-5} at 25 °C), meaning very little ammonia is present in an equilibrium mixture. Also, the rate of this reaction is relatively slow at low temperatures. To raise the yield of ammonia, the industrial process is designed to operate under conditions favoring product formation:

- High pressures (concentrations) of reactants are used, ~150−250 atm, to shift the equilibrium right, favoring product formation.
- Ammonia is continually removed (collected) from the equilibrium mixture during the process, lowering its concentration and also shifting the equilibrium right.
- Although low temperatures favor product formation for this exothermic process, the reaction rate at low temperatures is inefficiently slow. A catalyst is used to accelerate the reaction to reasonable rates at relatively moderate temperatures (400−500 °C).

A diagram illustrating a typical industrial setup for production of ammonia via the Haber-Bosch process is shown in Figure 14.4.3

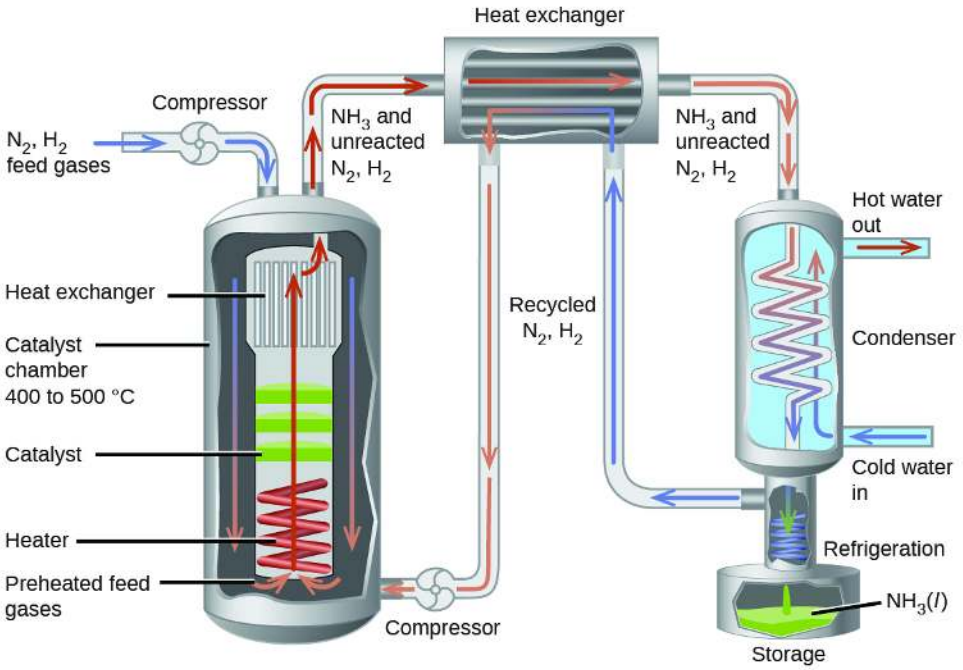


Figure 14.4.3: The figure shows a typical industrial setup for the commercial production of ammonia by the Haber-Bosch process. The process operates under conditions that stress the chemical equilibrium to favor product formation.

This page titled 14.4: Shifting Equilibria - Le Chatelier's Principle is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.3: Shifting Equilibria - Le Chatelier's Principle** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

14.5: Equilibrium Calculations

Learning Objectives

By the end of this section, you will be able to:

- Identify the changes in concentration or pressure that occur for chemical species in equilibrium systems
- Calculate equilibrium concentrations or pressures and equilibrium constants, using various algebraic approaches

Having covered the essential concepts of chemical equilibria in the preceding sections of this chapter, this final section will demonstrate the more practical aspect of using these concepts and appropriate mathematical strategies to perform various equilibrium calculations. These types of computations are essential to many areas of science and technology—for example, in the formulation and dosing of pharmaceutical products. After a drug is ingested or injected, it is typically involved in several chemical equilibria that affect its ultimate concentration in the body system of interest. Knowledge of the quantitative aspects of these equilibria is required to compute a dosage amount that will solicit the desired therapeutic effect.

Many of the useful equilibrium calculations that will be demonstrated here require terms representing changes in reactant and product concentrations. These terms are derived from the stoichiometry of the reaction, as illustrated by decomposition of ammonia:

$$2\,\mathrm{NH_3(g)} \rightleftharpoons \mathrm{N_2(g)} + 3\,\mathrm{H_2(g)}$$

As shown earlier in this chapter, this equilibrium may be established within a sealed container that initially contains either NH_3 only, or a mixture of any two of the three chemical species involved in the equilibrium. Regardless of its initial composition, a reaction mixture will show the same relationships between changes in the concentrations of the three species involved, as dictated by the reaction stoichiometry (see also the related content on expressing reaction rates in the chapter on kinetics). For example, if the nitrogen concentration increases by an amount *x*:

$$\Delta\left[N_2\right] = +x$$

the corresponding changes in the other species concentrations are

$$\Delta\left[\mathrm{H_2}\right] = \Delta\left[\mathrm{N_2}\right]\left(\frac{3\ \mathrm{mol\ H_2}}{1\ \mathrm{mol\ N_2}}\right) = +3x$$

$$\Delta\left[\mathrm{NH_3}\right] = -\Delta\left[\mathrm{N_2}\right]\left(\frac{2\ \mathrm{mol\ NH_3}}{1\ \mathrm{mol\ N_2}}\right) = -2x$$

where the negative sign indicates a decrease in concentration.

Example 14.5.1: Determining Relative Changes in Concentration

Derive the missing terms representing concentration changes for each of the following reactions.

a. $\underset{x}{\mathrm{C_2H_2(g)}} + 2\,\underset{?}{\mathrm{Br_2(g)}} \rightleftharpoons \underset{?}{\mathrm{C_2H_2Br_4(g)}}$

b. $\underset{?}{\mathrm{I_2(aq)}} + \underset{?}{\mathrm{I^-(aq)}} \rightleftharpoons \underset{x}{\mathrm{I_3^-(aq)}}$

c. $\underset{x}{\mathrm{C_3H_8(g)}} + 5\,\underset{?}{\mathrm{O_2(g)}} \rightleftharpoons 3\,\underset{?}{\mathrm{CO_2(g)}} + 4\,\underset{?}{\mathrm{H_2O(g)}}$

Solution

a. $\underset{x}{\mathrm{C_2H_2(g)}} + 2\,\underset{2x}{\mathrm{Br_2(g)}} \rightleftharpoons \underset{-x}{\mathrm{C_2H_2Br_4(g)}}$

b. $\underset{-x}{\mathrm{I_2(aq)}} + \underset{-x}{\mathrm{I^-(aq)}} \rightleftharpoons \underset{x}{\mathrm{I_3^-(aq)}}$

c. $\underset{x}{\mathrm{C_3H_8(g)}} + 5\,\underset{5x}{\mathrm{O_2(g)}} \rightleftharpoons 3\,\underset{-3x}{\mathrm{CO_2(g)}} + 4\,\underset{-4x}{\mathrm{H_2O(g)}}$

Exercise 14.5.1

Complete the changes in concentrations for each of the following reactions:

a. $\underset{?}{2\,SO_2}(g) + \underset{x}{O_2}(g) \rightleftharpoons \underset{?}{2\,SO_3}(g)$

b. $\underset{?}{C_4H_8}(g) \rightleftharpoons \underset{-2x}{2\,C_2H_4}(g)$

c. $\underset{?}{4\,NH_3}(g) + \underset{x}{7\,O_2}(g) \rightleftharpoons \underset{?}{4\,NO_2}(g) + \underset{?}{6\,H_2O}(g)$

Answer

(a) 2*x*, *x*, −2*x;* (b) *x*, −2*x;* (c) 4*x*, 7*x*, −4*x*, −6*x* or −4*x*, −7*x*, 4*x*, 6*x*

Calculation of an Equilibrium Constant

The equilibrium constant for a reaction is calculated from the equilibrium concentrations (or pressures) of its reactants and products. If these concentrations are known, the calculation simply involves their substitution into the K expression, as was illustrated by Example 14.5.1. A slightly more challenging example is provided next, in which the reaction stoichiometry is used to derive equilibrium concentrations from the information provided. The basic strategy of this computation is helpful for many types of equilibrium computations and relies on the use of terms for the reactant and product concentrations *initially* present, for how they *change* as the reaction proceeds, and for what they are when the system reaches *equilibrium*. The acronym ICE is commonly used to refer to this mathematical approach, and the concentrations terms are usually gathered in a tabular format called an ICE table.

Example 14.5.2: Calculation of an Equilibrium Constant

Iodine molecules react reversibly with iodide ions to produce triiodide ions.

$$I_2(aq) + I^-(aq) \rightleftharpoons I_3^-(aq)$$

If a solution with the concentrations of I_2 and I^- both equal to 1.000×10^{-3} M before reaction gives an equilibrium concentration of I_2 of 6.61×10^{-4} M, what is the equilibrium constant for the reaction?

Solution

To calculate the equilibrium constants, equilibrium concentrations are needed for all the reactants and products:

$$K_C = \frac{[I_3^-]}{[I_2]\,[I^-]}$$

Provided are the initial concentrations of the reactants and the equilibrium concentration of the product. Use this information to derive terms for the equilibrium concentrations of the reactants, presenting all the information in an ICE table.

	I_2 + I^- ⇌ I_3^-		
Initial concentration (*M*)	1.000×10^{-3}	1.000×10^{-3}	0
Change (*M*)	$-x$	$-x$	$+x$
Equilibrium concentration (*M*)	$1.000 \times 10^{-3} - x$	$1.000 \times 10^{-3} - x$	x

At equilibrium the concentration of I_2 is 6.61×10^{-4} M so that

$$\begin{aligned} 1.000 \times 10^{-3} - x &= 6.61 \times 10^{-4} \\ x &= 1.000 \times 10^{-3} - 6.61 \times 10^{-4} \\ &= 3.39 \times 10^{-4} M \end{aligned}$$

The ICE table may now be updated with numerical values for all its concentrations:

	I_2 + I^- ⇌ I_3^-		
Initial concentration (*M*)	1.000×10^{-3}	1.000×10^{-3}	0
Change (*M*)	-3.39×10^{-4}	-3.39×10^{-4}	$+3.39 \times 10^{-4}$
Equilibrium concentration (*M*)	6.61×10^{-4}	6.61×10^{-4}	3.39×10^{-4}

Finally, substitute the equilibrium concentrations into the *K* expression and solve:

$$\begin{aligned} K_c &= \frac{\left[I_3^-\right]}{\left[I_2\right]\left[I^-\right]} \\ &= \frac{3.39 \times 10^{-4} M}{\left(6.61 \times 10^{-4} M\right)\left(6.61 \times 10^{-4} M\right)} = 776 \end{aligned}$$

? Exercise 14.5.2

Ethanol and acetic acid react and form water and ethyl acetate, the solvent responsible for the odor of some nail polish removers.

$$C_2H_5OH + CH_3CO_2H \rightleftharpoons CH_3CO_2C_2H_5 + H_2O$$

When 1 mol each of C_2H_5OH and CH_3CO_2H are allowed to react in 1 L of the solvent dioxane, equilibrium is established when $\frac{1}{3}$ mol of each of the reactants remains. Calculate the equilibrium constant for the reaction. (Note: Water is a solute in this reaction.)

Answer

$K_c = 4$

Calculation of a Missing Equilibrium Concentration

When the equilibrium constant and all but one equilibrium concentration are provided, the other equilibrium concentration(s) may be calculated. A computation of this sort is illustrated in the next example exercise.

✓ Example 14.5.3: Calculation of a Missing Equilibrium Concentration

Nitrogen oxides are air pollutants produced by the reaction of nitrogen and oxygen at high temperatures. At 2000 °C, the value of the K_c for the reaction,

$$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$$

is 4.1×10^{-4}. Calculate the equilibrium concentration of NO(*g*) in air at 1 atm pressure and 2000 °C. The equilibrium concentrations of N_2 and O_2 at this pressure and temperature are 0.036 M and 0.0089 M, respectively.

Solution

Substitute the provided quantities into the equilibrium constant expression and solve for [NO]:

$$K_c = \frac{[NO]^2}{[N_2]\,[O_2]}$$

$$[NO]^2 = K_c\,[N_2]\,[O_2]$$

$$[NO] = \sqrt{K_c\,[N_2]\,[O_2]}$$

$$= \sqrt{\left(4.1 \times 10^{-4}\right)(0.036)(0.0089)}$$

$$= \sqrt{1.31 \times 10^{-7}}$$

$$= 3.6 \times 10^{-4}$$

Thus [NO] is 3.6×10^{-4} mol/L at equilibrium under these conditions.

To confirm this result, it may be used along with the provided equilibrium concentrations to calculate a value for *K*:

$$K_c = \frac{[NO]^2}{[N_2]\,[O_2]}$$

$$= \frac{\left(3.6 \times 10^{-4}\right)^2}{(0.036)(0.0089)}$$

$$= 4.0 \times 10^{-4}$$

This result is consistent with the provided value for *K* within nominal uncertainty, differing by just 1 in the least significant digit's place.

? Exercise 14.5.3

The equilibrium constant K_c for the reaction of nitrogen and hydrogen to produce ammonia at a certain temperature is 6.00×10^{-2}. Calculate the equilibrium concentration of ammonia if the equilibrium concentrations of nitrogen and hydrogen are 4.26 *M* and 2.09 *M*, respectively.

Answer

1.53 mol/L

Calculation of Equilibrium Concentrations from Initial Concentrations

Perhaps the most challenging type of equilibrium calculation can be one in which equilibrium concentrations are derived from initial concentrations and an equilibrium constant. For these calculations, a four-step approach is typically useful:

1. Identify the direction in which the reaction will proceed to reach equilibrium.
2. Develop an ICE table.
3. Calculate the concentration changes and, subsequently, the equilibrium concentrations.
4. Confirm the calculated equilibrium concentrations.

The last two example exercises of this chapter demonstrate the application of this strategy.

✔ Example 14.5.4: Calculation of Equilibrium Concentrations

Under certain conditions, the equilibrium constant K_c for the decomposition of $PCl_5(g)$ into $PCl_3(g)$ and $Cl_2(g)$ is 0.0211. What are the equilibrium concentrations of PCl_5, PCl_3, and Cl_2 in a mixture that initially contained only PCl_5 at a concentration of 1.00 *M*?

Solution

Use the stepwise process described earlier.

Step 1.

Determine the direction the reaction proceeds.

The balanced equation for the decomposition of PCl_5 is

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

Because only the reactant is present initially $Q_c = 0$ and the reaction will proceed to the right.

Step 2:

Develop an ICE table.

	$PCl_5 \rightleftharpoons$	PCl_3 +	Cl_2
Initial concentration (*M*)	1.00	0	0
Change (*M*)	$-x$	$+x$	$+x$
Equilibrium concentration (*M*)	$1.00 - x$	x	x

Step 3.

Solve for the change and the equilibrium concentrations.

Substituting the equilibrium concentrations into the equilibrium constant equation gives

$$\begin{aligned} K_c &= \frac{\left[PCl_3\right]\left[Cl_2\right]}{\left[PCl_5\right]} = 0.0211 \\ &= \frac{(x)(x)}{(1.00 - x)} \\ 0.0211 &= \frac{(x)(x)}{(1.00 - x)} \\ 0.0211(1.00 - x)7 &= x^2 \\ x^2 + 0.0211x - 0.0211 &= 0 \end{aligned}$$

Appendix B shows an equation of the form $ax^2 + bx + c = 0$ can be rearranged to solve for x:

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

In this case, $a = 1$, $b = 0.0211$, and $c = -0.0211$. Substituting the appropriate values for a, b, and c yields:

$$\begin{aligned} x &= \frac{-0.0211 \pm \sqrt{(0.0211)^2 - 4(1)(-0.0211)}}{2(1)} \\ &= \frac{-0.0211 \pm \sqrt{\left(4.45 \times 10^{-4}\right) + \left(8.44 \times 10^{-2}\right)}}{2} \\ &= \frac{-0.0211 \pm 0.291}{2} \end{aligned}$$

The two roots of the quadratic are, therefore,

$$x = \frac{-0.0211 + 0.291}{2} = 0.135$$

and

$$x = \frac{-0.0211 - 0.291}{2} = -0.156$$

For this scenario, only the positive root is physically meaningful (concentrations are either zero or positive), and so x = 0.135 M.

The equilibrium concentrations are

$$[\mathrm{PCl_5}] = 1.00 - 0.135 = 0.87\ \mathrm{M}$$
$$[\mathrm{PCl_3}] = x = 0.135\ \mathrm{M} =$$
$$[\mathrm{Cl_2}] = x = 0.135\ \mathrm{M}$$

Step 4.

Confirm the calculated equilibrium concentrations.

Substitution into the expression for K_c (to check the calculation) gives

$$K_c = \frac{[\mathrm{PCl_3}]\,[\mathrm{Cl_2}]}{[\mathrm{PCl_5}]} = \frac{(0.135)(0.135)}{0.87} = 0.021$$

The equilibrium constant calculated from the equilibrium concentrations is equal to the value of K_c given in the problem (when rounded to the proper number of significant figures).

Exercise 14.5.4A

Acetic acid, CH_3CO_2H, reacts with ethanol, C_2H_5OH, to form water and ethyl acetate, $CH_3CO_2C_2H_5$.

$$\mathrm{CH_3CO_2H + C_2H_5OH \rightleftharpoons CH_3CO_2C_2H_5 + H_2O}$$

The equilibrium constant for this reaction with dioxane as a solvent is 4.0. What are the equilibrium concentrations for a mixture that is initially 0.15 M in CH_3CO_2H, 0.15 M in C_2H_5OH, 0.40 M in $CH_3CO_2C_2H_5$, and 0.40 M in H_2O?

Answer

$[CH_3CO_2H]$ = 0.18 M, $[C_2H_5OH]$ = 0.18 M, $[CH_3CO_2C_2H_5]$ = 0.37 M, $[H_2O]$ = 0.37 M

Exercise 14.5.4B

A 1.00-L flask is filled with 1.00 mole of H_2 and 2.00 moles of I_2. The value of the equilibrium constant for the reaction of hydrogen and iodine reacting to form hydrogen iodide is 50.5 under the given conditions. What are the equilibrium concentrations of H_2, I_2, and HI in moles/L?

$$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)}$$

Answer

$[H_2]$ = 0.06 M, $[I_2]$ = 1.06 M, [HI] = 1.88 M

Example 14.5.5: Calculation of Equilibrium Concentrations Using an Algebra-Simplifying Assumption

What are the concentrations at equilibrium of a 0.15 M solution of HCN?

$$\mathrm{HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq)} \qquad K_c = 4.9 \times 10^{-10}$$

Solution

Using "x" to represent the concentration of each product at equilibrium gives this ICE table.

	$HCN(aq) \rightleftharpoons$	$H^+(aq)$ +	$CN^-(aq)$
Initial concentration (*M*)	0.15	0	0
Change (*M*)	$-x$	$+x$	$+x$
Equilibrium concentration (*M*)	$0.15 - x$	x	x

Substitute the equilibrium concentration terms into the K_c expression

$$K_c = \frac{(x)(x)}{0.15 - x}$$

rearrange to the quadratic form and solve for *x*

$$x^2 + 4.9 \times 10^{-10} - 7.35 \times 10^{-11} = 0$$

$$x = 8.56 \times 10^{-6} M (3 \text{ sig. figs. }) = 8.6 \times 10^{-6} M (2 \text{ sig. figs. })$$

Thus $[H^+] = [CN^-] = x = 8.6 \times 10^{-6}$ M and $[HCN] = 0.15 - x = 0.15$ M.

Note in this case that the change in concentration is significantly less than the initial concentration (a consequence of the small *K*), and so the initial concentration experiences a negligible change:

$$\text{if } x \ll 0.15M, \text{ then } (0.15 - x) \approx 0.15$$

This approximation allows for a more expedient mathematical approach to the calculation that avoids the need to solve for the roots of a quadratic equation:

$$K_c = \frac{(x)(x)}{0.15 - x} \approx \frac{x^2}{0.15}$$

$$4.9 \times 10^{-10} = \frac{x^2}{0.15}$$

$$x^2 = (0.15)\left(4.9 \times 10^{-10}\right)$$

$$= 7.4 \times 10^{-11}$$

$$x = \sqrt{7.4 \times 10^{-11}}$$

$$= 8.6 \times 10^{-6} \text{ M}$$

The value of x calculated is, indeed, much less than the initial concentration

$$8.6 \times 10^{-6} \ll 0.15$$

and so the approximation was justified. If this simplified approach were to yield a value for *x* that did *not* justify the approximation, the calculation would need to be repeated without making the approximation.

Exercise 14.5.5

What are the equilibrium concentrations in a 0.25 *M* NH_3 solution?

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \quad K_c = 1.8 \times 10^{-5}$$

Answer

$[OH^-] = [NH_4^+] = 0.0021$ M

$[NH_3] = 0.25$ M

This page titled 14.5: Equilibrium Calculations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **13.4: Equilibrium Calculations** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

14.6: Key Terms

equilibrium | state of a reversible reaction in which the forward and reverse processes occur at equal rates

equilibrium constant (*K*) | value of the reaction quotient for a system at equilibrium; may be expressed using concentrations (K_c) or partial pressures (K_p)

heterogeneous equilibria | equilibria in which reactants and products occupy two or more different phases

homogeneous equilibria | equilibria in which all reactants and products occupy the same phase

law of mass action | when a reversible reaction has attained equilibrium at a given temperature, the reaction quotient remains constant

Le Chatelier's principle | an equilibrium subjected to stress will shift in a way to counter the stress and re-establish equilibrium

reaction quotient (*Q*) | mathematical function describing the relative amounts of reactants and products in a reaction mixture; may be expressed in terms of concentrations (Q_c) or pressures (Q_p)

reversible reaction | chemical reaction that can proceed in both the forward and reverse directions under given conditions

---

This page titled 14.6: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.5: Key Terms** is licensed CC BY 4.0.

14.7: Key Equations

$Q_c = \frac{[\mathrm{C}]^x[\mathrm{D}]^y}{[\mathrm{A}]^m[\mathrm{B}]}\frac{n}{}$	$\mathrm{forthereaction}\ m\mathrm{A}+n\mathrm{B} \rightleftharpoons x\mathrm{C}+$
$Q_P = \frac{(P_C)^x(P_D)^y}{(P_A)^m(P_B)}\frac{n}{}$	$\mathrm{forthereaction}\ m\mathrm{A}+n\mathrm{B} \rightleftharpoons x\mathrm{C}+$
$P = MRT$	
$K_c = Q_c$ at equilibrium	
$K_p = Q_p$ at equilibrium	
$K_P = K_c\,(RT)^{\Delta n}$	

This page titled 14.7: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

14.8: Summary

A reversible reaction is at equilibrium when the forward and reverse processes occur at equal rates. Chemical equilibria are dynamic processes characterized by constant amounts of reactant and product species.

The composition of a reaction mixture may be represented by a mathematical function known as the reaction quotient, *Q*. For a reaction at equilibrium, the composition is constant, and *Q* is called the equilibrium constant, *K*.

A homogeneous equilibrium is an equilibrium in which all components are in the same phase. A heterogeneous equilibrium is an equilibrium in which components are in two or more phases.

Systems at equilibrium can be disturbed by changes to temperature, concentration, and, in some cases, volume and pressure. The system's response to these disturbances is described by Le Chatelier's principle: An equilibrium system subjected to a disturbance will shift in a way that counters the disturbance and re-establishes equilibrium. A catalyst will increase the rate of both the forward and reverse reactions of a reversible process, increasing the rate at which equilibrium is reached but not altering the equilibrium mixture's composition (*K* does not change).

Calculating values for equilibrium constants and/or equilibrium concentrations is of practical benefit to many applications. A mathematical strategy that uses initial concentrations, changes in concentrations, and equilibrium concentrations (and goes by the acronym ICE) is useful for several types of equilibrium calculations.

This page titled 14.8: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **13.7: Summary** is licensed CC BY 4.0.

14.9: Exercises

These are homework exercises to accompany the Textmap created for "General Chemistry: Principles and Modern Applications " by Petrucci et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Q14.1

Looking at the following, write the balanced equation and Kc expression for each reversible reaction:

a. the dissociation of water:
b. the reaction between gaseous CO and NO_2:
c. the addition of Hydrogen Iodide to water:

Q14.2

Balance the equilibrium equation and determine the corresponding K_p expression for each reversible reaction.

a. $N_2O_4(g) \rightleftharpoons NO(g)$
b. $Ag_2O(s) \rightleftharpoons Ag(s) + O_2(g)$
c. $Cs_2(g) + H_2(g) \rightleftharpoons CH_4(g) + H_2S(g)$

S14.2

a.) N2O4(g) ↔ **2**NO(g)

Kp = $(P_{NO})^2/(P_{N2O4})^1$

b.) Ag2O(s) ↔ **2**Ag(s) + **(1/2)**O2(g)

Kp = $(P_{Ag})^2$ x $(P_{O2})^{1/2}/(P_{Ag2O})$

c.) Cs2(g) + **4**H2(g) ↔ CH4(g) +**2** H2S(g)

Kp = $(P_{CH4})^1$ x $(P_{H2S})^2/(P_{Cs2})^1$ x $(P_{H2})^4$

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/The_Equilibrium_Constant/Calculating_An_Equilibrium_Concentration_From_An_Equilibrium_Constant/Writing_Equilibrium_Constant_Expressions_Involving_Gases/Gas_Equilibrium_Constants%3a_Kc_And_Kp

Q14.2

Based on the following, write a balanced equation and kp expression for each of the following:

a. $N_2(g) + H_2(g) \rightleftharpoons NH_3(g)$
b. The thermal decomposition of $CaCO_3$
c. The reaction of CO_2 with a sodium hydroxide solution

Q14.3

Give the K_c expressions for the following reactions:

a. $2HI(g) + O_2(g) \rightleftharpoons 2HIO_2(g)$
b. $ClNO_2(g) + NO(g) \rightleftharpoons NO_2(g) + ClNO(g)$
c. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

Q14.33

A mixture consisting of 0.201 mol H_2, and 0.201 mol of I_2 is brought to equilibrium at 500°C, in a 4.5 L flask. What are the equilibrium amounts of H_2, I_2, and HI? $K_c = 50.2$ at 500°C

S14.33

Video Solution

0.201 mol H2 /4.5 L = 0.0447 M H2 (First calculate molarity)

0.201 mol I2/4.5 L = 0.0447 M I2

$$H_2 + I_2 \rightleftharpoons 2HI \tag{14.9.1}$$

(Balance equation and set up ICE table)

ICE TABLE	H_2	I_2	HI
I	0.0447M	0.0447 M	0
C	-X	-X	+2X
E	0.0447-X	0.0447-X	2X

$$K_c = \frac{2X}{(0.0447 - X)^2} \tag{14.9.2}$$

$$50.2 = \frac{2X}{(0.0447 - X)^2} \tag{14.9.3}$$

(use quad formula)

X=0.0264

[H2] = 0.0183 M

[I2] = 0.0183 M

[HI] = 0.0528 M

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Ionization_Constants/Calculating_Equilibrium_Concentrations

Q14.43

Cadmium metal is added to 0.500 L of 1 M an aqueous solution of $[Cr^{3+}]$. What are the concentrations of the different ionic species at equilibrium for the oxidation of Cd into $Cd^{2+}_{(aq)}$:

$$2Cr^{3+}_{(aq)} + Cd_{(s)} \rightleftharpoons 2Cr^{2+}_{(aq)} + Cd^{2+}_{(aq)} \tag{14.9.4}$$

with $K_c = 0.288$.

Solution

The ICE table is first constructed

ICE	$[Cr^{3+}_{(aq)}]$ (M)	$Cr^{2+}_{(aq)}$ (M)	$Cd^{2+}_{(aq)}$ (M)
I	1	0	0
C	$-2x$	$+2x$	$+x$
E	$1-2x$	$2x$	x

Now construct the relationship for the equilibrium constant (do not include solid Cd in equilibrium):

$$K_c = \frac{[Cr^{2+}_{(aq)}]^2[Cd^{2+}_{(aq)}]}{[Cr^{3+}_{(aq)}]^2} \tag{1}$$

$$= \frac{(2x)^2(x)}{(1-2x)^2}$$

$$= \frac{4x^3}{(1-2x)^2} \tag{2}$$

Alternatively

$$K_c(1-2x)^2 = 4x^3 \tag{3}$$

$$K_c - 4K_cx + 4K_cx^2 = 4x^3 \tag{14.9.5}$$

$$4x^3 - 4K_cx^2 + 4K_cx - K_c = 0 \tag{4}$$

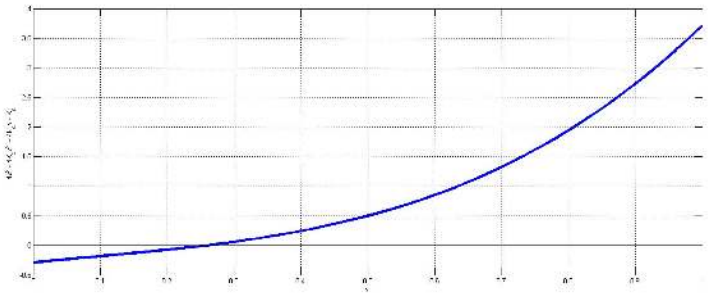
This is unfortunately is a cubic equation and can be solved analytically, but not easily. It is typically solved either graphically or numerically. Before solving numerical, we can evaluate one alternative by assuming that x is small ($x/2 << 1$), then the equation 2 can be approximated and the cubic equation becomes solvable:

$$= \frac{4x^3}{(1-2x)^2} \approx 4x^3 \tag{14.9.6}$$

$$x \approx \sqrt[3]{\frac{K_c}{4}} \tag{14.9.7}$$

$$x \approx \sqrt[3]{\frac{0.288}{4}} \approx 0.4160 \tag{5}$$

Unfortunately, this value for x is clearly not small and hence the approximate is incorrect. Hence, the solution requires a numerical approach to identify the roots of equation 4. By plotting $4x^3 - 4K_cx^2 + 4K_cx - K_c$ vs. x, the physical solution (i.e., root between $x = 0$ and $x = 1$) can be resolved graphically (below).



.

From this graph, is clear that $x \approx 0.2571$, which is decently far off from the crude approximation in equation 5.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Ionization_Constants/Calculating_Equilibrium_Concentrations

Q14.71

Use Le Chatelier's Principal to explain what happens to the following reaction when you *increase* the temperature. Why does this happen?

N2(g) + 3 H2(g) ↔ 2 NH3(g) ΔH = −92kJ/mol

S14.71

Increasing the temperature will cause the equilibrium to shift towards the reactants (left). This is because the reaction is exothermic and releases heat. So if you add heat, more N2(g) and H2(g) will be created and balance out the reaction.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/Le_Chatelier's_Principle

Q14.84

For the synthesis of ammonia at 500 K,

$$N_2(g) + 3H_2 \rightleftharpoons 2NH_3(g) \tag{14.9.8}$$

with $K_p = 11.08 \times 10^{-2}$. Assume that N_2 and H_2 are mixed in the mole ratio 1:3 and that the total pressure is maintained at 1.00 atm. What is the mole percent NH_3 at equilibrium? *Hint: Use this equation* → Kp = $((X_{NH3})^2/(X_{N2})(X_{H2})^2)$ x $(1/(P_{tot})^2)$

S14.84

11.08 x 10^-2 = $((X_{NH3})^2/(1)(3)^2$ x $(1/1)^2)$

X_{NH3} = .998 mol

0.998 mol NH3/6 tot mol x100 = **16.63% mol NH3**

http://chemwiki.ucdavis.edu/Physical_Chemistry/Chemical_Equilibrium/Calculating_an_Equilibrium_Concentration_from_the_Equilibrium_Constant/Calculating_an_Equilibrium_Constant%2c_Kp%2c_with_Partial_Pressures

Q14.29

Consider a mixture of 2.7 mol O_2, 3.2 mol SO_2, and 2.1 mol SO_3. Can this mixture be maintained in a 7.2 L container where Kc=20?

Q14.31

In the reaction $2ClO_2(g)+O_2(g)$<-->$2ClO_3(g)$, 0.241 mol ClO_2, 0.197 mol O_2, 0.670 mol ClO_3 are simultaneously placed into a 1.7 L vessel at a temperature of 1000K.

a. If Kc=2.53E2 is the mixture at equilibrium?
b. If it is not in equilibrium then in what direction will the reaction shift?

Q14.33

Consider a mixture of 0.263 mol H_2, and .263 mol Br_2that is broughtto equilibrium at 500 degrees Celsius in a 4.0 L flask. What are concentrations at equilibrium of H_2, Br_2, and HBr?

Q14.35

For the reversible reaction between CO and CO_2 in the gas phase

$$CO(g) + O_2(g) \rightleftharpoons CO_2(g) \tag{14.9.9}$$

If 0.352 mol of CO were added to 0.067 mol of CO_2 in a 3 L flask at 668 K, how many moles of $O_2(g)$ will be present at equilibrium? The K_c of this reaction is 1.2×10^3 at 668 K.

Q14.37

Equilibrium is reached in a 2.5 L flask at 800 °C for the reaction of

$$CaCO_3(s) \rightleftharpoons CaCO(s) + CO_2(g) \tag{14.9.10}$$

with $K_c = 2.68 \times 10^{-3}$

a. How many moles of $CaCO_3(s)$ is present at equilibrium if 0.715 mols $CaCO_3(g)$ is introduced into an empty flask?

Q14.39

Consider the reaction,$Ca(OH)_2$ + 2 CH_3COOH= $Ca(CH_3COO)_2$ + $2H_2O$; Kc=4.0

A reaction occurs in a mixture of 19.3 g $Ca(OH)_2$, 21.7 g CH_3COOH, 41.2 g $Ca(CH_3COO)_2$, and 70.1 g H_2O.

a. In what direction will the reaction proced?
b. How many grams of each will be present at equilibrium?

Q14.41

Given, the thermal decomposition of $CaCO_3(g)$ → $CaO(g)$ + $CO_2(g)$, Kc=4.84 @300K

if 0.362 mol $CaCO_3$ dissociatesin a 2.47 L flask at 300 K what will be the TOTAL Pressure at equilibrium?

Q14.43

write the equilibrium expression for the reaction [citrate(aq)]=.0027 M, [aconitate(aq)]=4.31E-5, and [H_2O]= 53.6M. Determine whether the reaction is at equilibrium / or which direction it will proceed if not at equilibrium.

Citrate(aq)ßàaconitate(aq)+H_2O(l) K=0.031

Q14.45

Based on the equation HNO3 (aq) --> H+ (aq) + NO3- (aq) ; K=3.21E11, if 2.0 M HNO3 is added to a .5 L flask then what are the resulting ion concentrations at equilibrium?

Q14.53

Based on the following equilibrium reaction 2H(g)+O(g)ßàH$_2$O(g) ; K=35

What is the partial pressure of Oxygen?

Q14.55

Given the following situations, determine what will happen to the amount of H$_2$that will be present in an equilibrium mixture in the reaction 3Fe(s)+4H$_2$O(g)<-->Fe$_3$O$_4$(s)+4H$_2$(g) ; ΔH˚=-150 kJ

a. Decreasing the temperature of the mixture
b. Introducing more Fe(s)
c. Halving the volume of the container
d. Adding a catalyst

63

Consider the reaction 4H$_2$(g)+2O$_2$(g)<-->2H$_2$O(g) ΔH˚=271 kJ

what affect does decreasing the temperature have on

a. The equilibrium production of NO(g)
b. The rate at which the reaction takes place

Q14.65

Based on the reaction N$_2$(g)+O$_2$(g)<-->2NO(g) what will happen to the Pressure of N$_2$ if the volume of the mixture were to be increased by decreasing the pressure by half?

Q14.71

A container of NOg) at 0.96 atm is brought into contact with 100 mL of a saturated aqueous solution with a NO$_2$ concentration of 5.53E-2 M. If the container is 1100 ml in volume at 30˚C, then a) what will be Kc for the equilibrium? And b) If .023 mol of ^{250}NO is added to the flask, then how many moles of ^{14}CO$_2$ (g) will form in the aqueous solution?

Q14.73

for the reaction V^{3+}(aq)+Cr^{2+}(aq)<-->V^{2+}(aq)+Cr^{3+}(aq) (Kc=7.2E2) with initial concentrations [V^{3+}]=[Cr^{2+}]=.0231 M and[V^{2+}]=[Cr^{3+}]=.260 M, what happens to the concentrations if the volume of the solution, 1L, is diluted to 2L with water?

Q14.75

Given the following generic reaction at 670 K what will be the total pressure at equilibrium if one starts with gas A at 0.837 atm?

$$2A(g) \rightleftharpoons 2B(g) + C(g) \tag{14.9.11}$$

with $K_p = 1.3 \times 10^{-4}$

Q14.77

Calculate the equilibrium amounts of the previous equation 2A<-->2B+C with a Kc= 1.5E2 if the initial [A]= 1.3 M.

Q14.79

based on the reaction 2SO$_2$(g)+O$_2$(g)<-->2SO$_3$(g) ΔH<0 , explain how each of the following will affect the amount of SO$_3$(g)

a. Increase pressure
b. Increase temperature
c. Increase the amount of O$_2$(g)

Q14.84

what will be the molar mass of the resulting gaseous mixture in the reaction $COCl_2(g)$<-->$CO(g)+Cl_2(g)$ given that Kp=4.44E-2 @122K and $P_{tot} = 2.5\ atm$?

Q14.102

given the following equilibrium equation, A+B<-->C with equal concentrations of A and C and a Kc=1000, what must the concentration of C be at equilibrium?

Q14.104

For the reaction \[2SO_2(g)+O_2(g) \rightletharpoons 2SO_3(g)\] with $K_p = 1.37 \times 10^{-2}$ at 1500˚C. What do you expect the relationship to be between K_c and K_p?

SolutionS

1) a) H20 <-->H^++OH^- ; kc=[H+][OH-]

b) Co(l)+NO2(l)<--> NO(l)+CO2(l) ; kc=[NO][CO2]/[CO][NO2]

c) HI(g)+H20(l) <-->H30+(aq)+I^-(aq) ; kc=[H30+][I-]/[HI]

2) a) $1N_2(g)+ 3H_2(g)$<-->$2NH_3(g)$ Kp= $P^2(NH_3)/P(N_2)P^3(H_2)$

b) $CaCO_3$ <--> CaO(s) + $CO_2(g)$ Kp= $P(CO_2)$

c) $CO_2(g)$ + NaOH(aq)<-->$NaHCO_3$(aq) Kp= $[NaHCO_3]/P(CO_2)[NaOH]$

3) a) Kc= $[HIO_2]^2/[HI]^2[O_2]$

b) Kc= $[NO_2][ClNO]/[ClNO_2][NO]$

c) Kc= $[HI]^2/[H_2][I_2]$

29) $2SO_2(g)$ + $O_2(g)$<--> $2SO_3(g)$

Kc=20=(2.1/7.2)/((2.7/7.2)*(3.2/7.2))=1.75<Kc=20

Therefore the mixture cannot be maintained indefinitely.

30) The reaction will proceed towards the products because Kc is much greater than one.

31)

a) Kc=2.53E2= (.67 mol/1.7 L)/((.241 mol/1.7 L)*(.197 mol/1.7 L)) = 23.99

23.99 is not equal to 2.53E2, therefor it is not at equilibrium.

b) Kc>Q therefore there will be a shift to the right, towards the products.

33) $H_2(g)+Br_2(g)$<-->$2HBr(g)$ Kc=50.2 @ 500 degrees C

[]I .263 .263 0

[]c -x -x +2x

[]e .263-x .263-x 0+2x 50.2= $((2x/4\ L)^2)/((.263-x/4\ L)^2)$

x= .205

$[H_2]_{eq}$=.263-.205= .058 M

$[Br_2]_{eq}$=.263-.205= .058 M

$[HBr]_{eq}$= 0+ .205 = .205 M

35) $CO(g)+O_2(g)$<-->$CO_2(g)$ Kc=$1.2*10^3$

[]I .352 0 .067

[]c +x +x -x

[]e .352+x x .067-x

Kc= 1.2E3=((.067-x/3L)/((.352+x/3 L)(x/3L))

X=5.15E-4= mols O_2

37)

a) $CaCO_3(g)$<-->$CaCO(g)+ CO_2(g)$ Kc=3.8E-2

[]I .731 .731 0

[]c -x +x +x

[]e .731-x .731+x x

Kc=3.8E-2=((x/2.5L)(.731+x/2.5L))/(.731-x/2.5L)

X=.268

$n(CaCO_3)$=.731-.268=.463 mols

n((CaCO)=.731+.268=.999 mols

$n(CO_2)$= x = .268 mols

b) $CaCO_3(g)$<-->$CaCO(g)+ CO_2(g)$ Kc=3.8E-2

[]I .715 0 0

[]c -x +x +x

[]e .715-x x x

Kc=3.8E-2=(x^2/2.5L)/(.715-x/2.5L)

X= .147

$n(CaCO_3)$=.715-.147=..568 mols

n((CaCO=.x= .147 mols

$n(CO_2)$= x = .147 mols

39)$Ca(OH)_2$ + 2 CH_3COOH= $Ca(CH_3COO)_2$ + $2H_2O$; Kc=4.0

a)

$n(Ca(OH)_2)$= 19.3g/74.093 g/mol=0.260 mols

$n(CH_3COOH)$=21.7g/60.05 g/mol=.361 mols

$n(Ca(CH_3COO)_2)$=41.2/158.1669 g/mol=.260 mols

$n(H_2O)$=70.1/18.02 g/mol=3.89 mol

Qc=((.260)(3.89))/((.361)(.260))= 10.78

Kc=4.0<10.78=Qc

Therefore the reaction will proceed to the left, towards reactants.

b)

$Ca(OH)_2$ + 2 CH_3COOH= $Ca(CH_3COO)_2$ + $2H_2O$; Kc=4.0

[]I .260 .361 .260 3.89

[]c +x +2x -x -2x

[]e .26+x .361+2x .26-x 3.89-2x

Kc=4.0=((.26-x)(3.89-2x))/((.26+x)(.361+2x))

X=0.184

Mass($Ca(OH)_2$)=.260+.184=.444(74.093)=32.9 g

Mass(CH_3COOH)=.361+2(.184)=.729(60.05)=43.77 g

Mass($Ca(CH_3COO)_2$)=.26-.184=.076(158.1669)=12.02 g

Mass(H2O)=3.89-2(.184)=3.522(18.02)=63.47 g

41) $CaCO_3(g) \rightarrow CaO(g) + CO_2(g)$, Kc=4.84

[]e .146-x x x

Kc=4.84=(x^2)/(.146-x)

X= .142

Equilibrium concentrations..

[$CaCO_3$]= .146-.142=0.004 M

[CaO]= .142 M

[CO_2]= .142 M

M_{tot}=.288

P_{tot}=(.288)(.08206)(300)=7.1

43)

Q=(4.31E-5)/.00217)=.0199

K>Q and therefore the reaction will proceed to the left (Not in initially in equilibrium)

45)

HNO3 is a strong acid and therefore dissociates into its ions completely. Hence the ion concentrations at equilibrium are [H+]=2M=[NO3-]

53)

P(O)= nRT/V=1mol x .08206 L atm $mol^{-1}K^{-1}$ atm x 600 K/1 =49.2 atm

55)

a) Decreasing the temperature of the mixture

-Because the reaction is exothermic heat can be thought of as a product and therefore, if the temperature decreases, then the amount of $H_2O(g)$ will increase.

b) Introducing more Fe(s)

-There should be no change in the amount of $H_2(g)$

c) Halving the volume of the container (increasing pressure)

-This will cause a shift from the side with the most moles of gas to the side with the least moles of gas. IN this case they are the same so there will be no change.

d) Adding a catalyst

-No change, catalysts only affect the speed at which the reaction occurs).

63)

a) Less $H_2O(g)$ formed at lower temperatures

b) The rate of the reaction slows with the decrease in temperature

65) The equilibrium will shift to the side with the most mols and therefore $P_{(N2)}$ should increase.

71)

a) [NO]=n/v=P/RT= .96/.0862*303=.039 M Kc=5.53E-2/.039 =1.418

NO(g)<-->NO(aq) Kc=1.418=((5.53E-2+x)/.1L)/((.062-x)/1L)

[]I .039 5.53E-2 x=.0029

+.023 -----

[]c -x +x

[]e (.062-x) 5.53E-2+x total mols (NO(g))=(1L)(.0591)=.0591 mol

total mols(NO(aq)=(.1L)(.0582)=.00582 mol

^{250}NO mol fraction= .01/.06492 x100= 15.4%=.154

mols ^{250}NO(g)=(.0591)(.154)=.0091

mols ^{250}NO(aq)=(.00582)(.154)= 8.96E-

Q14.73

The redox reaction in solution

$$V^{3+}(aq) + Cr^{2+}(aq) \rightleftharpoons V^{2+}(aq) + Cr^{3+}(aq) \tag{14.9.12}$$

has an equilibrium constant of $K_c = 7.2 \times 10^2$. If the reaction were initiated with with initial concentrations $[V^{3+}] = [Cr^{2+}] = 0.0231\ M$ and $[V^{2+}] = [Cr^{3+}] = 0.260\ M$, what happens to the concentrations if the volume of the solution, 1L, is diluted to 1.5L with water?

Equation: $V^{3+}(aq)+Cr^{2+}(aq)<-->V^{2+}(aq)+Cr^{3+}(aq)$ (Kc=7.2E2)

Equilib: .0231 .0231 .260 .260

Dilution: .01155 .01155 .13 .13

Change: -x -x +x +x

New eq: .01155-x .01155-x .13+x .13+x

7.2E2= $(.13+x)^2/(.01155-x)^2$

x=.0065

new concentrations :

$[V^{3+}]=[Cr^{2+}]$=.0231-.0065=.0166

$V^{2+}]=[Cr^{3+}]$=.260+.0065=.2665

Q14.75

$$2A(g) <--> 2B(g) + C(g) \tag{14.9.13}$$

$$Kp = 1.3E-4 \tag{14.9.14}$$

[]I .837 0 0

[]c -2x +2x +x

[]e (.837-2x) 2x x

Kp=1.3E-4=$((2x)^2(x))/(.837-2x)^2$

X=.8365 atm

Ptot= P(A)+P(B)+P(C)=(.837-2(.8365))+(2(.8365))+(.8365)=3.91atm

77) 2Aßà2B+C Kc= 1.5E2 [A]= 1.3 M

[]I 1.3 M 0 0

[]c -2x +2x +x

[]e 1.3-2x 2x x

Kc=1.5E-2=$(2x)^2(x)/(1.3-2x)^2$

X=.185

Equilibrium concentrations:

[A]=1.3-.185= 1.115 M

[B]=2(.185)= .37 M

[C]=.185 M

Q14.79

a) Increase pressure

-more SO_3 will be produced

b) Increase temperature

-SO_3 is reduced because heat is treated as a product in exothermic reactions

c) Increase the amount of $O_2(g)$

-more SO_3 will be produced in order to use up the excess oxygen

84)

$COCl_2(g)$<-->$CO(g)+Cl_2(g)$ given that Kp=4.44E-2 @122K

[]I P 0 0

[]c -x +x +x

[]e P-x x x

Kp=4.44E-2=(x*x)/(2.5-2x)

X=.292

Ptot=2.5 atm= P-x+x+x=P+x

P=2.5-x

$P(COCL_2)$=2.5-2x=1.916 atm

$P(CO)=P(CL_2)$=.292atm

Mavg= $(P(CO)*M(CO)/Ptot)+(P(Cl_2)*M(Cl_2)/Ptot)+(P(COCl_2)*M(COCl_2)/Ptot)$=

Mavg= 87.36 g/mol

Q102

A+BßàC Kc=1000, what must the concentration of C be at equilibrium?

$K=[C]^2/[A]^2[B]$ =1000

[B]=1/1000=.001 M

104) because $Kp=Kc(RT)^{\Delta n}$ we can determine that Kp>Kc

Q1

Based on these descriptions, write a balanced equation and the corresponding expression for each reversible reaction.

a. Carbonyl fluoride decomposes into gaseous carbon dioxide and gaseous carbon tetra fluoride.
b. Copper metal displaces silver (I) ion from aqueous solution, producing silver metal and an aqueous solution of copper (II) ion.
c. Peroxodisulfate ion oxidizes iron (II) ion to iron (III) ion in aqueous solution and is itself reduced to sulfate ion.

Q2

Based on these descriptions, write a balanced equation and the corresponding Kp expression for each reversible reaction.

a. Oxygen gas oxidizes gaseous ammonia to gaseous nitrogen and water vapor.
b. Hydrogen gas reduces gaseous nitrogen dioxide to gaseous ammonia and water vapor.

c. Nitrogen gas reacts with the solid sodium carbon- ate and carbon to produce solid sodium cyanide and carbon monoxide gas.

Q3

Write equilibrium constant expressions, K_c, for the reactions.

Q29

Can a mixture of 2.2 mol O_2, 3.6 mol SO_2, and 1.8 mol SO_3 be maintained indefinitely in a 7.2 L flask at a temperature at which Kc = 100 in this reaction?

Q30

Is a mixture of 0.0205 mol NO_2 (g) and 0.750 mol N_2O_4 (g) in a 5.25 L flask at 25 °C, at equilibrium? If not, in which direction will the reaction proceed toward products or reactants?

Q31

In the reaction 2 SO_2(g) + O_2 (g) à 2 SO_3(g), 0.455 mol SO_2, 0.183 mol O_2, and 0.568 mol SO_3 are introduced simultaneously into a 1.90 L vessel at 1000 K.

a. If Kc = 2.8 * 102, is this mixture at equilibrium?
b. If not, in which direction will a net change occur?

Q33

A mixture consisting of 0.150 mol H_2 and 0.150 mol I_2 is brought to equilibrium at 445 °C, in a 3.25 L flask. What are the equilibrium amounts of H_2, I_2, and HI?

Q35

Starting with 0.3500 mol CO (g) and 0.05500 mol $COCl_2$ (g) in a 3.050 L flask at 668 K, how many moles of Cl_2 (g) will be present at equilibrium?

Q37

Equilibrium is established in a 2.50 L flask at 250 degrees Celsius for the reaction

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) \tag{14.9.15}$$

with $K_c = 3.8 \times 10^{-2}$. How many moles of PCl_5, PCl_3, and Cl_2 are present at equilibrium, if

a. 0.550 mol each of PCl_5 and PCl_3 are initially introduced into the flask?
b. 0.610 mol PCl_5 alone is introduced into the flask?

Q39

In the following reaction, k_c=4.0.

$C_2H_5OH+CH_3COOH==CH_3COOC_2H_5+H_2O$

A reaction is allowed to occur in a mixture of 17.2 g C_2H_5OH, 23.8 g CH_3COOH, 48.6 g $CH_3COOC_2H_5$, and 71.2 g H2O.

a. In what direction will a net change occur?
b. How many grams of each substance will be present at equilibrium?

Q41

Formamide, used in the manufacture of pharmaceuticals, dyes, and agricultural chemicals, decomposes at high temperatures.

$$HCONH_2(g) \rightleftharpoons NH_3(g) + CO(g) \tag{14.9.16}$$

Kc = 4.84 at 400 K

If 0.186 mol $HCONH_2$(g) dissociates in a 2.16 L flask at 400 K, what will be the total pressure at equilibrium?

Q43

Cadmium metal is added to 0.350 L of an aqueous solution in which [Cr^{3+}] = 1.00 M. What are the concentrations of the different ionic species at equilibrium? What is the minimum mass of cadmium metal required to establish this equilibrium?

Q45

One sketch below represents an initial non-equilibrium mixture in the reversible reaction

SO_2 (g) + Cl_2 (g) == SO_2Cl_2 (g)

Kc = 4.0

Which of the other three sketches best represents an equilibrium mixture? Explain.

Q53

1.00 mol each of CO and Cl2 are introduced into an evacuated 1.75 L flask and the following equilibrium is established at 668 K.

CO (g) + Cl_2 (g) à $COCl_2$ (g)

Kp = 22.5

For this equilibrium, calculate (a) the partial pressure of $COCl_2$ (g); (b) the total gas pressure.

Q55

Continuous removal of one of the products of a chemical reaction has the effect of causing the reaction to go to completion. Explain this fact in terms of Le Chatelier's principle.

63

What effect does increasing the volume of the system have on the equilibrium condition in each of the following reactions?

Q71

Explain why the percent of molecules that dissociate into atoms in reactions of the type I_2(g) == 2I (g) always increases with an increase in temperature.

Q73

Refer to Example 15-13. Suppose that 0.100 L of the equilibrium mixture is diluted to 0.250 L with water. What will be the new concentrations when equilibrium is re-established?

Q75

Starting with SO_3 (g) at 1.00 atm, what will be the total pressure when equilibrium is reached in the following reaction at 700 K?

77

Derive, by calculation, the equilibrium amounts of SO_2, O_2, and SO_3 listed in (a) Figure 15-6(c); (b) Figure 15-7(b).

79

One of the key reactions in the gasification of coal is the methanation reaction, in which methane is produced from synthesis gas a mixture of CO and $\mathrm{H_2}$.

$$\mathrm{CO(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)}$$

$\Delta H = -230kJ$ and $K_c = 190$ at 1000 K

a. Is the equilibrium conversion of synthesis gas to methane favored at higher or lower temperatures? Higher or lower pressures?
b. Assume you have 4.00 mol of synthesis gas with a 3:1 mol ratio of $\mathrm{H_2(g)}$ to $\mathrm{CO(g)}$ in a 15.0 L flask. What will be the mole fraction of $\mathrm{CH_4(g)}$ at equilibrium at 1000 K?

84.

For the synthesis of ammonia at 500 K,

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \tag{14.9.17}$$

with $K_p = 9.06 \times 10^{-2}$. Assume that N_2 and H_2 are mixed in the molar ratio of $1 : 3$ in a 1 liter flask, respectively and that the total pressure is maintained at 1.00 atm. What is the mole percent NH_3 at equilibrium?

102.

Equilibrium is established in the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ at a temperature where Kc = 100. If the number of moles of SO_3 (g) in the equilibrium mixture is the same as the number of moles of SO_2 (g), (a) the number of moles of O_2 (g) is also equal to the number of moles of SO_2 (g); (b) the number of moles of O_2 (g) is half the number of moles of SO_2; (c) [O2] may have any of several values; (d) [O2] = 0.010 M.

104.

For the reaction

$$(2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g) \tag{14.9.18}$$

$K_c = 1.8 \times 10^{-6}$ at 184 °C.

At 184 °C, the value of K_c for the reaction

$$NO(g) + 3/2O_2(g) \rightleftharpoons NO_2(g) \tag{14.9.19}$$

is:

a. 0.9 * 10^6;
b. 7.5 * 10^2;
c. 5.6 * 10^5;
d. 2.8 * 10^5.

Solutions

S1

S2

Video Solution

S3

29. Cannot be maintained indefinitely in a 7.2L flask.

30. Shift forward to products.

31. No. Forward to product.

33. I_2: 0.0330mol

HI: 0.234mol

35. 4.0*10^{-4} mol

37. (a) PCl_3: 0.623mol

PCl_5: 0.478mol

(b) PCl_3: 0.20mol

PCl_5 0.41mol

39. (a) towards left

(b) 68g

41. 5.58atm

43. 10.1g

45. b

53. (a) 30.14 atm

(b) 32.4 atm

55. Continuous removal of the product of a chemical reaction has the effect of causing the reaction to go to completion that is which makes Q less than K and reaction shift the right as result and change in pressure of a constant temperature equilibrium mixture.

63. (a) Volume

(b) No effect

(c) More reactions

71. The percent of molecules that dissociates into atoms on reaction of the I_2 and 2I (g) always increase in temperature.

73. 2.98

75. 1.016 atm

S77.

a. 0.27
b. 0.085mol

S79

- Video Solution (Part A)
- Video Solution (Part B)

a. low temperatures and high pressures favor conversation of synthesis gas to methane.
b. 0.134

S84

~~28.06%~~

102

d

104. 750 (b)

1

Write a balanced equation that describes sodium chloride splits into Sodium and Chloride Ions.

$$NaCl \rightarrow Na^{+} + Cl^{-} \tag{14.9.20}$$

2

Write a balanced equation that describes the decomposition of Hydrochloric Acid into its Ions.

$$HCl \rightarrow H^{+} + Cl^{-} \tag{14.9.21}$$

3

Write equilibrium constant expressions, K_c, for the reactions

a. $C + 2H_2 \rightleftharpoons CH_4$
b. $2Na + Cl_2 \rightleftharpoons 2NaCl$
c. $NaHCO_3 \rightleftharpoons Na + HCO_3$

S3

(a) $K_c = [CH_4]/[H_2]^2$

(b)$K_c = [NaCl]^2/[Cl_2]$

(c) $K_c = [HCO_3]/[NaHCO_3]$

29.

There are two gases in a 4.5 L vessel. The mass of C_6H_{16} is 6.68g and the mass of $C_3H_8$is 2.28g. What is the value of Kc for this reaction.

S29

Mol C_6H_{16}= 6.68g (1mol C_6H_{16}/88.22g C_6H_{16})=0.0757mol

[C_6H_{16}]=0.0757mol/4.5L=0.0168M

Mol C_3H_8= 2.28g C_3H_8 x 1 mol C_3H_8/44.11g C_3H_8=0.0517mol

[C_3H_8]= 0.0517/4.5L= 0.0115M

Kc=[C_3H_8]/[C_6H_{16}]= (0.0115)/(0.0168)=0.685

30.

Is a mixture of 1.30 L of water and 3.15 mol of HCN and .001 mol of CN- in equilibrium? If not, in which direction will the shift proceed, towards products or reactants?

S30

(Ka (HCN) = 6.2e-10)

[HCN] = 3.15 mol / 1.30 L = 2.42 M [CN-] = [H+] = 7.69e-4 M

Q = [CN-] * [H+] / [HCN] = (7.69e-4)^2 / (2.42) = 2.45e-7

Q>K... therefore... reaction shifts left, towards the reactants.

31.

For the reaction in equilibrium,

$$H_2SO_4 + H_2O \rightleftharpoons HSO_4^- + H_3O^+ \tag{14.9.22}$$

if the K_a value for H_2SO_4 is 2.42×10^{-5}, the molar concentration of H_2SO_4 is 0.056 M, and the pH of the solution is 4.5, what is the concentration of HSO_4^-?

S31

[H_3O+] = 10^-4.5 = 3.16e-5 M

2.42e-5 = [HSO_4-] * (3.16e-5) / (.056)

[HSO_4-] = .0429 M

33

In an exothermic reaction increasing the temperature shifts the equilibrium in which direction?

S33

Left, more heat is more products so it shifts towards reactants.

35

If the number of moles remains constant, a change in volume does what to the equilibrium of a system?

S35

Nothing, K_{eq} also remains constant.

37

Using the above equilibrium reaction at 25 degrees Celsius with $K_c = 7.1 \times 10^2$

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \tag{14.9.23}$$

a. find the molarity of HI, assuming an equilibrium concentration of 0.76 M I_2, and 0.32 M H_2

b. find the molarity of I_2, assuming an equilibrium concentration of 2.36 M HI, and 0.42 M H_2

S37

a) Kc= [HI]^2 / [I2] x [H2]

7.1 x 10^2 = [HI]^2 / (.76 M x .32 M)

=13.14 M HI

b) Kc= [HI]^2 / [I2] x [H2]

7.1 x 10^2= (2.36M)^2 / (.42) x [I2]

(2.36M)^2 / ((.42) x (7.1 x 10^2) = [I2]

=0.018 M I2

39

In the following equilibrium reaction, $K_p = 3.16$:

$$2PO_4 \rightleftharpoons 2PO_3 + O_2 \tag{14.9.24}$$

A reaction occurs at 273 K in a 1.0 L container with 4.56 g of PO_4, 6.54 g of 2 PO_3 and 1.32 g of O_2.

a. In what direction will a net change occur?;
b. What will be the partial pressures of each substance at equilibrium?

S39

A)

n = (4.56 g PO_4)(1 mol / 94.97 g) = 0.0480 mol PO_4

n = (6.54 g PO_3)(1 mol / 78.97 g) = .0828 mol PO_3

n = (1.32 g O_2)(1 mol / 32 g) = .0413 mol O_2

PV=nRT

P_{PO4} = (.0480 mol)(.08206)(273 K) / (1.0 L) = 1.08 atm

P_{PO3} = (.0828 mol)(.08206)(273 K) / (1.0 L) = 1.85 atm

P_{O2} = (.0413 mol)(.08206)(273 K) / (1.0 L) = .925 atm

Q = (1.85)^2 * (.925) / (1.08)^2 = 2.71

2.71<3.16... Q<K... therefore... reaction shifts right, towards products.

B)

K = 3.16 = (1.85 + 2x)^2 * (.925+x) / (1.08 – 2x)^2

x = 0.0218

P_{PO4} = 1.08 – 2(.0218) = 1.04 atm

P_{PO3} = 1.85 + 2(.0218) = 1.89 atm

P_{O2} = .925 + .0218 = .947 atm

41.

Dinitrogen tetroxide, (N_2O_4) in combination with specific forms of hydrazine can burn on contact without a separate ignition source, make it popular as a component of rocket fuel. At high temperatures the equilibrium is:

$$N_2O_4 \rightleftharpoons 2NO_2 \tag{14.9.25}$$

What are the partial pressures at equilibrium at 83 Celsius and 1.1 atm, if the density of the two equilibrium products is 2.44 g/l?

S41

Knowns:

- p=1.1atm
- W= average molar mass = ?
- R= .08206 atm-L/mol-K
- T= 83+273K = 356K

Using the periodic table:

- N2O4 is 92 g/mol
- NO2 is 46 g/mol

Calculations:

p * W = density * RT

1.1 atm *W = (2.44 g/L) * (.08206 atm-L/mol-K) * (356K)

W = 64.8 g/mol

If X is the mol fraction of N2O4, then 1 - X is the mol fraction of NO2

92 g/mol * X + 46 g/mol * (1-X) = 64.8 g/mol

X = .4087

and

1-X = .5913

p(N2O4) = 0.4087 *1.1atm = .4496 atm

p(NO2) = 0.5913*1.1atm = .6504 atm

Therefore, the partial pressure of N2O4 is .4496 atm and the partial pressure of NO2 is .6504 atm.

43

What will a catalyst do to the composition of an equilibrium mixture?

Nothing, catalysts effect time, but once at equilibrium a system won't take effect to a catalyst.

45.

For the reaction

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad (14.9.26)$$

Kc = 50.3 at 731 K. Equal amounts (1.00 M each of the three components) are added to a beaker, and then the temperature is raised to 731 K.

Calculate the equilibrium concentrations.

S45

H2(g) + I2(g) = 2 HI(g)
1.00 1.00 1.00 initial concentration
-x -x 2x amount of change

1-x 1-x 1+2x equilibrium concentration

Kc=50.3 = [HI]2/[H2][I2] = (1+2x)2/(1-x)(1-x)

Therefore (1+4x+4x2)= 50.3*(1-2x-x2)=50.3-100.6x-50.3x2

-46.3x2+104.6x - 49.3 = 0

Solving the quadratic equation: The roots are: 1.589 and .670

Since x cannot be greater than 1, the only solution is x = .670.

Therefore, the equilibrium concentrations are:

[H2] = 1-.670 = .33

[I2] = 1-.670 = .33

[HI] = 1+2*.670 = 2.34

53

Determine the partial pressure of the Hydrogen in the following reaction held within a 2L flask

$$H^+ + Br^- \rightleftharpoons HBr \tag{14.9.27}$$

P[Br] = (nRT)/((V)) = ((1mol)(0.08206)*(668K)/(2L) = 27.4atm

55.

What is Le Chatelier's principle?

Answer:

Le Chatelier's principle is when a new equilibrium has occurred due to a chance of change in temperature or pressure on a chemical equilibrium. It can be known as the Equilibrium Law.

63

What affect does decreasing the volume of the system have on the equilibrium condition in each of the following reactions:

a. $2H_2(g) + C(s) \rightleftharpoons CH_4(g)$
b. $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$
c. $2N_2O(g) + O_2(g) \rightleftharpoons 4NO(g)$

S63

a. shifts right
b. shifts left
c. shifts left

65

Hb(aq) + O2(aq) àHb:O2(aq) ΔH<0

How does decreasing temperature effect this system?

S65

Because the reaction is endothermic, we know that decreasing the temperature adds more to the reactants side, shifting the equilibrium towards the products.

69.

The standard enthalpy of reaction for the combustion of ethanol is 1,418kJ/mole. Write the equation:

$$C_2H_5OH(l) + 3O_2(g) \rightleftharpoons 2CO_2(g) + 3H_2O(l) \tag{14.9.28}$$

How does pressure affect the equilibrium?

As 3 molecules of O2 give 2 molecules of CO2, increasing pressure drives the equilibrium to the right side of the reaction, as well as decrease the volume and increase the number of moles.

77

In a reversible reaction,

$$2Na + Cl_2 \rightleftharpoons 2NaCl \tag{14.9.29}$$

initial conditions indicate 2 mol Na and 1 mol Cl_2. What is the expected amount of NaCl at equilibrium?

Less than 2mols, because the Cl_2 is the limiting reagent at 1mol, but the theoredical yield of NaCl is 2mols, so it can never hit that point unless more Cl is added.

79.

One of the key reactions in the gasification of coal is the ethanation reaction, in which ethane is produced from synthesis gas – a mixture of CO and H_2.

2CO(g) + 2H2(g) àßC_2H_4(g) + 2H_2O (g)

DH= - 230 kJ; K_c= 178 at 1200K

1. Is the equilibrium conversion of synthesis gas to methane favored at higher or lower temperatures? Higher or lower pressures? Products are favored with higher temperatures and lower pressures.

90.

Consider the following reaction:

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g) \tag{14.9.30}$$

An equilibrium mixture at 345 Kelvin in a 6.50 L vessel contains:

- 1.35 mol PCl_3,
- 0.450 mol Cl_2, and
- 3.20 mol PCl_5.

The equilibrium constant for the reaction is $K_p = 1.21$. If 0.50 mol of PCl_5 are removed from the system, what will the pressures of each species be once the reaction returns to equilibrium?

S90

PV=nRT P = nRT / V

P_{PCl3} = (1.35 mol)(.08206)(345 K) / (6.50 L) = 5.88 atm

P_{Cl2} = (.450 mol)(.08206)(345 K) / (6.50 L) = 1.96 atm

$P_{PCl5=}$ (3.20 mol)(.08206)(345 K) / (6.50 L) = 13.9 atm

$P_{PCl5=}$ (3.20 - 0.50 mol)(.08206)(345 K) / (6.50 L) = 11.8 atm

1.21 = (11.8 + x) / (5.88 – x)(1.96 – x) x = 0.210

P_{PCl3} = 5.88 – 0.210 = 5.67 atm

P_{Cl2}= 1.96 – 0.210 = 1.75 atm

P_{PCl5} = 11.8 + 0.210 = 12.01 atm

103.

For the oxidation of ethanol:

$$C_2H_5OH(l) + 3O_2(g) \rightleftharpoons 2CO_2(g) + 3H_2O(l) \tag{14.9.31}$$

How does pressure affect the equilibrium?

S103

Answer: As 3 molecules of O2 give 2 molecules of CO2, increasing pressure drives the equilibrium to the right side of the equation, as well as decreasing the volume and the number of moles.

Q1

Based on these descriptions, write a balanced equation and the corresponding K_c expression for each reversible reaction:

a. Sulfur trioxide, $SO_{3(g)}$, decomposes into gaseous oxygen and gaseous sulfur dioxide.
b. Zinc metal displaces silver(I) ion from aqueous solution, producing silver mean and an aqueous solution of zinc(II) ion.
c. Oxalate ion, $C_2O_4^-$, oxidizes copper(I) ion to copper(II) ion in aqueous solution and is itself reduced to carbon dioxide.

Solution

a)

$$2SO_{3(g)} \rightleftharpoons O_{2(g)} + 2SO_{2(g)} \tag{14.9.32}$$

$$K_c = \frac{[SO_2]^2[O_2]}{[SO_3]^2} \tag{14.9.33}$$

b)

$$Zn_{(s)} + 2Ag^+_{(aq)} \rightleftharpoons 2Ag_{(s)} + Zn^{2+}_{(aq)} \tag{14.9.34}$$

$$K_c = \frac{[Zn^{2+}]}{[Ag^+]^2} \tag{14.9.35}$$

c)

$$C_2O^-_{4(aq)} + 2Cu^+_{(aq)} \rightleftharpoons 2CO_{2(aq)} + [(Cu)]^(2+)_{(aq)} \tag{14.9.36}$$

$$K_c = \frac{[CO_2]^2[Cu]^{2+}}{[C_2O^-_4][Cu^+]^2} \tag{14.9.37}$$

Q2

Based on these descriptions, write a balanced equation and the corresponding K_p expression for each reversible reaction:

a. $KO_{2(s)} + CO_{2(g)} \rightleftharpoons KCO_3 + O_{2(g)}$
b. $HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)}$
c. $HCl_{(g)} + I_{2(g)} \rightleftharpoons HI_{(g)} + Cl_{2(g)}$

Solution

a)

$$4KO_{2(s)} + 2CO_{2(g)} \rightleftharpoons 2KCO_3 + 3O_{2(g)} \tag{14.9.38}$$

$$K_c = \frac{[O_2]^3}{[CO_2]^2} \tag{14.9.39}$$

b)

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)} \tag{14.9.40}$$

$$K_c = \frac{[H_3O^+][F^-]}{[HF]} \tag{14.9.41}$$

c)

$$2HCl_{(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)} + Cl_{2(g)} \tag{14.9.42}$$

$$K_c = \frac{[HI]^2[Cl_2]}{[HCl]} \tag{14.9.43}$$

Q3

Write equilibrium constant expressions, K_c, for the reactions:

a. $2CO_{(g)} + O_{2(g)} \rightleftharpoons 2CO_{2(g)}$
b. $Fe^{2+}_{(aq)} + 2H_2O_{(l)} \rightleftharpoons FeOH^{2+}_{(aq)} + H_3O^+_{(aq)}$
c. $HNO_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NO^-_{2(aq)}$

Solutions

$$K_c = \frac{[CO_2]^2}{[CO]^2[O_2]} \tag{14.9.44}$$

$$K_c = \frac{[Fe(OH)][H_3O^+]}{[Fe^{3+}]} \tag{14.9.45}$$

$$K_c = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} \tag{14.9.46}$$

29

Can a mixture of 3.1 mol CO, 3.9 mol O_2, and 2.2 mol CO_2 be maintained indefinitely in a $9.2\ L$ container at a temperature at which $K_c = 95$ in this reaction (and why or why not):

$$2CO_{(g)} + O_{2(g)} \rightleftharpoons 2CO_{2(g)} \tag{14.9.47}$$

Solution

Q_C = ([[([(CO)]_2])]^2)/([([CO})]^2 [O_2]) = ((([((2.2 mol [(CO)]_2)/(9.2 L)))]^2)/(((3.1 mol CO)/(9.2 L))^2 ((3.9 mol O_2)/(9.2 L))) = 1.19

1.19 < K_c = 95. Thus the mixture described cannot be maintained indefinitely. Because Q_C < K_c, the reaction will proceed toward the products until the equilibrium is established.

30

Is a mixture of 0.0196 mol $CO_{2(g)}$ and a 0.815 mol $C_2O_{4(g)}$ in a 6.15 L flask at 30°C at equilibrium? If not, in which direction will the reaction proceed: toward the reactants or products?

$$C_2O_{4(g)} \rightleftharpoons 2CO_{2(g)} \tag{14.9.48}$$

with

$$K_c = 7.66 \times 10^{-5} \tag{14.9.49}$$

Solution

$$Q = \frac{[CO_{2(g)}]^2}{[C_2O_{4(g)}]} \tag{14.9.50}$$

$$= \frac{\left(\frac{0.0196\ mol\ CO_2}{6.15\ L}\right)^2}{\frac{0.815\ mol\ C_2O_4}{6.15\ L}} = 7.66 \times 10^{-5} \tag{14.9.51}$$

$\therefore Q = K$

The reaction is already at equilibrium.

31

In the reaction:

$$2H_{2(g)} + O_{2(g)} \rightleftharpoons 2H_2O_{(l)} \tag{14.9.52}$$

if 0.05 mol H_2 and 0.02 mol O_2, 20 mol of H_2O are introduced simultaneously into a 1.0 L vessel.

a. If $K_c = 1.4 \times 10^{83}$ is the mixture initially at equilibrium.
b. If not, in which direction will a net change occur?

Solution

a)

$$K_c = \frac{1}{[H_{2(g)}]^2[O_{2(g)}]} = 1.4 \times 10^{83} \tag{14.9.53}$$

$$Q = \frac{1}{(0.05)^2(0.02)} = 4 \times 10^5 \tag{14.9.54}$$

$\therefore Q \neq K_c$

The mixture is initially not at equilibrium.

b) The reaction will experience a net change in the direction of the products because since $Q < K$.

33

A mixture consisting of 0.5 mol Cu^{2+} and 0.5 mol Sn^2+ is brought to equilibrium in a 1 L flask. What is the equilibrium concentration of Sn^{4+}. The reaction proceeds as follows:

$$2Cu^{2+}_{(aq)} + Sn^{2+}_{(aq)} \rightleftharpoons 2Cu^{+} + Sn^{4+} \tag{14.9.55}$$

wth $K_c = 1.48$

Solution:

An ICE table is necessary to solve this problem.

ICE	$2Cu^{2+}_{(aq)}$	$Sn^{2+}_{(aq)}$	$\rightleftharpoons$	$2Cu^{+}$	Sn^{4+}
Initial	0.5 mol	0.5 mol		0 mol	0 mol
Change	-2x	-x		+2x	+x
Equilibrium	0.5 mol - 2x	0.5 mol - x		+2x	+x

$$K_c = \frac{(x^2)(x)}{(0.5-2x)^2(0.5-x)} = 1.48 \tag{14.9.56}$$

$$x = 0.1752 \tag{14.9.57}$$

$$\therefore [Sn^{4+}] = 0.1752 \tag{14.9.58}$$

35

Starting with 0.5 mol of $PCl_{5(g)}$ and 1 mol of $Cl_{2(g)}$, in a 2.0 L flask, how many moles of $PCl_{3(g)}$ will be present at equilibrium?

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)} \tag{14.9.59}$$

with

$$K_c = 3.8 \times 10^{-2} \tag{14.9.60}$$

Solution:

Reaction is already balanced:

$$PCl_{5(g)} = PCl_{3(g)} + Cl_{2(g)} \tag{14.9.61}$$

An ICE table is necessary to solve this problem (constructed in terms of molarity, not mol).

ICE	$PCl_{5(g)}$	$\rightleftharpoons$	$PCl_{3(g)}$	$Cl_{2(g)}$
Initial	0.25 mol		0 mol	0.5 mol
Change	-x		+x	+x
Equilibrium	0.25 mol - x		+x	0.5 mol +x

$$K_c = 3.8 \times 10^{-2} \tag{14.9.62}$$

$$3.8 \times 10^{-2} = \frac{x(0.5+x)}{0.25-x} \tag{14.9.63}$$

$$x = 0.0335 \tag{14.9.64}$$

$$\therefore [PCl_3] = 0.0335\ M \tag{14.9.65}$$

37

Equilibrium is established in a 3.0 L flask for the following reaction.

$$2H_2S_{(g)} + CH_{4(g)} \rightleftharpoons 4H_{2(g)} + Cs_{2(g)} \tag{14.9.66}$$

with

$$K_c = 5.27 \times 10^{-8} \tag{14.9.67}$$

How many moles of $CS_{2(g)}$ and $CH_{4(g)}$ are present at equilibrium in the two initial situations:

a. If 3 mol of $H_2S_{(g)}$ and $CH_{4(g)}$ are initially introduced into the flask.
b. If 4 mol of $4H_{2(g)}$ and $CS_{2(g)}$ are initially introduced into the flask.

Solution:

The reaction is already balanced

$$2H_2S_{(g)} + CH_{4(g)} \rightleftharpoons 4H_{2(g)} + CS_{2(g)} \tag{14.9.68}$$

Scenario A)

An ICE table is necessary to solve this problem.

ICE	$2H_2S_{(g)}$	$CH_{4(g)}$	$\rightleftharpoons$	$4H_{2(g)}$	$CS_{2(g)}$
Initial	1 M	1 M		0 M	0 M
Change	-x	-x		+x	+x
Equilibrium	1 M - 2x	1 M - x		+x	+x

5.27 \times 10^{-8} = \dfrac{(x^5)}{(1-x)^3

\(x= .0343= [CS_2]\]

\(1 - 0.0343 = 0.9657 = [CH_4]\]

Scenario B)

An ICE table is necessary to solve this problem.

ICE	$2H_2S_{(g)}$	$CH_{4(g)}$	$\rightleftharpoons$	$4H_{2(g)}$	$CS_{2(g)}$
Initial	0	0		4/3 M	4/3 M
Change	+x	+x		-x	-x
Equilibrium	+x	+x		4/3 M - x	4/3 M - x

$$5.27 \times 10^{-8} = \frac{(4/3-x)^5}{x^3} \tag{14.9.69}$$

$$x = 1.2925 = [CH_4] \tag{14.9.70}$$

$$4/3 - 1.2925 = 0.04083 = [CS_2] \tag{14.9.71}$$

39

The following reaction occurs in the presence of 10 grams of $CH_{4(g)}$, 15 grams of $C_2H_{2(g)}$, and 2 grams of $H_{2(g)}$.

$$2CH_{4(g)} \rightleftharpoons C_2H_{2(g)} + 3H_{2(g)} \tag{14.9.72}$$

with

$$K_c = 0.154 \tag{14.9.73}$$

What direction will a net change occur and now many grams of each substance will be present at equilibrium?

Solution:

First one must find the moles of each substance.

- 10 grams CH_4/16 grams/mol CH_4= 0.625 mol CH_4
- 15 grams C_2H_2/26 grams/mol= 0.5769 mol C_2H_2
- 2 grams H_2/2 grams/mol= 1 mol H_2

$$Q = \frac{0.5769X1^3}{(0.625)^2} = 1.454 > 0.154 \tag{14.9.74}$$

Since $Q > K_c$, the reaction shifts to the left.

$$2CH_{4(g)} \rightleftharpoons C_2H_{2(g)} + 3H_{2(g)} \tag{14.9.75}$$

An ICE table is necessary to solve this problem (in terms of moles).

ICE	$2CH_{4(g)}$	$\rightleftharpoons$	$C_2H_{2(g)}$	$3H_{2(g)}$
Initial	0.625		0.5769	1
Change	+2x		-x	-3x
Equilibrium	0.625 + 2x		0.5769 - x	1 - 3x

$$0.154 = \frac{(0.5769 - x)(1 - x)^3}{(0.625 + x)^3} \tag{14.9.76}$$

$x = 0.277$

- $[CH_4] = 0.625 + 0.277 = 0.902$
- $[C_2H_2] = 0.299$
- $[H_2] = 0.722$

41

Gaseous hydrogen iodide is placed in a closed container, where it partially decomposes.

$$2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)} \tag{14.9.77}$$

with

$$K_c = 5.5 \times 10^5 \ @ \ 698 \ K \tag{14.9.78}$$

If 0.172 mol $HI_{(g)}$ is initially added to a 1.67 L flask at 698 K, what will be the total partial pressure at equilibrium?

Solution

- $[HI]_i = \frac{0.172 \ mol}{1.67 \ L} = 0.103 \ M$

An ICE table is necessary to solve this problem (in terms of mole).

ICE	$2HI_{(g)}$	$\rightleftharpoons$	$H_{2(g)}$	$I_{2(g)}$

Initial	0.172 M		0	0
Change	-2x		+x	+x
Equilibrium	0.172 M - 2x		+x	+x

$$K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{x^2}{(0.172 - x)^2} = 5.5 \times 10^5 \tag{14.9.79}$$

$$x^2 = 94600 - 5.5 \times 10^5 x0 = x^2 + 5.5 \times 10^5 (x - 94600) \tag{14.9.80}$$

Solving the quadratic equation:

$$x = -5.5 \times 10)^5 \pm \sqrt{\frac{(3.03 \times 10^1 1) + 378400)}{2}} = 227M, -550227M \tag{14.9.81}$$

[total] = [H2] + [I2] + [2HI] = x+x+ 0.172-x = 0.172 + 227 ≈ 227 M

Ptot = (227 mol)/L× (0.08206 L atm)/(mol K)× 698K=1.3 ×[(10)]^4 atm

55

The continuous removal of reactants of a chemical reaction has the effect of causing the reaction to go backwards. Explain this in terms of Le Châtlier's principle.

Solution

Continuous removal of the reactants has the effect of decreasing the concentrations of the reactants below their equilibrium values. Therefore, the equilibrium system is disturbed by the removal of the reactant; the system will remedy this by trying to re-establish the equilibrium to the left to generate more reactants.

63

How would decreasing the volume of the following systems affect their equilibrium?

Decreasing the volume of a system causes the equilibrium to shift towards the side whose sum of the stoichiometric coefficients of only the species in a gaseous state is smaller. If Δn_gas is greater than 0, it shifts to the left. If Δn_gas is less than 0, it shifts to the right.

a) CH_4 (g)+ 2 H_2 O (g) \rightleftharpoonsCO_2 (g)+ 4H_2 (g)

Δn_gas is 2, Since there are 3 moles of reactants and 5 moles of products, the reaction shifts to the left, towards the reactants.

b) CaCO_3 (s)\rightleftharpoonsCaO(s)+CO_2 (g)

Δn_gas is 1, since we only count the moles of gas, there is only one mole of gas in the products. Thus, the reaction shifts towards the reactants.

c) C(s)+O_2 (g)\rightleftharpoonsCO_2 (g)

Δn_gas is 0, since we only count the number of moles of species in their gaseous state. Thus, there is no shift in equilibrium despite change in volume.

71

Explain why dissociation diatomic elements from molecules to elements, such as the reaction O_2 (g)\rightleftharpoons2O(g), will always increase with an increase in temperature.

First, lets think about this logically. Breaking bonds, especially strong bonds in diatomic elements, requires energy. When you add more energy to the system, there is a higher amount of energy, and more bonds being broken.

Second, lets consider the equilibrium of this reaction. If we add heat to the reactants, the equilibrium shifts to the right, towards the products.

O_2 (g)+heat\rightleftharpoons2O(g)

→ → reaction shifts to products

1

Write a balanced equation & K_c expression for each reaction.

a) CH_4 undergoes combustion to become gaseous carbon dioxide and liquid water

$$CH_4 + 2O_2 \leftrightarrow CO_2 + 2H_2O \; ; \; K_c = [CO_2]/ ([CH_4][O_2]^2)$$

b) $Cl_2O_{5\,(g)}$ decomposes into Cl_2 and O_2

$$2Cl_2O_{5\,(g)} \leftrightarrow 2Cl_2 + 5O_2 \; ; \; K_c = \{[Cl_2]^2[O_2]^5\}/[N_2O_5]^2$$

c) N_2O_5 decomposes into O_2 and NO_2

$$2N_2O_5 \leftrightarrow O_2 + 4NO_2 \; ; \; K_c = \{[NO_2]^4[O_2]\}/[N_2O_5]^2$$

2

Based on the descriptions, write a balanced equation and its K_p expression for each reaction

a) $SO_{2\,(g)}$ is converted to sulphur trioxide

$$2SO_{2(g)} + O_{2(g)} \leftrightarrow 2SO_{3(g)} \; ; \; K_p = (SO_3)^2/[(SO_2)^2(O_2)]$$

b) N_2 and H_2 are converted to NH_3

$$N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)} \; ; \; K_p = (NH_3)^2/[N_2)(H_2)^3$$

c) Solid carbon is heated with steam to produce hydrogen gas and carbon monoxide

$$C_{(s)} + H_2O_{(g)} \leftrightarrow H_{2(g)} + CO_{(g)} \; ; \; K_p = [(H_2)(CO)]/(H_2O)$$

3

Write the K_c expression for:

a) $CaCO_{3(s)} \leftrightarrow CaO_{(s)} + CO_{2(g)}$

$K_c = [CO_2]$

b) $N_2 + O_2 \leftrightarrow 2NO$

$K_c = [NO]^2/ \{[N_2][O_2]\}$

c) $2NOCl \leftrightarrow 2NO + Cl_2$

$K_c = \{[NO]^2[Cl_2]\}/[NOCl]^2$

29

Can a mixture of 1.5 mol CH_3, 3 mol O_2, 2 mol CO_2, and 0.5 mol H_2O be maintained in a 5 L flask (K_c = 175)? Explain.

-No. Q < K, so the reaction will shift to the right.

30

Is the following reaction in an 8 L flask at 273 K at equilibrium? Explain.

$$BaCl_{2(aq)} + Na_2SO_{4(aq)} \leftrightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$$

-No. Q > K, so the reactants are favored.

31

If 0.68 mol CH_3OH, 0.54 mol CO, and 0.94 H_2 are all added to a 5 L container at 500 K (K_c = 150), is the mixture at equilibrium? If not, which way will it shift?

No. Q < K, so it will shift to the right.

33

A mixture of 0.28 mol H_2, 0.28 mol CO_2, 0.22 mol CO, and 0.22 mol H_2O is brough to equilibrium at 562 K in a 3.5 L flask. What are the equilibrium amounts of each? (K_c = 9.84)

$$CO_{2(g)} + H_{2(g)} \leftrightarrow CO_{(g)} + H_2O_{(l)}$$

H_2: 0.05657 mol

CO_2: 0.05657 mol

CO: 0.080 mol

H_2O: n/a

35

Starting with 0.217 mol $NO_{2(g)}$ and 0.423 mol $O_{2(g)}$ in a 2 L flask at 500 K, how many moles of $NO_{(g)}$ will be present at equilibrium? ($K_c = 1.4 \times 10^{-3}$)

$$NO + O_2 \leftrightarrow NO_2$$

- Moles NO: 0.008

37

How many moles of each substance will there be at equilibrium if 1 mol each of CO_2 and COF_2 are added?

$2COF_{2(g)} \leftrightarrow CO_{2(g)} + CF_{4(g)}$ at 300 K in a 4 L flask ($Kc = 3.4 \times 10^{-3}$)

COF_2: 0.215 mol

CO_2: 0.215 mol

CF_4: 0.305 mol

39.

$CS_{2(g)} + 4H_{2(g)} \leftrightarrow CH_{4(g)} + 2H_2S_{(g)}$ (K_c= 3.5)

a) Which way will the reaction shift?

-To the right

b) How many moles of each will be present at equilibrium?

- 0.45, 0.3, 0.15, respectively

41

If 1.5 mols of PCl_5 dissociate in a 2 L flask, what is the total equilibrium pressure at 375 K? (K_c = 3.0)

$$PCl_{5(g)} \leftrightarrow Cl_{2(g)} + PCl_{3(aq)}$$

- 69.2 atm

43

In the reaction below, what are the concentrations of Cr^{3+} and Cr^{2+}?

$$Pb_{(s)} + 2Cr^{3+}_{(aq)} \leftrightarrow Pb^{2+}_{(aq)} + 2Cr^{2+}_{(aq)}$$

a) Cr^{3+}: 2

b) Cr^{2+}: 2

45.

One sketch below represents an initial nonequilibrium mixture in the reversible reaction:

$$SO_2\ (g) + Cl_2\ (g) \rightleftharpoons SO_2Cl_2\ (g) \qquad (14.9.82)$$

with

$$K_c = 4.0 \tag{14.9.83}$$

Which of the other three sketches best represents an equilibrium mixture? Explain.

Solution

$$[SO_2Cl_2] = 4.0 * [Cl_2] * [SO_2] \tag{14.9.84}$$

represents the best due to it having numbers of SO2Cl2,SO4,Cl2 molecules that are consistent with Kc. (b) has 12 SO2Cl2 molecules per unit volume, 1 Cl2 molecules per unit volume, 3 SO2 molecules per unit volume. They are the requisite number of each type of molecules needed to generate the excepted Kc.

53. Problem

1.11mol each of CO and Cl_2 are introduced into an evacuated 2.97L flask, and the following equilibrium is established at 771 K.

$$CO\ (g) + Cl_2\ (g) \rightleftharpoons COCl_2\ (g) \tag{14.9.85}$$

with

Kp=22.5

Calculate **(a)** the partial pressure of $COCl_{2(}g)$ **(b)** total gas pressure

Solution:

(a) P(CO)=P(Cl2)=nRT/V=1.11mol*0.08206Latmmol^-1K^-1*771K/2.97K=23.6atm

Equation: CO(g)+ Cl2(g)à COCl2 Kp=22.5

Initial: 23.6atm 23.6atm 0atm

Changes: -xatm -xatm +xatm

Equil: 23.6-xatm 23.6-xatm xatm

Kp=P(COCl2)/P(CO)P(Cl2)=22.5 x/(23.6-x)^2=x/(557.0-47.2x+x^2)

22.5(557.0-47.2x+x^2)=x=12532.5-1062x+22.5x^2=x

12532.5-1063x+22.5x^2=0

x=(-b+/-(b^2-4ac)^(1/2))/2a=(-(-1063)+/-((-1063)^2-4(22.5)(12532.5))^(1/2))/2(22.5)=69.04(too large),22.62

P(CO)=P(Cl2)=23.6atm-22.62atm=.98atm P(COCl2)=22.62 atm

(a) Ptotal=P(CO)+P(Cl2)+P(COCl2)=.98atm+.98atm+22.62atm=24.58atm

55. Problem

Explain how Le Chatelier's pirinciple supports the fact that the continuous removal of one of the products of a chemical reaction has the effect of causing the reaction to go to completion.

Solution:

Makes Q less than K and reaction shift the right as result and change in pressure of a constant temp equilibrium mixture.

63. Problem

If the volume increases the system, what effect does it have on the equilibrium condition?

a. $C(s) + H_2O\ (g) \rightleftharpoons CO(g) + H_2\ (g)$
b. $4NH_3\ (g) + 5O_2\ (g) \rightleftharpoons 4NO(g) + 6\ H_2O(g)$
c. $Ca(OH)_2\ (s) + CO_2\ (g) \rightleftharpoons CaCO_3(s) + H_2O(g)$

Solution

a. It shifts the right
b. Shifts to the right
c. No effect

71. Problem

Why the % molecules dissociate into atoms in reactions of I2(g)à2I(g) always raises with a. raise temp?

Solution:

Increasing temp adds heat and decreasing temp takes away the heat. When the temp raises, heat is absorbs increases and helps in breaking molecules into atoms. Lower temp, lower the % of dissociation and higher temp, higher % of dissociation.

73. Problem

Refer to Example 14-13. Suppose that 0.1L of the equilibrium mixture is diluted to 0.25L with water. What will be the new concentrations when equilibrium is re-established?

Solution:

Eq: Ag+(aq)+ Fe2+(aq)à Fe3+(aq)+ Ag(s) Kc=2.98

Initi: 0.31M 0.21M 0.19M

Ch: 0.12M 0.084M 0.076M

Equil: +xM +xM -xM

New: 0.12+x 0.0084+x 0.076-x

Kc=[Fe3+]/[Ag+][Fe2+]=2.98=0.076-x/(0.12+x)(0.084+x)

29.8(0.12+x)(0.084+x)=0.076-x

0.076-x=0.03+0.61x+2.98x^2 2.98x^2+1.16x-0.046=0

x=(-16.1+/-(2.59+0.55)^(1/2))/5.96=0.027,-0.57

[Fe2+]=0.084-0.57=-0.49M

[Fe2+]=0.084+0.027=0.111M [Ag+]=0.12+0.027=0.15M

[Fe3+]=0.076-0.027=0.049M

Kc=0.049M/(0.111M*0.15M)=2.94

75. Problem

Started with $SO_3(g)$ at 1.00 atm, what will be the total pressure when equilibrium is reached in the following reaction at 700 K?

$$2\ SO_3\ (g) \rightleftharpoons 2\ SO_2\ (g) + O_2\ (g) \tag{14.9.86}$$

with

$$K_p = 1.65 * 10^{-5} \tag{14.9.87}$$

Solution:

Pressure: 2SO3(g)à 2SO2(g) + O2(g)

Initi: 1.00atm 0 0

Change: -2x +2x +x

Equil: 1.00-2x 2x x

Kp=(PSO2)^2(PO2)/(PSO3)^2

1.65*10^-5=(2x)^2(x)/(1.00-2x)^2

x=P02=0.0155atm

PSO2=2x=2(0.0155)=0.0310atm

PSO3=1.00-2x=[1.00-2(0.0155)]=[1.00-0.0310]=0.969atm

Ptot=PO2+PSO2+PSO3=(0.0115+0.0310+0.969) atm=1.016atm

77. Problem

Calculate the amount of SO2, O2, SO3 listed on **(a)** Figure 14-6(c); **(b)** Figure 14-7(b)

Solution:

(a)2SO3(g)à2SO2(g)+O2(g) Kc=1/2.8*10^2

I: 1.68mol 0.32 0.16

C: -2x +2x x

E: 1.68-2x 0.32+2x 0.16+x

3.57*10^-3=((((0.32+2x)^2)/10)*((0.16+x)/10)))/((1.68-2x)^2)/10

X=0.10755M

[SO3]=1.68-2(0.10755)=1.46mol

[SO2]=0.32+2(0.10755)=0.54mol

[O2]=0.16+0.10755=0.27mol

(b) 2SO2(g)+O2(g)à 2SO3(g) Kc=2.8*10^2

I: 0.32mol 0.16 0.68

C: -2x -x +2x

E: 0.32-2x 0.16-x 0.68+2x

Kc=[SO3]^2/[SO2]^2[O2]

2.8*10^2=((0.68+2x)^2)/(0.32-2x)^2(0.16-x)

X=0.0749M

[SO3]=0.68+2(0.0749)=0.83mol

[SO2]=0.32-2(0.0749)=0.17mol

[O2]=0.16-0.0749=0.085mol

79. Problem

Methanation reaction: methane is produced from a mixture of CO and H2.

CO(s)+3H2(g)àCH4(g)+H2O(g) change=-230kJ Kc=190 at 1000K

(a) Is the reaction favored at higher or lower temp? Higher or lower pressure? **(b)** 4mol of synthesis gas with a 3:1 ratio of H2(g) to CO(g) in a 15.0L flask. What is the mole fraction of CH4 at 1000k?

Solution:

(a) Reaction is exothermic. Lowering temp, shifts right and increase conversion of synthesis gas to methane. High pressure, shifts right and increase conversion of synthesis gas to methane

(b) CO(s) + 3H2(g)àCH4(g)+H2O(g)

I: 1mol/15L 3mol/15L 0 0

=0.0667M =0.2M

C: -x -3x +x +x

E: 0.0667-x 0.2-3x x x

Kc=[CH4][H20]/[CO][H2]^3

190=(x)((x)/(0.0667-x)(0.2-3x)^3

x=0.0424M

Mole fraction of CH4=equilibrium mol of CH4/total number of mol at

equilibrium

=0.0243mol/(0.0424+0.0424+0.0243+0.0728)=0.134mol

84. **Problem:**

Synthesis of ammounia at 500K, N2+(g)+3H2(g)à2NH3(g), Kp=8.71*10^-

2. N2 and H2 are mixed in a 1:7 ratio and total pressure is 1.00atm.Find the

mole % NH3 at equilibrium.

Solution:

Kp=(PNH3)^2/(PN2)(PH2)^3

= ((XNH3)^2/(XN2)(XH2)^3)*(1/(Ptotal)^2)

= ((XNH3)^2/1*(7)^2)*1/(1^2)

= 8.71*10^-2

(xNH3)^2/49=8.71*10^-2

(xNH3)^2=4.2679

xNH3=2.07

%Mol of NH3=mol/total mol*100%=2.07/1+7+2.07*(100%)=20.52%

102. Problem

Equilibrium is established in the reaction 2SO2(g)+O2(g)à2SO3(g) at a Temp where Kc=123. If the number of mol of SO3(g) in the equilibrium mixture is the same as the number of mol of SO2(g) **(a)** the number of mol of O2(g) is also equal to the number of mol of SO2(g) **(b)** the number of mol of O2(g) is half the number of mol of SO2 **(c)** [O2] may have any of several values **(d)** [O2]=0.008M

Solution

Kc=[SO3]^2/[SO2]^2[02]

= (x/V)^2/(x/V)^2[O2]

Kc=1/[O2]

123=1/[O2]

[O2]=0.008M

D is the correct answer

104. Problem

2NO2(g)à2NO(g)+O2(g), Kc=3.19*10^-12 at 184C. At 184C, the value of Kc for this reaction is **(a)** 0.9*10^6 **(b)** 7.5*10^2 **(c)** 5.6*10^5 **(d)** 2.8*10^5.

Solution

K'c=[NO2]/[NO][O2]^(1/2)

= 1/(Kc)^(1/2)

= 1/(3.19*10^-12)^(1/2)

=1/1.79*10^-6

= 5.6*10^5 (c) is correct

Q14.1

based on these descriptions, write a balanced equation for each reversible reaction.

a. Oxygen gas oxidizes gaseous ammonia vapor to gaseous nitrogen and water vapor.

b. Hydrogen gas reduces gaseous nitrogen dioxide to gaseous ammonia and water vapor.
c. Liquid acetone is in equilibrium with its vapor.

S31

a. 3 O2(g) + 4 NH3(g) ↔ 2 N2(g) + 6 H2O(g)
b. 7 H2(g) + 2 NO2(g) ↔ 2 NH3(g) + 4 H2O(g)
c. (CH3)2CO(l) ↔ (CH3)2CO(g)

Q13

For the following reactions, determine the equilibrium constant expression in terms of concentration.

a. a) N2O4(g) ⇄ 2NO2(g)
b. b)NH3(aq) + H2O(ℓ) ⇄ NH4+(aq) + OH-(aq)
c. c)C(s) + H2O(g) ⇄ H2(g) + CO(g)

S13

a)Kc = [NO2]2/[N2O4]

b)KC=[NH4+][OH-]/[NH3]

c)KC=[H2][CO]/[H2O]

Q25

At a certain temperature 1L flask with .298mol PCl3(g) and 8.70*10^-3molPCl5(g). After the system reached equilibrium, 2*10^-3molCl2(g) was found in the flask.

Calculate equilibrium concentrations and the value of Kc and Kp assuming 298K.

PCl5 ⇄ PCl3 + Cl2

[i]8.7*10^-3mol .298 0

[c]-x +x +x=2*10^-3

[e]8.7*10^-3mol-x .298+x 2*10^-3

Kc=(.002)(.3)/.0067=.0896

Kp=Kc(RT)Δn

Kp=(.0821)(298)2-1

Kp=2.19

Q26

At 2000 K, Kc=0.154 for the reaction N2(g) + 3H2(g)⇄2NH3(g).

a. Is the conversion of N2 (g) to 2NH3(g) favored at high or low pressures?
b. If the equilibrium mixture is transferred from a 1.00 L flask to a 2.00L flask, will the number of moles of N2(g) increase, decrease, or remain unchanged?
c. If H2 is added, which direction would the reaction shift?

S26

a. High because the reactants have more moles than the products
b. Increase
c. Right

Q29

In the following reaction, which direction will the reaction favor if the initial concentration of O2and CO2 is 4 M, and Kc is 1.16 at 298 K?

4KO2(s) + 2CO2(g) ⇄ 2KCO3(s) + 3O2(g)

S29

Q= [O2]3 / [CO2]2

Q= 43 / 42= 4K<Q= reaction will shift left

Q49

At 338 K, pure PCl5 gas is present in a flask at a pressure of 26.7 kPa. At 473 K this is completely dissociated into PCl3 gas and Cl2 gas. Calculate the pressure in the flask at 473 K. No volume is specified, so take a convenient value, say V = 1.00 L. The equation for the reaction that occurs at 473 K is

S49

PCl5(g) ® PCl3(g) + Cl2(g)

At 338 K, moles of PCl5 gas = PV/(RT)

= (26.7 x 1.00)/(8.314 x 338) = 9.50 x 10-3 mol.

At 473 K, moles of gas present = 2 x 9.50 x 10-3 = 0.0190 mol.

Then, using the Ideal Gas Equation, the new pressure can be deduced.

P = nRT/V

= 0.0190 x 8.314 x 473/1.00 = 74.7 kPa

Q57

Predict the effects on the amount of NH_3 in the following reaction

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} \tag{14.9.88}$$

with $\Delta H = -46.11$

when

a. temperature is lowered
b. introducing more H2
c. doubling the vol. of the container.

S57

a. amount will increase because this is an exothermic reaction
b. amount will increase to obtain equilibrium
c. amount will decrease to produce more moles of gas to fill the larger container

Q59

If the system is heated which direction will the equilibrium shift?

a. 2CO2(g)⇔ 2CO(g)+O2(g) ΔH°=566kj
b. C2H4(g) + H2(g) ⇔ C2H6(g)

S59

a. shifts to the right
b. shifts to the left

Q61

If the volume of an equilibrium mixture of N2(g), H2 (g), and NH3 (g) is reduced by doubling the pressure, will PNH3 have increased, decreased, or remained the same when equilibrium is re established? Explain.

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)} \tag{14.9.89}$$

Increase because the number of moles on the product side is less than the amount of moles on the reactant side, therefore, if there is an increase in pressure, the reaction will move toward the side with less moles.

Q63

If we increased the pressure for the following reactions, what would the effect on equilibrium be?

a. 2A (g) + B (s) ⇄ 4C (g)
b. 5A (g) + 2B(l) ⇄ C(s) + 4D(g)
c. A(l) + 4B(g) + 2C (g) ⇄ 3D(s) + 5E(g)

S63

Pressure is the inverse of volume, so when we increase pressure, we are decreasing volume. A will result in a shift left, B will shift right, and C will shift left.

This page titled 14.9: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

CHAPTER OVERVIEW

15: Acid-Base Equilibria - A More Detailed Look

In this chapter, we delve into the fascinating world of acids and bases, focusing on the dissociation of weak acids and bases and the use of ICE tables to solve equilibrium problems to quantify the concentrations of acids and bases in solutions. Understanding these fundamental concepts is crucial for mastering more complex topics in chemistry.

Faculty Feature



We highlight the work of Dr. Frank Wania. He is a renowned expert in environmental chemistry, particularly in the study of the environmental fate and transport of organic contaminants. His research is pivotal in understanding how chemicals interact with and move through the environment, which often involves investigating the dissociation properties of various compounds. Dr. Wania's work emphasizes the importance of basic physical-chemical properties, such as vapor pressure and the air/water Henry's law constant. These properties are essential for predicting how contaminants will behave in different environmental contexts, mirroring the principles you will explore in this chapter about the dissociation of acids and bases. Through a combination of fieldwork, laboratory experimentation, and model simulations, Dr. Wania's research provides valuable insights into contaminant enrichment processes and helps inform strategies for mitigating environmental pollution.

Read more about Dr. Frank Wania and his research here: https://www.utsc.utoronto.ca/labs/wania/

As you study the principles of weak acids and bases and their dissociation, consider how these concepts are applied in real-world environmental chemistry scenarios. Dr. Wania's investigations exemplify the critical role that a thorough understanding of chemical properties plays in addressing environmental challenges.

15.1: Introduction
15.2: Arrhenius Theory- A Brief Review
15.3: Brønsted-Lowry Theory of Acids and Bases
15.4: Self-Ionization of Water and the pH Scale
15.5: Strong Acids and Strong Bases
15.6: Weak Acids and Weak Bases
15.7: Polyprotic Acids
15.8: Ions as Acids and Bases
15.9: Molecular Structure and Acid-Base Behavior
15.10: Lewis Acids and Bases
15.11: Common-Ion Effect in Acid-Base Equilibria
15.12: Buffer Solutions
15.13: Acid-Base Indicators
15.14: Neutralization Reactions and Titration Curves

15.15: Solutions of Salts of Polyprotic Acids
15.16: Acid-Base Equilibrium Calculations- A Summary
15.17: Key Terms
15.18: Key Equations
15.19: Summary
15.20: Exercises

This page titled 15: Acid-Base Equilibria - A More Detailed Look is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

15.1: Introduction

Liquid water is essential to life on our planet, and chemistry involving the characteristic ions of water, H^+ and OH^-, is widely encountered in nature and society. As introduced in another chapter of this text, acid-base chemistry involves the transfer of hydrogen ions from donors (acids) to acceptors (bases). These H^+ transfer reactions are reversible, and the equilibria established by acid-base systems are essential aspects of phenomena ranging from sinkhole formation (Figure 15.1.1) to oxygen transport in the human body. This chapter will further explore acid-base chemistry with an emphasis on the equilibrium aspects of this important reaction class.



Figure 15.1.1: Sinkholes such as this are the result of reactions between acidic groundwaters and basic rock formations, like limestone. (credit: modification of work by Emil Kehnel)

This page titled 15.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **14.0: Introduction** is licensed CC BY 4.0.

15.2: Arrhenius Theory- A Brief Review

Acids and bases have been known for a long time. When Robert Boyle characterized them in 1680, he noted that acids dissolve many substances, change the color of certain natural dyes (for example, they change litmus from blue to red), and lose these characteristic properties after coming into contact with alkalis (bases). In the eighteenth century, it was recognized that acids have a sour taste, react with limestone to liberate a gaseous substance (now known to be CO_2), and interact with alkalis to form neutral substances. In 1815, Humphry Davy contributed greatly to the development of the modern acid-base concept by demonstrating that hydrogen is the essential constituent of acids. Around that same time, Joseph Louis Gay-Lussac concluded that acids are substances that can neutralize bases and that these two classes of substances can be defined only in terms of each other. The significance of hydrogen was reemphasized in 1884 when Svante Arrhenius defined an acid as a compound that dissolves in water to yield hydrogen cations (now recognized to be hydronium ions) and a base as a compound that dissolves in water to yield hydroxide anions.

Acids and bases are common solutions that exist everywhere. Almost every liquid that we encounter in our daily lives consists of acidic and basic properties, with the exception of water. They have completely different properties and are able to neutralize to form H_2O, which will be discussed later in a subsection. Acids and bases can be defined by their physical and chemical observations (Table 15.2.1).

Table 15.2.1: General Properties of Acids and Bases

ACIDS	BASES
produce a piercing pain in a wound.	give a slippery feel.
taste sour.	taste bitter.
are colorless when placed in phenolphthalein (an indicator).	are pink when placed in phenolphthalein (an indicator).
are red on blue litmus paper (a pH indicator).	are blue on red litmus paper (a pH indicator).
have a pH<7.	have a pH>7.
produce hydrogen gas when reacted with metals.	
produce carbon dioxide when reacted with carbonates.	
Common examples: Lemons, oranges, vinegar, urine, sulfuric acid, hydrochloric acid	Common Examples: Soap, toothpaste, bleach, cleaning agents, limewater, ammonia water, sodium hydroxide.

Acids and bases in aqueous solutions will conduct electricity because they contain dissolved ions. Therefore, acids and bases are **electrolytes**. Strong acids and bases will be strong electrolytes. Weak acids and bases will be weak electrolytes. This affects the amount of conductivity.

The Arrhenius Definition of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. The Arrhenius definition of acid-base reactions is a development of the "hydrogen theory of acids". It was used to provide a modern definition of acids and bases, and followed from Arrhenius's work with Friedrich Wilhelm Ostwald in establishing the presence of ions in aqueous solution in 1884. This led to Arrhenius receiving the Nobel Prize in Chemistry in 1903.

An Arrhenius acid is a compound that increases the concentration of H^+ ions that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

In this reaction, hydrochloric acid (HCl) dissociates into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

The Arrhenius definitions of acidity and alkalinity are restricted to aqueous solutions and refer to the concentration of the solvated ions. Under this definition, pure H_2SO_4 or HCl dissolved in toluene are not acidic, despite the fact that both of these acids will donate a proton to toluene. In addition, under the Arrhenius definition, a solution of sodium amide ($NaNH_2$) in liquid ammonia is not alkaline, despite the fact that the amide ion (NH_2^-) will readily deprotonate ammonia. Thus, the Arrhenius definition can only describe acids and bases in an aqueous environment.

Limitation of the Arrhenius Definition of Acids and Bases

The Arrhenius definition can **only** describe acids and bases in an aqueous environment.

In chemistry, acids and bases have been defined differently by three sets of theories: One is the Arrhenius definition defined above, which revolves around the idea that acids are substances that ionize (break off) in an aqueous solution to produce hydrogen (H^+) ions while bases produce hydroxide (OH^-) ions in solution. The other two definitions are discussed in detail alter in the chapter and include the Brønsted-Lowry definition the defines acids as substances that donate protons (H^+) whereas bases are substances that accept protons and the Lewis theory of acids and bases states that acids are electron pair acceptors while bases are electron pair donors.

Contributors and Attributions

- Anonymous

This page titled 15.2: Arrhenius Theory- A Brief Review is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.1: Acids and Bases - A Brief Review** is licensed CC BY-NC-SA 3.0.

15.3: Brønsted-Lowry Theory of Acids and Bases

Learning Objectives

- Identify acids, bases, and conjugate acid-base pairs according to the Brønsted-Lowry definition
- Write equations for acid and base ionization reactions
- Use the ion-product constant for water to calculate hydronium and hydroxide ion concentrations
- Describe the acid-base behavior of amphiprotic substances

The Arrhenius definition of an acid as a compound that dissolves in water to yield hydronium ions (H_3O^+) and a base as a compound that dissolves in water to yield hydroxide ions (OH^-). This definition is not wrong; it is simply limited. We extended the definition of an acid or a base using the more general definition proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry. Their definition centers on the proton, H^+. A proton is what remains when a normal hydrogen atom, ${}^1_1H$, loses an electron. A compound that donates a proton to another compound is called a Brønsted-Lowry acid, and a compound that accepts a proton is called a Brønsted-Lowry base. An acid-base reaction is the transfer of a proton from a proton donor (acid) to a proton acceptor (base). In a subsequent chapter of this text we will introduce the most general model of acid-base behavior introduced by the American chemist G. N. Lewis.

Brønsted-Lowry Defintions

- A compound that **donates** a proton to another compound is called a Brønsted-Lowry acid.
- A compound that **accepts** a proton is called a Brønsted-Lowry base.



Definitions of Acids and Bases: https://youtu.be/r8reN0CSIHw

Acids may be compounds such as HCl or H_2SO_4, organic acids like acetic acid (CH_3COOH) or ascorbic acid (vitamin C), or H_2O. Anions (such as HSO_4^-, $H_2PO_4^-$, HS^-, and HCO_3^-) and cations (such as H_3O^+, NH_4^+, and $[Al(H_2O)_6]^{3+}$) may also act as acids. Bases fall into the same three categories. Bases may be neutral molecules (such as H_2O, NH_3, and CH_3NH_2), anions (such as OH^-, HS^-, HCO_3^-, CO_3^{2-}, F^-, and PO_4^{3-}), or cations (such as $[Al(H_2O)_5OH]^{2+}$). The most familiar bases are ionic compounds such as NaOH and $Ca(OH)_2$, which contain the hydroxide ion, OH^-. The hydroxide ion in these compounds accepts a proton from acids to form water:

$$H^+ + OH^- \rightarrow H_2O \tag{15.3.1}$$

We call the product that remains after an acid donates a proton the **conjugate base** of the acid. This species is a base because it can accept a proton (to re-form the acid):

$$\text{acid} \rightleftharpoons \text{proton} + \text{conjugate base} \tag{15.3.2}$$

$$HF \rightleftharpoons H^+ + F^-$$

$$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$$

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$$

$$NH_4^+ \rightleftharpoons H^+ + NH_3$$

We call the product that results when a base accepts a proton the base's **conjugate acid**. This species is an acid because it can give up a proton (and thus re-form the base):

$$\text{base} + \text{proton} \rightleftharpoons \text{conjugate acid} \tag{15.3.3}$$

$$OH^- + H^+ \rightleftharpoons H_2O$$

$$H_2O + H^+ \rightleftharpoons H_3O^+$$

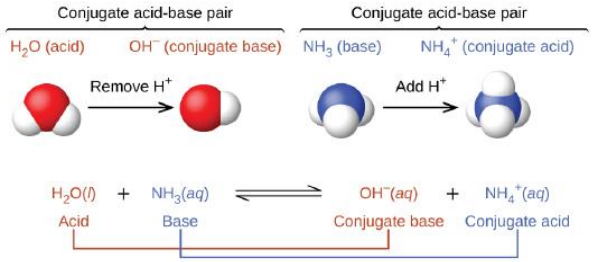
$$NH_3 + H^+ \rightleftharpoons NH_4^+$$

$$S^{2-} + H^+ \rightleftharpoons HS^-$$

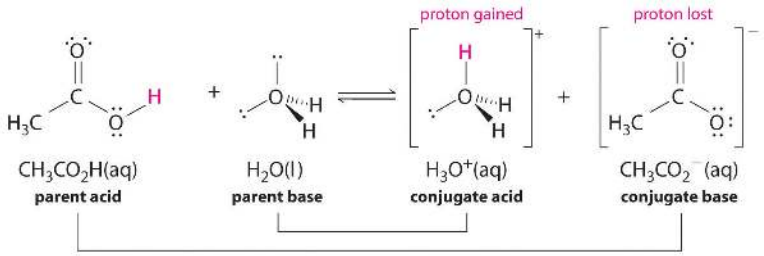
$$CO_3^{2-} + H^+ \rightleftharpoons HCO_3^-$$

$$F^- + H^+ \rightleftharpoons HF$$

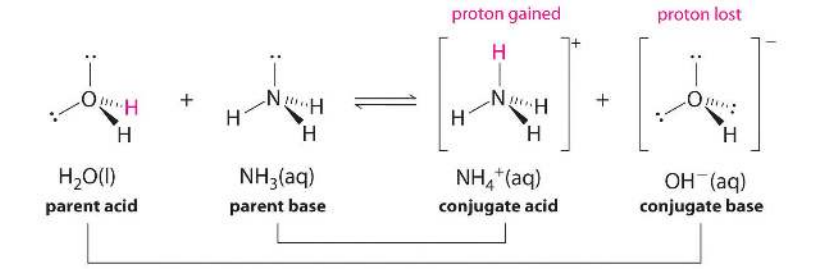
In these two sets of equations, the behaviors of acids as proton donors and bases as proton acceptors are represented in isolation. In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, OH^-, and the conjugate acid of ammonia, NH_4^+:



Similarly, in the reaction of acetic acid with water, acetic acid donates a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base. Once again, we have two conjugate acid–base pairs: the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and the parent base and its conjugate acid (H_3O^+/H_2O).



In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-.



Conjugate Acid-Base Pairs: https://youtu.be/pPrp3xEQef4

Amphiprotic Species

Like water, many molecules and ions may either gain or lose a proton under the appropriate conditions. Such species are said to be amphiprotic. Another term used to describe such species is amphoteric, which is a more general term for a species that may act either as an acid or a base by any definition (not just the Brønsted-Lowry one). Consider for example the bicarbonate ion, which may either donate or accept a proton as shown here:

$$\mathrm{HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)}$$

$$\mathrm{HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)}$$

Water is the most important amphiprotic species. It can form both the hydronium ion, $\mathrm{H_3O^+}$, and the hydroxide ion, $\mathrm{OH^-}$ when it undergoes **autoionization**:

$$\mathrm{2\,H_2O_{(l)} \rightleftharpoons H_3O^+(aq) + OH^-(aq)}$$

Example 15.3.1: The Acid-Base Behavior of an Amphoteric Substance

Write separate equations representing the reaction of $\mathrm{HSO_3^-}$

a. as an acid with $\mathrm{OH^-}$
b. as a base with HI

Solution

a. $\mathrm{HSO_3^-(aq) + OH^-(aq) \rightleftharpoons SO_3^{2-}(aq) + H_2O(l)}$
b. $\mathrm{HSO_3^-(aq) + HI(aq) \rightleftharpoons H_2SO_3(aq) + I^-(aq)}$

Exercise 15.3.1

Write separate equations representing the reaction of $\mathrm{H_2PO_4^-}$

a. as a base with HBr
b. as an acid with $\mathrm{OH^-}$

Answer a

$$H_2PO_4^-(aq) + HBr(aq) \rightleftharpoons H_3PO_4(aq) + Br-(aq)$$

Answer b

$$H_2PO_4^-(aq) + OH^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H_2O_{(l)}$$

Key Concepts and Summary

A compound that can donate a proton (a hydrogen ion) to another compound is called a Brønsted-Lowry acid. The compound that accepts the proton is called a Brønsted-Lowry base. The species remaining after a Brønsted-Lowry acid has lost a proton is the conjugate base of the acid. The species formed when a Brønsted-Lowry base gains a proton is the conjugate acid of the base. Thus, an acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form both the hydronium ion, H_3O^+, and the hydroxide ion, OH^- when it undergoes autoionization:

$$2\,H_2O_{(l)} \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Glossary

acid ionization
reaction involving the transfer of a proton from an acid to water, yielding hydronium ions and the conjugate base of the acid

amphiprotic
species that may either gain or lose a proton in a reaction

amphoteric
species that can act as either an acid or a base

autoionization
reaction between identical species yielding ionic products; for water, this reaction involves transfer of protons to yield hydronium and hydroxide ions

base ionization
reaction involving the transfer of a proton from water to a base, yielding hydroxide ions and the conjugate acid of the base

Brønsted-Lowry acid
proton donor

Brønsted-Lowry base
proton acceptor

conjugate acid
substance formed when a base gains a proton

conjugate base
substance formed when an acid loses a proton

ion-product constant for water (K_w)
equilibrium constant for the autoionization of water

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 15.3: Brønsted-Lowry Theory of Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.2: Brønsted-Lowry Theory of Acids and Bases** is licensed CC BY-NC-SA 4.0.

15.4: Self-Ionization of Water and the pH Scale

Learning Objectives

- To understand the autoionization reaction of liquid water.
- To know the relationship among pH, pOH, and pK_w.

Because of its highly polar structure, liquid water can act as either an acid (by donating a proton to a base) or a base (by using a lone pair of electrons to accept a proton). For example, when a strong acid such as HCl dissolves in water, it dissociates into chloride ions (Cl^-) and protons (H^+). The proton, in turn, reacts with a water molecule to form the hydronium ion (H_3O^+):

$$\underset{aicd}{\mathrm{HCl(aq)}} + \underset{base}{\mathrm{H_2O(l)}} \rightarrow \underset{acid}{\mathrm{H_3O^+(aq)}} + \underset{base}{\mathrm{Cl^-(aq)}} \tag{15.4.1}$$

In this reaction, HCl is the acid, and water acts as a base by accepting an $\mathrm{H^+}$ ion. The reaction in Equation 15.4.1 is often written in a simpler form by removing $\mathrm{H_2O}$ from each side:

$$\mathrm{HCl(aq)} \rightarrow \mathrm{H^+(aq)} + \mathrm{Cl^-(aq)} \tag{15.4.2}$$

In Equation 15.4.2, the hydronium ion is represented by $\mathrm{H^+(aq)}$, although free $\mathrm{H^+}$ ions do not exist in liquid water as this reaction demonstrates:

$$\mathrm{H^+(aq)} + \mathrm{H_2O(l)} \rightarrow \mathrm{H_3O^+(aq)} \tag{15.4.3}$$

Water can also act as an acid, as shown in Equation 15.4.4. In this equilibrium reaction, H_2O donates a proton to NH_3, which acts as a base:

$$\underset{aicd}{\mathrm{H_2O(aq)}} + \underset{base}{\mathrm{NH_3(aq)}} \rightleftharpoons \underset{acid}{\mathrm{NH_4^+(aq)}} + \underset{base}{\mathrm{OH^-(aq)}} \tag{15.4.4}$$

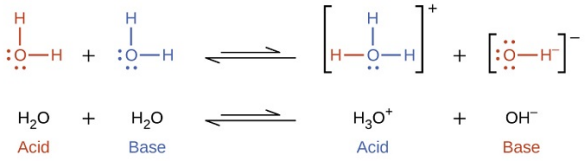
Water is thus termed **amphiprotic**, meaning that it can behave as either an acid or a base, depending on the nature of the other reactant. Notice that Equation 15.4.4 is an equilibrium reaction as indicated by the double arrow and hence has an equilibrium constant associated with it.

The Ion-Product Constant of Pure Water

Because water is amphiprotic, one water molecule can react with another to form an OH^- ion and an H_3O^+ ion in an autoionization process:

$$2\,\mathrm{H_2O(l)} \rightleftharpoons \mathrm{H_3O^+(aq)} + \mathrm{OH^-(aq)} \tag{15.4.5}$$

In pure water, a very small fraction of water molecules donate protons to other water molecules to form hydronium ions and hydroxide ions:



The equilibrium constant K for this reaction can be written as follows:

$$K_a = \frac{a_{H_3O^+} \cdot a_{OH^-}}{a^2_{H_2O}} \approx \frac{[H_3O^+][HO^-]}{(1)^2} = [H_3O^+][HO^-] \tag{15.4.6}$$

where a is the activity of a species. Because water is the solvent, and the solution is assumed to be dilute, the activity of the water is approximated by the activity of pure liquid water, which is defined as having a value of 1. The activity of each solute is approximated by the molarity of the solute.

Note

It is a common error to claim that the molar concentration of the solvent is in some way involved in the equilibrium law. This error is a result of a misunderstanding of solution thermodynamics. For example, it is often claimed that Ka = $K_{eq}[H_2O]$ for aqueous solutions. This equation is incorrect because it is an erroneous interpretation of the correct equation Ka = $K_{eq}(a_{H_2O})$. Because a_{H_2O} = 1 for a dilute solution, $K_a = K_{eq}(1)$, or $K_a = K_{eq}$.

In this reaction, one water molecule acts as an acid and one water molecule acts as a base. Thus, this reaction actually can be designated as the K_a of water and as the K_b of water. It is most common, however, to designate this reaction and the associated law of mass action as the K_w of water:

$$K_w = [H_3O^+][HO^-] \tag{15.4.7}$$

When pure liquid water is in equilibrium with hydronium and hydroxide ions at 25 °C, the concentrations of the hydronium ion and the hydroxide ion are equal:

$$[H_3O^+] = [OH^-] = 1.003 \times 10^{-7}\ M \tag{15.4.8}$$

Thus the number of dissociated water molecules is very small indeed, approximately 2 ppb.

Substituting the values for $[H_3O^+]$ and $[OH^-]$ at 25°C into this expression

$$K_w = (1.003 \times 10^{-7})(1.003 \times 10^{-7}) = 1.006 \times 10^{-14} \tag{15.4.9}$$

The slight ionization of pure water is reflected in the small value of the equilibrium constant; at 25 °C.Thus, to three significant figures, $K_w = 1.01 \times 10^{-14}\ M$ *at room temperature*. Like any other equilibrium constant, K_w varies with temperature, ranging from 1.15×10^{-15} at 0°C to 4.99×10^{-13} at 100°C.

In pure water, the concentrations of the hydronium ion and the hydroxide ion are equal, and the solution is therefore neutral. If $[H_3O^+] > [OH^-]$, however, the solution is acidic, whereas if $[H_3O^+] < [OH^-]$, the solution is basic. For an aqueous solution, the H_3O^+ concentration is a quantitative measure of acidity: the higher the H_3O^+ concentration, the more acidic the solution. Conversely, the higher the OH^- concentration, the more basic the solution. In most situations that you will encounter, the H_3O^+ and OH^- concentrations from the dissociation of water are so small ($1.003 \times 10^{-7} M$) that they can be ignored in calculating the H_3O^+ or OH^- concentrations of solutions of acids and bases, but this is not always the case.

Example 15.4.1: Ion Concentrations in Pure Water

What are the hydronium ion concentration and the hydroxide ion concentration in pure water at 25 °C?

Solution

The autoionization of water yields the same number of hydronium and hydroxide ions. Therefore, in pure water, $[\mathrm{H_3O^+}] = [\mathrm{OH^-}]$. At 25 °C:

$$\begin{aligned} K_\mathrm{w} &= [\mathrm{H_3O^+}][\mathrm{OH^-}] \\ &= [\mathrm{H_3O^+}]^2 \\ &= [\mathrm{OH^-}]^2 \\ &= 1.0 \times 10^{-14} \end{aligned}$$

So:

$$[\mathrm{H_3O^+}] = [\mathrm{OH^-}] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7}\ M$$

The hydronium ion concentration and the hydroxide ion concentration are the same, and we find that both equal $1.0 \times 10^{-7}\ M$.

Exercise 15.4.1

The ion product of water at 80 °C is 2.4×10^{-13}. What are the concentrations of hydronium and hydroxide ions in pure water at 80 °C?

Answer

$$[H_3O^+] = [OH^-] = 4.9 \times 10^{-7}\ M$$

It is important to realize that the autoionization equilibrium for water is established in all aqueous solutions. Adding an acid or base to water will not change the position of the equilibrium. Example 15.4.2 demonstrates the quantitative aspects of this relation between hydronium and hydroxide ion concentrations.

Example 15.4.2: The Inverse Proportionality of Hydronium and Hydroxide Concentrations

A solution of carbon dioxide in water has a hydronium ion concentration of $2.0 \times 10^{-6}\ M$. What is the concentration of hydroxide ion at 25 °C?

Solution

We know the value of the ion-product constant for water at 25 °C:

$$2\,H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

and

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Thus, we can calculate the missing equilibrium concentration.

Rearrangement of the k_w expression yields that $[OH^-]$ is directly proportional to the inverse of $[H_3O^+]$:

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-6}} = 5.0 \times 10^{-9}$$

The hydroxide ion concentration in water is reduced to $5.0 \times 10^{-9}\ M$ as the hydrogen ion concentration increases to $2.0 \times 10^{-6}\ M$. This is expected from Le Chatelier's principle; the autoionization reaction shifts to the left to reduce the stress of the increased hydronium ion concentration and the $[OH^-]$ is reduced relative to that in pure water.

A check of these concentrations confirms that our arithmetic is correct:

$$K_w = [H_3O^+][OH^-] = (2.0 \times 10^{-6})(5.0 \times 10^{-9}) = 1.0 \times 10^{-14}$$

Exercise 15.4.2

What is the hydronium ion concentration in an aqueous solution with a hydroxide ion concentration of 0.001 M at 25 °C?

Answer

$$[H_3O^+] = 1 \times 10^{-11} M$$



Self-Ionization of Water (Kw): https://youtu.be/RMpO0rqUnFg

The Relationship among pH, pOH, and pK_w

The pH scale is a concise way of describing the H_3O^+ concentration and hence the acidity or basicity of a solution. Recall that pH and the H^+ (H_3O^+) concentration are related as follows:

$$pH = -\log_{10}[\mathrm{H}^+] \tag{15.4.10}$$

$$[\mathrm{H}^+] = 10^{-pH} \tag{15.4.11}$$

Because the scale is logarithmic, a pH difference of 1 between two solutions corresponds to a difference of a factor of 10 in their hydronium ion concentrations. Recall also that the pH of a neutral solution is 7.00 ($[H_3O^+] = 1.0 \times 10^{-7}\ M$), whereas acidic solutions have pH < 7.00 (corresponding to $[\mathrm{H_3O^+}] > 1.0 \times 10^{-7}$) and basic solutions have pH > 7.00 (corresponding to $[\mathrm{H_3O^+}] < 1.0 \times 10^{-7}$).

Similar notation systems are used to describe many other chemical quantities that contain a large negative exponent. For example, chemists use an analogous pOH scale to describe the hydroxide ion concentration of a solution. The pOH and $[\mathrm{OH^-}]$ are related as follows:

$$pOH = -\log_{10}[\mathrm{OH}^-] \tag{15.4.12}$$

$$[\mathrm{OH}^-] = 10^{-pOH} \tag{15.4.13}$$

The constant K_w can also be expressed using this notation, where $pK_w = -\log\ K_w$.

Because a neutral solution has $[OH^-] = 1.0 \times 10^{-7}$, the pOH of a neutral solution is 7.00. Consequently, the sum of the pH and the pOH for a neutral solution at 25°C is 7.00 + 7.00 = 14.00. We can show that the sum of pH and pOH is equal to 14.00 for any aqueous solution at 25°C by taking the negative logarithm of both sides of Equation ???:

$$-\log K_w = pK_w \tag{15.4.14}$$

$$= -\log([H_3O^+][OH^-]) \tag{15.4.15}$$

$$= (-\log[H_3O^+]) + (-\log[OH^-]) \tag{15.4.16}$$

$$= pH + pOH \tag{15.4.17}$$

Thus at any temperature, $pH + pOH = pK_w$, so at 25°C, where $K_w = 1.0 \times 10^{-14}$ and $pH + pOH = 14.00$. More generally, the pH of any neutral solution is half of the pK_w at that temperature. The relationship among pH, pOH, and the acidity or basicity of a solution is summarized graphically in Figure 15.4.1 over the common pH range of 0 to 14. Notice the inverse relationship between the pH and pOH scales.

> For any neutral solution, pH + pOH = 14.00 (at 25°C) with pH=pOH=7.

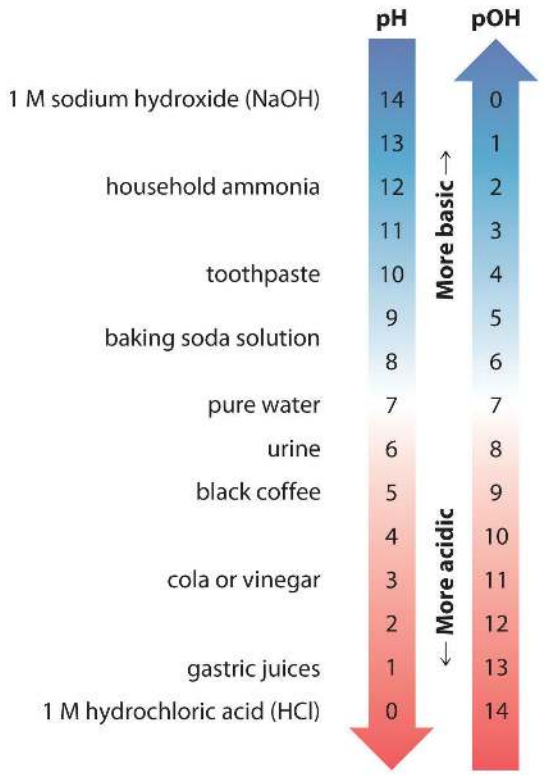


Figure 15.4.1: The Inverse Relationship between the pH and pOH Scales. As pH decreases, $[H^+]$ and the acidity increase. As pOH increases, $[OH^-]$ and the basicity decrease. Common substances have pH values that range from extremely acidic to extremely basic. (CC BY; OpenStax)

Example 15.4.3

The k_w for water at 100°C is 4.99×10^{-13}. Calculate pK_w for water at this temperature and the pH and the pOH for a neutral aqueous solution at 100°C. Report pH and pOH values to two decimal places.

Given: K_w

Asked for: pK_w, pH, and pOH

Strategy:

A. Calculate pKw by taking the negative logarithm of K_w.
B. For a neutral aqueous solution, $[H_3O^+] = [OH^-]$. Use this relationship and Equation ??? to calculate $[H_3O^+]$ and $[OH^-]$. Then determine the pH and the pOH for the solution.

Solution:

A Because pK_w is the negative logarithm of Kw, we can write

$$pK_w = -\log K_w = -\log(4.99 \times 10^{-13}) = 12.302$$

The answer is reasonable: K_w is between 10^{-13} and 10^{-12}, so pK_w must be between 12 and 13.

B Equation ??? shows that $K_w = [H_3O^+][OH^-]$. Because $[H_3O^+] = [OH^-]$ in a neutral solution, we can let $x = [H_3O^+] = [OH^-]$:

$$\begin{aligned} K_w &= [\mathrm{H_3O^+}][\mathrm{OH^-}] \\ &= (x)(x) \\ &= x^2 \\ x &= \sqrt{K_w} \\ &= \sqrt{4.99 \times 10^{-13}} \\ &= 7.06 \times 10^{-7}\ M \end{aligned}$$

Because x is equal to both $[\mathrm{H_3O^+}]$ and $[\mathrm{OH^-}]$,

$$\begin{aligned} pH = pOH\ &= -\log(7.06 \times 10^{-7}) \\ &= 6.15 \text{ (to two decimal places)} \end{aligned}$$

We could obtain the same answer more easily (without using logarithms) by using the pK_w. In this case, we know that $pK_w = 12.302$, and from Equation 15.4.17, we know that $pK_w = pH + pOH$. Because $pH = pOH$ in a neutral solution, we can use Equation 15.4.17directly, setting $pH = pOH = y$. Solving to two decimal places we obtain the following:

$$pK_w = pH + pOH = y + y = 2y$$

$$y = \frac{pK_w}{2} = \frac{12.302}{2} = 6.15 = pH = pOH$$

? Exercise 15.4.3

Humans maintain an internal temperature of about 37°C. At this temperature, $K_w = 3.55 \times 10^{-14}$. Calculate pK_w and the pH and the pOH of a neutral solution at 37°C. Report pH and pOH values to two decimal places.

Answer

$pK_w = 13.45$ and $pH = pOH = 6.73$



Introduction to pH: https://youtu.be/pQOa3bb5YEE

Summary

- For any neutral solution, $pH + pOH = 14.00$ (at 25°C) and $pH = \frac{1}{2}pK_w$.
- Ion-product constant of liquid water:

$$K_w = [H_3O^+][OH^-]$$

- Definition of pH:

$$pH = -\log_{10}[H^+]$$

or

$$[H^+] = 10^{-pH}$$

- Definition of pOH:

$$pOH = -\log_{10}[OH^+]$$

or

$$[OH^-] = 10^{-pOH}$$

- Relationship among pH, pOH, and pK_w:

$$pK_w = pH + pOH$$

Water is amphiprotic: it can act as an acid by donating a proton to a base to form the hydroxide ion, or as a base by accepting a proton from an acid to form the hydronium ion (H_3O^+). The autoionization of liquid water produces OH^- and H_3O^+ ions. The equilibrium constant for this reaction is called the ion-product constant of liquid water (K_w) and is defined as $K_w = [H_3O^+][OH^-]$. At 25°C, K_w is 1.01×10^{-14}; hence $pH + pOH = pK_w = 14.00$. The process is endothermic, and so the extent of ionization and the resulting concentrations of hydronium ion and hydroxide ion **increase with temperature**. For example, at 100 °C, the value for K_w is approximately 5.1×10^{-13}, roughly 100-times larger than the value at 25 °C.

Glossary

acid ionization

reaction involving the transfer of a proton from an acid to water, yielding hydronium ions and the conjugate base of the acid

autoionization

reaction between identical species yielding ionic products; for water, this reaction involves transfer of protons to yield hydronium and hydroxide ions

ion-product constant for water (K_w)

equilibrium constant for the autoionization of water

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 15.4: Self-Ionization of Water and the pH Scale is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.3: Self-Ionization of Water and the pH Scale** is licensed CC BY-NC-SA 4.0.
- **16.3: The Autoionization of Water** is licensed CC BY-NC-SA 3.0.

15.5: Strong Acids and Strong Bases

Learning Objectives

- Give the names and formulas of some strong acids and bases.
- Explain the pH scale, and convert pH and concentration of hydronium ions.
- Evaluate solution pH and pOH of strong acids or bases.

Acids and bases that are completely ionized when dissolved in water are called **strong acids** and **strong bases** There are only a few strong acids and bases, and everyone should know their names and properties. These acids are often used in industry and everyday life. The concentrations of acids and bases are often expressed in terms of pH, and as an educated person, you should have the skill to convert concentrations into pH and pOH. The pH is an indication of the hydrogen ion concentration, $[H^+]$.

Strong Acids

Strong acids are acids that are completely or nearly 100% ionized in their solutions; Table $15.5.1$ includes some common strong acids. Hence, the ionization in Equation 15.5.1 for a strong acid HA can be represented with a single arrow:

$$HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq) \tag{15.5.1}$$

Water is the base that reacts with the acid HA, A^- is the conjugate base of the acid HA, and the hydronium ion is the conjugate acid of water. By definition, a strong acid yields 100% of H_3O^+ and A^- when the acid ionizes in water. Table $15.5.1$ lists several strong acids.

Table 15.5.1: Some of the common strong acids and bases are listed here.

Strong Acids	Strong Bases
perchloric acid ($HClO_4$)	lithium hydroxide ($LiOH$)
hydrochloric acid (HCl)	sodium hydroxide ($NaOH$)
hydrobromic acid (HBr)	potassium hydroxide (KOH)
hydroiodic acid (HI)	calcium hydroxide ($Ca(OH)_2$)
nitric acid (HNO_3)	strontium hydroxide ($Sr(OH)_2$)
sulfuric acid (H_2SO_4)	barium hydroxide ($Ba(OH)_2$)

For a strong acid, $[H^+]$ = $[A^-]$ = concentration of acid if the concentration is much higher than $1 \times 10^{-7}\ M$. However, for a very dilute strong acid solution with concentration less than $1 \times 10^{-7}\ M$, the pH is dominated by the **autoionization of water**

$$H_2O \rightleftharpoons H^+ + OH^- \tag{15.5.2}$$

Example 15.5.1

Calculate the pH of a solution with $1.2345 \times 10^{-4}\ M HCl$, a strong acid.

Solution

The solution of a strong acid is completely ionized. That is, this equation goes to completion

$$HCl(aq) \longrightarrow H(aq) + Cl^-(aq)$$

Thus, $[H^+] = 1.2345 \times 10^{-4}$.

$$pH = -\log(1.2345 \times 10^{-4}) = 3.90851$$

Exercise 15.5.1

What is the pH for a solution containing 1.234 M $[\mathrm{HCl}]$?

Answer

pH = -0.0913

Example 15.5.2

Calculate the pH of a stock HCl solution that is 32% by mass HCl.

Solution

The density of such a solution is needed before we can calculate the pH. Since the density is not on the label, we need to find it from the Material Safety Data Sheet, which gives the specific gravity of 1.150. Thus, the amount of acid in 1.0 L is 1150 g.

$$\begin{aligned}\text{The amount of HCl} &= 1000 \times 1.150 \times 0.32 \\ &= 368 \text{ g} \left(\frac{1 \text{ mol}}{36.5 \text{ g}} \leftarrow \text{molar mass of HCl}\right) \\ &= 10.08 \text{ M} \\ &= [\mathrm{H^+}]\end{aligned}$$

$$\mathrm{pH} = -\log(10.08) = -1.003$$

Discussion

Yes, pH have negative values if $[\mathrm{H^+}] > 1.0$

Exercise 15.5.2

Check out the information on nitric acid, a strong acid, and calculate the pH of a stock nitric acid solution.

Example 15.5.3

Calculate the pH of a solution containing $1.00 \times 10^{-7}\ M$ of the strong acid HCl.

Solution

$[\mathrm{H^+}] = 1.0 \times 10^{-7}\ M$ from the strong acid, and if *x* is the amount from the ionization of water, then we have the equilibrium due to the **autoionization** of water. We can model this with an ICE table.

$$H_2O \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

ICE Table	$H_2O_{(l)}$	$H^+_{(aq)}$	$OH^-_{(aq)}$
Initial	-	1×10^{-7}	0
Change	-	+x	+x
Equilibrium	-	$1 \times 10^{-7} + x$	x

Recall that $K_w = [\mathrm{H^+}][\mathrm{OH^-}] = 1 \times 10^{-14}$, due to the ionization equilibrium of water in the solution:

$$\begin{aligned}(1.00 \times 10^{-7} + x)x &= 1 \times 10^{-14} \\ x^2 + 1.00 \times 10^{-7}x - 1.00 \times 10^{-14} &= 0\end{aligned}$$

Solving this equation for x from the quadratic equation results in

$$x = \frac{-1.00 \times 10^{-7} \pm \sqrt{1.00 \times 10^{-14} + (4)(1)(1.00 \times 10^{-14})}}{2}$$
$$= 0.61 \times 10^{-7}$$

only the additive root is physical (positive concentration); therefore

$$[\mathrm{H^+}] = (1.00 + 0.61) \times 10^{-7}\ M$$
$$\mathrm{pH} = -\log(1.61 \times 10^{-7})$$
$$= 6.79$$

Discussion

If you require only 1 significant figure, the pH is about 7.

Strong Bases

Strong bases are completely ionized in solution. Table 15.5.1 includes some common strong bases. For example, KOH dissolves in water in the reaction

$$\mathrm{KOH} \rightarrow \mathrm{K^+ + OH^-}$$

Relative to the number of strong acids, there are fewer number of strong bases and most are alkali hydroxides. Calcium hydroxide is considered a strong base, because it is completely, almost completely, ionized. However, the solubility of calcium hydroxide is very low. When $\mathrm{Ca(OH)_2}$ dissolves in water, the ionization reaction is as follows:

$$\mathrm{Ca(OH)_2} \rightarrow \mathrm{Ca^{2+} + 2\,OH^-}$$

Because of the stoichiometry of calcium hydroxide, upon dissociation, the concentration of $\mathrm{OH^-}$ will be twice the concentration of $\mathrm{Ca^{2+}}$:

$$[\mathrm{OH^-}] = 2[\mathrm{Ca^{2+}}]$$

✔ Example 15.5.4

Calculate the pOH of a solution containing $1.2345 \times 10^{-4}\ M\ \mathrm{Ca(OH)_2}$.

Solution

Based on complete ionization of

$$\mathrm{Ca(OH)_2} \rightarrow \mathrm{Ca^{+2} + 2\,OH^-}$$
$$[\mathrm{OH^-}] = 2 \times 1.2345 \times 10^{-4}$$
$$= 2.4690 \times 10^{-4}\ M$$
$$\mathrm{pOH} = -\log(2.4690 \times 10^{-4})$$
$$= 3.6074$$

? Exercise 15.5.4

The molar solubility of calcium hydroxide is 0.013 M $\mathrm{Ca(OH)_2}$. Calculate the pOH.

Answer

pOH = 1.58



Calculating pH in Strong Acid or Strong Base Solutions: https://youtu.be/NNTptn7hV2s

Questions

1. What is the pH of a solution containing 0.01 M HNO_3?
2. What is the pH of a solution containing 0.0220 M $Ba(OH)_2$? Give 3 significant figures.
3. Exactly 1.00 L solution was made by dissolving 0.80 g of $NaOH$ in water. What is $[H^+]$? (Atomic mass: Na, 23.0; O, 16.0; H, 1.0)
4. What is the pH for a solution which is 0.050 M HCl?
5. Which of the following is usually referred to as strong acid in water solution?

$$HF,\ HNO_2,\ H_2CO_3,\ H_2S,\ HSO_4^-,\ Cl^-,\ HNO_3,\ HCN$$

Solutions

1. ***Answer*** 2
   Hint...
   You do not need a calculator to evaluate $-\log(0.01) = 2$
2. ***Answer*** 12.64
   Hint...
   $Ba(OH)_2 \rightarrow Ba^{2+} + 2\ OH^-$
3. ***Answer*** 5.0×10^{-13}
   Hint...
   $[OH^-] = \dfrac{0.80}{40} = 0.020$ M; $[H^+] = \dfrac{1.0 \times 10^{-14}}{0.020} = 5 \times 10^{-13} M$. The pH is 12.30.
4. ***Answer*** 1.3
   Hint...
   This solution contains 1.83 g of HCl per liter. $[H^+] = 0.050$.
5. ***Answer*** HNO_3
   Consider...
   All others are weak acids

Contributors and Attributions

- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)

This page titled 15.5: Strong Acids and Strong Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.4: Strong Acids and Strong Bases** is licensed CC BY-NC-SA 4.0.

15.6: Weak Acids and Weak Bases

Learning Objectives

- To know the relationship between acid or base strength and the magnitude of K_a, K_b, pK_a, and pK_b.
- To understand the leveling effect.

The magnitude of the equilibrium constant for an ionization reaction can be used to determine the relative strengths of acids and bases. For example, the general equation for the ionization of a weak acid in water, where HA is the parent acid and A− is its conjugate base, is as follows:

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)} \tag{15.6.1}$$

The equilibrium constant for this dissociation is as follows:

$$K = \frac{[H_3O^+][A^-]}{[H_2O][HA]} \tag{15.6.2}$$

As we noted earlier, the concentration of water is essentially constant for all reactions in aqueous solution, so $[H_2O]$ in Equation 15.6.2 can be incorporated into a new quantity, the acid ionization constant (K_a), also called the acid dissociation constant:

$$K_a = K[H_2O] = \frac{[H_3O^+][A^-]}{[HA]} \tag{15.6.3}$$

Thus the numerical values of K and K_a differ by the concentration of water (55.3 M). Again, for simplicity, H_3O^+ can be written as H^+ in Equation 15.6.3. Keep in mind, though, that free H^+ does not exist in aqueous solutions and that a proton is transferred to H_2O in all acid ionization reactions to form H^3O^+. The larger the K_a, the stronger the acid and the higher the H^+ concentration at equilibrium.Like all equilibrium constants, acid–base ionization constants are actually measured in terms of the activities of H^+ or OH^-, thus making them **unitless**. The values of K_a for a number of common acids are given in Table 15.6.1.

Table 15.6.1: Values of K_a, pK_a, K_b, and pK_b for Selected Acids (HA and Their Conjugate Bases (A^-)

Acid	HA	K_a	pK_a	A^-	K_b	pK_b
hydroiodic acid	HI	2×10^9	−9.3	I^-	5.5×10^{-24}	23.26
sulfuric acid (1)*	H_2SO_4	1×10^2	−2.0	HSO_4^-	1×10^{-16}	16.0
nitric acid	HNO_3	2.3×10^1	−1.37	NO_3^-	4.3×10^{-16}	15.37
hydronium ion	H_3O^+	1.0	0.00	H_2O	1.0×10^{-14}	14.00
sulfuric acid (2)*	HSO_4^-	1.0×10^{-2}	1.99	SO_4^{2-}	9.8×10^{-13}	12.01
hydrofluoric acid	HF	6.3×10^{-4}	3.20	F^-	1.6×10^{-11}	10.80
nitrous acid	HNO_2	5.6×10^{-4}	3.25	$NO2^-$	1.8×10^{-11}	10.75
formic acid	HCO_2H	1.78×10^{-4}	3.750	HCO_2-	5.6×10^{-11}	10.25
benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}	4.20	$C_6H_5CO_2^-$	1.6×10^{-10}	9.80
acetic acid	CH_3CO_2H	1.7×10^{-5}	4.76	$CH_3CO_2^-$	5.8×10^{-10}	9.24
pyridinium ion	$C_5H_5NH^+$	5.9×10^{-6}	5.23	C_5H_5N	1.7×10^{-9}	8.77
hypochlorous acid	$HOCl$	4.0×10^{-8}	7.40	OCl^-	2.5×10^{-7}	6.60
hydrocyanic acid	HCN	6.2×10^{-10}	9.21	CN^-	1.6×10^{-5}	4.79
ammonium ion	NH_4^+	5.6×10^{-10}	9.25	NH_3	1.8×10^{-5}	4.75

*The number in parentheses indicates the ionization step referred to for a polyprotic acid.

Acid	HA	K_a	pK_a	A^-	K_b	pK_b
water	H_2O	1.0×10^{-14}	14.00	OH^-	1.00	0.00
acetylene	C_2H_2	1×10^{-26}	26.0	HC_2^-	1×10^{12}	−12.0
ammonia	NH_3	1×10^{-35}	35.0	NH_2^-	1×10^{21}	−21.0
*The number in parentheses indicates the ionization step referred to for a polyprotic acid.						

Weak bases react with water to produce the hydroxide ion, as shown in the following general equation, where B is the parent base and BH+ is its conjugate acid:

$$\mathrm{B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)} \tag{15.6.4}$$

The equilibrium constant for this reaction is the **base ionization constant** (K_b), also called the base dissociation constant:

$$K_b = K[H_2O] = \frac{[BH^+][OH^-]}{[B]} \tag{15.6.5}$$

Once again, the concentration of water is constant, so it does not appear in the equilibrium constant expression; instead, it is included in the K_b. The larger the K_b, the stronger the base and the higher the OH^- concentration at equilibrium. The values of K_b for a number of common weak bases are given in Table 15.6.2

Table 15.6.2: Values of K_b, pK_b, K_a, and pK_a for Selected Weak Bases (B) and Their Conjugate Acids (BH+)

Base	B	K_b	pK_b	BH^+	K_a	pK_a
hydroxide ion	OH^-	1.0	0.00*	H_2O	1.0×10^{-14}	14.00
phosphate ion	PO_4^{3-}	2.1×10^{-2}	1.68	HPO_4^{2-}	4.8×10^{-13}	12.32
dimethylamine	$(CH_3)_2NH$	5.4×10^{-4}	3.27	$(CH_3)_2NH_2^+$	1.9×10^{-11}	10.73
methylamine	CH_3NH_2	4.6×10^{-4}	3.34	$CH_3NH_3^+$	2.2×10^{-11}	10.66
trimethylamine	$(CH_3)_3N$	6.3×10^{-5}	4.20	$(CH_3)_3NH^+$	1.6×10^{-10}	9.80
ammonia	NH_3	1.8×10^{-5}	4.75	NH_4^+	5.6×10^{-10}	9.25
pyridine	C_5H_5N	1.7×10^{-9}	8.77	$C_5H_5NH^+$	5.9×10^{-6}	5.23
aniline	$C_6H_5NH_2$	7.4×10^{-10}	9.13	$C_6H_5NH_3^+$	1.3×10^{-5}	4.87
water	H_2O	1.0×10^{-14}	14.00	H_3O^+	1.0*	0.00
*As in Table 15.6.1.						

There is a simple relationship between the magnitude of K_a for an acid and K_b for its conjugate base. Consider, for example, the ionization of hydrocyanic acid (HCN) in water to produce an acidic solution, and the reaction of CN^- with water to produce a basic solution:

$$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-_{(aq)} \tag{15.6.6}$$

$$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HCN_{(aq)} \tag{15.6.7}$$

The equilibrium constant expression for the ionization of HCN is as follows:

$$K_a = \frac{[H^+][CN^-]}{[HCN]} \tag{15.6.8}$$

The corresponding expression for the reaction of cyanide with water is as follows:

$$K_b = \frac{[OH^-][HCN]}{[CN^-]} \tag{15.6.9}$$

If we add Equations 15.6.6 and 15.6.7, we obtain the following (recall that the equilibrium constant for the sum of two reactions is the product of the equilibrium constants for the individual reactions):

$$\cancel{HCN_{(aq)}} \rightleftharpoons H^+_{(aq)} + \cancel{CN^-_{(aq)}} \quad K_a = [H^+]\,[\cancel{CN^-}]/\,[\cancel{HCN}] \tag{15.6.10}$$

$$\cancel{CN^-_{(aq)}} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + \cancel{HCN_{(aq)}} \quad K_b = [OH^-]\,[\cancel{HCN}]/\,[\cancel{CN^-}] \tag{15.6.11}$$

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \quad K = K_a \times K_b = [H^+][OH^-] \tag{15.6.12}$$

In this case, the sum of the reactions described by K_a and K_b is the equation for the autoionization of water, and the product of the two equilibrium constants is K_w:

$$K_aK_b = K_w \tag{15.6.13}$$

Thus if we know either K_a for an acid or K_b for its conjugate base, we can calculate the other equilibrium constant for any conjugate acid–base pair.

Just as with pH, pOH, and pK_w, we can use negative logarithms to avoid exponential notation in writing acid and base ionization constants, by defining pK_a as follows:

$$pKa = -\log_{10} K_a \tag{15.6.14}$$

$$K_a = 10^{-pK_a} \tag{15.6.15}$$

and pK_b as

$$pK_b = -\log_{10} K_b \tag{15.6.16}$$

$$K_b = 10^{-pK_b} \tag{15.6.17}$$

Similarly, Equation 15.6.13, which expresses the relationship between K_a and K_b, can be written in logarithmic form as follows:

$$pK_a + pK_b = pK_w \tag{15.6.18}$$

At 25°C, this becomes

$$pK_a + pK_b = 14.00 \tag{15.6.19}$$

The values of pK_a and pK_b are given for several common acids and bases in Table 15.6.1 and Table 15.6.2, respectively, and a more extensive set of data is provided in Tables E1 and E2. Because of the use of negative logarithms, smaller values of pK_a correspond to larger acid ionization constants and hence stronger acids. For example, nitrous acid (HNO_2), with a pK_a of 3.25, is about a 1000 times stronger acid than hydrocyanic acid (HCN), with a pK_a of 9.21. Conversely, smaller values of pK_b correspond to larger base ionization constants and hence stronger bases.

The relative strengths of some common acids and their conjugate bases are shown graphically in Figure 16.5. The conjugate acid–base pairs are listed in order (from top to bottom) of increasing acid strength, which corresponds to decreasing values of pK_a. This order corresponds to decreasing strength of the conjugate base or increasing values of pK_b. At the bottom left of Figure 15.6.2 are the common strong acids; at the top right are the most common strong bases. Notice the inverse relationship between the strength of the parent acid and the strength of the conjugate base. Thus the conjugate base of a strong acid is a very weak base, and the conjugate base of a very weak acid is a strong base.

> *The conjugate base of a strong acid is a weak base and vice versa.*

We can use the relative strengths of acids and bases to predict the direction of an acid–base reaction by following a single rule: an acid–base equilibrium always favors the side with the weaker acid and base, as indicated by these arrows:

$$\text{stronger acid} + \text{stronger base} \rightleftharpoons \text{weaker acid} + \text{weaker base}$$

In an acid–base reaction, the proton always reacts with the stronger base.

For example, hydrochloric acid is a strong acid that ionizes essentially completely in dilute aqueous solution to produce H_3O^+ and Cl^-; only negligible amounts of HCl molecules remain undissociated. Hence the ionization equilibrium lies virtually all the way to the right, as represented by a single arrow:

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow\rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)} \tag{15.6.20}$$

In contrast, acetic acid is a weak acid, and water is a weak base. Consequently, aqueous solutions of acetic acid contain mostly acetic acid molecules in equilibrium with a small concentration of H_3O^+ and acetate ions, and the ionization equilibrium lies far to the left, as represented by these arrows:

$$\mathrm{CH_3CO_2H_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3CO^-_{2\,(aq)}}$$

Similarly, in the reaction of ammonia with water, the hydroxide ion is a strong base, and ammonia is a weak base, whereas the ammonium ion is a stronger acid than water. Hence this equilibrium also lies to the left:

$$H_2O_{(l)} + NH_{3(aq)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

> *All acid–base equilibria favor the side with the weaker acid and base. Thus the proton is bound to the stronger base.*

Example 15.6.1: Butyrate and Dimethylammonium Ions

a. Calculate K_b and pK_b of the butyrate ion ($\mathrm{CH_3CH_2CH_2CO_2^-}$). The pK_a of butyric acid at 25°C is 4.83. Butyric acid is responsible for the foul smell of rancid butter.
b. Calculate K_a and pK_a of the dimethylammonium ion ($(CH_3)_2NH_2^+$). The base ionization constant K_b of dimethylamine ($(CH_3)_2NH$) is 5.4×10^{-4} at 25°C.

Given: pK_a and K_b

Asked for: corresponding K_b and pK_b, K_a and pK_a

Strategy:

The constants K_a and K_b are related as shown in Equation 15.6.13. The pK_a and pK_b for an acid and its conjugate base are related as shown in Equation 15.6.18 and Equation 15.6.19. Use the relationships $pK = -\log K$ and $K = 10^{-pK}$ (Equations 15.6.14and 15.6.16) to convert between K_a and pK_a or K_b and pK_b.

Solution:

We are given the pK_a for butyric acid and asked to calculate the K_b and the pK_b for its conjugate base, the butyrate ion. Because the pK_a value cited is for a temperature of 25°C, we can use Equation 15.6.19: pK_a + pK_b = pKw = 14.00. Substituting the pK_a and solving for the pK_b,

$$\begin{aligned} 4.83 + pK_b &= 14.00 \\ pK_b &= 14.00 - 4.83 \\ &= 9.17 \end{aligned}$$

Because $pK_b = -\log K_b$, K_b is $10^{-9.17} = 6.8 \times 10^{-10}$.

In this case, we are given K_b for a base (dimethylamine) and asked to calculate K_a and pK_a for its conjugate acid, the dimethylammonium ion. Because the initial quantity given is K_b rather than pK_b, we can use Equation 15.6.13: $K_aK_b = K_w$. Substituting the values of K_b and K_w at 25°C and solving for K_a,

$$\begin{aligned} K_a(5.4 \times 10^{-4}) &= 1.01 \times 10^{-14} \\ K_a &= 1.9 \times 10^{-11} \end{aligned}$$

Because pK_a = −log K_a, we have $pK_a = -\log(1.9 \times 10^{-11}) = 10.72$. We could also have converted K_b to pK_b to obtain the same answer:

$$\begin{aligned} pK_b &= -\log(5.4 \times 10^{-4}) \\ &= 3.27 \end{aligned}$$

$$\begin{aligned} pKa + pK_b &= 14.00 \\ pK_a &= 10.73 \\ K_a &= 10^{-pK_a} \\ &= 10^{-10.73} \\ &= 1.9 \times 10^{-11} \end{aligned}$$

If we are given any one of these four quantities for an acid or a base (K_a, pK_a, K_b, or pK_b), we can calculate the other three.

? Exercise 15.6.1: Lactic Acid

Lactic acid ($CH_3CH(OH)CO_2H$) is responsible for the pungent taste and smell of sour milk; it is also thought to produce soreness in fatigued muscles. Its pK_a is 3.86 at 25°C. Calculate K_a for lactic acid and pK_b and K_b for the lactate ion.

Answer

$K_a = 1.4 \times 10^{-4}$ for lactic acid; pK_b = 10.14 and $K_b = 7.2 \times 10^{-11}$ for the lactate ion

Solutions of Strong Acids and Bases: The Leveling Effect

You will notice in Table 15.6.1 that acids like H_2SO_4 and HNO_3 lie above the hydronium ion, meaning that they have pK_a values less than zero and are stronger acids than the H_3O^+ ion. Recall that the acidic proton in virtually all oxoacids is bonded to one of the oxygen atoms of the oxoanion. Thus nitric acid should properly be written as $HONO_2$. Unfortunately, however, the formulas of oxoacids are almost always written with hydrogen on the left and oxygen on the right, giving HNO_3 instead. In fact, all six of the common strong acids that we encountered in in section 15.5 have pK_a values less than zero, which means that they have a greater tendency to lose a proton than does the H_3O^+ ion. Conversely, the conjugate bases of these strong acids are weaker bases than water. Consequently, the proton-transfer equilibria for these strong acids lie far to the right, and adding any of the common strong acids to water results in an essentially stoichiometric reaction of the acid with water to form a solution of the H_3O^+ ion and the conjugate base of the acid.

Although K_a for HI is about 108 greater than K_a for HNO_3, the reaction of either HI or HNO_3 with water gives an essentially stoichiometric solution of H_3O^+ and I− or NO_3^-. In fact, a 0.1 M aqueous solution of any strong acid actually contains 0.1 M H_3O^+, regardless of the identity of the strong acid. This phenomenon is called the leveling effect: any species that is a stronger acid than the conjugate acid of water (H_3O^+) is leveled to the strength of H_3O^+ in aqueous solution because H_3O^+ is the strongest acid that can exist in equilibrium with water. Consequently, it is impossible to distinguish between the strengths of acids such as HI and HNO3 in aqueous solution, and an alternative approach must be used to determine their relative acid strengths.

One method is to use a solvent such as anhydrous acetic acid. Because acetic acid is a stronger acid than water, it must also be a weaker base, with a lesser tendency to accept a proton than H_2O. Measurements of the conductivity of 0.1 M solutions of both HI and HNO_3 in acetic acid show that HI is completely dissociated, but HNO_3 is only partially dissociated and behaves like a weak acid in this solvent. This result clearly tells us that HI is a stronger acid than HNO_3. The relative order of acid strengths and approximate K_a and pK_a values for the strong acids at the top of Table 15.6.1 were determined using measurements like this and different nonaqueous solvents.

Note: Leveling Effect

In *aqueous solutions*, H_3O^+ is the strongest acid and OH^- is the strongest base that can exist in equilibrium with H_2O.

The leveling effect applies to solutions of strong bases as well: In aqueous solution, any base stronger than OH^- is leveled to the strength of OH^- because OH^- is the strongest base that can exist in equilibrium with water. Salts such as K_2O, $NaOCH_3$ (sodium methoxide), and $NaNH_2$ (sodamide, or sodium amide), whose anions are the conjugate bases of species that would lie

below water in Table 15.6.2 are all strong bases that react essentially completely (and often violently) with water, accepting a proton to give a solution of OH^- and the corresponding cation:

$$K_2O(s) + H_2O(l) \longrightarrow 2\,OH^-(aq) + 2\,K^+(aq)$$

$$NaOCH_3(s) + H_2O(l) \longrightarrow OH^-(aq) + Na^+(aq) + CH_3OH(aq)$$

$$NaNH_2(s) + H_2O(l) \longrightarrow OH^-(aq) + Na^+(aq) + NH_3(aq)$$

Other examples that you may encounter are potassium hydride (KH) and organometallic compounds such as methyl lithium (CH_3Li).



Calculating the pH of Weak Acids and Weak Bases: https://youtu.be/zr1V1THJ5P0

Summary

Two species that differ by only a proton constitute a conjugate acid–base pair. The magnitude of the equilibrium constant for an ionization reaction can be used to determine the relative strengths of acids and bases. For an aqueous solution of a weak acid, the dissociation constant is called the acid ionization constant (Ka). Similarly, the equilibrium constant for the reaction of a weak base with water is the base ionization constant (Kb). For any conjugate acid–base pair, $K_aK_b = K_w$. Smaller values of pK_a correspond to larger acid ionization constants and hence stronger acids. Conversely, smaller values of pK_b correspond to larger base ionization constants and hence stronger bases. At 25°C, $pK_a + pK_b = 14.00$. Acid–base reactions always proceed in the direction that produces the weaker acid–base pair. No acid stronger than H_3O^+ and no base stronger than OH^- can exist in aqueous solution, leading to the phenomenon known as the leveling effect.

Key Takeaways

- Acid–base reactions always contain two conjugate acid–base pairs.
- Each acid and each base has an associated ionization constant that corresponds to its acid or base strength.

Key Equations

- Acid ionization constant:

$$K_a = K[H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$$

- Base ionization constant:

$$K_b = K[H_2O] = \frac{[BH^+][OH^-]}{[B]}$$

- Relationship between K_a and K_b of a conjugate acid–base pair:

$$K_aK_b = K_w$$

- Definition of pK_a:

$$pKa = -\log_{10} K_a$$

$$K_a = 10^{-pK_a}$$

- Definition of pK_b:

$$pK_b = -\log_{10} K_b$$

$$K_b = 10^{-pK_b}$$

- Relationship between pK_a and pK_b of a conjugate acid–base pair:

$$pK_a + pK_b = pK_w$$

$$pK_a + pK_b = 14.00 \text{ at } 25^\circ \text{C}$$

This page titled 15.6: Weak Acids and Weak Bases is shared under a CC BY-NC-SA 3.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.5: Weak Acids and Weak Bases** by Anonymous is licensed CC BY-NC-SA 3.0.

15.7: Polyprotic Acids

Learning Objectives

- Extend previously introduced equilibrium concepts to acids and bases that may donate or accept more than one proton

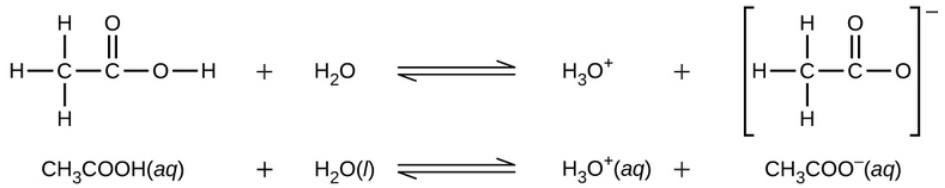
We can classify acids by the number of protons per molecule that they can give up in a reaction. Acids such as HCl, HNO_3, and HCN that contain one ionizable hydrogen atom in each molecule are called monoprotic acids. Their reactions with water are:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

$$HNO_3(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + NO_3^-(aq)$$

$$HCN(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + CN^-(aq)$$

Even though it contains four hydrogen atoms, acetic acid, CH_3CO_2H, is also monoprotic because only the hydrogen atom from the carboxyl group ($-COOH$) reacts with bases:



This image contains two equilibrium reactions. The first shows a C atom bonded to three H atoms and another C atom. The second C atom is double bonded to an O atom and also forms a single bond to another O atom. The second O atom is bonded to an H atom. There is a plus sign and then the molecular formula H subscript 2 O. An equilibrium arrow follows the H subscript 2 O. To the right of the arrow is H subscript 3 O superscript positive sign. There is a plus sign. The final structure shows a C atom bonded the three H atoms and another C atom. This second C atom is double bonded to an O atom and single bonded to another O atom. The entire structure is in brackets and a superscript negative sign appears outside the brackets. The second reaction shows C H subscript 3 C O O H (a q) plus H subscript 2 O (l) equilibrium arrow H subscript 3 O (a q) plus C H subscript 3 C O O superscript negative sign (a q).

Similarly, monoprotic bases are bases that will accept a single proton.

Diprotic Acids

Diprotic acids contain two ionizable hydrogen atoms per molecule; ionization of such acids occurs in two steps. The first ionization always takes place to a greater extent than the second ionization. For example, sulfuric acid, a strong acid, ionizes as follows:

- **The first ionization is**

$$H_2SO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$$

with $K_{a1} > 10^2$; *complete dissociation*.

- **The second ionization is**

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

with $K_{a2} = 1.2 \times 10^{-2}$.

This stepwise ionization process occurs for all polyprotic acids. When we make a solution of a weak diprotic acid, we get a solution that contains a mixture of acids. Carbonic acid, H_2CO_3, is an example of a weak diprotic acid. The first ionization of carbonic acid yields hydronium ions and bicarbonate ions in small amounts.

- **First Ionization**

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

with

$$K_{H_2CO_3} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = 4.3 \times 10^{-7}$$

The bicarbonate ion can also act as an acid. It ionizes and forms hydronium ions and carbonate ions in even smaller quantities.

- **Second Ionization**

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$$

with

$$K_{HCO_3^-} = \frac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]} = 4.7 \times 10^{-11}$$

$K_{H_2CO_3}$ is larger than $K_{HCO_3^-}$ by a factor of 10^4, so H_2CO_3 is the dominant producer of hydronium ion in the solution. This means that little of the HCO_3^- formed by the ionization of H_2CO_3 ionizes to give hydronium ions (and carbonate ions), and the concentrations of H_3O^+ and HCO_3^- are practically equal in a pure aqueous solution of H_2CO_3.

If the first ionization constant of a weak diprotic acid is larger than the second by a factor of at least 20, it is appropriate to treat the first ionization separately and calculate concentrations resulting from it before calculating concentrations of species resulting from subsequent ionization. This can simplify our work considerably because we can determine the concentration of H_3O^+ and the conjugate base from the first ionization, then determine the concentration of the conjugate base of the second ionization in a solution with concentrations determined by the first ionization. To perform this type of calculation, we will once again be using ICE Tables as the main tool.

Example 15.7.1: Ionization of a Diprotic Acid

When we buy soda water (carbonated water), we are buying a solution of carbon dioxide in water. The solution is acidic because CO_2 reacts with water to form carbonic acid, H_2CO_3. What are $[H_3O^+]$, $[HCO_3^-]$, and $[CO_3^{2-}]$ in a saturated solution of CO_2 with an initial $[H_2CO_3]$ = 0.033 *M*?

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq) \quad K_{a1} = 4.3 \times 10^{-7} \qquad \text{(equilibrium step 1)}$$

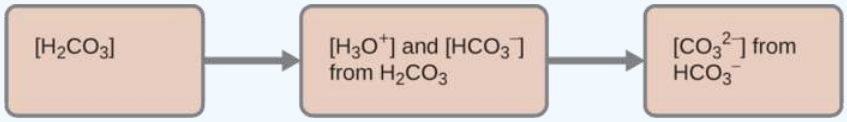
$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq) \quad K_{a2} = 4.7 \times 10^{-11} \qquad \text{(equilibrium step 2)}$$

Solution

As indicated by the ionization constants, H_2CO_3 is a much stronger acid than HCO_3^-, so H_2CO_3 is the dominant producer of hydronium ion in solution. Thus there are two parts in the solution of this problem:

1. Using the customary four steps, we determine the concentration of H_3O^+ and HCO_3^- produced by ionization of H_2CO_3.
2. Then we determine the concentration of CO_3^{2-} in a solution with the concentration of H_3O^+ and HCO_3^- determined in (1).

To summarize:



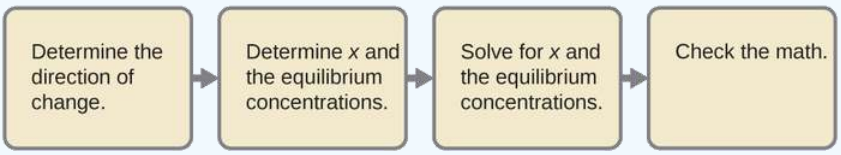
Four tan rectangles are shown that are connected with right pointing arrows. The first is labeled “left bracket H subscript 2 C O subscript 3 right bracket.” The second is labeled “left bracket H subscript 3 O superscript plus right bracket and left bracket H C O subscript 3 superscript negative right bracket from H subscript 2 C O subscript 3.” The third is labeled “left bracket C O subscript 3 superscript 2 negative right bracket from H C O subscript 3 superscript negative.”

1. First Ionization: Determine the concentrations of H_3O^+ *and* HCO_3^-.

Since equilibrium step 1 is has a much bigger $K_{a1} = 4.3 \times 10^{-7}$ than $K_{a2} = 4.7 \times 10^{-11}$ for equilibrium step 2, we can safely ignore the second ionization step and focus only on the first step (but address it in next part of problem).

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq) \quad K_{a1} = 4.3 \times 10^{-7}$$

As for the ionization of any other weak acid:



Four tan rectangles are shown that are connected with right pointing arrows. The first is labeled “Determine the direction of change.” The second is labeled “Determine x and the equilibrium concentrations.” The third is labeled “Solve for x and the equilibrium concentrations.” The fourth is labeled “Check the math.”

An abbreviated table of changes and concentrations shows:

Abbreviated table of changes and concentrations

ICE Table	$H_2CO_3(aq)$	$H_2O(l)$	$H_3O^+(aq)$	$HCO_3^-(aq)$
Initial (M)	$0.033\ M$	-	0	0
Change (M)	$-x$	-	$+x$	$+x$
Equilibrium (M)	$0.033\ M - x$	-	x	x

Substituting the equilibrium concentrations into the equilibrium constant gives us:

$$K_{H_2CO_3} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{(x)(x)}{0.033 - x} = 4.3 \times 10^{-7}$$

Solving the preceding equation making our standard assumptions gives:

$$x = 1.2 \times 10^{-4}$$

Thus:

$$[H_2CO_3] = 0.033\ M$$

$$[H_3O^+] = [HCO_3^-] = 1.2 \times 10^{-4}\ M$$

2. Second Ionization: Determine the concentration of CO_3^{2-} in a solution at equilibrium.

Since the equilibrium step 1 is has a much bigger K_a than equilibrium step 2, we can the equilibrium conditions calculated from first part of example as the initial conditions for an ICER Table for the equilibrium step 2:

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$$

ICER Table for the equilibrium step 2:

ICE Table	$HCO_3^-(aq)$	$H_2O(l)$	$H_3O^+(aq)$	$CO_3^{2-}(aq)$
Initial (M)	$1.2 \times 10^{-4}\ M$	-	$1.2 \times 10^{-4}\ M$	0
Change (M)	$-y$	-	$+y$	$+y$
Equilibrium (M)	$1.2 \times 10^{-4}\ M - y$	-	$1.2 \times 10^{-4}\ M + y$	y

$$K_{HCO_3^-} = \frac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]} = \frac{(1.2 \times 10^{-4}\ M + y)(y)}{(1.2 \times 10^{-4}\ M - y)}$$

To avoid solving a quadratic equation, we can assume $y \ll 1.2 \times 10^{-4}\ M$ so

$$K_{HCO_3^-} = 4.7 \times 10^{-11} \approx \frac{(1.2 \times 10^{-4}\ M)(y)}{(1.2 \times 10^{-4}\ M)}$$

Rearranging to solve for y

$$y \approx \frac{(4.7 \times 10^{-11})(1.2 \times 10^{-4}\ M)}{1.2 \times 10^{-4}\ M}$$

$$[\mathrm{CO_3^{2-}}] = y \approx 4.7 \times 10^{-11}$$

To summarize:

In part 1 of this example, we found that the $\mathrm{H_2CO_3}$ in a 0.033-M solution ionizes slightly and at equilibrium $[\mathrm{H_2CO_3}] = 0.033\ M$, $[\mathrm{H_3O^+}] = 1.2 \times 10^{-4}$, and $[\mathrm{HCO_3^-}] = 1.2 \times 10^{-4}\ M$. In part 2, we determined that $[\mathrm{CO_3^{2-}}] = 5.6 \times 10^{-11}\ M$.

? Exercise 15.7.2: Hydrogen Sulfide

The concentration of H_2S in a saturated aqueous solution at room temperature is approximately 0.1 M. Calculate $[\mathrm{H_3O^+}]$, $[\mathrm{HS^-}]$, and $[\mathrm{S^{2-}}]$ in the solution:

$$\mathrm{H_2S}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{HS^-}(aq) \quad K_{\mathrm{a1}} = 8.9 \times 10^{-8}$$

$$\mathrm{HS^-}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{S^{2-}}(aq) \quad K_{\mathrm{a2}} = 1.0 \times 10^{-19}$$

Answer

$[\mathrm{H_2S}] = 0.1M$, $[\mathrm{H_3O^+}] = [HS^-] = 0.0001\ M$, $[S^{2-}] = 1 \times 10^{-19}\ M$

We note that the concentration of the sulfide ion is the same as K_{a2}. This is due to the fact that each subsequent dissociation occurs to a lesser degree (as acid gets weaker).

Triprotic Acids

A triprotic acid is an acid that has three dissociable protons that undergo stepwise ionization: Phosphoric acid is a typical example:

- **The first ionization is**

$$\mathrm{H_3PO_4}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{H_2PO_4^-}(aq)$$

with $K_{\mathrm{a1}} = 7.5 \times 10^{-3}$.

- **The second ionization is**

$$\mathrm{H_2PO_4^-}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{HPO_4^{2-}}(aq)$$

with $K_{\mathrm{a2}} = 6.2 \times 10^{-8}$.

- **The third ionization is**

$$\mathrm{HPO_4^{2-}}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{PO_4^{3-}}(aq)$$

with $K_{\mathrm{a3}} = 4.2 \times 10^{-13}$.

As with the diprotic acids, the differences in the ionization constants of these reactions tell us that in each successive step the degree of ionization is significantly weaker. This is a general characteristic of polyprotic acids and successive ionization constants often differ by a factor of about 10^5 to 10^6. This set of three dissociation reactions may appear to make calculations of equilibrium concentrations in a solution of $\mathrm{H_3PO_4}$ complicated. However, because the successive ionization constants differ by a factor of 10^5 to 10^6, the calculations can be broken down into a series of parts similar to those for diprotic acids.

Polyprotic bases can accept more than one hydrogen ion in solution. The carbonate ion is an example of a diprotic base, since it can accept up to two protons. Solutions of alkali metal carbonates are quite alkaline, due to the reactions:

$$\mathrm{H_2O}(l) + \mathrm{CO_3^{2-}}(aq) \rightleftharpoons \mathrm{HCO_3^-}(aq) + \mathrm{OH^-}(aq)$$

and

$$\mathrm{H_2O}(l) + \mathrm{HCO_3^-}(aq) \rightleftharpoons \mathrm{H_2CO_3}(aq) + \mathrm{OH^-}(aq)$$

Summary

An acid that contains more than one ionizable proton is a polyprotic acid. The protons of these acids ionize in steps. The differences in the acid ionization constants for the successive ionizations of the protons in a polyprotic acid usually vary by roughly five orders of magnitude. As long as the difference between the successive values of K_a of the acid is greater than about a factor of 20, it is appropriate to break down the calculations of the concentrations of the ions in solution into a series of steps.

$$e = mc^2 \tag{15.7.1}$$

Glossary

diprotic acid
: acid containing two ionizable hydrogen atoms per molecule. A diprotic acid ionizes in two steps

diprotic base
: base capable of accepting two protons. The protons are accepted in two steps

monoprotic acid
: acid containing one ionizable hydrogen atom per molecule

stepwise ionization
: process in which an acid is ionized by losing protons sequentially

triprotic acid
: acid that contains three ionizable hydrogen atoms per molecule; ionization of triprotic acids occurs in three steps

---

This page titled 15.7: Polyprotic Acids is shared under a CC BY license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **14.5: Polyprotic Acids** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

15.8: Ions as Acids and Bases

Learning Objectives

- Extend previously introduced equilibrium concepts to acids and bases that may donate or accept more than one proton

We can also use the relative strengths of conjugate acid–base pairs to understand the acid–base properties of solutions of salts. A neutralization reaction can be defined as the reaction of an acid and a base to produce a salt and water. That is, another cation, such as Na^+, replaces the proton on the acid. An example is the reaction of CH_3CO_2H, a weak acid, with NaOH, a strong base:

$$\underset{acid}{CH_3CO_2H(l)} + \underset{base}{NaOH(s)} \xrightarrow{H_2O} H_2O\underset{salt}{CH_3CO_2Na(aq)} + \underset{water}{H_2O(l)} \tag{15.8.1}$$

Depending on the acid–base properties of its component ions, however, a salt can dissolve in water to produce a neutral solution, a basic solution, or an acidic solution.

When a salt such as $NaCl$ dissolves in water, it produces $Na^+_{(aq)}$ and $Cl^-_{(aq)}$ ions. Using a Lewis approach, the Na^+ ion can be viewed as an acid because it is an electron pair acceptor, although its low charge and relatively large radius make it a very weak acid. The Cl^- ion is the conjugate base of the strong acid HCl, so it has essentially no basic character. Consequently, dissolving $NaCl$ in water has no effect on the pH of a solution, and the solution remains neutral.

Now let's compare this behavior to the behavior of aqueous solutions of potassium cyanide and sodium acetate. Again, the cations (K^+ and Na^+) have essentially no acidic character, but the anions (CN^- and $CH_3CO_2^-$) are weak bases that can react with water because they are the conjugate bases of the weak acids HCN and acetic acid, respectively.

$$CN^-(aq) + H_2O(l) \rightleftharpoons HCN(aq) + OH^-(aq) \tag{15.8.2}$$

$$CH_3CO_2^-(aq) + H_2O(l) \rightleftharpoons CH_3CO_2H(aq) + OH^-(aq) \tag{15.8.3}$$

Neither reaction proceeds very far to the right as written because the formation of the weaker acid–base pair is favored. Both HCN and acetic acid are stronger acids than water, and hydroxide is a stronger base than either acetate or cyanide, so in both cases, the equilibrium lies to the left. Nonetheless, each of these reactions generates enough hydroxide ions to produce a basic solution. For example, the pH of a 0.1 M solution of sodium acetate or potassium cyanide at 25°C is 8.8 or 11.1, respectively. From Table E1 and E2, we can see that CN^- is a stronger base ($pK_b = 4.79$) than acetate ($pK_b = 9.24$), which is consistent with KCN producing a more basic solution than sodium acetate at the same concentration.

In contrast, the conjugate acid of a weak base should be a weak acid. For example, ammonium chloride and pyridinium chloride are salts produced by reacting ammonia and pyridine, respectively, with HCl. As you already know, the chloride ion is such a weak base that it does not react with water. In contrast, the cations of the two salts are weak acids that react with water as follows:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq) \tag{15.8.4}$$

$$C_5H_5NH^+(aq) + H_2O(l) \rightleftharpoons C_5H_5NH(aq) + H_3O^+(aq) \tag{15.8.5}$$

Figure 15.8.2 shows that H_3O^+ is a stronger acid than either NH_4^+ or $C_5H_5NH^+$, and conversely, ammonia and pyridine are both stronger bases than water. The equilibrium will therefore lie far to the left in both cases, favoring the weaker acid–base pair. The H_3O^+ concentration produced by the reactions is great enough, however, to decrease the pH of the solution significantly: the pH of a 0.10 M solution of ammonium chloride or pyridinium chloride at 25°C is 5.13 or 3.12, respectively.

What happens with aqueous solutions of a salt such as ammonium acetate, where both the cation and the anion can react separately with water to produce an acid and a base, respectively? According to Figure 16.10, the ammonium ion will lower the pH, while according to Figure 16.9, the acetate ion will raise the pH. This particular case is unusual, in that the cation is as strong an acid as the anion is a base (pKa ≈ pKb). Consequently, the two effects cancel, and the solution remains neutral. With salts in which the cation is a stronger acid than the anion is a base, the final solution has a pH < 7.00. Conversely, if the cation is a weaker acid than the anion is a base, the final solution has a pH > 7.00.



Ions as Acids and Bases: https://youtu.be/XYAGNonPSow

Acidic Metal Ions

Solutions of simple salts of metal ions can also be acidic, even though a metal ion cannot donate a proton directly to water to produce H_3O^+. Instead, a metal ion can act as a Lewis acid and interact with water, a Lewis base, by coordinating to a lone pair of electrons on the oxygen atom to form a hydrated metal ion (Figure 15.8.1*a*). A water molecule coordinated to a metal ion is more acidic than a free water molecule for two reasons. First, repulsive electrostatic interactions between the positively charged metal ion and the partially positively charged hydrogen atoms of the coordinated water molecule make it easier for the coordinated water to lose a proton.

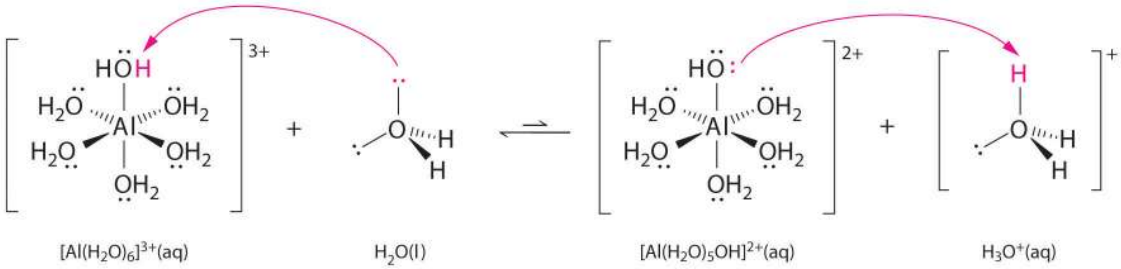
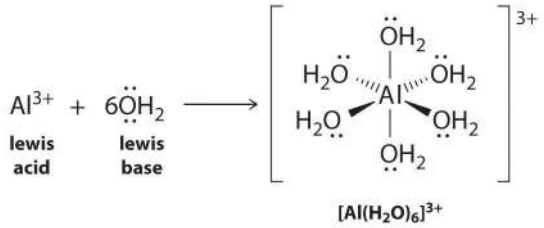
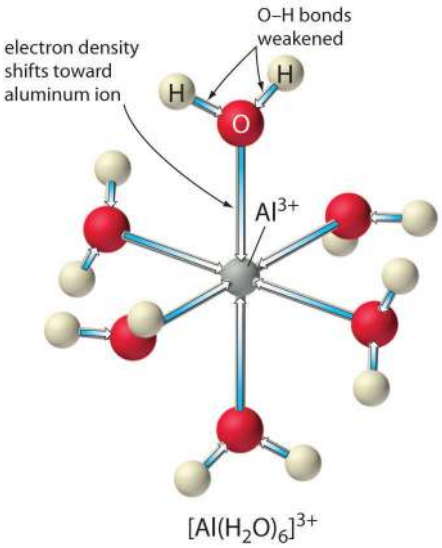


Figure 15.8.1

Second, the positive charge on the Al^{3+} ion attracts electron density from the oxygen atoms of the water molecules, which decreases the electron density in the O–H bonds, as shown in Figure 15.8.1*b* With less electron density between the O atoms and the H atoms, the O–H bonds are weaker than in a free H_2O molecule, making it easier to lose a H^+ ion.



(a)



(b)

Figure 15.8.2: Effect of a Metal Ion on the Acidity of Water (a) Reaction of the metal ion Al^{3+} with water to form the hydrated metal ion is an example of a Lewis acid–base reaction. (b) The positive charge on the aluminum ion attracts electron density from the oxygen atoms, which shifts electron density away from the O–H bonds. The decrease in electron density weakens the O–H bonds in the water molecules and makes it easier for them to lose a proton.

The magnitude of this effect depends on the following two factors (Figure 15.8.3):

1. The charge on the metal ion. A divalent ion (M^{2+}) has approximately twice as strong an effect on the electron density in a coordinated water molecule as a monovalent ion (M^{+}) of the same radius.
2. The radius of the metal ion. For metal ions with the same charge, the smaller the ion, the shorter the internuclear distance to the oxygen atom of the water molecule and the greater the effect of the metal on the electron density distribution in the water molecule.

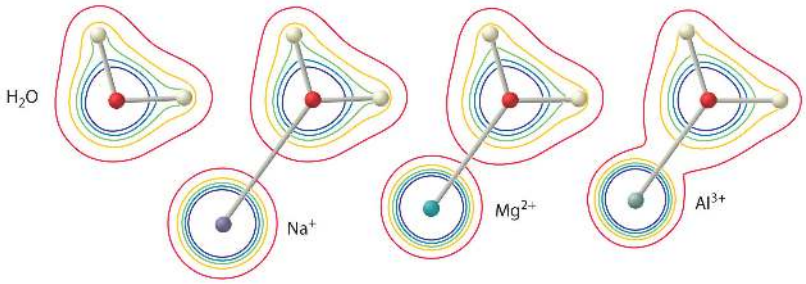


Figure 15.8.3: The Effect of the Charge and Radius of a Metal Ion on the Acidity of a Coordinated Water Molecule. The contours show the electron density on the O atoms and the H atoms in both a free water molecule (left) and water molecules coordinated to Na^{+}, Mg^{2+}, and Al^{3+} ions. These contour maps demonstrate that the smallest, most highly charged metal ion (Al^{3+}) causes the greatest decrease in electron density of the O–H bonds of the water molecule. Due to this effect, the acidity of hydrated metal ions increases as the charge on the metal ion increases and its radius decreases.

Thus aqueous solutions of small, highly charged metal ions, such as Al^{3+} and Fe^{3+}, are acidic:

$$[Al(H_2O)_6]^{3+}_{(aq)} \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}_{(aq)} + H^+_{(aq)} \tag{15.8.6}$$

The $[Al(H_2O)_6]^{3+}$ ion has a pK_a of 5.0, making it almost as strong an acid as acetic acid. Because of the two factors described previously, the most important parameter for predicting the effect of a metal ion on the acidity of coordinated water molecules is the charge-to-radius ratio of the metal ion. A number of pairs of metal ions that lie on a diagonal line in the periodic table, such as Li^+ and Mg^{2+} or Ca^{2+} and Y^{3+}, have different sizes and charges, but similar charge-to-radius ratios. As a result, these pairs of metal ions have similar effects on the acidity of coordinated water molecules, and they often exhibit other significant similarities in chemistry as well.

Acidic Metal Ions

Solutions of small, highly charged metal ions in water are acidic.

Reactions such as those discussed in this section, in which a salt reacts with water to give an acidic or basic solution, are often called hydrolysis reactions. Using a separate name for this type of reaction is unfortunate because it suggests that they are somehow different. In fact, hydrolysis reactions are just acid–base reactions in which the acid is a cation or the base is an anion; they obey the same principles and rules as all other acid–base reactions.

Hydrolysis

A hydrolysis reaction is an acid–base reaction.

Example 15.8.1

Predict whether aqueous solutions of these compounds are acidic, basic, or neutral.

a. KNO_3
b. $CrBr_3 \cdot H_2O$
c. Na_2SO_4

Given: compound

Asked for: acidity or basicity of aqueous solution

Strategy:

A. Assess the acid–base properties of the cation and the anion. If the cation is a weak Lewis acid, it will not affect the pH of the solution. If the cation is the conjugate acid of a weak base or a relatively highly charged metal cation, however, it will react with water to produce an acidic solution.
B. f the anion is the conjugate base of a strong acid, it will not affect the pH of the solution. If, however, the anion is the conjugate base of a weak acid, the solution will be basic.

Solution:

a

A. The K^+ cation has a small positive charge (+1) and a relatively large radius (because it is in the fourth row of the periodic table), so it is a very weak Lewis acid.
B. The NO_3- anion is the conjugate base of a strong acid, so it has essentially no basic character (Table 16.1). Hence neither the cation nor the anion will react with water to produce H^+ or OH^-, and the solution will be neutral.

b.

A. The Cr^{3+} ion is a relatively highly charged metal cation that should behave similarly to the Al^{3+} ion and form the [Cr(H2O)6]3+ complex, which will behave as a weak acid:

$$Cr(H_2O)_6]^{3+}_{(aq)} \rightleftharpoons Cr(H_2O)_5(OH)]^{2+}_{(aq)} + H^+_(aq) \tag{15.8.7}$$

B. The Br^- anion is a very weak base (it is the conjugate base of the strong acid HBr), so it does not affect the pH of the solution. Hence the solution will be acidic.

c.

A. The Na^+ ion, like the K^+, is a very weak acid, so it should not affect the acidity of the solution.
B. In contrast, SO_4^{2-} is the conjugate base of HSO_4^-, which is a weak acid. Hence the SO_4^{2-} ion will react with water as shown in Figure 16.6 to give a slightly basic solution.

? Exercise 15.8.1

Predict whether aqueous solutions of the following are acidic, basic, or neutral.

a. KI
b. $Mg(ClO_4)_2$
c. $NaHS$

Answer a

neutral

Answer b

acidic

Answer c

basic (due to the reaction of HS^- with water to form H_2S and OH^-)

Summary

A salt can dissolve in water to produce a neutral, a basic, or an acidic solution, depending on whether it contains the conjugate base of a weak acid as the anion (A^-), the conjugate acid of a weak base as the cation (BH^+), or both. Salts that contain small, highly charged metal ions produce acidic solutions in water. The reaction of a salt with water to produce an acidic or a basic solution is called a hydrolysis reaction.

This page titled 15.8: Ions as Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.7: Ions as Acids and Bases** is licensed CC BY-NC-SA 4.0.

15.9: Molecular Structure and Acid-Base Behavior

Learning Objectives

- To understand how molecular structure affects the strength of an acid or base.

We have seen that the strengths of acids and bases vary over many orders of magnitude. In this section, we explore some of the structural and electronic factors that control the acidity or basicity of a molecule.

Bond Strengths

In general, the stronger the A– H or B– H^+ bond, the less likely the bond is to break to form H^+ ions and thus the less acidic the substance. This effect can be illustrated using the hydrogen halides:

Relative Acid Strength	HF	HCl	HBr	HI
H–X Bond Energy (kJ/mol)	570	432	366	298
pKa	3.20	−6.1	−8.9	−9.3

The trend in bond energies is due to a steady decrease in overlap between the 1s orbital of hydrogen and the valence orbital of the halogen atom as the size of the halogen increases. The larger the atom to which H is bonded, the weaker the bond. Thus the bond between H and a large atom in a given family, such as I or Te, is weaker than the bond between H and a smaller atom in the same family, such as F or O. As a result, acid strengths of binary hydrides increase as we go down a column of the periodic table. For example, the order of acidity for the binary hydrides of Group 16 elements is as follows, with pK_a values in parentheses:

$$H_2O(14.00 = pK_w) < H_2S(7.05) < H_2Se(3.89) < H_2Te(2.6) \tag{15.9.1}$$

Stability of the Conjugate Base

Whether we write an acid–base reaction as $AH \rightleftharpoons A^- + H^+$ or as $BH^+ \rightleftharpoons B + H^+$, the conjugate base ($A^-$ or B) contains one more lone pair of electrons than the parent acid (AH or BH^+). Any factor that stabilizes the lone pair on the conjugate base favors dissociation of H^+ and makes the parent acid a stronger acid. Let's see how this explains the relative acidity of the binary hydrides of the elements in the second row of the periodic table. The observed order of increasing acidity is the following, with pKa values in parentheses:

$$CH_4(\ 50) \ll NH_3(\ 36) < H_2O(14.00) < HF(3.20) \tag{15.9.2}$$

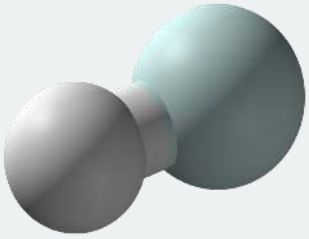
Consider, for example, the compounds at both ends of this series: methane and hydrogen fluoride. The conjugate base of CH_4 is CH_3^-, and the conjugate base of HF is F^-. Because fluorine is much more electronegative than carbon, fluorine can better stabilize the negative charge in the F^- ion than carbon can stabilize the negative charge in the CH3− ion. Consequently, HF has a greater tendency to dissociate to form H^+ and F^- than does methane to form H^+ and CH_3^-, making HF a much stronger acid than CH_4.

The same trend is predicted by analyzing the properties of the conjugate acids. For a series of compounds of the general formula HE, as the electronegativity of E increases, the E–H bond becomes more polar, favoring dissociation to form E^- and H^+. Due to both the increasing stability of the conjugate base and the increasing polarization of the E–H bond in the conjugate acid, acid strengths of binary hydrides increase as we go from left to right across a row of the periodic table.

Acid strengths of binary hydrides increase as we go down a column or from left to right across a row of the periodic table.

the strongest acid Known: The hydrohelium Cation

The stornger acid, the weaker the covalent bond to a hydrogen atom. So the strongest acid possible is the molecule with the weakest bond. That is the hydrohelium (1+) cation, HeH^+, which is a positively charged ion formed by the reaction of a proton with a helium atom in the gas phase. It was first produced in the laboratory in 1925 and is isoelectronic with molecular hydrogen (\ce{H2}). It is the strongest known acid, with a proton affinity of 177.8 kJ/mol.



Ball and stick model of the hydrohelium ion. (CC BY-SA 3.0; CCoil).

HeH^+ cannot be prepared in a condensed phase, as it would protonate any anion, molecule or atom with which it were associated. However it is possible to estimate a *hypothetical* aqueous acidity using Hess's law:

$HHe^+(g)$	→	$H^+(g)$	+ He(g)	+178 kJ/mol
$HHe^+(aq)$	→	$HHe^+(g)$		+973 kJ/mol
$H^+(g)$	→	$H^+(aq)$		−1530 kJ/mol
He(g)	→	He(aq)		+19 kJ/mol
$HHe^+(aq)$	→	$H^+(aq)$	+ He(aq)	−360 kJ/mol

A free energy change of dissociation of −360 kJ/mol is equivalent to a pK_a of −63.

It has been suggested that HeH^+ should occur naturally in the interstellar medium, but it has not yet been detected.

Inductive Effects

Atoms or groups of atoms in a molecule other than those to which H is bonded can induce a change in the distribution of electrons within the molecule. This is called an inductive effect, and, much like the coordination of water to a metal ion, it can have a major effect on the acidity or basicity of the molecule. For example, the hypohalous acids (general formula HOX, with X representing a halogen) all have a hydrogen atom bonded to an oxygen atom. In aqueous solution, they all produce the following equilibrium:

$$HOX_{(aq)} \rightleftharpoons H^+_{(aq)} + OX^-(aq) \tag{15.9.3}$$

The acidities of these acids vary by about three orders of magnitude, however, due to the difference in electronegativity of the halogen atoms:

HOX	Electronegativity of X	pKa
HOCl	3.0	7.40
HOBr	2.8	8.55
HOI	2.5	10.5

As the electronegativity of X increases, the distribution of electron density within the molecule changes: the electrons are drawn more strongly toward the halogen atom and, in turn, away from the H in the O–H bond, thus weakening the O–H bond and allowing dissociation of hydrogen as H^+.

The acidity of oxoacids, with the general formula $HOXO_n$ (with n = 0−3), depends strongly on the number of terminal oxygen atoms attached to the central atom X. As shown in Figure 15.9.1, the K_a values of the oxoacids of chlorine increase by a factor of about 10^4 to 10^6 with each oxygen as successive oxygen atoms are added. The increase in acid strength with increasing number of terminal oxygen atoms is due to both an inductive effect and increased stabilization of the conjugate base.

> *Any inductive effect that withdraws electron density from an O–H bond increases the acidity of the compound.*

Because oxygen is the second most electronegative element, adding terminal oxygen atoms causes electrons to be drawn away from the O–H bond, making it weaker and thereby increasing the strength of the acid. The colors in Figure 15.9.1 show how the electrostatic potential, a measure of the strength of the interaction of a point charge at any place on the surface of the molecule, changes as the number of terminal oxygen atoms increases. In Figure 15.9.1 and Figure 15.9.2, blue corresponds to low electron densities, while red corresponds to high electron densities. The oxygen atom in the O–H unit becomes steadily less red from $HClO$ to $HClO_4$ (also written as $HOClO_3$, while the H atom becomes steadily bluer, indicating that the electron density on the O–H unit decreases as the number of terminal oxygen atoms increases. The decrease in electron density in the O–H bond weakens it, making it easier to lose hydrogen as H^+ ions, thereby increasing the strength of the acid.

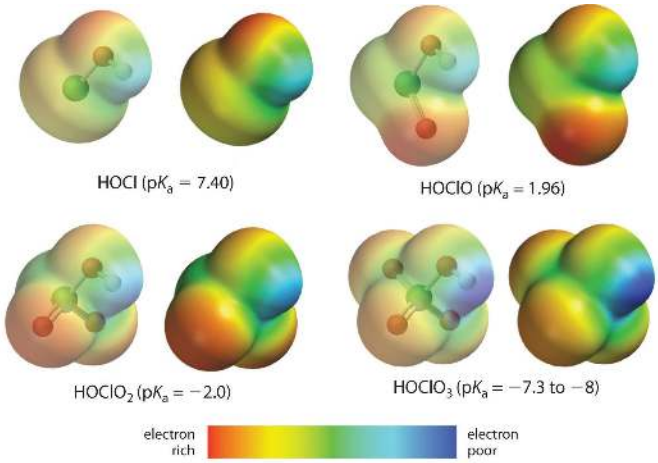
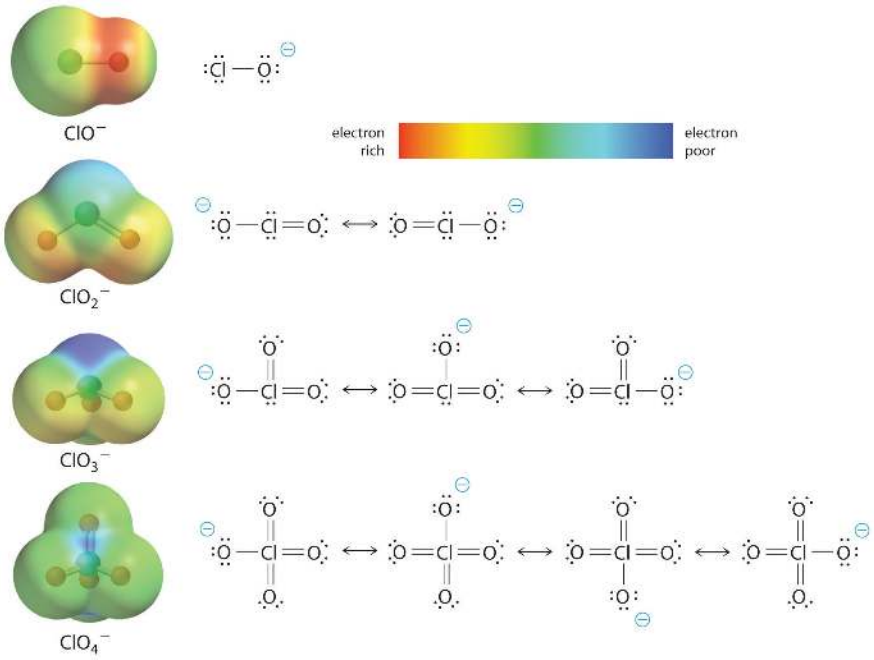


Figure 15.9.1: The Relationship between the Acid Strengths of the Oxoacids of Chlorine and the Electron Density on the O–H Unit. These electrostatic potential maps show how the electron density on the O–H unit decreases as the number of terminal oxygen atoms increases. Blue corresponds to low electron densities, whereas red corresponds to high electron densities. Source: Chlorine oxoacids pKa values from J. R. Bowser, Inorganic Chemistry (Pacific Grove, CA: Brooks-Cole,1993).

At least as important, however, is the effect of delocalization of the negative charge in the conjugate base. As shown in Figure 15.9.2, the number of resonance structures that can be written for the oxoanions of chlorine increases as the number of terminal oxygen atoms increases, allowing the single negative charge to be delocalized over successively more oxygen atoms.

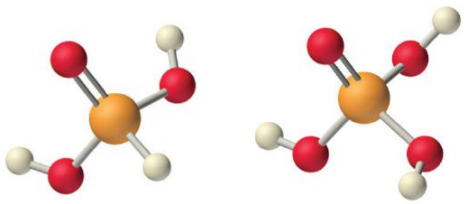
Electron delocalization in the conjugate base increases acid strength.

The electrostatic potential plots in Figure 15.9.2 demonstrate that the electron density on the terminal oxygen atoms decreases steadily as their number increases. The oxygen atom in ClO^- is red, indicating that it is electron rich, and the color of oxygen progressively changes to green in ClO_4^+, indicating that the oxygen atoms are becoming steadily less electron rich through the series. For example, in the perchlorate ion (ClO_4^-), the single negative charge is delocalized over all four oxygen atoms, whereas in the hypochlorite ion (OCl^-), the negative charge is largely localized on a single oxygen atom (Figure 15.9.2). As a result, the perchlorate ion has no localized negative charge to which a proton can bind. Consequently, the perchlorate anion has a much lower affinity for a proton than does the hypochlorite ion, and perchloric acid is one of the strongest acids known.



As the number of terminal oxygen atoms increases, the number of resonance structures that can be written for the oxoanions of chlorine also increases, and the single negative charge is delocalized over more oxygen atoms. As these electrostatic potential plots demonstrate, the electron density on the terminal oxygen atoms decreases steadily as their number increases. As the electron density on the oxygen atoms decreases, so does their affinity for a proton, making the anion less basic. As a result, the parent oxoacid is more acidic.

Similar inductive effects are also responsible for the trend in the acidities of oxoacids that have the same number of oxygen atoms as we go across a row of the periodic table from left to right. For example, H_3PO_4 is a weak acid, H_2SO_4 is a strong acid, and $HClO_4$ is one of the strongest acids known. The number of terminal oxygen atoms increases steadily across the row, consistent with the observed increase in acidity. In addition, the electronegativity of the central atom increases steadily from P to S to Cl, which causes electrons to be drawn from oxygen to the central atom, weakening the O–H bond and increasing the strength of the oxoacid.



Careful inspection of the data in Table 15.9.1 shows two apparent anomalies: carbonic acid and phosphorous acid. If carbonic acid (H_2CO_3) were a discrete molecule with the structure $(HO)_2C{=}O$, it would have a single terminal oxygen atom and should be comparable in acid strength to phosphoric acid (H_3PO_4), for which pKa1 = 2.16. Instead, the tabulated value of pK_{a1} for carbonic acid is 6.35, making it about 10,000 times weaker than expected. As we shall see, however, H_2CO_3 is only a minor component of the aqueous solutions of CO_2 that are referred to as carbonic acid. Similarly, if phosphorous acid (H_3PO_3) actually had the structure $(HO)_3P$, it would have no terminal oxygen atoms attached to phosphorous. It would therefore be expected to be about as strong an acid as $HOCl$ (pKa = 7.40). In fact, the pK_{a1} for phosphorous acid is 1.30, and the structure of phosphorous acid is $(HO)_2P(=O)H$ with one H atom directly bonded to P and one P=O bond. Thus the pKa1 for phosphorous acid is similar to that of other oxoacids with one terminal oxygen atom, such as H_3PO_4. Fortunately, phosphorous acid is the only common oxoacid in which a hydrogen atom is bonded to the central atom rather than oxygen.

Table 15.9.1: Values of pKa for Selected Polyprotic Acids and Bases

Polyprotic Acids	Formula	pK_{a1}	pK_{a2}	pK_{a3}
carbonic acid*	"H_2CO_3"	6.35	10.33	
citric acid	$HO_2CCH-2C(OH)(CO_2H)CH_2CO_2H$	3.13	4.76	6.40
malonic acid	$HO-2CCH_2CO_2H$	2.85	5.70	
oxalic acid	HO_2CCO_2H	1.25	3.81	
phosphoric acid	H_3PO_4	2.16	7.21	12.32
phosphorous acid	H_3PO_3	1.3	6.70	
succinic acid	$HO_2CCH_2CH_2CO_2H$	4.21	5.64	
sulfuric acid	H_2SO_4	−2.0	1.99	
sulfurous acid*	"H_2SO_3"	1.85	7.21	
Polyprotic Bases	**Formula**	pK_{b1}	pK_{b2}	
ethylenediamine	$H_2N(CH_2)_2NH_2$	4.08	7.14	
piperazine	$HN(CH_2CH_2)_2NH$	4.27	8.67	
propylenediamine	$H_2N(CH_2)_3NH_2$	3.45	5.12	

*H_2CO_3 and H_2SO_3 are at best minor components of aqueous solutions of $CO_{2(g)}$ and $SO_{2(g)}$, respectively, but such solutions are commonly referred to as containing carbonic acid and sulfurous acid, respectively.

Inductive effects are also observed in organic molecules that contain electronegative substituents. The magnitude of the electron-withdrawing effect depends on both the nature and the number of halogen substituents, as shown by the pKa values for several acetic acid derivatives:

$$pK_aCH_3CO_2H4.76 < CH_2ClCO_2H2.87 < CHCl_2CO_2H1.35 < CCl_3CO_2H0.66 < CF_3CO_2H0.52$$

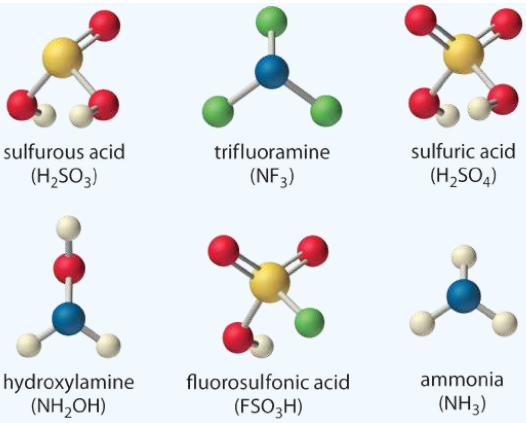
As you might expect, fluorine, which is more electronegative than chlorine, causes a larger effect than chlorine, and the effect of three halogens is greater than the effect of two or one. Notice from these data that inductive effects can be quite large. For instance, replacing the $-CH_3$ group of acetic acid by a $-CF_3$ group results in about a 10,000-fold increase in acidity!

Example 15.9.1

Arrange the compounds of each series in order of increasing acid or base strength.

a. sulfuric acid [H_2SO_4, or $(HO)_2SO_2$], fluorosulfonic acid (FSO_3H, or FSO_2OH), and sulfurous acid [H_2SO_3, or $(HO)_2SO$]
b. ammonia (NH_3), trifluoramine (NF_3), and hydroxylamine (NH_2OH)

The structures are shown here.



Given: series of compounds

Asked for: relative acid or base strengths

Strategy:

Use relative bond strengths, the stability of the conjugate base, and inductive effects to arrange the compounds in order of increasing tendency to ionize in aqueous solution.

Solution:

Although both sulfuric acid and sulfurous acid have two –OH groups, the sulfur atom in sulfuric acid is bonded to two terminal oxygen atoms versus one in sulfurous acid. Because oxygen is highly electronegative, sulfuric acid is the stronger acid because the negative charge on the anion is stabilized by the additional oxygen atom. In comparing sulfuric acid and fluorosulfonic acid, we note that fluorine is more electronegative than oxygen. Thus replacing an –OH by –F will remove more electron density from the central S atom, which will, in turn, remove electron density from the S–OH bond and the O–H bond. Because its O–H bond is weaker, FSO_3H is a stronger acid than sulfuric acid. The predicted order of acid strengths given here is confirmed by the measured pKa values for these acids:

$$pKaH_2SO_31.85 < H_2SO_4^{-2} < FSO_3H - 10$$

The structures of both trifluoramine and hydroxylamine are similar to that of ammonia. In trifluoramine, all of the hydrogen atoms in NH3 are replaced by fluorine atoms, whereas in hydroxylamine, one hydrogen atom is replaced by OH. Replacing the three hydrogen atoms by fluorine will withdraw electron density from N, making the lone electron pair on N less available to bond to an H^+ ion. Thus NF_3 is predicted to be a much weaker base than NH_3. Similarly, because oxygen is more electronegative than hydrogen, replacing one hydrogen atom in NH_3 by OH will make the amine less basic. Because oxygen is less electronegative than fluorine and only one hydrogen atom is replaced, however, the effect will be smaller. The predicted order of increasing base strength shown here is confirmed by the measured pK_b values:

$$pK_bNF_3 — << NH_2OH8.06 < NH_34.75$$

Trifluoramine is such a weak base that it does not react with aqueous solutions of strong acids. Hence its base ionization constant has never been measured.

? Exercise 15.9.1

Arrange the compounds of each series in order of

a. decreasing acid strength: H_3PO_4, $CH_3PO_3H_2$, and $HClO_3$.
b. increasing base strength: CH_3S^-, OH^-, and CF_3S^-.

Answer a

$HClO-3 > CH_3PO_3H_2 > H_3PO_4$

Answer a

$CF_3S^- < CH_3S^- < OH^-$

Summary

Inductive effects and charge delocalization significantly influence the acidity or basicity of a compound. The acid–base strength of a molecule depends strongly on its structure. The weaker the A–H or B–H+ bond, the more likely it is to dissociate to form an H^+ ion. In addition, any factor that stabilizes the lone pair on the conjugate base favors the dissociation of H^+, making the conjugate acid a stronger acid. Atoms or groups of atoms elsewhere in a molecule can also be important in determining acid or base strength through an inductive effect, which can weaken an O– H bond and allow hydrogen to be more easily lost as H^+ ions.

Contributors and Attributions

- Anonymous

This page titled 15.9: Molecular Structure and Acid-Base Behavior is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.8: Molecular Structure and Acid-Base Behavior** is licensed CC BY-NC-SA 4.0.

15.10: Lewis Acids and Bases

Learning Objectives

Make sure you thoroughly understand the following essential ideas which have been presented. It is especially important that you know the precise meanings of all the highlighted terms in the context of this topic.

- Write the equation for the proton transfer reaction involving a Brønsted-Lowry acid or base, and show how it can be interpreted as an **electron-pair transfer** reaction, clearly identifying the donor and acceptor.
- Give an example of a **Lewis acid-base reaction** that does not involve protons.
- Write equations illustrating the behavior of a given **non-aqueous acid-base system**.

The Brønsted-Lowry proton donor-acceptor concept has been one of the most successful theories of Chemistry. But as with any such theory, it is fair to ask if this is not just a special case of a more general theory that could encompass an even broader range of chemical science. In 1916, G.N. Lewis of the University of California proposed that the *electron pair* is the dominant actor in acid-base chemistry. The Lewis theory did not become very well known until about 1923 (the same year that Brønsted and Lowry published their work), but since then it has been recognized as a very powerful tool for describing chemical reactions of widely different kinds and is widely used in organic and inorganic chemistry. The Brønsted–Lowry concept of acids and bases defines a base as any species that can accept a proton, and an acid as any substance that can donate a proton. Lewis proposed an alternative definition that focuses on *pairs of electrons* instead. According to Lewis:

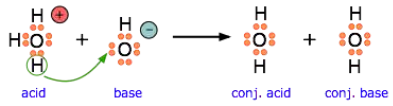
- An *acid* is a substance that **accepts** a pair of electrons, and in doing so, forms a covalent bond with the entity that supplies the electrons.
- A *base* is a substance that **donates** an unshared pair of electrons to a recipient species with which the electrons can be shared.

Note

In modern chemistry, electron donors are often referred to as nucleophiles, while acceptors are electrophiles.

Lewis Acid-Base Neutralization Involving Electron-Pair Transfer

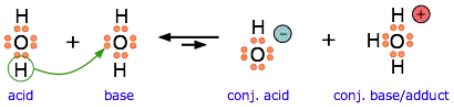
Just as any Arrhenius acid is also a Brønsted acid, any Brønsted acid is also a Lewis acid, so the various acid-base concepts are all "upward compatible". Although we do not really need to think about electron-pair transfers when we deal with ordinary aqueous-solution acid-base reactions, it is important to understand that it is the opportunity for electron-pair sharing that enables proton transfer to take place.



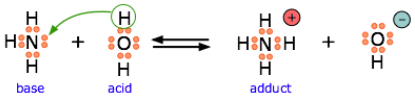
This equation for a simple acid-base neutralization shows how the Brønsted and Lewis definitions are really just different views of the same process. Take special note of the following points:

- The arrow shows the movement of a proton from the hydronium ion to the hydroxide ion.
- Note that the electron-pairs themselves do not move; they remain attached to their central atoms. The electron pair on the base is "donated" to the acceptor (the proton) only in the sense that it ends up being *shared* with the acceptor, rather than being the exclusive property of the oxygen atom in the hydroxide ion.
- Although the hydronium ion is the nominal Lewis acid here, it does not itself accept an electron pair, but acts merely as the source of the proton that coordinates with the Lewis base.

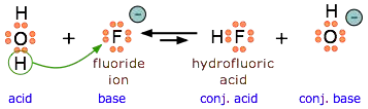
The point about the electron-pair remaining on the donor species is especially important to bear in mind. For one thing, it distinguishes a *Lewis acid-base reaction* from an *oxidation-reduction reaction*, in which a physical transfer of one or more electrons from donor to acceptor does occur. The product of a Lewis acid-base reaction is known formally as an "adduct" or "complex", although we do not ordinarily use these terms for simple proton-transfer reactions such as the one in the above example. Here, the proton combines with the hydroxide ion to form the "adduct" H_2O. The following examples illustrate these points for some other proton-transfer reactions that you should already be familiar with.



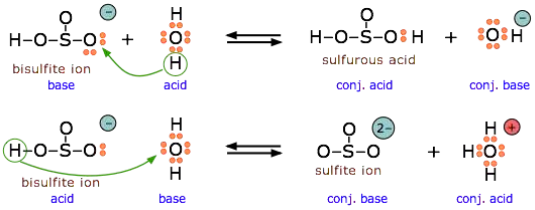
Another example, showing the autoprotolysis of water. Note that the conjugate base is also the adduct.



Ammonia is both a Brønsted and a Lewis base, owing to the unshared electron pair on the nitrogen. The reverse of this reaction represents the *hydrolysis* of the ammonium ion.



Because HF is a weak acid, fluoride salts behave as bases in aqueous solution. As a Lewis base, F^- accepts a proton from water, which is transformed into a hydroxide ion.

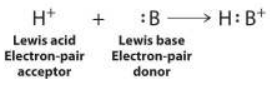


The bisulfite ion is *amphiprotic* and can act as an electron donor or acceptor.

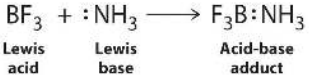
> **Note**
>
> All Brønsted–Lowry bases (proton acceptors), such as OH^-, H_2O, and NH_3, are also electron-pair donors. Thus the Lewis definition of acids and bases does not contradict the Brønsted–Lowry definition. Rather, it expands the definition of acids to include substances other than the H^+ ion.

Lewis Acid-Base Neutralization without Transferring Protons

Electron-deficient molecules, such as BCl_3, contain less than an octet of electrons around one atom and have a strong tendency to gain an additional pair of electrons by reacting with substances that possess a lone pair of electrons. Lewis's definition, which is less restrictive than either the Brønsted–Lowry or the Arrhenius definition, grew out of his observation of this tendency. A general Brønsted–Lowry acid–base reaction can be depicted in Lewis electron symbols as follows:

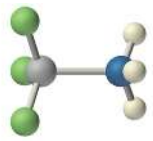


The proton (H^+), which has no valence electrons, is a Lewis acid because it accepts a lone pair of electrons on the base to form a bond. The proton, however, is just one of many electron-deficient species that are known to react with bases. For example, neutral compounds of boron, aluminum, and the other Group 13 elements, which possess only six valence electrons, have a very strong tendency to gain an additional electron pair. Such compounds are therefore potent Lewis acids that react with an electron-pair donor such as ammonia to form an acid–base adduct, a new covalent bond, as shown here for boron trifluoride (BF_3):



The bond formed between a Lewis acid and a Lewis base is a *coordinate covalent bond* because both electrons are provided by only one of the atoms (N, in the case of $F_3B{:}NH_3$). After it is formed, however, a coordinate covalent bond behaves like any other

covalent single bond.



Boron trifluoride-ammonia adduct

Species that are very weak Brønsted–Lowry bases can be relatively strong Lewis bases. For example, many of the group 13 trihalides are highly soluble in ethers (R–O–R′) because the oxygen atom in the ether contains two lone pairs of electrons, just as in H_2O. Hence the predominant species in solutions of electron-deficient trihalides in ether solvents is a Lewis acid–base adduct. A reaction of this type is shown in Figure 15.10.1for boron trichloride and diethyl ether:

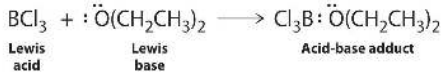
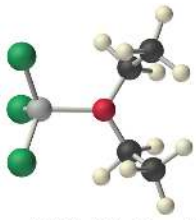


Figure 15.10.1 *Lewis Acid/Base reaction of boron trichloride and diethyl ether reaction*

Note

- Electron-deficient molecules (those with less than an octet of electrons) are Lewis acids.
- The acid-base behavior of many compounds can be explained by their Lewis electron structures.

Many molecules with multiple bonds can act as Lewis acids. In these cases, the Lewis base typically donates a pair of electrons to form a bond to the central atom of the molecule, while a pair of electrons displaced from the multiple bond becomes a lone pair on a terminal atom.



Boron trichloride-diethyl ether adduct

Figure 15.10.2. The highly electronegative oxygen atoms pull electron density away from carbon, so the carbon atom acts as a Lewis acid. Arrows indicate the direction of electron flow.

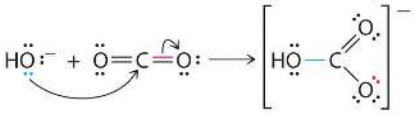


Figure 15.10.2 *Lewis Acid/Base reaction of the hydroxide ion with carbon dioxide*

Example 15.10.1

Identify the acid and the base in each Lewis acid–base reaction.

a. $BH_3 + (CH_3)_2S \rightarrow H_3B{:}S(CH_3)_2$
b. $CaO + CO_2 \rightarrow CaCO_3$
c. $BeCl_2 + 2\ Cl^- \rightarrow BeCl_4^{2-}$

Given: reactants and products

Asked for: identity of Lewis acid and Lewis base

Strategy:

In each equation, identify the reactant that is electron deficient and the reactant that is an electron-pair donor. The electron-deficient compound is the Lewis acid, whereas the other is the Lewis base.

Solution:

a. In BH_3, boron has only six valence electrons. It is therefore electron deficient and can accept a lone pair. Like oxygen, the sulfur atom in $(CH_3)_2S$ has two lone pairs. Thus $(CH_3)_2S$ donates an electron pair on sulfur to the boron atom of BH_3. The Lewis base is $(CH_3)_2S$, and the Lewis acid is BH_3.
b. As in the reaction shown in Equation 8.21, CO_2 accepts a pair of electrons from the O^{2-} ion in CaO to form the carbonate ion. The oxygen in CaO is an electron-pair donor, so CaO is the Lewis base. Carbon accepts a pair of electrons, so CO_2 is the Lewis acid.
c. The chloride ion contains four lone pairs. In this reaction, each chloride ion donates one lone pair to $BeCl_2$, which has only four electrons around Be. Thus the chloride ions are Lewis bases, and $BeCl_2$ is the Lewis acid.

? Exercise 15.10.1

Identify the acid and the base in each Lewis acid–base reaction.

a. $(CH_3)_2O + BF_3 \rightarrow (CH_3)_2O{:}BF_3$
b. $H_2O + SO_3 \rightarrow H_2SO_4$

Answer

a. Lewis base: $(CH_3)_2O$; Lewis acid: BF_3
b. Lewis base: H_2O; Lewis acid: SO_3

? Exercise 15.10.2

Here are several more examples of Lewis acid-base reactions that **cannot** be accommodated within the Brønsted or Arrhenius models. Identify the Lewis acid and Lewis base in each reaction.

a. $Al(OH)_3 + OH^- \rightarrow Al(OH)_4^-$
b. $SnS_2 + S^{2-} \rightarrow SnS_3^{2-}$
c. $Cd(CN)_2 + 2CN^- \rightarrow Cd(CN)_4^{2+}$
d. $AgCl + 2NH_3 \rightarrow Ag(NH_3)_2^+ + Cl^-$
e. $Fe^{2+} + NO \rightarrow Fe(NO)^{2+}$
f. $[Ni^{2+} + 6NH_3 \rightarrow Ni(NH_3)_5^{2+}$

Non-Aqueous Acid/Base Systems

We ordinarily think of Brønsted-Lowry acid-base reactions as taking place in aqueous solutions, but this need not always be the case. A more general view encompasses a variety of acid-base *solvent systems*, of which the *water system* is only one (Table 15.10.1). Each of these has as its basis an amphiprotic solvent (one capable of undergoing autoprotolysis), in parallel with the familiar case of water.

The limiting acid in a given solvent is the solvonium ion, such as H_3O^+ (hydronium) ion in water. An acid which has more of a tendency to donate a hydrogen ion than the limiting acid will be a strong acid in the solvent considered, and will exist mostly or entirely in its dissociated form. Likewise, the limiting base in a given solvent is the solvate ion, such as OH^- (hydroxide) ion, in water. A base which has more affinity for protons than the limiting base cannot exist in solution, as it will react with the solvent.

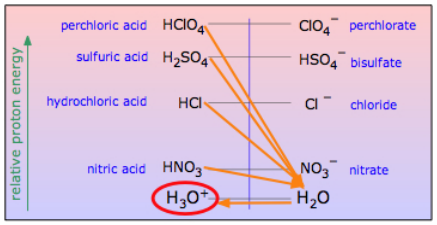


Figure } \PageIndex{3}\): *Use of non-aqueous solvents allows the study of strong acids that are hindered by the "leveling effect" of the solvent.*

The ammonia system is one of the most common non-aqueous system in Chemistry. Liquid ammonia boils at –33° C, and can conveniently be maintained as a liquid by cooling with dry ice (–77° C). It is a good solvent for substances that also dissolve in water, such as ionic salts and organic compounds since it is capable of forming hydrogen bonds. Bases can exist in solution in

liquid ammonia which cannot exist in aqueous solution: this is the case for any base which is **stronger** than the hydroxide ion, but **weaker** than the amide ion NH_2^-. The limiting acid in liquid ammonia is the ammonium ion, which has a pK_a value in water of 9.25. The limiting base is the amide ion, NH_2^-. This is a stronger base than the hydroxide ion and so cannot exist in aqueous solution. The pK_a value of ammonia is estimated to be approximately 33. Any acid which is a stronger acid than the ammonium ion will be a strong acid in liquid ammonia. This is the case for acetic acid, which is completely dissociated in liquid ammonia solution. The addition of pure acetic acid and the addition of ammonium acetate have exactly the same effect on a liquid ammonia solution: the increase in its acidity: in practice, the latter is preferred for safety reasons.

Note

As with OH^- and H_3O^+ in water, the strongest acid and base in NH_3 is dictated by the corresponding autoprotolysis reaction of the solvent:

$$2NH_3 \rightleftharpoons NH_4^+ + NH_2^-$$

One use of non-aqueous acid-base systems is to examine the relative strengths of the strong acids and bases, whose strengths are "leveled" by the fact that they are all totally converted into H_3O^+ or OH^- ions in water. By studying them in appropriate non-aqueous solvents which are poorer acceptors or donors of protons, their relative strengths can be determined. Many familiar substances can serve as the basis of protonic solvent systems (Table 15.10.1).

Table 15.10.1: Popular Solvent systems

solvent	autoprotolysis reaction	pK_a
water	$2H_2O \rightleftharpoons H_3O^+ + OH^-$	14
ammonia	$2NH_3 \rightleftharpoons NH_4^+ + NH_2^-$	33
acetic acid	$2CH_3COOH \rightleftharpoons CH_3COOH_2^+ + CH_3COO^-$	13
ethanol	$2C_2H_5OH \rightleftharpoons C_2H_5OH_2^+ + C_2H_5O^-$	19
hydrogen peroxide	$2HO-OH \rightleftharpoons HO-OH_2^+ + HO-O^-$	13
hydrofluoric acid	$2HF \rightleftharpoons H^2F^+ + F^-$	10
sulfuric acid	$2H_2SO_4 \rightleftharpoons H3SO_4^+ + HSO_4^-$	3.5

The extreme case is a **superacid**, a medium in which the hydrogen ion is only very weakly solvated. The classic example is a mixture of antimony pentafluoride and liquid hydrogen fluoride:

$$SbF_5 + HF \rightleftharpoons H^+ + SbF_6^- \tag{15.10.1}$$

The limiting base, the hexfluoroantimonate anion SbF_6^-, is so weakly attracted to the hydrogen ion that virtually any other base will bind more strongly: hence, this mixture can be used to protonate organic molecules which would not be considered bases in other solvents. It should noted that pH is undefined in aprotic solvents, which assumes presence of hydronium ions. In other solvents, the concentration of the respective solvonium/solvate ions should be used (e.g., $[NH_4^+]$ and $[NH_2^-]$ in $NH_{3(l)}$.

Contributors and Attributions

- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Wikipedia

This page titled 15.10: Lewis Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.9: Lewis Acids and Bases** is licensed CC BY-NC-SA 4.0.

15.11: Common-Ion Effect in Acid-Base Equilibria

Learning Objectives

- Recognize common ions from various salts, acids, and bases.
- Calculate concentrations involving common ions.
- Calculate ion concentrations involving chemical equilibrium.

The common-ion effect is used to describe the effect on an equilibrium involving a substance that adds an ion that is a part of the equilibrium.

Introduction

The solubility products K_{sp}'s are equilibrium constants in heterogeneous equilibria (i.e., between two different phases). If several salts are present in a system, they all ionize in the solution. If the salts contain a common cation or anion, these salts contribute to the concentration of the common ion. Contributions from all salts must be included in the calculation of concentration of the common ion. For example, a solution containing sodium chloride and potassium chloride will have the following relationship:

$$[\mathrm{Na^+}] + [\mathrm{K^+}] = [\mathrm{Cl^-}] \tag{15.11.1}$$

Consideration of *charge balance* or *mass balance* or both leads to the same conclusion.

Common Ions

When NaCl and KCl are dissolved in the same solution, the $\mathrm{Cl^-}$ ions are *common* to both salts. In a system containing NaCl and KCl, the $\mathrm{Cl^-}$ ions are common ions.

$$\mathrm{NaCl} \rightleftharpoons \mathrm{Na^+} + \mathrm{Cl^-}$$

$$\mathrm{KCl} \rightleftharpoons \mathrm{K^+} + \mathrm{Cl^-}$$

$$\mathrm{CaCl_2} \rightleftharpoons \mathrm{Ca^{2+}} + 2\mathrm{Cl^-}$$

$$\mathrm{AlCl_3} \rightleftharpoons \mathrm{Al^{3+}} + 3\mathrm{Cl^-}$$

$$\mathrm{AgCl} \rightleftharpoons \mathrm{Ag^+} + \mathrm{Cl^-}$$

For example, when AgCl is dissolved into a solution already containing NaCl (actually $\mathrm{Na^+}$ and $\mathrm{Cl^-}$ ions), the $\mathrm{Cl^-}$ ions come from the ionization of both AgCl and NaCl. Thus, $[\mathrm{Cl^-}]$ differs from $[\mathrm{Ag^+}]$. The following examples show how the concentration of the common ion is calculated.

Example 15.11.1

What are $[\mathrm{Na^+}]$, $[\mathrm{Cl^-}]$, $[\mathrm{Ca^{2+}}]$, and $[\mathrm{H^+}]$ in a solution containing 0.10 M each of NaCl, $\mathrm{CaCl_2}$, and HCl?

Solution

Due to the conservation of ions, we have

$$[\mathrm{Na^+}] = [\mathrm{Ca^{2+}}] = [\mathrm{H^+}] = 0.10\ \mathrm{M}$$

but

$$\begin{aligned}[\mathrm{Cl^-}] = &\ 0.10\ (\text{due to NaCl}) \\ &+ 0.20\ (\text{due to } \mathrm{CaCl_2}) \\ &+ 0.10\ (\text{due to HCl}) \\ = &\ 0.40\ \mathrm{M}\end{aligned}$$

Exercise 15.11.1

John poured 10.0 mL of 0.10 M NaCl, 10.0 mL of 0.10 M KOH, and 5.0 mL of 0.20 M HCl solutions together and then he made the total volume to be 100.0 mL. What is $[\mathrm{Cl^-}]$ in the final solution?

Solution

$$[\mathrm{Cl^-}] = \frac{0.1\ \mathrm{M} \times 10\ \mathrm{mL} + 0.2\ \mathrm{M} \times 5.0\ \mathrm{mL}}{100.0\ \mathrm{mL}} = 0.020\ \mathrm{M}$$

Le Châtelier's Principle states that if an equilibrium becomes unbalanced, the reaction will shift to restore the balance. If a common ion is added to a weak acid or weak base equilibrium, then the equilibrium will shift towards the reactants, in this case the weak acid or base.

Example 15.11.2: Solubility of Lead Chloride

Consider the lead(II) ion concentration in this *saturated* solution of $PbCl_2$. The balanced reaction is

$$PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

Defining s as the concentration of dissolved lead(II) chloride, then:

$$[Pb^{2+}] = s$$

$$[Cl^-] = 2s$$

These values can be substituted into the solubility product expression, which can be solved for s:

$$K_{sp} = [Pb^{2+}][Cl^-]^2 \tag{15.11.2}$$

$$= s \times (2s)^2 \tag{15.11.3}$$

$$1.7 \times 10^{-5} = 4s^3 \tag{15.11.4}$$

$$s^3 = \frac{1.7 \times 10^{-5}}{4} \tag{15.11.5}$$

$$= 4.25 \times 10^{-6} \tag{15.11.6}$$

$$s = \sqrt[3]{4.25 \times 10^{-6}} \tag{15.11.7}$$

$$= 1.62 \times 10^{-2}\ mol\ dm^{-3} \tag{15.11.8}$$

The concentration of lead(II) ions in the solution is 1.62 x 10^{-2} M. Consider what happens if sodium chloride is added to this saturated solution. Sodium chloride shares an ion with lead(II) chloride. The chloride ion is **common** to both of them; this is the origin of the term "common ion effect".

Look at the original equilibrium expression again:

$$PbCl_2\ (s) \rightleftharpoons Pb^{2+}\ (aq) + 2Cl^-\ (aq)$$

What happens to that equilibrium if extra chloride ions are added? According to *Le* Châtelier, the position of equilibrium will shift to counter the change, in this case, by removing the chloride ions by making extra solid lead(II) chloride.

Of course, the concentration of lead(II) ions in the solution is so small that only a tiny proportion of the extra chloride ions can be converted into solid lead(II) chloride. The lead(II) chloride becomes even *less soluble*, and the concentration of lead(II) ions in the solution *decreases*. This type of response occurs with any sparingly soluble substance: it is less soluble in a solution which contains any ion which it has in common. This is the common ion effect.

Example 15.11.3

If an attempt is made to dissolve some lead(II) chloride in some 0.100 M sodium chloride solution instead of in water, what is the equilibrium concentration of the lead(II) ions this time? As before, define s to be the concentration of the lead(II) ions.

$$[Pb^{2+}] = s \tag{15.11.9}$$

The calculations are different from before. This time the concentration of the chloride ions is governed by the concentration of the sodium chloride solution. The number of ions coming from the lead(II) chloride is going to be tiny compared with the 0.100 M coming from the sodium chloride solution.

In calculations like this, it can be assumed that the concentration of the common ion is entirely due to the other solution. This simplifies the calculation.

So we assume:

$$[Cl^-] = 0.100\ M \tag{15.11.10}$$

The rest of the mathematics looks like this:

$$\begin{aligned} K_{sp} &= [Pb^{2+}][Cl^-]^2 \\ &= s \times (0.100)^2 \\ 1.7 \times 10^{-5} &= s \times 0.00100 \end{aligned}$$

therefore:

$$\begin{aligned} s &= \frac{1.7 \times 10^{-5}}{0.0100} \\ &= 1.7 \times 10^{-3}\ \text{M} \end{aligned} \tag{15.11.11}$$

Finally, compare that value with the simple saturated solution:

Original solution:

$$[Pb^{2+}] = 0.0162\ M \tag{15.11.12}$$

Solution in 0.100 M NaCl solution:

$$[Pb^{2+}] = 0.0017\ M \tag{15.11.13}$$

The concentration of the lead(II) ions has decreased by a factor of about 10. If more concentrated solutions of sodium chloride are used, the solubility decreases further.



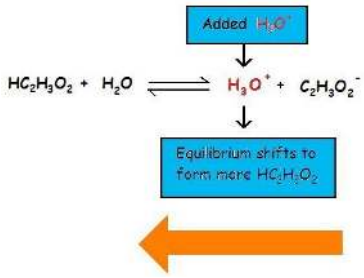
A Video Discussing Finding the Solubility of a Salt: Finding the Solubility of a Salt(opens in new window) [youtu.be]

Common Ion Effect with Weak Acids and Bases

Adding a common ion prevents the weak acid or weak base from ionizing as much as it would without the added common ion. The common ion effect suppresses the ionization of a weak acid by adding more of an ion that is a product of this equilibrium.

> *Adding a common ion to a system at equilibrium affects the equilibrium composition, but not the ionization constant.*

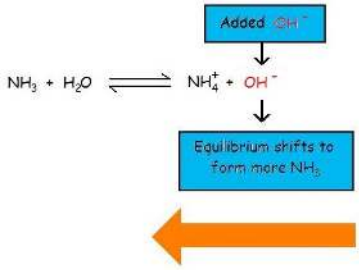
The common ion effect of H_3O^+ on the ionization of acetic acid



HC2H3O2 reactions with H2O to produce H3O plus and C2H3 O2 minus. If more H3O plus is added to the reaction, the equilibrium would shift to form more HC2H3O2.

When a strong acid supplies the common ion H_3O^+ the equilibrium shifts to form more $HC_2H_3O_2$.

The common ion effect suppresses the ionization of a weak base by adding more of an ion that is a product of this equilibrium. Now consider the common ion effect of OH^- on the ionization of ammonia



When a strong base supplies the common ion OH^- the equilibrium shifts to form more NH_3.

NH3 reacts with H2 O to produce NH4 plus and OH minus. If more OH minus is added to the reaction, the equilibrium would shifts to form more NH3.

Adding the common ion of hydroxide shifts the reaction towards the left to decrease the stress (in accordance with Le Chatelier's Principle), forming more reactants. This decreases the reaction quotient, because the reaction is being pushed towards the left to reach equilibrium. The equilibrium constant, $K_b = 1.8 \times 10^{-5}$, does not change. The reaction is put out of balance, or equilibrium.

$$Q_a = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

At first, when more hydroxide is added, the quotient is greater than the equilibrium constant. The reaction then shifts right, causing the denominator to increase, decreasing the reaction quotient and pulling towards equilibrium and causing Q to decrease towards K.

Common Ion Effect on Solubility

When a slightly soluble ionic compound is added to water, some of it dissolves to form a solution, establishing an equilibrium between the pure solid and a solution of its ions. For the dissolution of calcium phosphate, one of the two main components of kidney stones, the equilibrium can be written as follows, with the solid salt on the left:

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3\,Ca^{2+}(aq) + 2\,PO_4^{3-}(aq) \tag{15.11.14}$$

As you will discover in more advanced chemistry courses, basic anions, such as S^{2-}, PO_4^{3-}, and CO_3^{2-}, react with water to produce OH^- and the corresponding protonated anion. Consequently, their calculated molarities, assuming no protonation in aqueous solution, are only approximate. The equilibrium constant for the dissolution of a sparingly soluble salt is the **solubility product (K_{sp})** of the salt. Because the concentration of a pure solid such as $Ca_3(PO_4)_2$ is a constant, it does not appear explicitly in the equilibrium constant expression. The equilibrium constant expression for the dissolution of calcium phosphate is therefore

$$K = \frac{[Ca^{2+}]^3[PO_4^{3-}]^2}{[Ca_3(PO_4)_2]} \tag{15.11.15}$$

$$[Ca_3(PO_4)_2]K = K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2 \tag{15.11.16}$$

At 25°C and pH 7.00, K_{sp} for calcium phosphate is 2.07×10^{-33}, indicating that the concentrations of Ca^{2+} and PO_4^{3-} ions in solution that are in equilibrium with solid calcium phosphate are very low. The values of K_{sp} for some common salts vary dramatically for different compounds (Table E3). Although K_{sp} is not a function of pH in Equation 15.11.15, changes in pH can affect the solubility of a compound.

The solubility product expression tells us that the equilibrium concentrations of the cation and the anion are inversely related. That is, as the concentration of the anion increases, the maximum concentration of the cation needed for precipitation to occur decreases—and vice versa—so that K_{sp} is constant. **Consequently, the solubility of an ionic compound depends on the concentrations of other salts that contain the same ions.** This dependency is another example of the common ion effect where adding a common cation or anion shifts a solubility equilibrium in the direction predicted by Le Chatelier's principle. As a result, the solubility of any sparingly soluble salt is almost always decreased by the presence of a soluble salt that contains a common ion.

Consider, for example, the effect of adding a soluble salt, such as $CaCl_2$, to a saturated solution of calcium phosphate [$Ca_3(PO_4)_2$]. We have seen that the solubility of $Ca_3(PO_4)_2$ in water at 25°C is 1.14×10^{-7} M ($K_{sp} = 2.07 \times 10^{-33}$). Thus a saturated solution of $Ca_3(PO_4)_2$ in water contains

- $3 \times (1.14 \times 10^{-7}\ M) = 3.42 \times 10^{-7} M\ of\ Ca^{2+}$
- $2 \times (1.14 \times 10^{-7} M) = 2.28 \times 10^{-7} M\ of\ PO_4^{3-}$

according to the stoichiometry shown in Equation 15.11.15 (neglecting hydrolysis to form HPO_4^{2-}). If $CaCl_2$ is added to a saturated solution of $Ca_3(PO_4)_2$, the Ca^{2+} ion concentration will increase such that [Ca^{2+}] > 3.42×10^{-7} M, making $Q > K_{sp}$. The only way the system can return to equilibrium is for the reaction in Equation 15.11.15 to proceed to the left, resulting in precipitation of $Ca_3(PO_4)_2$. This will decrease the concentration of both Ca^{2+} and PO_4^{3-} until $Q = K_{sp}$.

Note

Adding a common ion decreases solubility, as the reaction shifts toward the left to relieve the stress of the excess product. Adding a common ion to a dissociation reaction causes the equilibrium to shift left, toward the reactants, causing precipitation.

Example 15.11.5

Consider the reaction:

$$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$$

What happens to the solubility of $PbCl_2$(s) when 0.1 M NaCl is added?

Solution

$$K_{sp} = 1.7 \times 10^{-5}$$

$$Q_{sp} = 1.8 \times 10^{-5}$$

Identify the common ion: Cl^-

Notice: $Q_{sp} > K_{sp}$ The addition of NaCl has caused the reaction to shift out of equilibrium because there are more dissociated ions. Typically, solving for the molarities requires the assumption that the solubility of $PbCl_2$ is equivalent to the concentration of Pb^{2+} produced because they are in a 1:1 ratio.

Because K_{sp} for the reaction is 1.7×10^{-5}, the overall reaction would be $(s)(2s)^2 = 1.7\times10^{-5}$. Solving the equation for s gives $s = 1.62\times10^{-2}$ M. The coefficient on Cl^- is 2, so it is assumed that twice as much Cl^- is produced as Pb^{2+}, hence the '2s.' The solubility equilibrium constant can be used to solve for the molarities of the ions at equilibrium.

The molarity of Cl^- added would be 0.1 M because Na^+ and Cl^- are in a 1:1 ration in the ionic salt, NaCl. Therefore, the overall molarity of Cl^- would be 2s + 0.1, with 2s referring to the contribution of the chloride ion from the dissociation of lead chloride.

$$Q_{sp} = [Pb^{2+}][Cl^-]^2 \tag{15.11.17}$$
$$1.8 \times 10^{-5} = (s)(2s + 0.1)^2 \tag{15.11.18}$$
$$s = [Pb^{2+}] \tag{15.11.19}$$
$$= 1.8 \times 10^{-3} M \tag{15.11.20}$$
$$2s = [Cl^-] \tag{15.11.21}$$
$$\approx 0.1 M \tag{15.11.22}$$

Notice that the molarity of Pb^{2+} is lower when NaCl is added. The equilibrium constant remains the same because of the increased concentration of the chloride ion. To simplify the reaction, it can be assumed that $[Cl^-]$ is approximately 0.1M since the formation of the chloride ion from the dissociation of lead chloride is so small. The reaction quotient for $PbCl_2$ is greater than the equilibrium constant because of the added Cl^-. This therefore shift the reaction left towards equilibrium, causing precipitation and lowering the current solubility of the reaction. Overall, the solubility of the reaction decreases with the added sodium chloride.

Exercise 15.11.5

Calculate the solubility of silver carbonate in a 0.25 M solution of sodium carbonate. The solubility of silver carbonate in pure water is 8.45×10^{-12} at 25°C.

Answer

2.9×10^{-6} M (versus 1.3×10^{-4} M in pure water)



A Video Discussing the Common Ion Effect in Solubility Products: The Common Ion Effect in Solubility Products (opens in new window) [youtu.be]

References

1. Harwood, William S., F. G. Herring, Jeffry D. Madura, and Ralph H. Petrucci. General Chemistry Principles and Modern Applications. 9th ed. New Jersey: Prentice Hall, 2007.

Contributors and Attributions

- Emmellin Tung, Mahtab Danai (UCD)
- Jim Clark (ChemGuide)
- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)

This page titled 15.11: Common-Ion Effect in Acid-Base Equilibria is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.1: The Common-Ion Effect** is licensed CC BY-NC-SA 3.0.

https://chem.libretexts.org/@go/page/499087



https://chem.libretexts.org/@go/page/499087

15.12: Buffer Solutions

Learning Objectives

- To understand how adding a common ion affects the position of an acid–base equilibrium.
- To know how to use the Henderson-Hasselbalch approximation to calculate the pH of a buffer.

Buffers are solutions that maintain a relatively constant pH when an acid or a base is added. They therefore protect, or "buffer," other molecules in solution from the effects of the added acid or base. Buffers contain either a weak acid (HA) and its conjugate base (A^-) or a weak base (B) and its conjugate acid (BH^+), and they are critically important for the proper functioning of biological systems. In fact, every biological fluid is buffered to maintain its physiological pH.

The Common Ion Effect: Weak Acids Combined with Conjugate Bases

To understand how buffers work, let's look first at how the ionization equilibrium of a weak acid is affected by adding either the conjugate base of the acid or a strong acid (a source of H^+). Le Chatelier's principle can be used to predict the effect on the equilibrium position of the solution. A typical buffer used in biochemistry laboratories contains acetic acid and a salt such as sodium acetate. The dissociation reaction of acetic acid is as follows:

$$\mathrm{CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)} \tag{15.12.1}$$

and the equilibrium constant expression is as follows:

$$K_a = \frac{[\mathrm{H^+}][\mathrm{CH_3COO^-}]}{[\mathrm{CH_3CO_2H}]} \tag{15.12.2}$$

Sodium acetate ($\mathrm{CH_3CO_2Na}$) is a strong electrolyte that ionizes completely in aqueous solution to produce $\mathrm{Na^+}$ and $\mathrm{CH_3CO_2^-}$ ions. If sodium acetate is added to a solution of acetic acid, Le Chatelier's principle predicts that the equilibrium in Equation 15.12.1will shift to the left, consuming some of the added $\mathrm{CH_3COO^-}$ and some of the $\mathrm{H^+}$ ions originally present in solution.

Because $\mathrm{Na^+}$ is a **spectator ion**, it has no effect on the position of the equilibrium and can be ignored. The addition of sodium acetate produces a new equilibrium composition, in which $[\mathrm{H^+}]$ is less than the initial value. Because $[\mathrm{H^+}]$ has decreased, the pH will be higher. Thus adding a salt of the conjugate base to a solution of a weak acid increases the pH. This makes sense because sodium acetate is a base, and adding any base to a solution of a weak acid should increase the pH.

If we instead add a strong acid such as HCl to the system, $[\mathrm{H^+}]$ increases. Once again the equilibrium is temporarily disturbed, but the excess $\mathrm{H^+}$ ions react with the conjugate base ($\mathrm{CH_3CO_2^-}$), whether from the parent acid or sodium acetate, to drive the equilibrium to the left. The net result is a new equilibrium composition that has a lower $[\mathrm{CH_3CO_2^-}]$ than before. In both cases, only the equilibrium composition has changed; the ionization constant K_a for acetic acid remains the same. Adding a strong electrolyte that contains one ion in common with a reaction system that is at equilibrium, in this case $\mathrm{CH_3CO_2^-}$, will therefore shift the equilibrium in the direction that reduces the concentration of the common ion. The shift in equilibrium is via the common ion effect.

> *Adding a common ion to a system at equilibrium affects the equilibrium composition, but not the ionization constant.*

Example 15.12.1

A 0.150 M solution of formic acid at 25°C (pKa = 3.75) has a pH of 2.28 and is 3.5% ionized.

a. Is there a change to the pH of the solution if enough solid sodium formate is added to make the final formate concentration 0.100 M (assume that the formic acid concentration does not change)?
b. What percentage of the formic acid is ionized if 0.200 M HCl is added to the system?

Given: solution concentration and pH, pK_a, and percent ionization of acid; final concentration of conjugate base or strong acid added

Asked for: pH and percent ionization of formic acid

Strategy:

A. Write a balanced equilibrium equation for the ionization equilibrium of formic acid. Tabulate the initial concentrations, the changes, and the final concentrations.
B. Substitute the expressions for the final concentrations into the expression for Ka. Calculate $[H^+]$ and the pH of the solution.
C. Construct a table of concentrations for the dissociation of formic acid. To determine the percent ionization, determine the anion concentration, divide it by the initial concentration of formic acid, and multiply the result by 100.

Solution:

A Because sodium formate is a strong electrolyte, it ionizes completely in solution to give formate and sodium ions. The Na^+ ions are spectator ions, so they can be ignored in the equilibrium equation. Because water is both a much weaker acid than formic acid and a much weaker base than formate, the acid–base properties of the solution are determined solely by the formic acid ionization equilibrium:

$$HCO_2H(aq) \rightleftharpoons HCO_2^-(aq) + H^+(aq)$$

The initial concentrations, the changes in concentration that occur as equilibrium is reached, and the final concentrations can be tabulated.

Final Concentration

ICE	$[HCO_2H(aq)]$	$[H^+(aq)]$	$[HCO_2^-(aq)]$
Initial	0.150	1.00×10^{-7}	0.100
Change	−x	+x	+x
Equilibrium	(0.150 − x)	x	(0.100 + x)

B We substitute the expressions for the final concentrations into the equilibrium constant expression and make our usual simplifying assumptions, so

$$\begin{aligned} K_a = \frac{[H^+][HCO_2^-]}{[HCO_2H]} &= \frac{(x)(0.100+x)}{0.150-x} \\ &\approx \frac{x(0.100)}{0.150} \\ &\approx 10^{-3.75} \\ &\approx 1.8 \times 10^{-4} \end{aligned}$$

Rearranging and solving for x,

$$\begin{aligned} x &= (1.8 \times 10^{-4}) \times \frac{0.150\ M}{0.100\ M} \\ &= 2.7 \times 10^{-4} \\ &= [H^+] \end{aligned}$$

The value of x is small compared with 0.150 or 0.100 M, so our assumption about the extent of ionization is justified. Moreover,

$$K_a C_{HA} = (1.8 \times 10^{-4})(0.150) = 2.7 \times 10^{-5}$$

which is greater than 1.0×10^{-6}, so again, our assumption is justified. The final pH is:

$$pH = -\log(2.7 \times 10^{-4}) = 3.57$$

compared with the initial value of 2.29. Thus adding a salt containing the conjugate base of the acid has increased the pH of the solution, as we expect based on Le Chatelier's principle; the stress on the system has been relieved by the consumption of H^+ ions, driving the equilibrium to the left.

C Because HCl is a strong acid, it ionizes completely, and chloride is a spectator ion that can be neglected. Thus the only relevant acid–base equilibrium is again the dissociation of formic acid, and initially the concentration of formate is zero. We can construct a table of initial concentrations, changes in concentration, and final concentrations.

$$HCO_2H(aq) \leftrightharpoons H^+(aq) + HCO_2^-(aq)$$

initial concentrations, changes in concentration, and final concentrations

	$[HCO_2H(aq)]$	$[H^+(aq)]$	$[HCO_2^-(aq)]$
initial	0.150	0.200	0
change	−x	+x	+x
final	(0.150 − x)	(0.200 + x)	x

To calculate the percentage of formic acid that is ionized under these conditions, we have to determine the final $[\mathrm{HCO_2^-}]$. We substitute final concentrations into the equilibrium constant expression and make the usual simplifying assumptions, so

$$K_a = \frac{[H^+][HCO_2^-]}{[HCO_2H]} = \frac{(0.200+x)(x)}{0.150-x} \approx \frac{x(0.200)}{0.150} = 1.80 \times 10^{-4}$$

Rearranging and solving for x,

$$\begin{aligned} x &= (1.80 \times 10^{-4}) \times \frac{0.150\ M}{0.200\ M} \\ &= 1.35 \times 10^{-4} = [HCO_2^-] \end{aligned}$$

Once again, our simplifying assumptions are justified. The percent ionization of formic acid is as follows:

$$\text{percent ionization} = \frac{1.35 \times 10^{-4}\ M}{0.150\ M} \times 100\% = 0.0900\%$$

Adding the strong acid to the solution, as shown in the table, decreased the percent ionization of formic acid by a factor of approximately 38 (3.45%/0.0900%). Again, this is consistent with Le Chatelier's principle: adding $\mathrm{H^+}$ ions drives the dissociation equilibrium to the left.

? Exercise 15.12.1

A 0.225 M solution of ethylamine ($\mathrm{CH_3CH_2NH_2}$ with $pK_b = 3.19$) has a pH of 12.08 and a percent ionization of 5.4% at 20°C. Calculate the following:

a. the pH of the solution if enough solid ethylamine hydrochloride ($\mathrm{EtNH_3Cl}$) is added to make the solution 0.100 M in $\mathrm{EtNH_3^+}$
b. the percentage of ethylamine that is ionized if enough solid NaOH is added to the original solution to give a final concentration of 0.050 M NaOH

Answer a

11.16

Answer b

1.3%



A Video Discussing the Common Ion Effect: The Common Ion Effecr(opens in new window) [youtu.be]

The Common Ion Effect: Weak Bases Combined with Conjugate Acids

Now let's suppose we have a buffer solution that contains equimolar concentrations of a weak base (B) and its conjugate acid (BH^+). The general equation for the ionization of a weak base is as follows:

$$B(aq) + H_2O(l) \leftrightharpoons BH^+(aq) + OH^-(aq) \tag{15.12.3}$$

If the equilibrium constant for the reaction as written in Equation 15.12.3is small, for example $K_b = 10^{-5}$, then the equilibrium constant for the reverse reaction is very large: $K = \dfrac{1}{K_b} = 10^5$. Adding a strong base such as OH^- to the solution therefore causes the equilibrium in Equation 15.12.3to shift to the left, consuming the added OH^-. As a result, the OH^- ion concentration in solution remains relatively constant, and the pH of the solution changes very little. Le Chatelier's principle predicts the same outcome: when the system is stressed by an increase in the OH^- ion concentration, the reaction will proceed to the left to counteract the stress.

If the pK_b of the base is 5.0, the pK_a of its conjugate acid is

$$pK_a = pK_w - pK_b = 14.0 – 5.0 = 9.0.$$

Thus the equilibrium constant for ionization of the conjugate acid is even smaller than that for ionization of the base. The ionization reaction for the conjugate acid of a weak base is written as follows:

$$BH^+(aq) + H_2O(l) \leftrightharpoons B(aq) + H_3O^+(aq) \tag{15.12.4}$$

Again, the equilibrium constant for the reverse of this reaction is very large: K = $1/K_a$ = 10^9. If a strong acid is added, it is neutralized by reaction with the base as the reaction in Equation 15.12.4shifts to the left. As a result, the H^+ ion concentration does not increase very much, and the pH changes only slightly. In effect, a buffer solution behaves somewhat like a sponge that can absorb H^+ and OH^- ions, thereby preventing large changes in pH when appreciable amounts of strong acid or base are added to a solution.

Buffers are characterized by the pH range over which they can maintain a more or less constant pH and by their buffer capacity, the amount of strong acid or base that can be absorbed before the pH changes significantly. Although the useful pH range of a buffer depends strongly on the chemical properties of the weak acid and weak base used to prepare the buffer (i.e., on K), its buffer capacity depends solely on the concentrations of the species in the buffered solution. The more concentrated the buffer solution, the greater its buffer capacity. As illustrated in Figure 15.12.1, when $NaOH$ is added to solutions that contain different concentrations of an acetic acid/sodium acetate buffer, the observed change in the pH of the buffer is inversely proportional to the concentration of the buffer. If the buffer capacity is 10 times larger, then the buffer solution can absorb 10 times more strong acid or base before undergoing a significant change in pH.

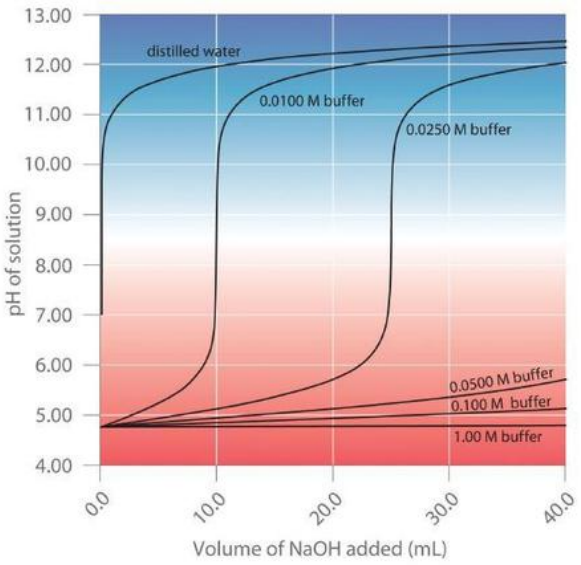


Figure 15.12.1: Effect of Buffer Concentration on the Capacity of a Buffer. (CC BY-SA-NC; Anonymous by request)
Graph of pH of solution against volume of NaOH added. Six different buffer concentrations are plotted.

A buffer maintains a relatively constant pH when acid or base is added to a solution. The addition of even tiny volumes of 0.10 M $NaOH$ to 100.0 mL of distilled water results in a very large change in pH. As the concentration of a 50:50 mixture of sodium acetate/acetic acid buffer in the solution is increased from 0.010 M to 1.00 M, the change in the pH produced by the addition of the same volume of $NaOH$ solution decreases steadily. For buffer concentrations of at least 0.500 M, the addition of even 25 mL of the $NaOH$ solution results in only a relatively small change in pH.

Calculating the pH of a Buffer

The pH of a buffer can be calculated from the concentrations of the weak acid and the weak base used to prepare it, the concentration of the conjugate base and conjugate acid, and the pK_a or pK_b of the weak acid or weak base. The procedure is analogous to that used in Example 15.12.1 to calculate the pH of a solution containing known concentrations of formic acid and formate.

An alternative method frequently used to calculate the pH of a buffer solution is based on a rearrangement of the equilibrium equation for the dissociation of a weak acid. The simplified ionization reaction is $HA \leftrightharpoons H^+ + A^-$, for which the equilibrium constant expression is as follows:

$$K_a = \frac{[H^+][A^-]}{[HA]} \tag{15.12.5}$$

This equation can be rearranged as follows:

$$[H^+] = K_a \frac{[HA]}{[A^-]} \tag{15.12.6}$$

Taking the logarithm of both sides and multiplying both sides by −1,

$$-\log[H^+] = -\log K_a - \log\left(\frac{[HA]}{[A^-]}\right) \tag{15.12.7}$$

$$= -\log K_a + \log\left(\frac{[A^-]}{[HA]}\right) \tag{15.12.8}$$

Replacing the negative logarithms in Equation 15.12.8,

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) \tag{15.12.9}$$

or, more generally,

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right) \tag{15.12.10}$$

Equation 15.12.9 and Equation 15.12.10 are both forms of the Henderson-Hasselbalch approximation, named after the two early 20th-century chemists who first noticed that this rearranged version of the equilibrium constant expression provides an easy way to calculate the pH of a buffer solution. In general, the validity of the Henderson-Hasselbalch approximation may be limited to solutions whose concentrations are at least 100 times greater than their K_a values.

There are three special cases where the Henderson-Hasselbalch approximation is easily interpreted without the need for calculations:

- $[base] = [acid]$: Under these conditions,

  $$\frac{[base]}{[acid]} = 1$$

  in Equation 15.12.10. Because $\log 1 = 0$,

  $$pH = pK_a$$

  regardless of the actual concentrations of the acid and base. Recall that this corresponds to the midpoint in the titration of a weak acid or a weak base.
- $[base]/[acid] = 10$: In Equation 15.12.10, because $\log 10 = 1$,

  $$pH = pK_a + 1.$$
- $[base]/[acid] = 100$: In Equation 15.12.10, because $\log 100 = 2$,

  $$pH = pK_a + 2.$$

Each time we increase the [base]/[acid] ratio by 10, the pH of the solution increases by 1 pH unit. Conversely, if the [base]/[acid] ratio is 0.1, then pH = pK_a − 1. Each additional factor-of-10 decrease in the [base]/[acid] ratio causes the pH to decrease by 1 pH unit.

> If [base] = [acid] for a buffer, then pH = pK_a. Changing this ratio by a factor of 10 either way changes the pH by ±1 unit.

Example 15.12.2

What is the pH of a solution that contains

a. 0.135 M HCO_2H and 0.215 M HCO_2Na? (The pK_a of formic acid is 3.75.)
b. 0.0135 M HCO_2H and 0.0215 M HCO_2Na?
c. 0.119 M pyridine and 0.234 M pyridine hydrochloride? (The pK_b of pyridine is 8.77.)

Given: concentration of acid, conjugate base, and pK_a; concentration of base, conjugate acid, and pK_b

Asked for: pH

Strategy:

Substitute values into either form of the Henderson-Hasselbalch approximation (Equations 15.12.9 or 15.12.10) to calculate the pH.

Solution:

According to the Henderson-Hasselbalch approximation (Equation 15.12.9), the pH of a solution that contains both a weak acid and its conjugate base is

$$pH = pK_a + \log([A-]/[HA]).$$

A

Inserting the given values into the equation,

$$pH = 3.75 + \log\left(\frac{0.215}{0.135}\right)$$
$$= 3.75 + \log 1.593$$
$$= 3.95$$

This result makes sense because the $[A^-]/[HA]$ ratio is between 1 and 10, so the pH of the buffer must be between the pK_a (3.75) and $pK_a + 1$, or 4.75.

B

This is identical to part (a), except for the concentrations of the acid and the conjugate base, which are 10 times lower. Inserting the concentrations into the Henderson-Hasselbalch approximation,

$$pH = 3.75 + \log\left(\frac{0.0215}{0.0135}\right)$$
$$= 3.75 + \log 1.593$$
$$= 3.95$$

This result is identical to the result in part (a), which emphasizes the point that the pH of a buffer depends only on the ratio of the concentrations of the conjugate base and the acid, not on the magnitude of the concentrations. Because the $[A^-]/[HA]$ ratio is the same as in part (a), the pH of the buffer must also be the same (3.95).

C

In this case, we have a weak base, pyridine (Py), and its conjugate acid, the pyridinium ion (HPy^+). We will therefore use Equation 15.12.10, the more general form of the Henderson-Hasselbalch approximation, in which "base" and "acid" refer to the appropriate species of the conjugate acid–base pair. We are given [base] = [Py] = 0.119 M and $[acid] = [HPy^+] = 0.234\ M$. We also are given $pK_b = 8.77$ for pyridine, but we need pK_a for the pyridinium ion. Recall from Equation 16.23 that the pK_b of a weak base and the pK_a of its conjugate acid are related:

$$pK_a + pK_b = pK_w.$$

Thus pK_a for the pyridinium ion is $pK_w - pK_b = 14.00 - 8.77 = 5.23$. Substituting this pK_a value into the Henderson-Hasselbalch approximation,

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$
$$= 5.23 + \log\left(\frac{0.119}{0.234}\right)$$
$$= 5.23\ -0.294$$
$$= 4.94$$

Once again, this result makes sense: the $[B]/[BH^+]$ ratio is about 1/2, which is between 1 and 0.1, so the final pH must be between the pK_a (5.23) and $pK_a - 1$, or 4.23.

? Exercise 15.12.2

What is the pH of a solution that contains

a. 0.333 M benzoic acid and 0.252 M sodium benzoate?
b. 0.050 M trimethylamine and 0.066 M trimethylamine hydrochloride?

The pK_a of benzoic acid is 4.20, and the pK_b of trimethylamine is also 4.20.

Answer a

4.08

Answer b

9.68



A Video Discussing Using the Henderson Hasselbalch Equation: Using the Henderson Hasselbalch Equation(opens in new window) [youtu.be] (opens in new window)

The Henderson-Hasselbalch approximation ((Equation 15.12.9) can also be used to calculate the pH of a buffer solution after adding a given amount of strong acid or strong base, as demonstrated in Example 15.12.3

Example 15.12.3

The buffer solution in Example 15.12.2contained 0.135 M HCO_2H and 0.215 M HCO_2Na and had a pH of 3.95.

a. What is the final pH if 5.00 mL of 1.00 M HCl are added to 100 mL of this solution?
b. What is the final pH if 5.00 mL of 1.00 M $NaOH$ are added?

Given: composition and pH of buffer; concentration and volume of added acid or base

Asked for: final pH

Strategy:

A. Calculate the amounts of formic acid and formate present in the buffer solution using the procedure from Example 15.12.1. Then calculate the amount of acid or base added.
B. Construct a table showing the amounts of all species after the neutralization reaction. Use the final volume of the solution to calculate the concentrations of all species. Finally, substitute the appropriate values into the Henderson-Hasselbalch approximation (Equation 15.12.10) to obtain the pH.

Solution:

The added HCl (a strong acid) or NaOH (a strong base) will react completely with formate (a weak base) or formic acid (a weak acid), respectively, to give formic acid or formate and water. We must therefore calculate the amounts of formic acid and formate present after the neutralization reaction.

A We begin by calculating the millimoles of formic acid and formate present in 100 mL of the initial pH 3.95 buffer:

$$100\ \cancel{mL}\left(\frac{0.135\,mmol\ HCO_2H}{\cancel{mL}}\right) = 13.5\,mmol\,HCO_2H$$

$$100\ \cancel{mL}\left(\frac{0.215\,mmol\ HCO_2^-}{\cancel{mL}}\right) = 21.5\,mmol\,HCO_2^-$$

The millimoles of H^+ in 5.00 mL of 1.00 M HCl is as follows:

$$5.00\ \cancel{mL}\left(\frac{1.00\,mmol\ H^+}{\cancel{mL}}\right) = 5\,mmol\,H^+$$

B Next, we construct a table of initial amounts, changes in amounts, and final amounts:

$$\mathrm{HCO}^{2-}(\mathrm{aq}) + \mathrm{H}^{+}(\mathrm{aq}) \rightleftharpoons \mathrm{HCO_2H(aq)}$$

initial amounts, changes in amounts, and final amounts:

	$HCO^{2-}(aq)$	$H^{+}(aq)$	$HCO_2H(aq)$
Initial	21.5 mmol	5.00 mmol	13.5 mmol
Change	−5.00 mmol	−5.00 mmol	+5.00 mmol
Final	16.5 mmol	~0 mmol	18.5 mmol

The final amount of H^+ in solution is given as "~0 mmol." For the purposes of the stoichiometry calculation, this is essentially true, but remember that the point of the problem is to calculate the final $[H^+]$ and thus the pH. We now have all the information we need to calculate the pH. We can use either the lengthy procedure of Example 15.12.1 or the Henderson–Hasselbach approximation. Because we have performed many equilibrium calculations in this chapter, we'll take the latter approach. The Henderson-Hasselbalch approximation requires the concentrations of HCO_2^- and HCO_2H, which can be calculated using the number of millimoles (n) of each and the total volume (VT). Substituting these values into the Henderson-Hasselbalch approximation (Equation 15.12.10):

$$\begin{aligned} pH &= pK_a + \log\left(\frac{[HCO_2^-]}{[HCO_2H]}\right) \\ &= pK_a + \log\left(\frac{n_{HCO_2^-}/V_f}{n_{HCO_2H}/V_f}\right) \\ &= pK_a + \log\left(\frac{n_{HCO_2^-}}{n_{HCO_2H}}\right) \end{aligned}$$

Because the total volume appears in both the numerator and denominator, it cancels. We therefore need to use only the ratio of the number of millimoles of the conjugate base to the number of millimoles of the weak acid. So

$$\begin{aligned} pH &= pK_a + \log\left(\frac{n_{HCO_2^-}}{n_{HCO_2H}}\right) \\ &= 3.75 + \log\left(\frac{16.5\ mmol}{18.5\ mmol}\right) \\ &= 3.75\ -0.050 = 3.70 \end{aligned}$$

Once again, this result makes sense on two levels. First, the addition of HClhas decreased the pH from 3.95, as expected. Second, the ratio of HCO_2^- to HCO_2H is slightly less than 1, so the pH should be between the pK_a and pK_a − 1.

A The procedure for solving this part of the problem is exactly the same as that used in part (a). We have already calculated the numbers of millimoles of formic acid and formate in 100 mL of the initial pH 3.95 buffer: 13.5 mmol of HCO_2H and 21.5 mmol of HCO_2^-. The number of millimoles of OH^- in 5.00 mL of 1.00 M $NaOH$ is as follows:

B With this information, we can construct a table of initial amounts, changes in amounts, and final amounts.

$$\mathrm{HCO_2H(aq)} + \mathrm{OH}^{-}(\mathrm{aq}) \rightleftharpoons \mathrm{HCO_2^-(aq)} + \mathrm{H_2O(l)}$$

initial amounts, changes in amounts, and final amounts

	$HCO_2H(aq)$	OH^-	$HCO_2^-(aq)$
Initial	13.5 mmol	5.00 mmol	21.5 mmol
Change	−5.00 mmol	−5.00 mmol	+5.00 mmol
Final	8.5 mmol	~0 mmol	26.5 mmol

The final amount of OH^- in solution is not actually zero; this is only approximately true based on the stoichiometric calculation. We can calculate the final pH by inserting the numbers of millimoles of both HCO_2^- and HCO_2H into the simplified Henderson-Hasselbalch expression used in part (a) because the volume cancels:

$$\begin{aligned} pH &= pK_a + \log\left(\frac{n_{HCO_2^-}}{n_{HCO_2H}}\right) \\ &= 3.75 + \log\left(\frac{26.5\ mmol}{8.5\ mmol}\right) \\ &= 3.75 + 0.494 = 4.24 \end{aligned}$$

Once again, this result makes chemical sense: the pH has increased, as would be expected after adding a strong base, and the final pH is between the pK_a and pK_a + 1, as expected for a solution with a HCO_2^-/HCO_2H ratio between 1 and 10.

Exercise 15.12.3

The buffer solution from Example 15.12.2 contained 0.119 M pyridine and 0.234 M pyridine hydrochloride and had a pH of 4.94.

a. What is the final pH if 12.0 mL of 1.5 M NaOH are added to 250 mL of this solution?
b. What is the final pH if 12.0 mL of 1.5 M HCl are added?

Answer a

5.30

Answer b

4.42

Only the amounts (in moles or millimoles) of the acidic and basic components of the buffer are needed to use the Henderson-Hasselbalch approximation, not their concentrations.



A Video Discussing the Change in pH with the Addition of a Strong Acid to a Buffer: The Change in pH with the Addition of a Strong Acid to a Buffer(opens in new window) [youtu.be]



The Change in pH with the Addition of a Strong Base to a Buffer:

The Change in pH with the Addition of a Strong Base to a Buffer (opens in new window) [youtu.be]

The results obtained in Example $15.12.3$ and its corresponding exercise demonstrate how little the pH of a well-chosen buffer solution changes despite the addition of a significant quantity of strong acid or strong base. Suppose we had added the same amount of HCl or $NaOH$ solution to 100 mL of an unbuffered solution at pH 3.95 (corresponding to 1.1×10^{-4} M HCl). In this case, adding 5.00 mL of 1.00 M HCl would lower the final pH to 1.32 instead of 3.70, whereas adding 5.00 mL of 1.00 M $NaOH$ would raise the final pH to 12.68 rather than 4.24. (Try verifying these values by doing the calculations yourself.) Thus the presence of a buffer significantly increases the ability of a solution to maintain an almost constant pH.

> *The most effective buffers contain **equal** concentrations of an acid and its conjugate base.*

A buffer that contains approximately equal amounts of a weak acid and its conjugate base in solution is equally effective at neutralizing either added base or added acid. This is shown in Figure $15.12.2$ for an acetic acid/sodium acetate buffer. Adding a given amount of strong acid shifts the system along the horizontal axis to the left, whereas adding the same amount of strong base shifts the system the same distance to the right. In either case, the change in the ratio of $CH_3CO_2^-$ to CH_3CO_2H from 1:1 reduces the buffer capacity of the solution.

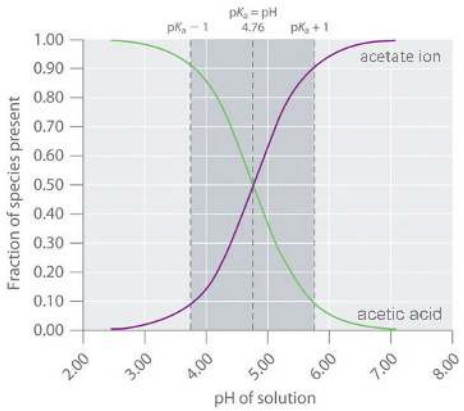


Figure 15.12.2: Distribution Curve Showing the Fraction of Acetic Acid Molecules and Acetate Ions as a Function of pH in a Solution of Acetic Acid. The pH range over which the acetic acid/sodium acetate system is an effective buffer (the darker shaded region) corresponds to the region in which appreciable concentrations of both species are present (pH 3.76–5.76, corresponding to $pH = pK_a \pm 1$). (CC BY-SA-NC; Anonymous by request)

Graph of mole fraction against pH of solution. The green line is acetic acid and the purple line is the acetate ion.



A Video Discussing The Buffer Region: The Buffer Region (opens in new window) [youtu.be]

The Relationship between Titrations and Buffers

There is a strong correlation between the effectiveness of a buffer solution and the titration curves discussed in Section 16.5. Consider the schematic titration curve of a weak acid with a strong base shown in Figure 15.12.3 As indicated by the labels, the region around pK_a corresponds to the midpoint of the titration, when approximately half the weak acid has been neutralized. This portion of the titration curve corresponds to a buffer: it exhibits the smallest change in pH per increment of added strong base, as shown by the nearly horizontal nature of the curve in this region. The nearly flat portion of the curve extends only from approximately a pH value of 1 unit less than the pK_a to approximately a pH value of 1 unit greater than the pK_a, which is why buffer solutions usually have a pH that is within ±1 pH units of the pK_a of the acid component of the buffer.

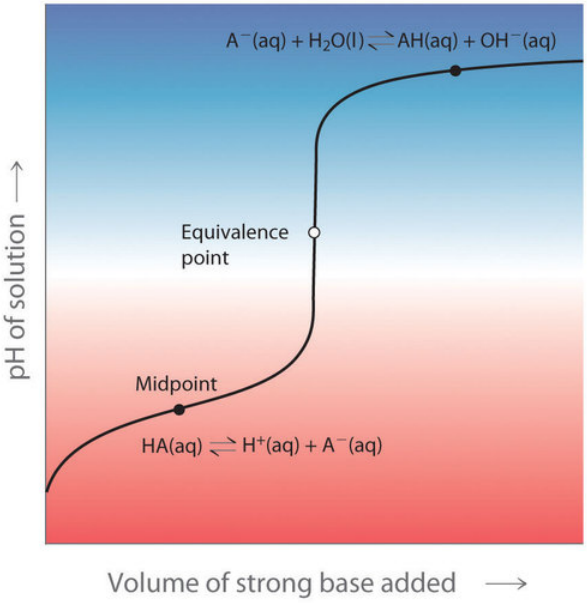


Figure 15.12.3: The Relationship between Titration Curves and Buffers. (CC BY-SA-NC; Anonymous by request)

Titration curves graph pH of solution against volume of strong base added. On the curve we look for the midpoints and equivalence points.

This schematic plot of pH for the titration of a weak acid with a strong base shows the nearly flat region of the titration curve around the midpoint, which corresponds to the formation of a buffer. At the lower left, the pH of the solution is determined by the equilibrium for dissociation of the weak acid; at the upper right, the pH is determined by the equilibrium for reaction of the conjugate base with water.

In the region of the titration curve at the lower left, before the midpoint, the acid–base properties of the solution are dominated by the equilibrium for dissociation of the weak acid, corresponding to K_a. In the region of the titration curve at the upper right, after the midpoint, the acid–base properties of the solution are dominated by the equilibrium for reaction of the conjugate base of the

weak acid with water, corresponding to K_b. However, we can calculate either K_a or K_b from the other because they are related by K_w.

Blood: A Most Important Buffer

Metabolic processes produce large amounts of acids and bases, yet organisms are able to maintain an almost constant internal pH because their fluids contain buffers. This is not to say that the pH is uniform throughout all cells and tissues of a mammal. The internal pH of a red blood cell is about 7.2, but the pH of most other kinds of cells is lower, around 7.0. Even within a single cell, different compartments can have very different pH values. For example, one intracellular compartment in white blood cells has a pH of around 5.0.

Because no single buffer system can effectively maintain a constant pH value over the entire physiological range of approximately pH 5.0 to 7.4, biochemical systems use a set of buffers with overlapping ranges. The most important of these is the CO_2/HCO_3^- system, which dominates the buffering action of blood plasma.

The acid–base equilibrium in the CO_2/HCO_3^- buffer system is usually written as follows:

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \tag{15.12.11}$$

with $K_a = 4.5 \times 10^{-7}$ and $pK_a = 6.35$ at 25°C. In fact, Equation 15.12.11 is a grossly oversimplified version of the CO_2/HCO_3^- system because a solution of CO_2 in water contains only rather small amounts of H_2CO_3. Thus Equation 15.12.11does not allow us to understand how blood is actually buffered, particularly at a physiological temperature of 37°C.

As shown in Equation 15.12.12, CO_2 is in equilibrium with H_2CO_3, but the equilibrium lies far to the left, with an H_2CO_3/CO_2 ratio less than 0.01 under most conditions:

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \tag{15.12.12}$$

with $K' = 4.0 \times 10^{-3}$ at 37°C. The true pK_a of carbonic acid at 37°C is therefore 3.70, not 6.35, corresponding to a K_a of 2.0×10^{-4}, which makes it a much stronger acid than Equation 15.12.11suggests. Adding Equation 15.12.11and Equation 15.12.12and canceling H_2CO_3 from both sides give the following overall equation for the reaction of CO_2 with water to give a proton and the bicarbonate ion:

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \tag{15.12.13}$$

with $K' = 4.0 \times 10^{-3} (37^\circ C)$

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \tag{15.12.14}$$

with $K_a = 2.0 \times 10^{-4} (37^\circ C)$

$$CO_2(aq) + H_2O(l) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \tag{15.12.15}$$

with $K = 8.0 \times 10^{-7} (37^\circ C)$

The K value for the reaction in Equation 15.12.15is the product of the true ionization constant for carbonic acid (K_a) and the equilibrium constant (K) for the reaction of $CO_2(aq)$ with water to give carbonic acid. The equilibrium equation for the reaction of CO_2 with water to give bicarbonate and a proton is therefore

$$K = \frac{[H^+][HCO_3^-]}{[CO_2]} = 8.0 \times 10^{-7} \tag{15.12.16}$$

The presence of a gas in the equilibrium constant expression for a buffer is unusual. According to **Henry's law**,

$$[CO_2] = kP_{CO_2}$$

where k is the Henry's law constant for CO_2, which is $3.0 \times 10^{-5}\ M/mmHg$ at 37°C. Substituting this expression for $[CO_2]$ in Equation 15.12.16,

$$K = \frac{[H^+][HCO_3^-]}{(3.0 \times 10^{-5}\ M/mmHg)(P_{CO_2})}$$

where P_{CO_2} is in mmHg. Taking the negative logarithm of both sides and rearranging,

$$pH = 6.10 + \log\left(\frac{[HCO_3^-]}{(3.0 \times 10^{-5} M/mm\ Hg)\ (P_{CO_2})}\right) \tag{15.12.17}$$

Thus the pH of the solution depends on both the CO_2 pressure over the solution and $[HCO_3^-]$. Figure 15.12.4 plots the relationship between pH and $[HCO_3^-]$ under physiological conditions for several different values of P_{CO_2}, with normal pH and $[HCO_3^-]$ values indicated by the dashed lines.

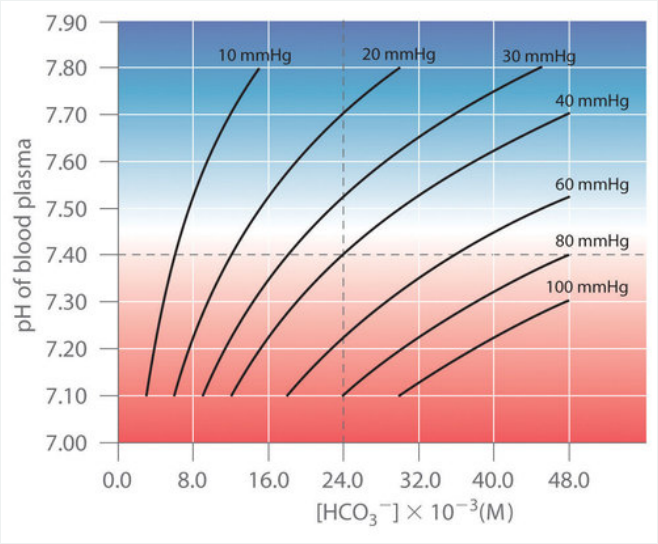


Figure 15.12.4: Buffering in Blood: pH versus $[HCO_3^-]$ Curves for Buffers with Different Values of P_{CO_2}. Only those combinations of pH and $[HCO_3^-]$ that lie on a given line are allowed for the particular value of P_{CO_2} indicated. Normal values of blood plasma pH and $[HCO_3^-]$ are indicated by dashed lines. (CC BY-SA-NC; Anonymous by request)

According to Equation 15.12.17, adding a strong acid to the CO_2/HCO_3^- system causes $[HCO_3^-]$ to decrease as HCO_3^- is converted to CO_2. Excess CO_2 is released in the lungs and exhaled into the atmosphere, however, so there is essentially no change in P_{CO_2}. Because the change in $[HCO_3^-]/P_{CO_2}$ is small, Equation 15.12.17predicts that the change in pH will also be rather small. Conversely, if a strong base is added, the OH^- reacts with CO_2 to form HCO_3^-, but CO_2 is replenished by the body, again limiting the change in both $[HCO_3^-]/P_{CO_2}$ and pH. The CO_2/HCO_3^- buffer system is an example of an open system, in which the total concentration of the components of the buffer change to keep the pH at a nearly constant value.

If a passenger steps out of an airplane in Denver, Colorado, for example, the lower P_{CO_2} at higher elevations (typically 31 mmHg at an elevation of 2000 m versus 40 mmHg at sea level) causes a shift to a new pH and $[HCO_3^-]$. The increase in pH and decrease in $[HCO_3^-]$ in response to the decrease in P_{CO_2} are responsible for the general malaise that many people experience at high altitudes. If their blood pH does not adjust rapidly, the condition can develop into the life-threatening phenomenon known as altitude sickness.



A Video Summary of the pH Curve for a Strong Acid/Strong Base Titration:

Summary of the pH Curve for a Strong Acid/Strong Base Titration(opens in new window) [youtu.be]

Summary

Buffers are solutions that resist a change in pH after adding an acid or a base. Buffers contain a weak acid (HA) and its conjugate weak base (A^-). Adding a strong electrolyte that contains one ion in common with a reaction system that is at equilibrium shifts the equilibrium in such a way as to reduce the concentration of the common ion. The shift in equilibrium is called the common ion effect. Buffers are characterized by their pH range and buffer capacity. The useful pH range of a buffer depends strongly on the chemical properties of the conjugate weak acid–base pair used to prepare the buffer (the K_a or K_b), whereas its buffer capacity depends solely on the concentrations of the species in the solution. The pH of a buffer can be calculated using the Henderson-Hasselbalch approximation, which is valid for solutions whose concentrations are at least 100 times greater than their K_a values. Because no single buffer system can effectively maintain a constant pH value over the physiological range of approximately 5 to 7.4, biochemical systems use a set of buffers with overlapping ranges. The most important of these is the CO_2/HCO_3^- system, which dominates the buffering action of blood plasma.

This page titled 15.12: Buffer Solutions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **17.2: Buffered Solutions** by Anonymous is licensed CC BY-NC-SA 3.0.

15.13: Acid-Base Indicators

Learning Objectives

- Explain the function and color changes of acid-base indicators
- Demonstrate how to select the proper indicator for a titration experiment
- Determine the acidic dissociation constants K_a or K_{ai} of indicators.

Certain organic substances change color in dilute solution when the hydronium ion concentration reaches a particular value. For example, phenolphthalein is a colorless substance in any aqueous solution with a hydronium ion concentration greater than 5.0×10^{-9} *M* (pH < 8.3). In more basic solutions where the hydronium ion concentration is less than 5.0×10^{-9} *M* (pH > 8.3), it is red or pink. Substances such as phenolphthalein, which can be used to determine the pH of a solution, are called acid-base indicators. Acid-base indicators are either weak organic acids or weak organic bases.

The equilibrium in a solution of the acid-base indicator methyl orange, a weak acid, can be represented by an equation in which we use HIn as a simple representation for the complex methyl orange molecule:

$$\underbrace{\mathrm{HIn}_{(aq)}}_{\text{red}} + \mathrm{H_2O}_{(l)} \rightleftharpoons \mathrm{H_3O^+}_{(aq)} + \underbrace{\mathrm{In^-}_{(aq)}}_{\text{yellow}} \tag{15.13.1}$$

$$K_a = \frac{[\mathrm{H_3O^+}][\mathrm{In^-}]}{[\mathrm{HIn}]} = 4.0 \times 10^{-4} \tag{15.13.2}$$

The anion of methyl orange, In^-, is yellow, and the nonionized form, HIn, is red. When we add acid to a solution of methyl orange, the increased hydronium ion concentration shifts the equilibrium toward the nonionized red form, in accordance with Le Chatelier's principle. If we add base, we shift the equilibrium towards the yellow form. This behavior is completely analogous to the action of buffers.

An indicator's color is the visible result of the ratio of the concentrations of the two species In^- and HIn. If most of the indicator (typically about 60−90% or more) is present as In^-, then we see the color of the In^- ion, which would be yellow for methyl orange. If most is present as HIn, then we see the color of the HIn molecule: red for methyl orange. For methyl orange, we can rearrange the equation for K_a and write:

$$\frac{[\mathrm{In^-}]}{[\mathrm{HIn}]} = \frac{[\text{substance with yellow color}]}{[\text{substance with red color}]} = \frac{K_a}{[\mathrm{H_3O^+}]} \tag{15.13.3}$$

This shows us how the ratio of $\frac{[\mathrm{In^-}]}{[\mathrm{HIn}]}$ varies with the concentration of hydronium ion. The above expression describing the indicator equilibrium can be rearranged:

$$\frac{[\mathrm{H_3O^+}]}{K_a} = \frac{[\mathrm{HIn}]}{[\mathrm{In^-}]} \tag{15.13.4}$$

$$\log\left(\frac{[\mathrm{H_3O^+}]}{K_a}\right) = \log\left(\frac{[\mathrm{HIn}]}{[\mathrm{In^-}]}\right) \tag{15.13.5}$$

$$\log([\mathrm{H_3O^+}]) - \log(K_a) = -\log\left(\frac{[\mathrm{In^-}]}{[\mathrm{HIn}]}\right) \tag{15.13.6}$$

$$-\mathrm{pH} + \mathrm{p}K_a = -\log\left(\frac{[\mathrm{In^-}]}{[\mathrm{HIn}]}\right) \tag{15.13.7}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\mathrm{In^-}]}{[\mathrm{HIn}]}\right) \text{ or } \mathrm{pH} = \mathrm{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right) \tag{15.13.8}$$

The last formula is the same as the Henderson-Hasselbalch equation, which can be used to describe the equilibrium of indicators.

When $[\mathrm{H_3O^+}]$ has the same numerical value as K_a, the ratio of $[\mathrm{In^-}]$ to [HIn] is equal to 1, meaning that 50% of the indicator is present in the red form (HIn) and 50% is in the yellow ionic form (In^-), and the solution appears orange in color. When the

hydronium ion concentration increases to 8×10^{-4} M (a pH of 3.1), the solution turns red. No change in color is visible for any further increase in the hydronium ion concentration (decrease in pH). At a hydronium ion concentration of 4×10^{-5} M (a pH of 4.4), most of the indicator is in the yellow ionic form, and a further decrease in the hydronium ion concentration (increase in pH) does not produce a visible color change. The pH range between 3.1 (red) and 4.4 (yellow) is the color-change interval of methyl orange; the pronounced color change takes place between these pH values.

Many different substances can be used as indicators, depending on the particular reaction to be monitored. For example, red cabbage juice contains a mixture of colored substances that change from deep red at low pH to light blue at intermediate pH to yellow at high pH (Figure 15.13.1). In all cases, though, a good indicator must have the following properties:

- The color change must be easily detected.
- The color change must be rapid.
- The indicator molecule must not react with the substance being titrated.
- To minimize errors, the indicator should have a pKin that is within one pH unit of the expected pH at the equivalence point of the titration.



Figure 15.13.1: Naturally Occurring pH Indicators in Red Cabbage Juice. Image curtesty of Wikipedia.

Red cabbage juice contains a mixture of substances whose color depends on the pH. Each test tube contains a solution of red cabbage juice in water, but the pH of the solutions varies from pH = 2.0 (far left) to pH = 11.0 (far right). At pH = 7.0, the solution is blue.

Synthetic indicators have been developed that meet these criteria and cover virtually the entire pH range. Figure 15.13.2shows the approximate pH range over which some common indicators change color and their change in color. In addition, some indicators (such as thymol blue) are polyprotic acids or bases, which change color twice at widely separated pH values.

Indicator	pK_{in}	Color change
Alizarin yellow R	11.0	yellow to red
Thymolphthalein	9.9	colorless to blue
Phenolphthalein	9.5	colorless to pink
Thymol blue	9.2	yellow to blue
Phenol red	7.4	yellow to red
Bromothymol blue	7.3	yellow to blue
Bromocresol purple	6.4	yellow to purple
Ethyl red	5.4	colorless to red
Resorcin blue	5.3	red to blue
Resazurin	5.1	orange to violet
Methyl red	5.0	red to yellow
Bromophenol blue	4.1	yellow to blue
Thymol blue	1.7	red to yellow
Malachite green	1.3	yellow to turquoise
Methyl violet	0.8	yellow to blue

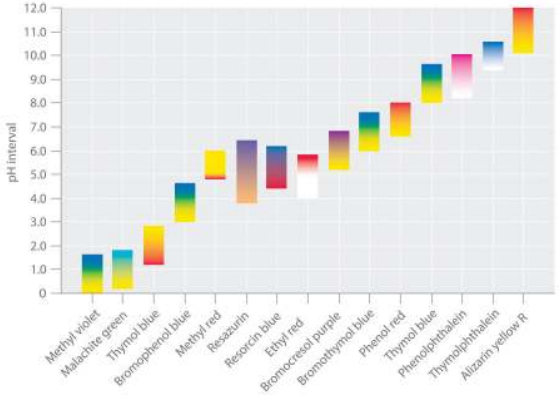


Figure 15.13.2: Some Common Acid–Base Indicators. Approximate colors are shown, along with pKin values and the pH range over which the color changes.

It is important to be aware that an indicator does not change color abruptly at a particular pH value; instead, it actually undergoes a pH titration just like any other acid or base. As the concentration of HIn decreases and the concentration of In^- increases, the color of the solution slowly changes from the characteristic color of HIn to that of In^-. As we will see in Section 16, the $[In^-]/[HIn]$ ratio changes from 0.1 at a pH one unit below pKin to 10 at a pH one unit above pKin. Thus most indicators change color over a pH range of about two pH units.

We have stated that a good indicator should have a pKin value that is close to the expected pH at the equivalence point. For a strong acid–strong base titration, the choice of the indicator is not especially critical due to the very large change in pH that occurs around the equivalence point. In contrast, using the wrong indicator for a titration of a weak acid or a weak base can result in relatively large errors, as illustrated in Figure 15.13.3 This figure shows plots of pH versus volume of base added for the titration of 50.0 mL of a 0.100 M solution of a strong acid (HCl) and a weak acid (acetic acid) with 0.100 M $NaOH$. The pH ranges over which two common indicators (methyl red, $pK_{in} = 5.0$, and phenolphthalein, $pK_{in} = 9.5$) change color are also shown. The horizontal bars indicate the pH ranges over which both indicators change color cross the HCl titration curve, where it is almost vertical. Hence both indicators change color when essentially the same volume of $NaOH$ has been added (about 50 mL), which corresponds to the equivalence point. In contrast, the titration of acetic acid will give very different results depending on whether methyl red or phenolphthalein is used as the indicator. Although the pH range over which phenolphthalein changes color is slightly greater than the pH at the equivalence point of the strong acid titration, the error will be negligible due to the slope of this portion of the titration curve. Just as with the HCl titration, the phenolphthalein indicator will turn pink when about 50 mL of $NaOH$ has been added to the acetic acid solution. In contrast, methyl red begins to change from red to yellow around pH 5, which is near the midpoint of the acetic acid titration, not the equivalence point. Adding only about 25–30 mL of $NaOH$ will therefore cause the methyl red indicator to change color, resulting in a huge error.

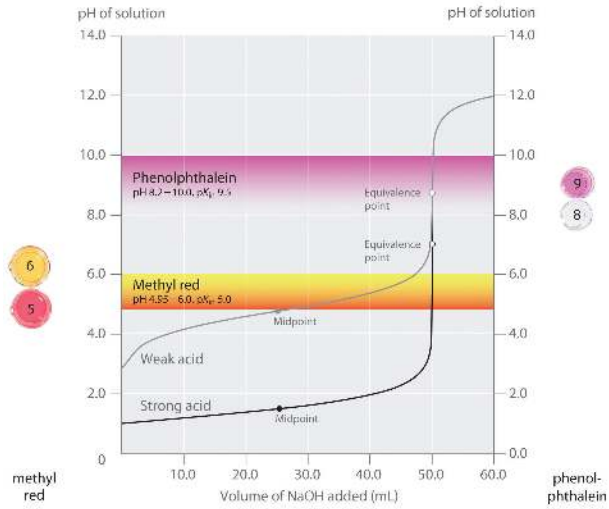


Figure 15.13.3: Choosing the Correct Indicator for an Acid–Base Titration

The graph shows the results obtained using two indicators (methyl red and phenolphthalein) for the titration of 0.100 M solutions of a strong acid (HCl) and a weak acid (acetic acid) with 0.100 M $NaOH$. Due to the steepness of the titration curve of a strong acid around the equivalence point, either indicator will rapidly change color at the equivalence point for the titration of the strong acid. In contrast, the pKin for methyl red (5.0) is very close to the pKa of acetic acid (4.76); the midpoint of the color change for methyl red occurs near the midpoint of the titration, rather than at the equivalence point.

In general, for titrations of strong acids with strong bases (and vice versa), any indicator with a pK in between about 4.0 and 10.0 will do. For the titration of a weak acid, however, the pH at the equivalence point is greater than 7.0, so an indicator such as phenolphthalein or thymol blue, with pKin > 7.0, should be used. Conversely, for the titration of a weak base, where the pH at the equivalence point is less than 7.0, an indicator such as methyl red or bromocresol blue, with pKin < 7.0, should be used.

Example 15.13.1

In the titration of a weak acid with a strong base, which indicator would be the best choice?

a. Methyl Orange
b. Bromocresol Green
c. Phenolphtalein

Solution

The correct answer is C. In the titration of a weak acid with a strong base, the conjugate base of the weak acid will make the pH at the equivalence point greater than 7. Therefore, you would want an indicator to change in that pH range. Both methyl

orange and bromocresol green change color in an acidic pH range, while phenolphtalein changes in a basic pH.

The existence of many different indicators with different colors and pKin values also provides a convenient way to estimate the pH of a solution without using an expensive electronic pH meter and a fragile pH electrode. Paper or plastic strips impregnated with combinations of indicators are used as “pH paper,” which allows you to estimate the pH of a solution by simply dipping a piece of pH paper into it and comparing the resulting color with the standards printed on the container (Figure 15.13.4).



Figure 15.13.4: pH Paper. pH paper contains a set of indicators that change color at different pH values. The approximate pH of a solution can be determined by simply dipping a paper strip into the solution and comparing the color to the standards provided.



pH Indicators: https://youtu.be/1IqzUa5lABs

Summary

Acid–base indicators are compounds that change color at a particular pH. They are typically weak acids or bases whose changes in color correspond to deprotonation or protonation of the indicator itself.

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- {template.ContribBoundless()}

This page titled 15.13: Acid-Base Indicators is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **17.3: Acid-Base Indicators** is licensed CC BY-NC-SA 4.0.

15.14: Neutralization Reactions and Titration Curves

Learning Objectives

- To calculate the pH at any point in an acid–base titration.

In an acid–base titration, a buret is used to deliver measured volumes of an acid or a base solution of known concentration (the titrant) to a flask that contains a solution of a base or an acid, respectively, of unknown concentration (the unknown). If the concentration of the titrant is known, then the concentration of the unknown can be determined. The following discussion focuses on the pH changes that occur during an acid–base titration. Plotting the pH of the solution in the flask against the amount of acid or base added produces a titration curve. The shape of the curve provides important information about what is occurring in solution during the titration.

Titrations of Strong Acids and Bases

Figure 15.14.1*a* shows a plot of the pH as 0.20 M HCl is gradually added to 50.00 mL of pure water. The pH of the sample in the flask is initially 7.00 (as expected for pure water), but it drops very rapidly as HCl is added. Eventually the pH becomes constant at 0.70—a point well beyond its value of 1.00 with the addition of 50.0 mL of HCl (0.70 is the pH of 0.20 M HCl). In contrast, when 0.20 M NaOH is added to 50.00 mL of distilled water, the pH (initially 7.00) climbs very rapidly at first but then more gradually, eventually approaching a limit of 13.30 (the pH of 0.20 M NaOH), again well beyond its value of 13.00 with the addition of 50.0 mL of NaOH as shown in Figure 15.14.1*b* As you can see from these plots, the titration curve for adding a base is the mirror image of the curve for adding an acid.

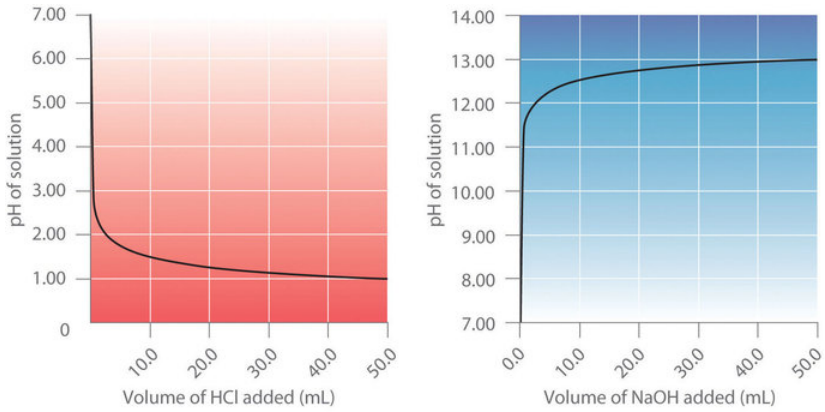


Figure 15.14.1: Solution pH as a Function of the Volume of a Strong Acid or a Strong Base Added to Distilled Water. (a) When 0.20 M HCl is added to 50.0 mL of distilled water, the pH rapidly decreases until it reaches a minimum at the pH of 0.20 M HCl. (b) Conversely, when 0.20 M NaOH is added to 50.0 mL of distilled water, the pH rapidly increases until it reaches a maximum at the pH of 0.20 M $NaOH$. (CC BY-SA-NC; Anonymous by request)

The graphs of water with strong acid added and water with strong base added are inverses of each other.

Suppose that we now add 0.20 M NaOH to 50.0 mL of a 0.10 M solution of HCl. Because HCl is a strong acid that is completely ionized in water, the initial $[H^+]$ is 0.10 M, and the initial pH is 1.00. Adding NaOH decreases the concentration of H+ because of the neutralization reaction (Figure 15.14.2*a*):

$$OH^- + H^+ \rightleftharpoons H_2O.$$

Thus the pH of the solution increases gradually. Near the equivalence point, however, the point at which the number of moles of base (or acid) added equals the number of moles of acid (or base) originally present in the solution, the pH increases much more rapidly because most of the H^+ ions originally present have been consumed. For the titration of a monoprotic strong acid (HCl) with a monobasic strong base (NaOH), we can calculate the volume of base needed to reach the equivalence point from the following relationship:

$$moles\ of\ base = (volume)_b(molarity)_b V_b M_b = moles\ of\ acid = (volume)_a(molarity)_a = V_a M_a \quad (15.14.1)$$

If 0.20 M NaOH is added to 50.0 mL of a 0.10 M solution of HCl, we solve for V_b:

$$V_b(0.20Me) = 0.025L = 25mL$$

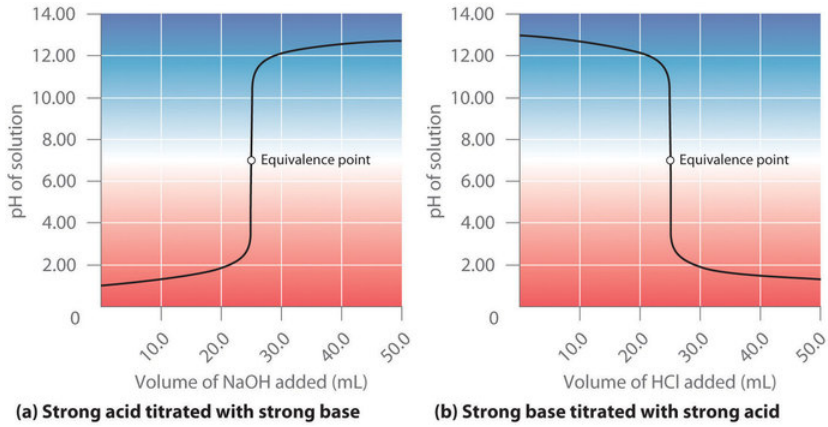


Figure 15.14.2: The Titration of (a) a Strong Acid with a Strong Base and (b) a Strong Base with a Strong Acid(a) As 0.20 M NaOH is slowly added to 50.0 mL of 0.10 M HCl, the pH increases slowly at first, then increases very rapidly as the equivalence point is approached, and finally increases slowly once more. (b) Conversely, as 0.20 M HCl is slowly added to 50.0 mL of 0.10 M $NaOH$, the pH decreases slowly at first, then decreases very rapidly as the equivalence point is approached, and finally decreases slowly once more. (CC BY-SA-NC; Anonymous by request)

The titration curves of strong acid titrated with strong base and strong base titrated with strong acid are inverses of each other.

At the equivalence point (when 25.0 mL of NaOH solution has been added), the neutralization is complete: only a salt remains in solution (NaCl), and the pH of the solution is 7.00. Adding more NaOH produces a rapid increase in pH, but eventually the pH levels off at a value of about 13.30, the pH of 0.20 M $NaOH$.

As shown in Figure $15.14.2b$, the titration of 50.0 mL of a 0.10 M solution of NaOH with 0.20 M HCl produces a titration curve that is nearly the mirror image of the titration curve in Figure $15.14.2a$. The pH is initially 13.00, and it slowly decreases as HCl is added. As the equivalence point is approached, the pH drops rapidly before leveling off at a value of about 0.70, the pH of 0.20 M HCl.

The titration of either a strong acid with a strong base or a strong base with a strong acid produces an S-shaped curve. The curve is somewhat asymmetrical because the steady increase in the volume of the solution during the titration causes the solution to become more dilute. Due to the leveling effect, the shape of the curve for a titration involving a strong acid and a strong base depends on only the concentrations of the acid and base, not their identities.

> The shape of the titration curve involving a strong acid and a strong base depends only on their concentrations, not their identities.

✔ Example 15.14.1: Hydrochloric Acid

Calculate the pH of the solution after 24.90 mL of 0.200 M NaOH has been added to 50.00 mL of 0.100 M HCl.

Given: volumes and concentrations of strong base and acid

Asked for: pH

Strategy:

A. Calculate the number of millimoles of H^+ and OH^- to determine which, if either, is in excess after the neutralization reaction has occurred. If one species is in excess, calculate the amount that remains after the neutralization reaction.
B. Determine the final volume of the solution. Calculate the concentration of the species in excess and convert this value to pH.

Solution

A Because 0.100 mol/L is equivalent to 0.100 mmol/mL, the number of millimoles of H^+ in 50.00 mL of 0.100 M HCl can be calculated as follows:

$$50.00\ \cancel{mL}\left(\frac{0.100\ mmol\ HCl}{\cancel{mL}}\right) = 5.00\ mmol\ HCl = 5.00\ mmol\ H^+$$

The number of millimoles of NaOH added is as follows:

$$24.90\ \cancel{mL}\left(\frac{0.200\ mmol\ NaOH}{\cancel{mL}}\right) = 4.98\ mmol\ NaOH = 4.98\ mmol\ OH^-$$

Thus H^+ is in excess. To completely neutralize the acid requires the addition of 5.00 mmol of OH^- to the HCl solution. Because only 4.98 mmol of OH^- has been added, the amount of excess H^+ is 5.00 mmol − 4.98 mmol = 0.02 mmol of H^+.

B The final volume of the solution is 50.00 mL + 24.90 mL = 74.90 mL, so the final concentration of H^+ is as follows:

$$[H^+] = \frac{0.02\ mmol\ H^+}{74.90\ mL} = 3 \times 10^{-4}\ M$$

Hence,

$$pH \approx -\log[H^+] = -\log(3 \times 10^{-4}) = 3.5$$

This is significantly less than the pH of 7.00 for a neutral solution.

Exercise 15.14.1

Calculate the pH of a solution prepared by adding $40.00\ mL$ of $0.237\ M\ HCl$ to $75.00\ mL$ of a $0.133 M$ solution of $NaOH$.

Answer

11.6



pH after the addition of 10 ml of Strong Base to a Strong Acid:

https://youtu.be/_cM1_-kdJ20 (opens in new window)



pH at the Equivalence Point in a Strong Acid/Strong Base Titration:

https://youtu.be/7POGDA5Ql2M

Titrations of Weak Acids and Bases

In contrast to strong acids and bases, the shape of the titration curve for a weak acid or a weak base depends dramatically on the identity of the acid or the base and the corresponding K_a or K_b. As we shall see, the pH also changes much more gradually around the equivalence point in the titration of a weak acid or a weak base. As you learned previously, $[\mathrm{H}^+]$ of a solution of a weak acid (HA) is not equal to the concentration of the acid but depends on both its pK_a and its concentration. Because only a fraction of a weak acid dissociates, $[\backslash(\mathrm{H}^+]$ is less than $[\mathrm{HA}]$. Thus the pH of a solution of a weak acid is greater than the pH of a solution of a strong acid of the same concentration.

Figure $15.14.3a$shows the titration curve for 50.0 mL of a 0.100 M solution of acetic acid with 0.200 M NaOH superimposed on the curve for the titration of 0.100 M HCl shown in part (a) in Figure $15.14.2$ Below the equivalence point, the two curves are very different. Before any base is added, the pH of the acetic acid solution is greater than the pH of the HCl solution, and the pH changes more rapidly during the first part of the titration. Note also that the pH of the acetic acid solution at the equivalence point is greater than 7.00. That is, at the equivalence point, the solution is basic. In addition, the change in pH around the equivalence point is only about half as large as for the HCl titration; the magnitude of the pH change at the equivalence point depends on the pK_a of the acid being titrated. Above the equivalence point, however, the two curves are identical. Once the acid has been neutralized, the pH of the solution is controlled only by the amount of excess NaOH present, regardless of whether the acid is weak or strong.

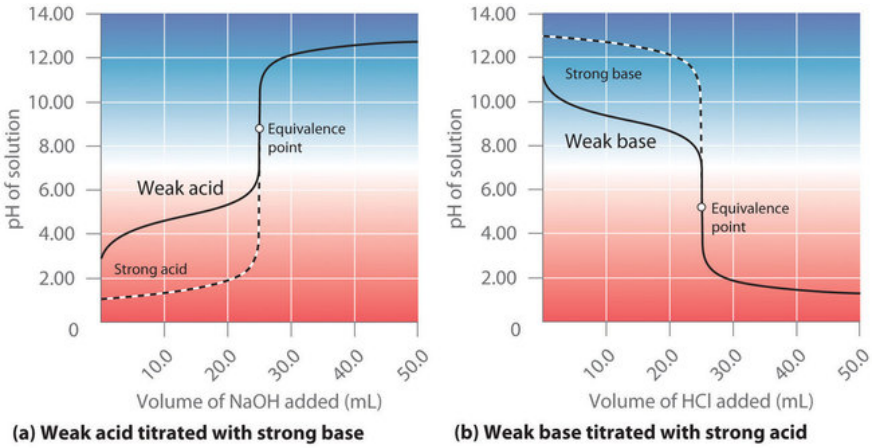


Figure 15.14.3: The Titration of (a) a Weak Acid with a Strong Base and (b) a Weak Base with a Strong Acid. (a) As 0.200 M NaOH is slowly added to 50.0 mL of 0.100 M acetic acid, the pH increases slowly at first, then increases rapidly as the equivalence point is approached, and then again increases more slowly. The corresponding curve for the titration of 50.0 mL of 0.100 M HCl with 0.200 M NaOH is shown as a dashed line. (b) As 0.200 M HCl is slowly added to 50.0 mL of 0.100 M NH3, the pH decreases slowly at first, then decreases rapidly as the equivalence point is approached, and then again decreases more slowly. The corresponding curve for the titration of 50.0 mL of 0.100 M NaOH with 0.200 M HCl is shown as a dashed line. (CC BY-SA-NC; Anonymous by request)

The titration curves of weak acid titrated with strong base and weak base titrated with strong acid are inverse of each other. These curves are shorter than the titration curves with strong acid and strong base.

> The shape of the titration curve of a weak acid or weak base depends heavily on their identities and the K_a or K_b.

The titration curve in Figure $15.14.3a$ was created by calculating the starting pH of the acetic acid solution before any NaOH is added and then calculating the pH of the solution after adding increasing volumes of $NaOH$. The procedure is illustrated in the following subsection and Example $15.14.2$ for three points on the titration curve, using the pK_a of acetic acid (4.76 at 25°C; $K_a = 1.7 \times 10^{-5}$.

Calculating the pH of a Solution of a Weak Acid or a Weak Base

As explained discussed, if we know K_a or K_b and the initial concentration of a weak acid or a weak base, we can calculate the pH of a solution of a weak acid or a weak base by setting up a ICE table (i.e, initial concentrations, changes in concentrations, and final concentrations). In this situation, the initial concentration of acetic acid is 0.100 M. If we define x as $[H^+]$ due to the dissociation of the acid, then the table of concentrations for the ionization of 0.100 M acetic acid is as follows:

$$CH_3CO_2H(aq) \rightleftharpoons H^+(aq) + CH_3CO_2^-$$

table of concentrations for the ionization of 0.100 M acetic acid

ICE	$[CH_3CO_2H]$	$[H^+]$	$[CH_3CO_2^-]$
initial	0.100	1.00×10^{-7}	0
change	−x	+x	+x
final	0.100 − x	x	x

In this and all subsequent examples, we will ignore $[H^+]$ and $[OH^-]$ due to the autoionization of water when calculating the final concentration. However, you should use Equation 16.45 and Equation 16.46 to check that this assumption is justified.

Inserting the expressions for the final concentrations into the equilibrium equation (and using approximations),

$$K_a = \frac{[H^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$
$$= \frac{(x)(x)}{0.100 - x}$$
$$\approx \frac{x^2}{0.100}$$
$$\approx 1.74 \times 10^{-5}$$

Solving this equation gives $x = [H^+] = 1.32 \times 10^{-3}\ M$. Thus the pH of a 0.100 M solution of acetic acid is as follows:

$$pH = -\log(1.32 \times 10^{-3}) = 2.879$$



pH at the Start of a Weak Acid/Strong Base Titration: https://youtu.be/AtdBKfrfJNg

Calculating the pH during the Titration of a Weak Acid or a Weak Base

Now consider what happens when we add 5.00 mL of 0.200 M NaOH to 50.00 mL of 0.100 M CH_3CO_2H (part (a) in Figure 15.14.3). Because the neutralization reaction proceeds to completion, all of the OH^- ions added will react with the acetic acid to generate acetate ion and water:

$$CH_3CO_2H_{(aq)} + OH^-_{(aq)} \rightarrow CH_3CO^-_{2\ (aq)} + H_2O_{(l)} \tag{15.14.2}$$

All problems of this type must be solved in two steps: a stoichiometric calculation followed by an equilibrium calculation. In the first step, we use the stoichiometry of the neutralization reaction to calculate the amounts of acid and conjugate base present in solution after the neutralization reaction has occurred. In the second step, we use the equilibrium equation to determine $[H^+]$ of the resulting solution.

Step 1

To determine the amount of acid and conjugate base in solution after the neutralization reaction, we calculate the amount of CH_3CO_2H in the original solution and the amount of OH^- in the NaOH solution that was added. The acetic acid solution contained

$$50.00\ \cancel{mL}(0.100\ mmol(CH_3CO_2H)/\ \cancel{mL}) = 5.00\ mmol(CH_3CO_2H)$$

The NaOH solution contained

5.00 mL=1.00 mmol $NaOH$

Comparing the amounts shows that CH_3CO_2H is in excess. Because OH^- reacts with CH_3CO_2H in a 1:1 stoichiometry, the amount of excess CH_3CO_2H is as follows:

5.00 mmol CH_3CO_2H − 1.00 mmol OH^- = 4.00 mmol CH_3CO_2H

Each 1 mmol of OH^- reacts to produce 1 mmol of acetate ion, so the final amount of $CH_3CO_2^-$ is 1.00 mmol.

The stoichiometry of the reaction is summarized in the following ICE table, which shows the numbers of moles of the various species, not their concentrations.

$$\mathrm{CH_3CO_2H(aq) + OH^-(aq) \rightleftharpoons CH_3CO_2^-(aq) + H_2O(l)}$$

ICE table

ICE	$[\mathrm{CH_3CO_2H}]$	$[\mathrm{OH^-}]$	$[\mathrm{CH_3CO_2^-}]$
initial	5.00 mmol	1.00 mmol	0 mmol
change	−1.00 mmol	−1.00 mmol	+1.00 mmol
final	4.00 mmol	0 mmol	1.00 mmol

This ICE table gives the initial amount of acetate and the final amount of OH^- ions as 0. Because an aqueous solution of acetic acid always contains at least a small amount of acetate ion in equilibrium with acetic acid, however, the initial acetate concentration is not actually 0. The value can be ignored in this calculation because the amount of $CH_3CO_2^-$ in equilibrium is insignificant compared to the amount of OH^- added. Moreover, due to the autoionization of water, no aqueous solution can contain 0 mmol of OH^-, but the amount of OH^- due to the autoionization of water is insignificant compared to the amount of OH^- added. We use the initial amounts of the reactants to determine the stoichiometry of the reaction and defer a consideration of the equilibrium until the second half of the problem.

Step 2

To calculate $[\mathrm{H^+}]$ at equilibrium following the addition of $NaOH$, we must first calculate $[\mathrm{CH_3CO_2H}]$ and $[\mathrm{CH_3CO_2^-}]$ using the number of millimoles of each and the total volume of the solution at this point in the titration:

$$final\ volume = 50.00\ mL + 5.00\ mL = 55.00\ mL$$

$$[CH_3CO_2H] = \frac{4.00\ mmol\ CH_3CO_2H}{55.00\ mL} = 7.27 \times 10^{-2}\ M$$

$$\left[CH_3CO_2^-\right] = \frac{1.00\ mmol\ CH_3CO_2^-}{55.00\ mL} = 1.82 \times 10^{-2}\ M$$

Knowing the concentrations of acetic acid and acetate ion at equilibrium and K_a for acetic acid (1.74×10^{-5}), we can calculate $[H^+]$ at equilibrium:

$$K_a = \frac{\left[CH_3CO_2^-\right][H^+]}{[CH_3CO_2H]}$$

$$\left[H^+\right] = \frac{K_a\,[CH_3CO_2H]}{\left[CH_3CO_2^-\right]} = \frac{\left(1.72 \times 10^{-5}\right)\left(7.27 \times 10^{-2}\ M\right)}{\left(1.82 \times 10^{-2}\right)} = 6.95 \times 10^{-5}\ M$$

Calculating $-\log[\mathrm{H^+}]$ gives

$$pH = -\log(6.95 \times 10^{-5}) = 4.158.$$

Comparing the titration curves for HCl and acetic acid in Figure 15.14.3*a*, we see that adding the same amount (5.00 mL) of 0.200 M NaOH to 50 mL of a 0.100 M solution of both acids causes a much smaller pH change for HCl (from 1.00 to 1.14) than for acetic acid (2.88 to 4.16). This is consistent with the qualitative description of the shapes of the titration curves at the beginning of this section. In Example 15.14.2 we calculate another point for constructing the titration curve of acetic acid.



pH Before the Equivalence Point of a Weak Acid/Strong Base Titration:

https://youtu.be/znpwGCsefXc

Example 15.14.2

What is the pH of the solution after 25.00 mL of 0.200 M NaOH is added to 50.00 mL of 0.100 M acetic acid?

Given: volume and molarity of base and acid

Asked for: pH

Strategy:

A. Write the balanced chemical equation for the reaction. Then calculate the initial numbers of millimoles of OH^- and CH_3CO_2H. Determine which species, if either, is present in excess.

B. Tabulate the results showing initial numbers, changes, and final numbers of millimoles.

C. If excess acetate is present after the reaction with OH^-, write the equation for the reaction of acetate with water. Use a tabular format to obtain the concentrations of all the species present.

D. Calculate K_b using the relationship $K_w = K_a K_b$. Calculate [OH−] and use this to calculate the pH of the solution.

Solution

A Ignoring the spectator ion (Na^+), the equation for this reaction is as follows:

$$CH_3CO_2H_{(aq)} + OH^-(aq) \rightarrow CH_3CO_2^-(aq) + H_2O(l)$$

The initial numbers of millimoles of OH^- and CH_3CO_2H are as follows:

$$25.00 \text{ mL}(0.200 \text{ mmol OH}-\text{mL}=5.00 \text{ mmol } OH-$$

$$50.00\ mL(0.100 CH_3CO_2HL = 5.00 mmol\ CH_3CO_2H$$

The number of millimoles of OH^- equals the number of millimoles of CH_3CO_2H, so neither species is present in excess.

B Because the number of millimoles of OH^- added corresponds to the number of millimoles of acetic acid in solution, this is the equivalence point. The results of the neutralization reaction can be summarized in tabular form.

$$CH_3CO_2H_{(aq)} + OH^-_{(aq)} \rightleftharpoons CH_3CO_2^-(aq) + H_2O(l)$$

results of the neutralization reaction

ICE	$[CH_3CO_2H]$	$[OH^-]$	$[CH_3CO_2^-]$
initial	5.00 mmol	5.00 mmol	0 mmol
change	−5.00 mmol	−5.00 mmol	+5.00 mmol
final	0 mmol	0 mmol	5.00 mmol

C Because the product of the neutralization reaction is a weak base, we must consider the reaction of the weak base with water to calculate [H+] at equilibrium and thus the final pH of the solution. The initial concentration of acetate is obtained from the neutralization reaction:

$$[\mathrm{CH_3CO_2}] = \frac{5.00\ mmol\ CH_3CO_2^-}{(50.00+25.00)\ mL} = 6.67\times10^{-2}\ M$$

The equilibrium reaction of acetate with water is as follows:

$$\mathrm{CH_3CO_2^-(aq) + H_2O(l) \rightleftharpoons CH_3CO_2H(aq) + OH^-(aq)}$$

The equilibrium constant for this reaction is

$$K_b = \frac{K_w}{K_a} \tag{15.14.3}$$

where K_a is the acid ionization constant of acetic acid. We therefore define x as $[\mathrm{OH^-}]$ produced by the reaction of acetate with water. Here is the completed table of concentrations:

$$H_2O_{(l)} + CH_3CO_{2(aq)}^- \rightleftharpoons CH_3CO_2H_{(aq)} + OH_{(aq)}^-$$

completed table of concentrations

	$[\mathrm{CH_3CO_2^-}]$	$[\mathrm{CH_3CO_2H}]$	$[\mathrm{OH^-}]$
initial	0.0667	0	1.00 × 10−7
change	−x	+x	+x
final	(0.0667 − x)	x	x

D We can obtain K_b by substituting the known values into Equation 15.14.3:

$$K_b = \frac{K_w}{K_a} = \frac{1.01\times10^{-14}}{1.74\times10^{-5}} = 5.80\times10^{-10} \tag{15.14.4}$$

Substituting the expressions for the final values from the ICE table into Equation 15.14.4and solving for x:

$$\frac{x^2}{0.0667} = 5.80\times10^{-10}$$

$$x = \sqrt{(5.80\times10^{-10})(0.0667)}$$

$$= 6.22\times10^{-6}$$

Thus $[OH^-] = 6.22\times10^{-6}\ M$ and the pH of the final solution is 8.794 (Figure 15.14.3*a*). As expected for the titration of a weak acid, the pH at the equivalence point is greater than 7.00 because the product of the titration is a base, the acetate ion, which then reacts with water to produce $\mathrm{OH^-}$.

? Exercise 15.14.2

Calculate the pH of a solution prepared by adding 45.0 mL of a 0.213 M HCl solution to 125.0 mL of a 0.150 M solution of ammonia. The pK_b of ammonia is 4.75 at 25°C.

Answer

9.23

As shown in part (b) in Figure 15.14.3, the titration curve for NH3, a weak base, is the reverse of the titration curve for acetic acid. In particular, the pH at the equivalence point in the titration of a weak base is less than 7.00 because the titration produces an acid.

The identity of the weak acid or weak base being titrated strongly affects the shape of the titration curve. Figure 15.14.4 illustrates the shape of titration curves as a function of the pK_a or the pK_b. As the acid or the base being titrated becomes weaker (its pK_a or pK_b becomes larger), the pH change around the equivalence point decreases significantly. With very dilute solutions, the curve becomes so shallow that it can no longer be used to determine the equivalence point.

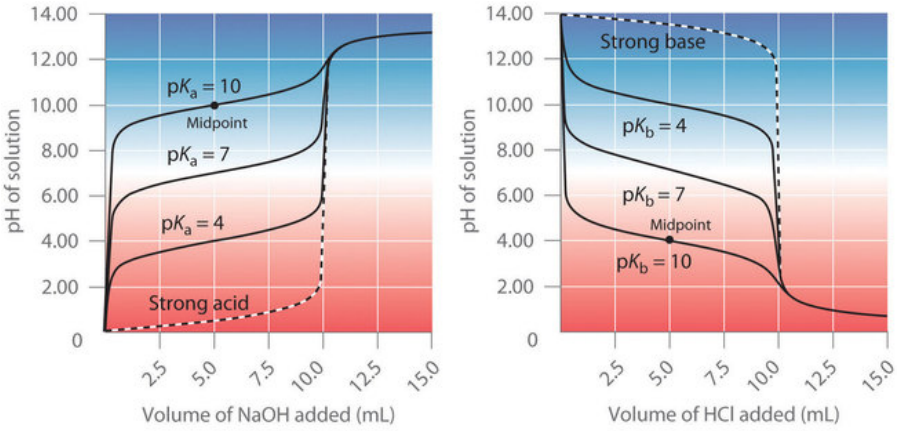


Figure 15.14.4: Effect of Acid or Base Strength on the Shape of Titration Curves. Unlike strong acids or bases, the shape of the titration curve for a weak acid or base depends on the pK_a or pK_b of the weak acid or base being titrated. (a) Solution pH as a function of the volume of 1.00 M NaOH added to 10.00 mL of 1.00 M solutions of weak acids with the indicated pK_a values. (b) Solution pH as a function of the volume of 1.00 M HCl added to 10.00 mL of 1.00 M solutions of weak bases with the indicated pK_b values. The shapes of the two sets of curves are essentially identical, but one is flipped vertically in relation to the other. Midpoints are indicated for the titration curves corresponding to pK_a = 10 and pK_b = 10. (CC BY-SA-NC; Anonymous by request)

The titration curves of weak acids with strong base and weak bases titrated with strong acid are inverses of each other. Three weak acids and three weak bases with pKa and pKb of 4, 7, and 10 are used.

One point in the titration of a weak acid or a weak base is particularly important: the midpoint of a titration is defined as the point at which exactly enough acid (or base) has been added to neutralize one-half of the acid (or the base) originally present and occurs halfway to the equivalence point. The midpoint is indicated in Figures 15.14.4*a* and 15.14.4*b* for the two shallowest curves. By definition, at the midpoint of the titration of an acid, [HA] = [A−]. Recall that the ionization constant for a weak acid is as follows:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

If $[HA] = [A^-]$, this reduces to $K_a = [H_3O^+]$. Taking the negative logarithm of both sides,

$$-\log K_a = -\log[H_3O+]$$

From the definitions of pK_a and pH, we see that this is identical to

$$pK_a = pH \tag{15.14.5}$$

Thus the pH at the midpoint of the titration of a weak acid is equal to the pK_a of the weak acid, as indicated in part (a) in Figure 15.14.4 for the weakest acid where we see that the midpoint for pK_a = 10 occurs at pH = 10. Titration methods can therefore be used to determine both the concentration and the pK_a (or the pK_b) of a weak acid (or a weak base).

> The pH at the midpoint of the titration of a weak acid is equal to the pK_a of the weak acid.

Titrations of Polyprotic Acids or Bases

When a strong base is added to a solution of a polyprotic acid, the neutralization reaction occurs in stages. The most acidic group is titrated first, followed by the next most acidic, and so forth. If the pK_a values are separated by at least three pK_a units, then the overall titration curve shows well-resolved “steps” corresponding to the titration of each proton. A titration of the triprotic acid H_3PO_4 with NaOH is illustrated in Figure 15.14.5 and shows two well-defined steps: the first midpoint corresponds to pK_a1, and the second midpoint corresponds to pK_{a2}. Because HPO_4^{2-} is such a weak acid, pK_{a3} has such a high value that the third step cannot be resolved using 0.100 M NaOH as the titrant.

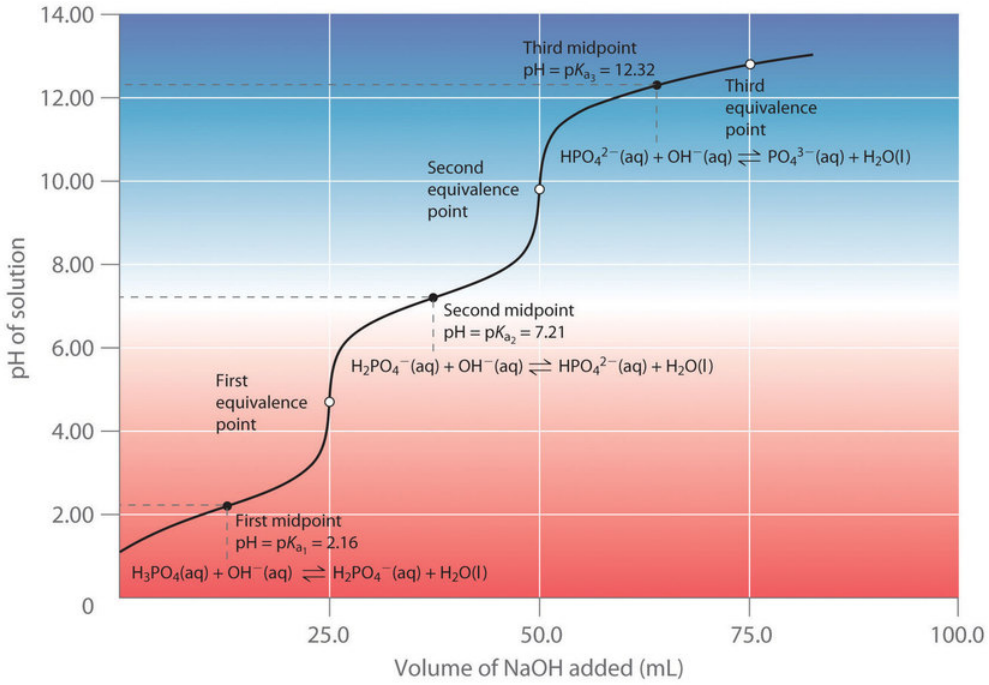
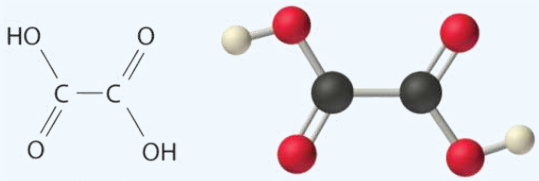


Figure 15.14.5: Titration Curve for Phosphoric Acid (H_3PO_4, a Typical Polyprotic Acid. The curve for the titration of 25.0 mL of a 0.100 M H_3PO_4 solution with 0.100 M NaOH along with the species in solution at each Ka is shown. Note the two distinct equivalence points corresponding to deprotonation of H_3PO_4 at pH ≈ 4.6 and $H_2PO_4^{2-}$ at pH ≈ 9.8. Because HPO_4^{2-} is a very weak acid, the third equivalence point, at pH ≈ 13, is not well defined. (CC BY-SA-NC; Anonymous by request)

The titration curve for the reaction of a polyprotic base with a strong acid is the mirror image of the curve shown in Figure 15.14.5 The initial pH is high, but as acid is added, the pH decreases in steps if the successive pK_b values are well separated. Table E1 lists the ionization constants and pK_a values for some common polyprotic acids and bases.

✔ Example 15.14.3

Calculate the pH of a solution prepared by adding 55.0 mL of a 0.120 M NaOH solution to 100.0 mL of a 0.0510 M solution of oxalic acid (HO_2CCO_2H), a diprotic acid (abbreviated as H_2ox). Oxalic acid, the simplest dicarboxylic acid, is found in rhubarb and many other plants. Rhubarb leaves are toxic because they contain the calcium salt of the fully deprotonated form of oxalic acid, the oxalate ion ($O_2CCO_2^{2-}$, abbreviated ox^{2-}).Oxalate salts are toxic for two reasons. First, oxalate salts of divalent cations such as Ca^{2+} are insoluble at neutral pH but soluble at low pH. As a result, calcium oxalate dissolves in the dilute acid of the stomach, allowing oxalate to be absorbed and transported into cells, where it can react with calcium to form tiny calcium oxalate crystals that damage tissues. Second, oxalate forms stable complexes with metal ions, which can alter the distribution of metal ions in biological fluids.



oxalic acid

Given: volume and concentration of acid and base

Asked for: pH

Strategy:

A. Calculate the initial millimoles of the acid and the base. Use a tabular format to determine the amounts of all the species in solution.

B. Calculate the concentrations of all the species in the final solution. Determine $[H+]$ and convert this value to pH.

Solution:

A Table E5 gives the pK_a values of oxalic acid as 1.25 and 3.81. Again we proceed by determining the millimoles of acid and base initially present:

$$100.00\ \cancel{mL}\left(\frac{0.510\ mmol\ H_2ox}{\cancel{mL}}\right) = 5.10\ mmol\ H_2ox$$

$$55.00\ \cancel{mL}\left(\frac{0.120\ mmol\ NaOH}{\cancel{mL}}\right) = 6.60\ mmol\ NaOH$$

The strongest acid (H_2ox) reacts with the base first. This leaves (6.60 − 5.10) = 1.50 mmol of OH^- to react with Hox−, forming ox^{2-} and H_2O. The reactions can be written as follows:

$$\underset{5.10\ mmol}{H_2ox} + \underset{6.60\ mmol}{OH^-} \rightarrow \underset{5.10\ mmol}{Hox^-} + \underset{5.10\ mmol}{H_2O}$$

$$\underset{5.10\ mmol}{Hox^-} + \underset{1.50\ mmol}{OH^-} \rightarrow \underset{1.50\ mmol}{ox^{2-}} + \underset{1.50\ mmol}{H_2O}$$

In tabular form,

Solutions to Example 17.3.3

	H_2ox	OH^-	Hox^-	ox^{2-}
initial	5.10 mmol	6.60 mmol	0 mmol	0 mmol
change (step 1)	−5.10 mmol	−5.10 mmol	+5.10 mmol	0 mmol
final (step 1)	0 mmol	1.50 mmol	5.10 mmol	0 mmol
change (step 2)	—	−1.50 mmol	−1.50 mmol	+1.50 mmol
final	0 mmol	0 mmol	3.60 mmol	1.50 mmol

B The equilibrium between the weak acid (Hox^-) and its conjugate base (ox^{2-}) in the final solution is determined by the magnitude of the second ionization constant, $K_{a2} = 10^{-3.81} = 1.6 \times 10^{-4}$. To calculate the pH of the solution, we need to know $[H^+]$, which is determined using exactly the same method as in the acetic acid titration in Example 15.14.2

$$\text{final volume of solution} = 100.0\,mL + 55.0\,mL = 155.0\,mL$$

Thus the concentrations of Hox^- and ox^{2-} are as follows:

$$\left[Hox^-\right] = \frac{3.60\ mmol\ Hox^-}{155.0\ mL} = 2.32 \times 10^{-2}\ M$$

$$\left[ox^{2-}\right] = \frac{1.50\ mmol\ ox^{2-}}{155.0\ mL} = 9.68 \times 10^{-3}\ M$$

We can now calculate [H+] at equilibrium using the following equation:

$$K_{a2} = \frac{\left[ox^{2-}\right]\left[H^+\right]}{\left[Hox^-\right]}$$

Rearranging this equation and substituting the values for the concentrations of Hox^- and ox^{2-},

$$\left[H^+\right] = \frac{K_{a2}\left[Hox^-\right]}{\left[ox^{2-}\right]} = \frac{\left(1.6 \times 10^{-4}\right)\left(2.32 \times 10^{-2}\right)}{\left(9.68 \times 10^{-3}\right)} = 3.7 \times 10^{-4}\ M$$

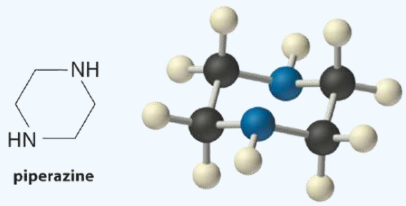
So

$$pH = -\log\left[H^+\right] = -\log\left(3.7 \times 10^{-4}\right) = 3.43$$

This answer makes chemical sense because the pH is between the first and second pK_a values of oxalic acid, as it must be. We added enough hydroxide ion to completely titrate the first, more acidic proton (which should give us a pH greater than pK_{a1}), but we added only enough to titrate less than half of the second, less acidic proton, with pK_{a2}. If we had added exactly enough hydroxide to completely titrate the first proton plus half of the second, we would be at the midpoint of the second step in the titration, and the pH would be 3.81, equal to pK_{a2}.

? Exercise 15.14.3: Piperazine

Piperazine is a diprotic base used to control intestinal parasites ("worms") in pets and humans. A dog is given 500 mg (5.80 mmol) of piperazine (pK_{b1} = 4.27, pK_{b2} = 8.67). If the dog's stomach initially contains 100 mL of 0.10 M HCl (pH = 1.00), calculate the pH of the stomach contents after ingestion of the piperazine.



Answer

pH=4.9

Indicators

In practice, most acid–base titrations are not monitored by recording the pH as a function of the amount of the strong acid or base solution used as the titrant. Instead, an acid–base indicator is often used that, if carefully selected, undergoes a dramatic color change at the pH corresponding to the equivalence point of the titration. Indicators are weak acids or bases that exhibit intense colors that vary with pH. The conjugate acid and conjugate base of a good indicator have very different colors so that they can be distinguished easily. Some indicators are colorless in the conjugate acid form but intensely colored when deprotonated (phenolphthalein, for example), which makes them particularly useful.

We can describe the chemistry of indicators by the following general equation:

$$\mathrm{HIn(aq)} \rightleftharpoons \mathrm{H^+(aq) + In^-(aq)}$$

where the protonated form is designated by HIn and the conjugate base by $\mathrm{In^-}$. The ionization constant for the deprotonation of indicator HIn is as follows:

$$K_{In} = \frac{[\mathrm{H^+}][\mathrm{In^-}]}{[\mathrm{HIn}]} \tag{15.14.6}$$

The pK_{in} (its pK_a) determines the pH at which the indicator changes color.

Many different substances can be used as indicators, depending on the particular reaction to be monitored. For example, red cabbage juice contains a mixture of colored substances that change from deep red at low pH to light blue at intermediate pH to yellow at high pH. Similarly, *Hydrangea macrophylla* flowers can be blue, red, pink, light purple, or dark purple depending on the soil pH (Figure 15.14.6). Acidic soils will produce blue flowers, whereas alkaline soils will produce pinkish flowers.

Irrespective of the origins, a good indicator must have the following properties:

- The color change must be easily detected.
- The color change must be rapid.
- The indicator molecule must not react with the substance being titrated.
- To minimize errors, the indicator should have a pK_{in} that is within one pH unit of the expected pH at the equivalence point of the titration.



Figure 15.14.6: Naturally occurring pH indicators can be found in red cabbage and Hydrangea macrophylla flowers. Red Cabbage image (CC BY-SA 3.0; KENPEI via Wikipedia) and Hydrangea macrophylla flowers (pixabay).

Synthetic indicators have been developed that meet these criteria and cover virtually the entire pH range. Figure 15.14.7shows the approximate pH range over which some common indicators change color and their change in color. In addition, some indicators (such as thymol blue) are polyprotic acids or bases, which change color twice at widely separated pH values.

Indicator	pK_{in}	Color change
Alizarin yellow R	11.0	yellow to red
Thymolphthalein	9.9	colorless to blue
Phenolphthalein	9.5	colorless to pink
Thymol blue	9.2	yellow to blue
Phenol red	7.4	yellow to red
Bromothymol blue	7.3	yellow to blue
Bromocresol purple	6.4	yellow to purple
Ethyl red	5.4	colorless to red
Resorcin blue	5.3	red to blue
Resazurin	5.1	orange to violet
Methyl red	5.0	red to yellow
Bromophenol blue	4.1	yellow to blue
Thymol blue	1.7	red to yellow
Malachite green	1.3	yellow to turquoise
Methyl violet	0.8	yellow to blue

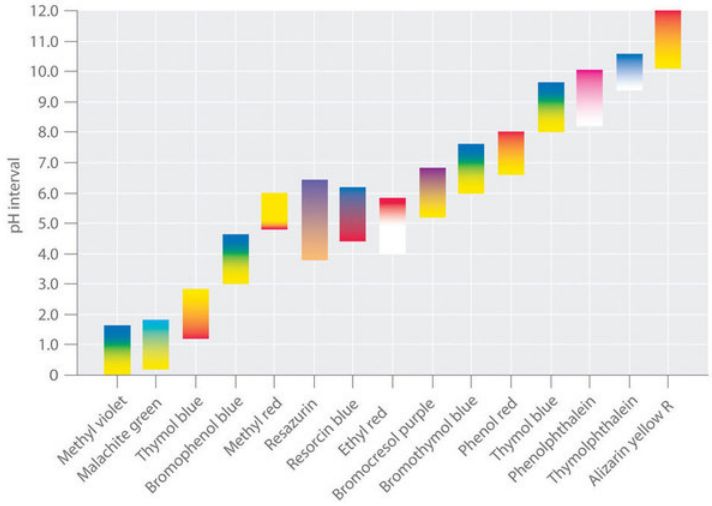


Figure 15.14.7: Some Common Acid–Base Indicators. Approximate colors are shown, along with pK_{in} values and the pH range over which the color changes. (CC BY-SA-NC; Anonymous by request)

It is important to be aware that an indicator does not change color abruptly at a particular pH value; instead, it actually undergoes a pH titration just like any other acid or base. As the concentration of HIn decreases and the concentration of In− increases, the color of the solution slowly changes from the characteristic color of HIn to that of In−. As we will see later, the [In−]/[HIn] ratio changes from 0.1 at a pH one unit below pKin to 10 at a pH one unit above pKin. Thus most indicators change color over a pH range of about two pH units.

We have stated that a good indicator should have a pKin value that is close to the expected pH at the equivalence point. For a strong acid–strong base titration, the choice of the indicator is not especially critical due to the very large change in pH that occurs around the equivalence point. In contrast, using the wrong indicator for a titration of a weak acid or a weak base can result in relatively large errors, as illustrated in Figure 15.14.8 This figure shows plots of pH versus volume of base added for the titration of 50.0 mL of a 0.100 M solution of a strong acid (HCl) and a weak acid (acetic acid) with 0.100 M $NaOH$. The pH ranges over which two common indicators (methyl red, $pK_{in} = 5.0$, and phenolphthalein, $pK_{in} = 9.5$) change color are also shown. The horizontal bars indicate the pH ranges over which both indicators change color cross the HCl titration curve, where it is almost vertical. Hence both indicators change color when essentially the same volume of NaOH has been added (about 50 mL), which corresponds to the equivalence point. In contrast, the titration of acetic acid will give very different results depending on whether methyl red or phenolphthalein is used as the indicator. Although the pH range over which phenolphthalein changes color is slightly greater than the pH at the equivalence point of the strong acid titration, the error will be negligible due to the slope of this portion of the titration curve. Just as with the HCl titration, the phenolphthalein indicator will turn pink when about 50 mL of NaOH has been added to the acetic acid solution. In contrast, methyl red begins to change from red to yellow around pH 5, which is near the

midpoint of the acetic acid titration, not the equivalence point. Adding only about 25–30 mL of NaOH will therefore cause the methyl red indicator to change color, resulting in a huge error.

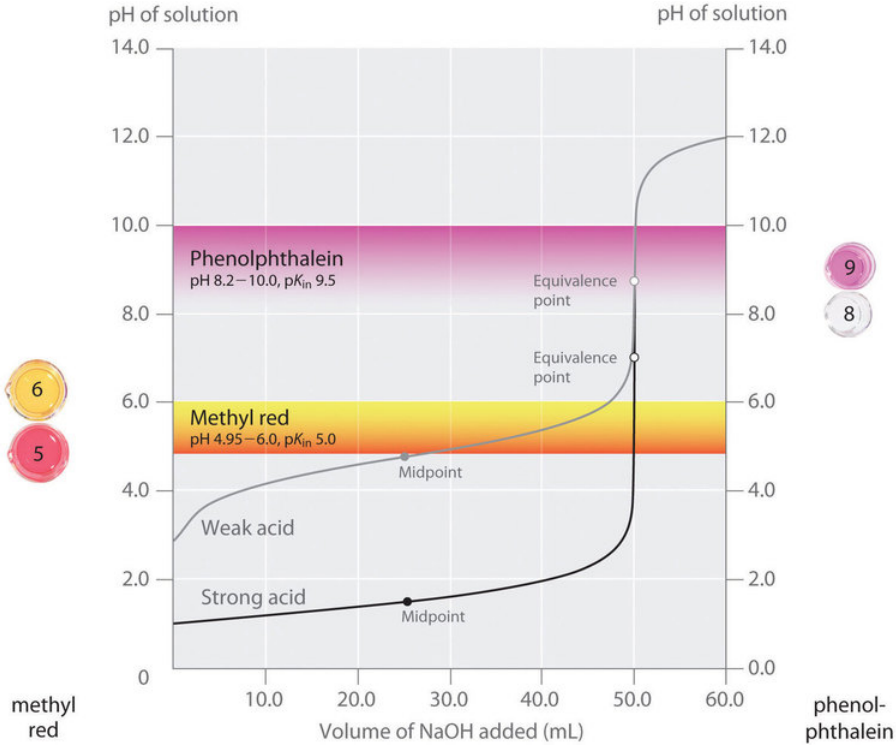


Figure 15.14.8: Choosing the Correct Indicator for an Acid–Base Titration. (CC BY-SA-NC; Anonymous by request)

If the pH of the solution is between 4.95 and 6 then methyl red should be used. If the pH is between 8,2 and 10 then phenolphthalein should be used.

The graph shows the results obtained using two indicators (methyl red and phenolphthalein) for the titration of 0.100 M solutions of a strong acid (HCl) and a weak acid (acetic acid) with 0.100 M $NaOH$. Due to the steepness of the titration curve of a strong acid around the equivalence point, either indicator will rapidly change color at the equivalence point for the titration of the strong acid. In contrast, the pKin for methyl red (5.0) is very close to the pK_a of acetic acid (4.76); the midpoint of the color change for methyl red occurs near the midpoint of the titration, rather than at the equivalence point.

In general, for titrations of strong acids with strong bases (and vice versa), any indicator with a pKin between about 4.0 and 10.0 will do. For the titration of a weak acid, however, the pH at the equivalence point is greater than 7.0, so an indicator such as phenolphthalein or thymol blue, with pKin > 7.0, should be used. Conversely, for the titration of a weak base, where the pH at the equivalence point is less than 7.0, an indicator such as methyl red or bromocresol blue, with pKin < 7.0, should be used.

The existence of many different indicators with different colors and pKin values also provides a convenient way to estimate the pH of a solution without using an expensive electronic pH meter and a fragile pH electrode. Paper or plastic strips impregnated with combinations of indicators are used as “pH paper,” which allows you to estimate the pH of a solution by simply dipping a piece of pH paper into it and comparing the resulting color with the standards printed on the container (Figure 15.14.9).

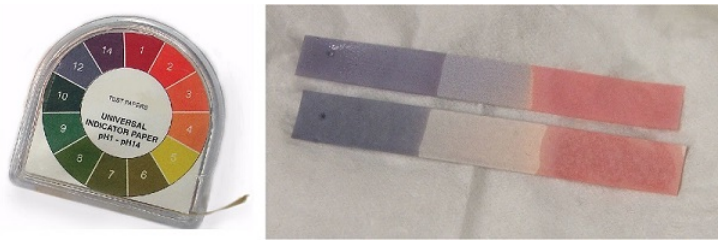


Figure 15.14.9: pH Paper. pH paper contains a set of indicators that change color at different pH values. The approximate pH of a solution can be determined by simply dipping a paper strip into the solution and comparing the color to the standards provided. (CC BY-SA-NC; Anonymous by request)



pH Indicators: pH Indicators(opens in new window) [youtu.be]

Summary and Takeaway

Plots of acid–base titrations generate titration curves that can be used to calculate the pH, the pOH, the pK_a, and the pK_b of the system. The shape of a titration curve, a plot of pH versus the amount of acid or base added, provides important information about what is occurring in solution during a titration. The shapes of titration curves for weak acids and bases depend dramatically on the identity of the compound. The equivalence point of an acid–base titration is the point at which exactly enough acid or base has been added to react completely with the other component. The equivalence point in the titration of a strong acid or a strong base occurs at pH 7.0. In titrations of weak acids or weak bases, however, the pH at the equivalence point is greater or less than 7.0, respectively. The pH tends to change more slowly before the equivalence point is reached in titrations of weak acids and weak bases than in titrations of strong acids and strong bases. The pH at the midpoint, the point halfway on the titration curve to the equivalence point, is equal to the pK_a of the weak acid or the pK_b of the weak base. Thus titration methods can be used to determine both the concentration and the pK_a (or the pK_b) of a weak acid (or a weak base). Acid–base indicators are compounds that change color at a particular pH. They are typically weak acids or bases whose changes in color correspond to deprotonation or protonation of the indicator itself.

This page titled 15.14: Neutralization Reactions and Titration Curves is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.3: Acid-Base Titrations** is licensed CC BY-NC-SA 3.0.

15.15: Solutions of Salts of Polyprotic Acids

Learning Objectives

- Predict and calculate the pH of a solution of salts of polyprotic acids

Salts can be thought of as being derived from the neutralization of an acid and a base. A salt formed from a strong acid and a strong base will not hydrolyze (e.g., extract a proton from water). When placed in water, these salts dissociate completely, and their ions remain uncombined in solution (e.g., the NaCl salt is formed from a strong acid (HCl) and a strong base (NaOH).

In discussing the titration of acids and bases, the acid or basic properties of the corresponding salts after a neutralization reaction must be addressed to calculate the final pH of a solution. To calculate the pH of a salt solution one needs to know the concentration of the salt solution, whether the salt is an acidic, basic, or neutral salt, the equation for the interaction of the ion with the water, the equilibrium expression for this interaction and the K_a or K_b value.

Note: The Hydrolysis of Salts

As a quick review, the general rules for the hydrolysis of monoprotic salts are:

1. If neither the cation/anion can affect the pH, the solution is neutral
2. If **only** the cation of the salt is acidic, the solution is acidic
3. If **only** the anion of the salt is basic, the solution will be basic
4. If a salt has a cation that is acidic and an anion that is basic, the pH of the solution is determined by the relative strengths of the acid and base
    - if $K_a = K_b$, then no net effect on pH
    - if $K_a > K_b$, then the solution is slightly acidic
    - if $K_b > K_b$, then the solution is slightly basic

Salts of Polyprotic Acids

Do not be intimidated by the salts of polyprotic acids. Yes they're bigger and "badder" then most other salts, but they can be handled the exact same way as other salts, just with a bit more math.

- All of the rules from above still apply, e.g., the pH of a salt of polyprotic acid will always be greater than 7.
- The same way that polyprotic acids lose H^+ stepwise, salts of polyprotic acids gain H^+ in the same manner, but in reverse order of the polyprotic acid.

Take for example dissociation of carbonic acid (H_2CO_3):

$$H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCO^-_{3(aq)} \tag{15.15.1}$$

with $K_{a1} = 2.5 \times 10^{-4}$

$$HCO^-_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CO^{2-}_{3(aq)} \tag{15.15.2}$$

with $K_{a2} = 5.61 \times 10^{-11}$. This means that when calculating the values for K_b of CO_3^{2-}, The K_b of the first hydrolysis reaction will be $K_{b1} = \dfrac{K_w}{K_{a2}}$ since it will go in the reverse order.

Example 15.15.1: $NaHSO_4$

What is the pH of a 0.100 M $NaHSO_4$ solution? Sulfuric acid is a strong acid, and the pK_{a_2} of HSO_4^- is 1.92.

Solution

First review the ions generates immediate upon dissociation of the salt.

- Na^+ is a non-acidic "spectator ion" that does not affect the acidity or basicity of a solution and can be ignored.
- The HSO_4^- anion will affect the pH

HSO_4^- is part of the dissociation chain of diprotic sulfuric acid (H_2SO_4) with the following steps

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^- \tag{15.15.3}$$

Because H_2SO_4 is a strong acid, this step goes to completion and no H_2SO_4 exists in the solution. The HSO_4^- ion is a weak acid and dissociates to a small extent:

$$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-} \tag{15.15.4}$$

with $K_{a2} = 1.0 \times 10^{-2}$ from Table E1. The equilibrium favors the reactants.

The $NaHSO4$ salt is completely ionized in its solution.

$$\mathrm{NaHSO_{4(s)} \rightarrow Na^+_{(sq)} + HSO_4{-}_{(sq)}} \tag{15.15.5}$$

The anion then further ionizes (Equation 15.15.4). We can construct an ICE table to solve for the final value of this ionization and hence $[H_3O^+]$

ICE Table	HSO_4^-	H_2O	H_3O^+	SO_4^{2-}
Initial	0.1 M	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.1 M -x	-	x	x

$$\begin{aligned} K_{a2} &= \frac{x^2}{0.100 - x} = 0.0100 \\ [\mathrm{H_3O^+}] &= x \\ &= \frac{-0.120 + \sqrt{0.012^2 + 4 \times 0.00120}}{2} \\ &= 0.0292 \text{ M} \end{aligned} \tag{15.15.6}$$

Thus,

$$\mathrm{pH} = -\log\ 0.0292 = 1.54$$

Note that while the dissociation is weak, the quadratic equation cannot be avoided since 0.0292 cannot be ignore when compared to 0.1 M.

? Exercise 15.15.1

Which of the following solutions are acidic, basic, or neutral?

a. $\mathrm{Na_2SO_4}$
b. $\mathrm{NaH_2PO_4}$
c. $\mathrm{Na_2HPO_4}$
d. $\mathrm{Na_3PO_4}$
e. $\mathrm{NaNO_3}$

✓ Example 15.15.2: Qualitatively Estimating the pH of a Polyprotic Salt Solution

Predict whether the Na_2HPO_4 salt will form an acidic or basic solution when dissolved in water. You will need values from Table E1 to address this.

Solution

First review the ions generates immediate upon dissociation of the salt.

- Na^+ is a non-acidic "spectator ion" that does not affect the acidity or basicity of a solution and can be ignored.

- The HPO_4^{2-} anion will affect the pH and does so via competing hydrolysis steps as discussed below.

From Table E1, the three equilibria values for the three acid/base reactions of the phosphate ion are:

$$H_3PO_{4(sq)} \rightleftharpoons H^+_{(aq)} + H_2PO^{2-}_{4(aq)} \tag{15.15.7}$$

with $K_{a1} = 6.9 \times 10^{-3}$

$$H_2PO^-_{4(sq)} \rightleftharpoons H^+_{(aq)} + HPO^{2-}_{4(aq)} \tag{15.15.8}$$

with $K_{a2} = 6.2 \times 10^{-8}$

$$HPO^{2-}_{4(sq)} \rightleftharpoons H^+_{(aq)} + PO^{3-}_{4(aq)} \tag{15.15.9}$$

with $K_{a3} = 4.8 \times 10^{-11}$

The low K_a value for the deprotonation of $HPO^{2-}_{4(sq)}$ means that it is a poor acid and better functions as a (Brønsted-Lowry) base:

$$HPO^{2-}_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_2PO^-_{4(aq)} + OH^-_{(aq)} \tag{15.15.10}$$

The two possible reactions that $HPO^{2-}_{4(aq)}$ undergoes are the acid reaction in Equation 15.15.8 and the basic reaction in Equation 15.15.10. Whether this ion will cause the solution to be acidic or basic depends on whether the K_a for Equation 15.15.8is bigger than the K_b for Equation 15.15.10.

The K_b of this reaction is related to the K_{a2} by

$$K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{1 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7} \text{ for } HPO_4^{2-}$$

Since $K_b > K_a$, the solution will be **basic**. To determine just how basic requires the use of an ICE table as discussed in Example 15.15.3

Solving the final pH for a Na_2HPO_4 salt solution is more complex than for the $NaHSO_4$, since the latter involves a single equilibrium. Two equilibria must be simultaneously solved for Na_2HPO_4. However, often time (but not always), they can be easily approximated given the relative pk_a describing the reversible reaction.

Example 15.15.3: Quantitatively Estimating the pH of a Polyprotic Salt Solution

Calculate the pH of the solution containing 3.875 g of Na_2HPO_4 that has been dissolved in a 250 mL of water.

Solution

First it is best to identify the molarity of Na_2HPO_4 that was initially generated (i.e., before any hydrolysis reactions happens).

$$c = \frac{3.875 \text{ } g/\text{molecular mass}}{\text{volume}} = \frac{3.875 \text{ } g/141.98 \text{ } g/mol}{250 \text{ } mL} = 0.11 \text{ } M$$

As discussed in Example 15.15.1, the low K_a value for the deprotonation of $HPO^{2-}_{4(sq)}$ means that it is a poor acid and instead functions as a base instead. From comparing the ionization constant in Example 15.15.1, the Na_2HPO_4 is a far better base than acid. If we ignore the acid properties of the ion and focus on the basic we get the following reaction

$$HPO^{2-}_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_2PO^-_{4(aq)} + OH^-_{(aq)} \tag{15.15.11}$$

with a K_b of 1.6×10^{-7}. Now do an ICE table to get the final $[H^+]$ concentration.

ICE Table	$HPO^{2-}_{4(aq)}$	$H_2O_{(l)}$	$H_2PO^-_{4(aq)}$	$OH^-_{(aq)}$

Initial	0.11 M	-	-	0
Change	-x	-	+x	+x
Equilibrium	0.11 M - x	-	x	x

$$K_b = \frac{[H_2PO^-_{4(aq)}][OH^-_{(aq)}]}{[HPO^{2-}_{4(aq)}]} = 1.6 \times 10^{-7} \tag{15.15.12}$$

$$\frac{x^2}{0.11\ M - x} = 1.6 \times 10^{-7} \tag{15.15.13}$$

if $x << 0.11\ M$ then $0.11\ M - x \approx 0.11\ M$ then

$$x = 1.3 \times 10^{-4} = [OH^-] \tag{15.15.14}$$

$$pOH = -\log_{10}[OH^-] = 3.8 \tag{15.15.15}$$

\[pH = pK_w - pOH]

Assuming that the reaction is at 25°, then $pK_w = 14$ and

$$pH = 14 - 3.8 = 10.2 \tag{15.15.16}$$

As expected, the solution is basic.

This page titled 15.15: Solutions of Salts of Polyprotic Acids is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.5: Solutions of Salts of Polyprotic Acids** is licensed CC BY-NC-SA 4.0.

15.16: Acid-Base Equilibrium Calculations- A Summary

This page titled 15.16: Acid-Base Equilibrium Calculations- A Summary is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

15.17: Key Terms

acid ionization | reaction involving the transfer of a proton from an acid to water, yielding hydronium ions and the conjugate base of the acid

acid ionization constant (K_a) | equilibrium constant for an acid ionization reaction

acid-base indicator | weak acid or base whose conjugate partner imparts a different solution color; used in visual assessments of solution pH

acidic | a solution in which $[H_3O^+] > [OH^-]$

amphiprotic | species that may either donate or accept a proton in a Bronsted-Lowry acid-base reaction

amphoteric | species that can act as either an acid or a base

autoionization | reaction between identical species yielding ionic products; for water, this reaction involves transfer of protons to yield hydronium and hydroxide ions

base ionization | reaction involving the transfer of a proton from water to a base, yielding hydroxide ions and the conjugate acid of the base

base ionization constant (K_b) | equilibrium constant for a base ionization reaction

basic | a solution in which $[H_3O^+] < [OH^-]$

Brønsted-Lowry acid | proton donor

Brønsted-Lowry base | proton acceptor

buffer | mixture of appreciable amounts of a weak acid-base pair the pH of a buffer resists change when small amounts of acid or base are added

buffer capacity | amount of an acid or base that can be added to a volume of a buffer solution before its pH changes significantly (usually by one pH unit)

color-change interval | range in pH over which the color change of an indicator is observed

conjugate acid | substance formed when a base gains a proton

conjugate base | substance formed when an acid loses a proton

diprotic acid | acid containing two ionizable hydrogen atoms per molecule

diprotic base | base capable of accepting two protons

Henderson-Hasselbalch equation | logarithmic version of the acid ionization constant expression, conveniently formatted for calculating the pH of buffer solutions

ion-product constant for water (K_w) | equilibrium constant for the autoionization of water

leveling effect | observation that acid-base strength of solutes in a given solvent is limited to that of the solvent's characteristic acid and base species (in water, hydronium and hydroxide ions, respectively)

monoprotic acid | acid containing one ionizable hydrogen atom per molecule

neutral | describes a solution in which $[H_3O^+] = [OH^-]$

oxyacid | ternary compound with acidic properties, molecules of which contain a central nonmetallic atom bonded to one or more O atoms, at least one of which is bonded to an ionizable H atom

percent ionization | ratio of the concentration of ionized acid to initial acid concentration expressed as a percentage

pH | logarithmic measure of the concentration of hydronium ions in a solution

pOH | logarithmic measure of the concentration of hydroxide ions in a solution

stepwise ionization | process in which a polyprotic acid is ionized by losing protons sequentially

titration curve | plot of some sample property (such as pH) versus volume of added titrant

triprotic acid | acid that contains three ionizable hydrogen atoms per molecule

---

This page titled 15.17: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **14.8: Key Terms** is licensed CC BY 4.0.

15.18: Key Equations

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ (at 25 °C)

$pOH = -\log[OH^-]$
$[H_3O^+] = \times 10^{-pH}$
$[OH^-] = \times 10^{-pOH}$
$pH + pOH = pK_w = 14.00$ at 25 °C

$K_a K_b = 1.0 \times 10^{-14} = K_w$

$pK_a = -\log K_a$
$pK_b = -\log K_b$

This page titled 15.18: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **14.9: Key Equations** is licensed CC BY 4.0.

15.19: Summary

A compound that can donate a proton (a hydrogen ion) to another compound is called a Brønsted-Lowry acid. The compound that accepts the proton is called a Brønsted-Lowry base. The species remaining after a Brønsted-Lowry acid has lost a proton is the conjugate base of the acid. The species formed when a Brønsted-Lowry base gains a proton is the conjugate acid of the base. Thus, an acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form both the hydronium ion, H_3O^+, and the hydroxide ion, OH^- when it undergoes autoionization:

The ion product of water, K_w is the equilibrium constant for the autoionization reaction:

Concentrations of hydronium and hydroxide ions in aqueous media are often represented as logarithmic pH and pOH values, respectively. At 25 °C, the autoprotolysis equilibrium for water requires the sum of pH and pOH to equal 14 for any aqueous solution. The relative concentrations of hydronium and hydroxide ion in a solution define its status as acidic ($[H_3O^+] > [OH^-]$), basic ($[H_3O^+] < [OH^-]$), or neutral ($[H_3O^+] = [OH^-]$). At 25 °C, a pH < 7 indicates an acidic solution, a pH > 7 a basic solution, and a pH = 7 a neutral solution.

The relative strengths of acids and bases are reflected in the magnitudes of their ionization constants; the stronger the acid or base, the larger its ionization constant. A reciprocal relation exists between the strengths of a conjugate acid-base pair: the stronger the acid, the weaker its conjugate base. Water exerts a leveling effect on dissolved acids or bases, reacting completely to generate its characteristic hydronium and hydroxide ions (the strongest acid and base that may exist in water). The strengths of the binary acids increase from left to right across a period of the periodic table ($CH_4 < NH_3 < H_2O < HF$), and they increase down a group ($HF < HCl < HBr < HI$). The strengths of oxyacids that contain the same central element increase as the oxidation number of the element increases ($H_2SO_3 < H_2SO_4$). The strengths of oxyacids also increase as the electronegativity of the central element increases [$H_2SeO_4 < H_2SO_4$].

The ions composing salts may possess acidic or basic character, ionizing when dissolved in water to yield acidic or basic solutions. Acidic cations are typically the conjugate partners of weak bases, and basic anions are the conjugate partners of weak acids. Many metal ions bond to water molecules when dissolved to yield complex ions that may function as acids.

An acid that contains more than one ionizable proton is a polyprotic acid. These acids undergo stepwise ionization reactions involving the transfer of single protons. The ionization constants for polyprotic acids decrease with each subsequent step; these decreases typically are large enough to permit simple equilibrium calculations that treat each step separately.

Solutions that contain appreciable amounts of a weak conjugate acid-base pair are called buffers. A buffered solution will experience only slight changes in pH when small amounts of acid or base are added. Addition of large amounts of acid or base can exceed the buffer capacity, consuming most of one conjugate partner and preventing further buffering action.

The titration curve for an acid-base titration is typically a plot of pH versus volume of added titrant. These curves are useful in selecting appropriate acid-base indicators that will permit accurate determinations of titration end points.

This page titled 15.19: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **14.10: Summary** is licensed CC BY 4.0.

15.20: Exercises

These are homework exercises to accompany the Textmap created for "General Chemistry: Principles and Modern Applications " by Petrucci et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Q15.1

At room temperature, a solution is made by adding 50mL of 0.325M HF $(Ka = 6.6 \times 10^{-4})$ with 10mL of 0.50M HCl. Calculate the following at equilibrium:

a. $[H_3O^+]$
b. $[OH^-]$
c. $[HF]$
d. $[Cl^-]$

Q15.2 - [Questionable, Please Recheck the question and the answer]

For a solution that is 0.138 M NH_3 $(Ka = 1.9 \times 10^{-5})$ and 0.101 M KN_3 (potassium azide) calculate the following:

a. $[H_3O^+]$
b. $[N_3^-]$
c. $[K^+]$
d. $[OH^-]$

Q15.3

Calculate the change in pH that results from adding

a. 0.150 mol $LiNO_2$ to 1.00 L of 0.150 M $HNO_{2\,(aq)}$
b. 0.150 mol $LiNO_3$ to 1.00 L of 0.150 M $HNO_{3\,(aq)}$

Why are these changes not the same? Explain your answer.

Q15.4

A solution of volume 50.0 mL contains 12.5 mmol $HClO_2$ and 7.25 mmol $KClO_2$.

a. What is the pH of this solution?
b. If 1.00 mmol $Ca(OH)_2$ is added to the solution, what will be the new pH?
c. If 2.00 mL of 6 M HCl is added to the original solution, what will be the new pH?

Q15.5

There are various procedures for preparing buffer solutions. To obtain a $pH = 4.60$, one such way is to mix 18.00 mL of 0.300 M C_5H_5N with 72.00 mL of 0.300 M $C_5H_5NH^+$.

a. Show by calculation that the pH of this solution is 4.60.
b. Would you expect the pH of this solution to remain the same if 100.00 mL of buffer solution were diluted to 500.00 mL? To 1.00 L? Explain.
c. What will be the pH of the original 100.00 mL of buffer solution if 0.40 mL of 1.00 M HBr was added to it?
d. What is the maximum volume of 1.00 M HBr that can be added to the original buffer solution so that the pH does not drop below 4.4?

Q16.6

Use the following data table to answer the following questions

Name	pK_{ind}	Acid Color	pH Range of Color Change	Base Color
Methyl violet		Yellow	0.0 – 1.6	Blue
Thymol blue		Red	1.2 – 2.8	Yellow
Methyl orange	3.7	Red	3.2 – 4.4	Yellow
Bromocresol green		Yellow	3.8 – 5.4	Blue
Methyl red		Red	4.8 – 6.0	Yellow
Litmus	6.5	Red	5.0 – 8.0	Blue
Bromothymol blue	7.9×10^{8}	Yellow	6.0 – 7.6	Blue
Thymol blue		Yellow	8.0 – 9.6	Blue
Phenolphthalein	9.3	Colorless	8.2 – 10.0	Pink
Thymolphthalein	1.0×10^{-10}	Colorless	9.4 – 10.6	Blue
Alizarin yellow R		Yellow	10.1 – 12.0	Red

Identify one indicator (each) that would change color in an acidic solution, a basic solution, and a neutral solution.

Q16.7

What color will a solution with bromphenol blue change to with the following species in solution

a. $Mg(OH)_2$
b. HCl
c. HF

Q16.8

What stoichiometric concentration of the indicated substance is required to obtain an aqueous solution with the pH value: $Na(OH)_2$ for $pH = 10.92$

Q16.9

A solution is prepared with 0.3 M NH_3 and 0.4 M NH_4Cl.

Show that this is a buffer solution.

Calculate the pH of this buffer solution. $K_b = 1.8 \times 10^{-5}$

What is the final pH if 1 L of 0.2 M $NaOH$ is added to 1 L of this buffer solution.

Q16.10

Hydroflouric acid, HF, is a stronger acid than water and when in an aqueous solution, the following reaction occurs.

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$

Three experiments were completed to obtain a pKa value for HF at 0^oC. A sample of HF was shaken and separated into two layers, one with 1-pentanol, and the other with water. Assume no ionization or disassociation occurs in the 1-pentanol layer.

- The first experiment showed that the water layer contained 5 times as much HF as the 1-pentanol layer did.
- The 2nd experiment involved adding .5 M $LiOH_{(aq)}$ and 1-pentanol. At equilibrium, the concentration of HF was 0.005 M in the 1- pentanol layer and 0.35 M in the aqueous layer.
- The third experiment showed that when .25 M $LiOH_{(aq)}$ was added, the concentrations in the 1-pentanol layer was .0015 in the 1-pentanol layer and .12 M in the aqueous layer. At 0^oC, $pK_w = 14.94$.

Find pK_a value of HF.

Q16.11

To increase the ionization of H_2S, which of the following compounds should be added to the solution?

a. $NaHSO_4$
b. HCl
c. Acetic acid (CH_3COOH)
d. $NaCN$

Q16.12

For the solution that is 0.330 M CH_3CH_2COOH $K_a = 1.3 \times 10^{-5}$ and 0.010M HI, calculate

(a) $[H_3O^+$

(b) $[OH^-]$

(c) $[CH_3CH_2COOH]$

(d) $[I^-]$

Q16.13

For a solution that is 0.654 M NH_3 and 0.202 M NH_4Cl, calculate :

(a) $[H_3O^+]$

(b) $[OH^-]$

(c) $[NH_4^+]$

d) $[Cl^-]$

Q16.14

Calculate the change in pH that results from adding (a) 0.300 mol $NaNO_2$ to 1.00 L of 0.200 M $HNO_{2(aq)}$; (b) 0.300 mol $NaNO_3$ to 1.00 L of 0.200M $HNO_{3(aq)}$.

Q16.15

A solution of volume 1L contains 0.500 mol $HCHO_2$ and 0.250 mol $NaHCOO$.

a. What is the pH of this solution?
b. If 0.1 mol $Ba(OH)_2$ is added to to the solution, what will be the new pH?
c. If 10.0 mL of 12 M HCl is added to the original solution what will be the new pH?

Q16.16

To obtain a $pH = 8.00$ you need to mix 35.00 mL of .220 M NH_3 with 65.00 mL of .250 M NH_4Cl.

a. Show the calculation that the pH of this solution is 8.00.
b. Would you expect the pH of this solution to remain at 8.00 if the 100 mL of the buffer were diluted to 1.00 L? 1000L? Explain.
c. What will be the pH of the original solution if 0.20 mL of 1.00 M HCl is added?

Q16.17

Use data from table

Indicator	K_{HIn}	Acid -> Anion
Bromphenol blue	1.4×10^{-4}	yellow -> blue
Bromcresol green	2.1×10^{-5}	yellow -> blue
Bromthymol blue	7.9×10^{-8}	yellow -> blue
2,4 – Dinitrophenol	1.3×10^{-4}	colorless -> yellow
Chlorphenol red	1.0×10^{-6}	yellow -> red
Thymolphthalein	1.0×10^{-10}	colorless -> blue

a) Which of these indicators change color in acidic, basic, or neutral point.

b) What is the approximate pH of a solution if bromocresol green indicator turns green? If chlorphenol red turns orange?

Q16.18

Phenol red indicator changes from yellow to red in pH ranges from 6.6 to 8.0. Without making detailed calculations, state what color the indicator will assume in each of the following solutions.

a. 0.10 M KOH
b. 0.10 M CH_3COOH.
c. 0.10M NH_4NO_3

d. 0.10M HBr
e. 0.10 M $NaCN$
f. 0.10 M CH_3COOH and 0.10 M $NaCH_3COO$.

Q16.19

Two aqueous solutions are mixed. 75 mL of .333 M H_2SO_4 and 25 mLof 0.05 M $NaOH$. What is the pH of the resulting solution?

Q16.20

Explain why the volume of 0.100 M $NaOH$ required to reach the equivalence point in the titration of 30.00 mL of 0.100 M HA is the same regardless of whether HA is a strong or a weak acid, yet the pH at the equivalence point is not the same.

Q16.21

Sketch a titration curve for each of the following three weak acids when titrated with 0.100 M $NaOH$.

a. 10.00 mL of 0.100 M HX; $Ka = 7.0 \times 10^{-3}$
b. 10.00 mL of 0.100 M HY; $Ka = 3.0 \times 10^{-4}$
c. 10.00 mL of 0.100 M HZ; $Ka = 2.0 \times 10^{-8}$

Q16.22

Is 0.20M $Na_2S_{(aq)}$ solution likely to be acidic, basic, or neutral? Explain.

Q16.23

Sodium phosphate, Na_3PO_4, is made by neutralizing phosphoric acid with sodium carbonate to obtain Na_2HPO_4. The Na_2HPO_4 is further neutralized to Na_3PO_4 with $NaOH$.

1. a) Write net ionic equations for these reactions.
2. b) Na_2CO_3 is a much cheaper base than $NaOH$. Why do you suppose that $NaOH$ must be used as well with Na_2CO_3 to produce Na_3PO_4?

Q16.24

The pH of a solution of 19.5 g malonic acid in 0.250 L is 1.47. The pH of a 0.300 M solution of sodium hydrogen malonate is 4.26. What are the values of K_{a1} and K_{a2} for malonic acid?

Q16.25

Using appropriate equilibrium constants but without doing detailed calculations, determine whether a solution can be simultaneously.

a. 0.10 M NH_3 and 0.10 M NH_4Cl, with $pH = 6.07$
b. 0.10 M $NaC_2H_3O_2$ and 0.058 M HI
c. 0.10 M KNO_2 and 0.25 M KNO_3
d. 0.050 M $Ba(OH)_2$ and 0.70 M NH_4Cl
e. 0.018 M C_6H_5COOH and 0.018 M NaC_6H_5COO with $pH = 4.10$

Q16.26

A buffer solution can be prepared by by starting with a weak acid, HA, and converting some of the weak acid to its salt by titration with a strong base. The fraction of the original acid that is converted to the salt is designated F.

1. a) Derive an equation similar to (16.2) but express in terms of F rather than concentrations.
2. b) What is the pH at the point in the titration of phenol, C_6H_5OH, at which $F = 0.27$ $(pK_a\ of\ phenol = 10.00)$?

Q16.27

Complete the derivation of equation (16.5). Then derive equation (16.6)

Q16.28

A solution is prepared that is 0.200 M CH_3COOH and 0.250M $NaHCOO$.

a. Show that this is a buffer solution
b. Calculate the pH of this solution.

Q16.29

For water at 0^oC, $pK_w = 14.94$. Find the approximate value of K_a for H_2O_2 at 0^oC.

Q16.30

To increase the ionization of formic acid, $HCOOH_{(aq)}$, which of the following should be added to the solution?

a. $NaCl$
b. $NaHCOO$
c. H_2SO_4
d. $NaHCO_3$

Q16.31

For a solution that is 0.345 M hydrosulfuric acid ($K_a = 1.0 \times 10^{-7}$) and a pH of 3.2. Find $[H_3O^+]$.

Q16.32

For a solution that is 0.35 M $HC_2H_3O_2$ and $C_2H_3O_2$ calculate:

a. pH
b. $[H_3O^+]$

Q16.33

Calculate the pH of the buffer that results from mixing 100.0 ml of .250 M $HCHO_2$ and 10.0 ml of 1 M $NaCHO_2$. $pKa = 3.744$

Q16.34

What is the pH of a 0.07M ammonia solution?

Q16.35

What are the approximate pHs of a solution for the following indicator strips at the specified color change?

a. methyl red indicator strip changes from orange to yellow
b. bromphenol indicator strip changes from yellow to purple
c. thymol blue indicator strip changes from yellow to blue

Q16.36

Azolitman indicator strips change from red to blue in the pH range from 4.5 to 8.3. What color will the indicator be for the following:

a. 0.000014 M HCl
b. 0.00014 M HCl
c. 0.0014 M HCl
d. 0.0000014 M HCl

Q16.37

Why do we usually try to keep the amount of indicator added relatively small when used as acid-base indicators?

Q16.38

Define the equivalence point in a titration curve.

Q16.39

What relationship do the molarity of Base added in a titration and molarity of either strong of weak acid share?

Q16.40

0.205 M $HClO$ reacts with water to form .020 $MClO^-$. The pH of the solution is 4.04. Determine K_a of $HClO$.

Q16.41

Titrate 50mL of 0.200M HNO_3 with 0.100M $NaOH$. Add 150mL $NaOH$. 15mmol HNO_3. And Find pOH and pH.

Q16.42

Is a solution of 0.23 M NaI likely to be acidic, basic, or pH neutral? Explain.

Q16.43

In a strong acid base titration, at what pH does the equivalence point always occur?

Q16.44

Find the pH for HCN at 0^oC if the pK_b of HCN is 4.8

$HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$

Q16.45

Can an aqueous solution of NH_3 and NH_4Cl form a pH of 6 with equal molarity of each added together?

Q16.46

A buffer solution contains 0.50M acetic acid ($Ka = 1.8 \times 10^{-5}$) and 0.50M sodium acetate. Find the pH of the solution.

Q16.47

You need bring the pH of .325 L of .452 M $NaOH$ to 7.00. What volume of which of the following solutions would you use: 5.00 M HBr or 5.00 M Na_2S?

Q16.48

What is the pH of a 0.025 M HCN solution? $K_a = 6.2 \times 10^{-10}$

Q16.49

Define the common Ion Effect and give an example of it.

Q16.50

If a solution contains 50 ml of 0.25M of NH_3 and 75 ml NH_4Cl, calculate the molarity of NH_4Cl.

Q16.51

If given two different solutions, one acid and one base, what is the best method to use to calculate the pH of the solution?

Q16.52

If a mixture of .005 mol $NaOH$ and .008 mol HCl has a volume of 50ml, what is range of pH of the solution?

Q16.53

If an original solution had a pH level of 6.05, in what direction would the pH shift (towards or away from equilibrium) if 50ml of HCl were added

Q16.54

Does Bromecresol Green (2.1×10^{-5}) change color in an acidic or basic solution?

Q16.55

If Chlorphenol Red indicator turns red at $ph = 6.0$, then will it change colors at 0.10 M KOH?

Q16.56

Two aqueous solutions are mixed together, 100 ml of 0.025M H_2SO_4 and a given amount of 0.03M $NaOH$. What is the total volume at equilibrium?

Q16.57

Why is the volume the same if 0.2 M HA in a solution if 0.2 M $NaOH$ is either a weak or strong acid?

Q16.58

Sketch a titration curve of a hypothetical weak acid when .1M of $NaOH$ is added. Assume 100ml at 0.1M of $C_2H_4O_2$.

Q16.59

With the titration of 50ml of 0.1M $NaOH$ with 0.1M HCl, solve for the volume of HCl

Q16.60

Is the solution 0.1M HCl likely to be acidic, basic, or neutral?

Q16.61

Define the term of stoichiometric concentration.

Q16.62

A buffer solution can be made by taking a weak acid and converting a portion to its salt. This is done by titration with what?

Q16.63

If a pH of an unknown substance is set at 5, would a strong acid or a strong base solution be needed for the substance to reach equilibrium?

Q16.64

A solution is made of 0.1M CH_3COOH and 0.1M $NaHCOO$. Calculate the range of pH of this solution.

Q16.65

In a solution of 0.153 M CH_3COOH ($K_a = 1.8 \times 10^{-5}$). Calculate $[H_3O^+]$.

Q16.66

In a Solution of 0.548M HCl ($K_a = 7.2 \times 10^{-4}$) and pure water, calculate pH.

Q16.67

In a 300 mL solution of 0.1 M $HClO_{2(aq)}$ calculate the pH change when adding 10 mL of 0.2 M $HCl_{(aq)}$ to the solution ($K_a = 1.1 \times 10^{-2}$).

Q16.68

A 55 mL solution is made up of 0.45 mmol $HClO_2$ and 6.0 mmol $KClO_2$. Calculate the pH.

Q16.69

In a buffer solution of 50 mL of 0.3 M H_3PO_4 and 25 mL of .3M NaH_2PO_4 with a pH of 1.82 what would the K_a of this solution be?

Q16.70

Between an indicator with a $K_{Hln} = 2 \times 10^{-4}$, an indicator with a $K_{Hln} = 2 \times 10^{-10}$, and an indicator with a $K_{Hln} = 2 \times 10^{-7}$, which would be used in a titration of strong acid and strong base to find the endpoint?

Q16.71

Bromophenol blue changes from yellow to blue-violet in the pH range 3-5. Which of the following titrations would this be best to use in: strong acid-strong base, strong acid-weak base, strong base-weak acid?

Q16.72

When 20mL of 0.05M KOH of is titrated with 0.05M HI, an indicator goes from yellow to red when 10mL of the weak acid is added. What is the pK_{Hln} of the indicator? Does this give the accurate endpoint of titration?

Q16.73

What is the pOH of a solution with 200mL of 0.007M HCl and 300mL of 0.02M $NaOH$?

Q16.74

Which will be greater, the volume of 0.2M $LiOH$ is needed to reach the equivalence point with 20mL of 0.2M HF or with 20mL of 0.2M CH_3COOH? What about their pH at the equivalence point?

Q16.75

Based on K_b, how do titrations curves of strong acids with weak bases differ?

Q16.76

Calculate the pH of the solution in the titration of 0.05M HBr with 40mL of 0.02M $CsOH$ when, 1) 5mL 2)10mL 3) 25mL of HBr is added.

Q16.77

Does 0.05M $NaHSO_4$ produce an acidic, basic, or neutral solution? (given that $pK_{a2} = 1.92$)

Q16.78

How can you obtain Na_3SO_4 from sulfuric acid and sodium carbonate? Write out equations.

Q16.79

Show by calculations that the stoichiometric concentration of CH_3COOH is 0.1 M is required to obtain an aqueous solution with the pH value of about 2.8 .

Q16.80

Determine the pH values of $NaH_2PO_{4\,(aq)}$ and $Na_2HPO_{4\,(aq)}$ when $pK_{a1} = 2.15$, $pK_{a2} = 7.20$, $pK_{a3} = 12.38$. Note that $H_2PO_4^-$ occurs simultaneously.

Q16.81

Define what the term simultaneously means in terms of hydrolysis. Determine whether a solution can be simultaneously (these don't need calculations).

a. 0.20M NH_3 and 0.20 NH_4Cl

b. 0.20 M $NaC_2H_3O_2$ and 0.116 M HI

c. 0.036 M C_6H_5COOH and 0.036 M NaC_6H_5COO

Q16.82

Determine the amount (in mL) of 6.0 M $NaOH$ that is required to adjust the pH of the initial buffer solution to a pH of 7.4.

Note that a student working with this buffer mixes together 400 mL of 1.80 M Na_2HPO_4 and 130 g of NaH_2PO_4.

Q16.83

What is the buffer range for a benzoic acid/ sodium benzoate buffer? $[Ka = 6.3 \times 10^{-5}]$.

Q16.84

Define what a common ion effect is when a strong acid supplies the common ion H_3O^+, the equilibrium shifts which way to form more CH_3COOH? When a strong basic supplies the common ion OH^-, the equilibrium shifts which way to form more NH_3?

Q16.85

What is the effect on the pH of adding 0.0060 mol HCl to 0.3 L of a buffer solution that is 0.250 M CH_3COOH and 0.560 M $NaCH_3COO$?

Q16.86 (Questionable – please check question and answer)

Find the pH of 2.0×10^{-8} M HCl. Note that water makes a significant contribution to the $[H_3O^+]$. ($K_w = [H_3O^+][OH^-]$)

Q16.87

Define what ionization energy is and determine which of the following would decrease the ionization of acidic acid, $CH_3COOH_{(aq)}$: $NaCl$, $NaHCOO$, H_2SO_4, and $NaHCO_3$.

Q16.88

0.45M Hypochlorous acid ($HOCl$) with a $K_a = 2.9 \times 10^{-8}$ is mixed with 0.001M HCl.

a) Find $[H_3O^+]$

b) Find $[OH^-]$

c) Find $[OCl^-]$

d) Find $[Cl^-]$

Q16.89

0.3M $C_5O_5H_5N$ with a $K_b = 1.5 \times 10^{-9}$ is mixed with 0.015M C_5H_5NHI

a) Find $[OH^-]$

b) Find $[C_5H_5NH^+]$

c) Find $[I^-]$

d) Find $[H_3O^+]$

Q16.90

A solution of volume 0.25 L contains 0.2 moles of HCN and .004 moles CN^-.

a) What is the pOH?

b) If 0.25M $NaOH$ is added, what is the pH?

c) If .076 L of 3M HNO_3 is added, what is the pH?

Q16.91

A manual says that in order to obtain a buffer with a pH of 7.00 you have to mix 32L of 0.644M $HOCl$ with .7L of 0.406M OCl^-.

a) Show by calculation that the pH of this solution is 7.00

b) Would the pH remain at 7.00 if 1 liter of H_2O was added to the solution?

c) What will the pH be of the original 1.02L solution if 0.15L of 0.8M $NaOH$ is added?

d) What is the maximum volume of 0.8M $NaOH$ that can be added to the 1.02L buffer solution so that the pH does not change above 7.7?

Q16.92

Bromocresol green ($K = 2.1 \times 10^{-5}$), as this indicator change from yellow to blue, the pH ranges from 4-5. Without making detailed calculations, state what color the indicator will assume in each of the following solutions.

a) 0.3M $NaOH$

b) 0.02M $HClO$

c) 0.3M NH_3

d) 0.25M HI

e) 0.4M Na_2CO_2

f) 0.2M HCO_3 and 0.2M \(CO_3^-) (Buffer)

Q16.93

Two liquid solutions are mixed, 25ml of 0.73M HI and 10ml of 0.29M $Ca(OH)_2$. What is the pOH of the resulting solution?

Q16.94

Explain why the volume of 0.25M HCl required to reach the equivalence point in a titration of 50ml of BOH is the same whether it is a strong or a weak base, yet the pOH at the equivalence point is not the same.

Q16.95

Is the solution that is 0.3M Na_3PO_4 likely to be acidic, basic, or pH neutral? Why?

Q16.96

For the solution that is 0.330 M CH_3CH_2COOH that has K_a value to be 1.3×10^{-5} and 0.01M HI, calculate

1. (a) $[H_3O^+]$
2. (b) $[OH^-]$
3. (c) $[CH_3CH_2COO^-]$
4. (d) $[I^-]$

Q16.97

For a solution that is 0.654 M NH_3 and 0.202 M NH_4Cl, calculate $[H_3O^+]$, $[OH^-]$, $[NH_4^+]$ with $K_b = 1.8 \times 10^{-5}$.

Q16.98

Calculate the change in pH that results from adding (a) 0.300 mol $NaNO_2$ to 1.00 L of 0.200 M $HNO_{2(aq)}$; (b) 0.300 mol $NaNO_3$ to 1.00 L of 0.200M $HNO_{3(aq)}$.

Q16.99

A solution of volume 1L contains 0.500 mol $HCHO_2$ and 0.250 mol $NaHCOO$.

a) What is the pH of this solution?

b) If 0.1 mol $Ba(OH)_2$ is added to to the solution, what will be the pH?

c) If 10.0 mL of 12 M HCl is added to the original solution what will be the pH?

Q16.100

To obtain a $pH = 8.00$ you need to mix 35.00 mL of .220 M NH_3 with 65.00 mL of .250 M NH_4Cl.

a) Show the calculation that the pH of this solution is 8.00.

b) Would you expect the pH of this solution to remain at $pH = 8.00$ if the 100 mL of the buffer were diluted to 1.00 L? 1000L? Explain.

c) What will be the pH of the solution if 0.20 mL of 1.00 M HCl is added?

Q16.101

What stoichiometric concentration of the indicated substance is required to obtain an aqueous solution with the pH value shown?

a) $Ba(OH)_2$ for $pH = 10.00$

b) CH_3COOH in 0.300 M $NaCH_3COO$ for $pH = 4.50$?

Q16.102

You are asked to bring the pH of 1.00 L of 0.500 M $NH_4Cl_{(aq)}$ to 7.00. How many drops (1 drop = 0.05 mL) of which of the following solutions would you use: 10.0 M HCl or 10.0 M NH_3.

Q16.103

A handbook lists various procedures for preparing buffer solutions. To obtain a pH of 9, the handbook says to mix 36.0 mL of 0.200 M NH_3 with 64 mL of 0.200 M NH_4Cl.

a.) Prove that $pH = 9$ is correct

b.) Would the pH remain at 9 if the 100 mL buffer solution was 200 L? Why or why not?

Q16.104

What are the main principal features of a titration curve for the titration of a strong acid with a strong base?

Q16.105

The pH of a solution of 25 g of malonic acid in 0.300 L is 2.68, with a K_{a1} value of 4.5×10^{-12}. What is $[HM^-]$ in the malonic acid solution?

Q16.106

Show the derivation of the Henderson-Hasselbalch equation.

Solutions

S1

$Moles\ of\ HCl\ added = Moles\ of\ H_{3}O^{+} = 0.01L(0.005M)=5\times 10^{-4}\ mol$

$Moles\ of\ HF\ added = (0.05L)(0.325M)=0.01625\ mol$

$Total\ Volume= 50mL + 10mL = 60mL = 0.06L$

$M_{H_{3}O^{+}\ initial}=\frac{5\times 10^{-4}\ mol}{0.06L}=0.00833M$

$M_{HF}=\frac{0.01625\ mol}{0.06L}= 0.271M$

$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} +H_3O^{+}_{(aq)}$

ICE Table	$HF_{(aq)}$	$H_2O_{(l)}$	$F^-_{(aq)}$	$H_3O^+_{(aq)}$
Initial	0.271 M	-	0	0.0083 M
Change	-x	-	+x	+x
Equilibrium	(0.271 –x) M	-	x	(0.00833 +x) M

$K_a= \dfrac{[H_3 O^+][F^-]}{[HF]} = 6.6 \times 10^{-4}= \dfrac{x(0.00833+x)}{0.271-x}$

Interpret the equation above we have:

$6.6\times 10^{-4}(0.271-x) = 0.00833x+ x^{2}$

$x^{2}+0.00899x-1.789\times10^{-4}=0$

$x=0.009615M$

a) $[H_{3}O^{+}]=0.00833+x=0.00833+0.009615=0.0179M$

b) At room temperature $pH + pOH = 14$,

$pH=-log[H_{3}O^{+}] = -log(0.0179)=1.75$

$pOH=-log[OH^{-}]=14-1.75=12.25$

Therefore, $[OH^{-}]=10^{-12.25}=5.623\times 10^{-13}M$

c)$[HF] = 0.271-x=0.271-0.096165=0.2614M$

d)$[Cl^{-}]=[HCl]=0.00833M$

S2

$NH_3 +H_2O \rightleftharpoons H_3O^+ + N_3^-$

ICE Table	NH_3	H_2O	H_3O^+	N_3^-
Initial	0.138 M	-	0	0.101 M
Change	-x	-	+x	+x
Equilibrium	(0.138-x)	-	x	0.101 +x

$K_a= \frac{[H_3 O^+][N_3^-]}{[NH_3]}=1.9\times 10^{-5}= \frac{x(0.101+x)}{(0.138-x)}\approx \frac{0.101x}{0.138}$

$x\approx 2.6 \times 10^{-7}$

The assumption that x<< 0.101 is correct. $[H_3O^+] = 2.6 \times 10^{-7}\ M$

b)

$[N_3^-]=(0.101+x)=(0.101+2.6 \times 10^{-7})\approx 0.101\ M$

c)

$[K^+]=[N_3^-]\approx 0.101\ M$

d)

$[OH^-]=\frac{K_w}{[H_3O^+]}=\frac{(1.0 \times 10^{-14})}{(2.6 \times 10^{-7})}=3.8 \times 10^{-8}\ M$

S3

a) First, determine the pH of 0.150 M HNO_2.

$HNO_{2(aq)} +H_2O_{(l)} \rightleftharpoons NO_{2(aq)}^- + H_3O^+_{(aq)}$

ICE Table	$HNO_{2(aq)}$	$H_2O_{(l)}$	$NO_{2(aq)}^-$	$H_3O_{(aq)}^+$
Initial	0.150	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.150-x	-	x	x

$K_a=\frac{[NO_2^-][H_3O^+]}{[HNO_2]}=7.2\times 10^{-4}=\frac{x^2}{0.150-x}$

Through a successful approximation:

$x \approx 0.01004\; M=[H_3O^+]$

Thus $pH \approx -\log(0.01004) = 2.0$

When 0.150 mol $LiNO_2$ is added to 1.00 L of a 0.150 M HNO_2, a solution with $[NO_2^-] = 0.150\ M = [HNO_2]$ is formed. To find the change in the pH we use the Henderson-Hasselbalch approximation at equilibrium

$pH=pK_a=-\log(7.2 \times 10^{-4}) =3.14$

Thus the addition has caused pH to increase by 1.14

b) $LiNO_3$ contributes nitrate ions to the solution. Since there is no molecular $HNO_{3(aq)}$ in equilibrium with the hydrogen and nitrate ions, there is no equilibrium to be shifted by the addition of the nitrate ions. Thus, the $[H_3O^+]$ and the pH are not affected by the addition of $LiNO_3$ to the solution. The pH changes are not the same because unlike the first solution, the second solution does not have an equilibrium to shift, just a total ionic strength.

S4

a) The pH of the solution can be determined by the Henderson-Hasselbalch equation.

$pH=pK_a+ log\frac{[ClO_2^-]}{[HClO_2]}=-log(1.1\times 10^{-2}) +log\frac{(\frac{7.25\ mmol}{50.0\ mL})}{(\frac{12.5\ mmol}{50.0\ mL})}$

$pH=1.96-0.24=1.72$

b)

$Ca(OH)_2 \rightarrow Ca^{2+}+2OH^{-}$

$Amount\ of\ added\ OH^- = 1.0\ mmol\ Ca(OH)_2 \times \frac{2OH^-}{Ca(OH)_2}=2.0\ mmol$

The additional OH^- reacts with the chlorous acid and produces chlorite ion.

$HClO_2 +OH^- \rightleftharpoons ClO_2^-+H_2O$

ICE Table	$HClO_2$	OH^-	ClO_2^-	H_2O
Initial	12.5 mmol	2 mmol	7.25 mmol	-
Change	- 2 mmol	- 2 mmol	+ 2 mmol	-
Equilibrium	10.5 mmol	0	9.25 mmol	-

$pH=pK_a + \log (\frac{[ClO_2^-]}{[HClO_2]})$

$=-log(1.1\times 10^{-2})+\log\frac{(\frac{9.25\ mmol}{50.0\ mL})}{(\frac{10.5\ mmol}{50.0\ mL})}$

$=1.96-0.055=1.90$

c)

$Amount\ of\ added\ H_3O^+= 2.00\ mL \times(6\ M)=12\ mmol\ H_3O^+$

The additional H_3O^+ reacts with the chlorite ion to produce chlorous acid.

$ClO_2^-+H_3O^+ \rightleftharpoons HClO_2 + H_2O$

ICE Table	ClO_2^-	H_3O^+	$HClO_2$	H_2O
Initial	7.25 mmol	12 mmol	12.5 mmol	-
Change	- 7.25 mmol	- 7.25 mmol	+ 7.25 mmol	-
Equilibrium	0	4.25 mmol	19.75 mmol	-

The buffer's capacity has been exceeded. The excess strong acid present determines the pH of the solution.

$[H_3O^+] = \frac{4.75 mmol}{50.0\ mL + 2.0\ mL} = 0.091\ M$

$pH=-log\ [H_3O^+]=-log(0.091)= 1.04 $

S5

a) Use the Henderson-Hasselbalch equation to determine the pH of the solution. The volume of the total solution is

$18.0\ mL + 72.0\ mL = 100.0\ mL$

$pKa=14-pKb=14-8.82=5.18$

$[C_5H_5N]=\frac{18.0\ mL\times 0.300M\ C_5H_5N}{100.0\ mL}=\frac{5.4\ mmol}{100.0\ mL}=0.054\ M$

$[C_5H_5NH^+]=\frac{72.0\ mL\times 0.300M}{100.0\ mL}=0.216$

$pH=pKa+log\frac{[C_5 H_5 N]}{[C_5 H_5 NH^+]}=5.18+log\frac{0.054\ M}{0.216\ M}=4.57\approx 4.60$

b) The solution has $[OH^-] = 10^{-9.43} = 3.7 \times 10^{-10}\ M$

The Henderson-Hasselbalch equation assumes that:

$[C_5 H_5 N]\ll1.5 \times 10^{-9}\ll [C_5 H_5 NH^+]$

If the solution is diluted to 500.00 mL, $[C_5H_5N] = 0.0108\ M$, and $[C_5H_5NH^+] = 4.3 \times 10^{-3}\ M$. Only one of the $[C_5H_5N]$ is consistent with this assumption. So the pH would not remain the same.

If the solution was diluted to 1.00ml, $[C_5H_5N] = 0.0054\ M$, and $[C_5H_5NH^+] = 2.2 \times 10^{-3}\ M$. Again, only the $[C_5H_5N]$ is consistent with the assumption, so in this case as well, the pH would not remain the same.

c) The 0.40 mL of added 1.00 HBr does not significantly change the volume of the solution. However, it does add $0.40\ mL \times 1.00\ M\ HBr = 0.40\ mmol\ H_3O^+$. This added H_3O^+ reacts with C_5H_5N, decreasing its amount from $5.4\ mmol\ C_5H_5N$ to $5.0\ mmol\ C_5H_5N$, and increasing the amount of $C_5H_5NH^+$ from $21.6\ mmol$ to $22\ mmol$.

$pH=5.18+log\frac{(\frac{5.00\ mmol\ C_5 H_5 N}{100.40\ mL})}{(\frac{22.00\ mmol C_5 H_5 NH^+}{100.40\ mL})}=4.54$

d) From part c) we see that the total volume of the solution does not affect the pOH of the solution as long as the Henderson-Hasselbalch equation is observed. Let x represent the number of mmols of H_3O^+ added, through $1.00\ M\ HBr$. This increases $C_5H_5NH^+$ and decrease C_5H_5N.

$pH=4.4=5.18+log\frac{5.40-x}{21.6+x}$

$log\frac{5.40-x}{21.6+x}=4.4-5.18=-0.78$

By inverting we get:

$\frac{21.6+x}{5.40-x}=10^{0.78}=6.03$

$21.6+x=6.03(5.4-x)=32.6-6.03x$

$x=\frac{32.6-21.6}{1.00+6.03}=1.6\ mmol\ H_3O^+$

$Vol\ 1.00\ M\ HBr=1.6\ mmol\ H_3O^+ \times \frac{1\ mmol\ HBr}{1\ mmol\ H_3O^+}\times \frac{1\ mL\ soln}{1.00\ mmol\ HBr}=1.6\ mL\ 1.00\ M\ HBr$

S6

Solve for the pK_{HIn} using the usual formula; $pK_{HIn} = log(K_{HIn})$

The pH that gives color change is within 1 pH of this value. For example, Bromphenol Blue is an acidic indicator because, since its pK value is 3.9, it has a color change range of between 2.9 and 4.9, all acidic pH values. Furthermore, Bromthymol Blue changes in neutral solutions, and Thymolphthalein changes in basic solutions.

What is the approximate pH of a solution that turns blue in Bromthymol Blue

As we learned from the first problem, the Brothymol Blue is changes in neutral pH range. So, we can safely say that the pH is somewhere between 6.1 and 8.1

S7

a. $Mg(OH)_2$: This solution is basic, so the acidic indicator will show its basic color, which is blue.
b. HCl: This solution is acidic, it will show its acidic color, yellow
c. HF: This is a weak, but will still show the yellow acidic color

S8

$[Na(OH)_2]\ for\ pH=10.92$

$pOH=14-10.92=3.08$

$[OH^-]=10^{-3.08}=8.32\times 10^{-4}$

$[Na(OH)_2]=\frac{8.32\times 10^{-4}\ mol\ OH^-}{1\ L}\times \frac{1\ mol\ [Na(OH)_2]}{2\ mol\ OH^-}=4.16\times 10^{-4}$

S9

A buffer solution will react with both acids and bases to prevent significant changes in pH. The chlorine is a spectator ion in the ammonium chloride and is therefore irrelevant for the purposes of this question.

$NH_3+H_3O^+\rightleftharpoons NH_4^+ + H_2O$

$NH_4^+ + OH^- \rightleftharpoons NH_3 + H_2O$

This is a buffer solution because based on these reactions that are occur, should either an acid or base be added.

We can use the Henderson-Hasselbach apporximation:

$pK_a=-log(1.8 \times 10^{-5})= 4.745$

$pK_b= 14- pK_a= 9.255$

$pH= 9.255 + log (\frac{0.3}{0.4})= 9.130$

To find the new pH, find the number of moles of each compound using the new volume.

- Moles $NH_3 = 0.3\ M \times 1\ L = 0.3\ mol$
- Moles of $NH_4^+ = 0.4\ M \times 1\ L = 0.4\ mol$
- Moles of $OH^- = 0.2\ M \times 1\ L = 0.2\ mol$

New number of moles using ice table and identifying limiting reactant:

- $0.2\ mol NH_4^+$
- $0\ OH^-$
- $0.1\ mol\ NH_3$

$pH= 9.255 + \log (\frac{0.1}{0.05})= 9.556$

S10

From experiments 1 and 2:

$[HF]+[F^-]= 0.35M$

$5\times (.005\ M) + [F^-] = 0.35M$

$[F^-]=0.325M$

$[HF]=0.025M$

$[H_3O^+]=10-(pK_w-pOH)$

$pOH= -log(0.5-0.325)= 0.757$

$[H_3O^+]=10-(14.94-.757)= 6.56\times 10^{-15}$

$K_a= \frac{[H_3O^+][F^-]}{[HF]}= 8.52989\times 10^{-14}$

$pK_a= 13.07$

S11

The answer is D. Since Na is a spectator ion, CN^- is left to react with the H_3O^+ created from the initial H_2S reaction. This causes the reaction to shift to the right, creating more HS- ions.

S12

Equation	CH_3CH_2COOH	+	$H_2O(l)$	↔	$CH_3CH_2COO^-$	+	H_3O^+
Initial	0.330 M		--		0		0.01M
Change	-x		--		+x		+x
Equilibrium	(0.330 M – x) M		--		x		(0.01 + x) M

$K_a = \frac{x \times (0.01 + x)}{(0.330\ M – x)} \approx \frac{0.01x}{0.330}$

$[CH_3CH_3COO^-] = 4.3 \times 10^{-4}\ M$

$[H_3O^+] = [HI^-] = 0.0104\ M$

$[OH^-] = \frac{K_w}{[H_3O^+]}$

https://chem.libretexts.org/@go/page/499096

$[OH^-] = 9.6 \times 10^{-13}\ M$

S13

$K_b = 1.8 \times 10^{-5}$

Equation	NH_3	+	$H_2O(l)$	↔	NH_4^+	+	OH^-
Initial	0.654 M		---		0.202 M		0
Change	-x		---		+x		+x
Equilibrium	(0.654 M – x) M		---		(0.202+ x) M		x

$K_b = x\times \frac{(0.202 + x)}{(0.654\ M – x)}\approx \frac{0.202x}{0.330}$

$[OH^-] = 2.9\times 10^{-5}\ M$

$[NH_4^+] = [Cl^-] = 0.202\ M$

$[OH^-] = \frac{K_w}{[H_3O^+]}$

$[H_3O^+] = 3.4 \times 10^{-10}\ M$

S14

a)

Equation	HNO_2	+	$H_2O(l)$	↔	HNO_2^-	+	H_3O^+
Initial	0.200 M		--		0		0
Change	-x		--		+x		+x
Equilibrium	(0.200 M – x) M		--		x		x

$K_a = 7.2\times 10^{-4}$

$K_a = \frac{x^2}{(0.200\ M – x)}\approx \frac{x^2}{0.200}$

$x = 0.012\ M$

$pH = -log[H_3O^+]$

$pH = 1.9$

$pH = pK_a – (-log[M])$

$\delta pH = 1.22$

b) Since there is no net change in equilibrium just a shift in ionic strength thus the pH of the second solution will be different.

S15

a)

$pH = pK_a +log \frac{[base]}{[acid]}$

$pH = 3.44$

b)

$OH^- = 0.1\ mol Ba(OH)_2 \times \frac{2\ mol\ OH^-}{ 1\ mol\ Ba(OH)_2} = 0.2\ mol\ OH^-$

Equation	$HCHO_2$	+	OH^-	↔	CHO_2^-	+	$H_2O_{(l)}$
Initial	0.5 mols		0.2 mols		0.250 mols		0

Change	-0.2 mols		-0.2 mols		0.2 mols		0
Equilibrium	0.3 mols		0		0.45 mols		0

$pH = pK_a +log \frac{[base]}{[acid]}$

$pH = 3.92$

c)

$H_3O^+ = 0.01L \times 12M\ HCl = 0.12\ mol\ H_3O^+$

Equation	CHO_2^-	+	H_3O^+	↔	$HCHO_2$	+	$H_2O_{(l)}$
Initial	0.25 mol				0.50 mol		--
Change	-0.25 mol		-0.25 mol		+0.25 mol		--
Equilibrium	0		0.12 mol		0.75 mol		--

$H_3O^+ = \frac{0.12\ mol}{(1L + 0.01 L)} = 0.119\ M$

$pH = 0.92$

S16

a)

$pH = pK_a +log\frac{[base]}{[acid]}$

$[NH_3] = 0.035\ L \times \frac{0.220\ M\ NH_3}{0.1\ L} = 0.077\ M$

$[NH_4^+] = 0.065\times \frac{0.250\ M}{0.1\ L} = 0.163 M$

$pH = 9.00$

b) The pH would change since adding more solution will dilute the concentration and thus reduce the pH to something near $pH = 7$.

c) pH should still be 9 since the amount of acid is so insignificant

S17

a)

Indicator	Changes color in
Bromphenol blue	Acidic
Bromcresol green	Acidic
Bromthymol blue	Neutral
2,4 – Dinitrophenol	Acidic
Chlorphenol red	Acidic
Thymolphthalein	Basic

b) Bromcresol green would be green at $pH = 4.7$

Chlorphenol red would be orange at $pH = 6.0$

S18

1. a) red
2. b) yellow
3. c) yellow

4. d) yellow
5. e) red
6. f) yellow

S19

$Amount\ H_3O^+ = 0.075\ L \times 0.333\ M\times \frac{2\ mol H_3O^+}{1\ mol\ H_2SO_4} = 0.05\ mol$

$Amount\ OH^- = 0.025\ L\times 0.05\ M \times \frac{1\ mol}{1\ mol} = 0.0013 mol$

$[H_3O^+] = \frac{mole\ H_3O^+}{\ L\ solution} = 0.487\ mole$

$pH = -log[H_3O^+]$

$pH = 0.31$

S20

Because the molarities of both solutions is the same all of the base will react with the acid. Depending on the strength of the acid at the equivalence point the amount of moles of acid equals the amount of moles of base therefore the volume doesn't change but the pH will depending on the type of acid or base.

S21

$K_a = \frac{x^2}{0.05}$

$x = [H_3O^+]$

$K_b = \frac{K_w}{K_a} =\frac{x^2}{0.05}$

a)

$[H_3O^+] = 0.019\ M$

$pH = 1.7$

$x = [OH^-] = \sqrt{0.05\times \frac{K_w}{K_a}}= 2.7\times 10^{-7}$

$pOH = 6.57$

$pH = 7.42$

b)

$pH = 8.11$

c)

$pH = 10.2$

S22

The solution should be basic due to the hydrolysis of the sulfide anion of a weak acid.

S23

a)

$H_3PO_{4(aq)}+CO_3^{2-} \rightleftharpoons H_2PO_{4(aq)}^-+HCO_{3(aq)}^-$

$ H_2PO_{4(aq)}^-+ CO_3^{2-} \rightleftharpoons HPO_{4(aq)}^{2-}+ HCO_{3(aq)}^-$

$HPO_{4(aq)}^{2-}+OH^-_{(aq)} \rightleftharpoons PO_{4(aq)}^{3-}+H_2O_{(l)}$

b) The inexpensive base is used in the first 2 ionizations and produces the bulk of the products. In the last ionization $NaOH$ must be used because the solution is not strong enough to pull the last hydrogen ion.

S24

$Malonic\ acid = 104.06 \frac{g}{mol}$

$Moles\ malonic\ acid = 19.5g\times \frac{1\ mol}{104.06\ g} = 0.187\ moles$

$Concentration = \frac{0.187\ moles}{0.250\ L} = 0.784 M$

Equation	H_2A	+	H_2O	↔	HA^-	+	H_3O^+
Initial	0.784 M		-		0		0
Change	-x		-		+x		+x
Equilibrium	(0.784 – x) M		-		x		x

$K_a = \frac{x^2}{(0.784 – x)}$

$pH = 1.47$

$[H_3O^+] = 10^{-1.47} = 0.034\ M$

$K_{a1} = 0.0016$

$K_{a2} = \frac{x^2}{0.300 – x}$

$pH = 4.26$

$[H_3O^+] = 10^{-4.26} = 5.5 \times 10^{-5}\ M$

$K_{a2}= 1\times 10^{-8}$

S25

1. a) No
2. b) No
3. c) Yes
4. d) Yes
5. e) Yes
6. e) No

S26

a)

$\frac{[conjugate base]}{[weak acid]} = \frac{(equil. Amount conj. Base)}{(equil. Amount weak acid)} = \frac{(F \times initial Amt. weak acid)}{(1–F)\times initial Amount weak acid} = \frac{F}{1 – F}$

$pH = pK_a + log\frac{F}{(1 – F)}$

b)

$pH = 10.00 + log \frac{0.27}{(1 – 0.23)} = 9.56$

S27

a. $for\ \ H_2PO_4^- : pH = \frac{1}{2}(pK_{a1} + pK_{a2})$

a. $for\ \ HPO_4^{2-} : pH = \frac{1}{2}(pK_{a2} + pK_{a3})$

S28

a)

$CHO_{2(aq)}^-+H_{3(aq)}O^+ \rightleftharpoons HCHO_{2(aq)}+H_2O_{(l)}$

$HC_2H_3O_{2(aq)} +OH^-_{(aq)} \rightleftharpoons H_2O_{(l)}+ C_2H_3O_{2(aq)}^-$

Since neither species effects the pH significantly we call this solution a buffer solution.

b)

$HCHO_{2(aq)} + H_2O_{(l)} \rightleftharpoons CHO_{2(aq)}^- + H_3O^+_{(aq)}$

$HC_2H_3O_{2(aq)}+H_2O_{(l)} \rightleftharpoons C_2H_3O_{2(aq)}^- + H_3O^+_{(aq)}$

$[Na^+] = 0.250$

$[OH^-] = 0$

$[H_3O^+] = x$

$[C_2H_3O_2^-] = y$

$[CHO_2^-] = z$

$[HC_2H_3O_2] = 0.200 – y$

$[HCHO_2] = 0.250 – z$

$K_A= 1.8\times 10^{-5} = \frac{(xy)}{0.150 – y}$

$K_F = 1.8\times 10^{-4}= \frac{xz}{0.250 – z}$

$y = \frac{0.150\times K_A}{K_A + x}$

$z = \frac{0.250\times K_F}{K_F + x}$

$[H_3O^+] = 4.0\times 10^{-5}\ M$

S29

$H_2O_{2(aq)}+H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)}+HO_{2(aq)}^-$

$[H_2O_2] + [HO_2^-] = 0.259\ M$

$[HO_2^-] = 0.235\ M$

$[H_2O_2] = 0.0242\ M$

$[H_3O^+] = 10^{-(pK_w-pOH)} = 7.7\times 10^{-14}$

$K_a = \frac{(7.7\times 10^{-14})(0.235)}{(0.0242)} = 7.4\times 10^{-13}$

$pK_a = 12.14$

$[H_2O_2] + [HO_2^-] = 0.123\ M$

$[HO_2^-] = 0.1096\ M$

$[H_2O_2] = 0.0134\ M$

$[H_3O^+] = 10^{-(pKw-pOH) }= 7.41\times 10^{-14}$

$K_a = \frac{(7.41\times 10^{-14})(0.1096)}{(0.0134)} = 6.06\times 10^{-13}$

$pK_a = 12.22$

$Average\ pK_a = 12.17$

S30

d, because it is a weak base and will shift the formic acid ionization equilibrium to the right.

S31

$pH=-log [H_3O^+]$

$3.2=-log[H_3O^+]$

$[H_3O^+]=6.31\times 10^{-4}$

S32

$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O_{2(aq)}^- + H_3O_{(aq)}^+$

ICE Table	$HC_2H_3O_{2(aq)}$	$H_2O_{(l)}$	$C_2H_3O_{2(aq)}^-$	$H_3O_{(aq)}^+$
Initial	0.35 M	-	0.62 M	0
Change	-x	-	+x	+x

Equilibrium	0.35-x	-	0.62+x	x

$K_a= 1.8 \times 10^{-5} = \frac{(0.62+x)(x)}{(0.35-x)}$

Assume: $x << 0.62$

$K_a=1.8 \times 10^{-5}=\frac{0.62x}{0.35}$

$x= 1.02 \times 10^{-5}=[H_3O^+]$

Definition of pH

$pH=-log[H_3O^+]$

$pH=-\log(1.02 \times 10^{-5}) = 4.99$

S33

$pH=pKa + log\frac{[base]}{[acid]}$

Solve for molarities:

$HCHO_2$ is 0.227 and $NaCHO2$ is 0.09 fill in were appropriate in formula and…

$pH = 4.14$

S34

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8\times 10^{-5}$

Since $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$,

$[NH_4^+] = [OH^-]$

$Let\ x = [OH^-] = [NH_4^+]$

$K_b = \frac{(x)(x)}{(7.00\times 10^{-2} – x)} = 1.8\times 10^{-5}$

Solving for x using the quadratic formula,

$x = 1.11\times 10^{-3}\ M\ OH^-$

Hence:

$ pOH = -log[OH^-] = 2.95$

$pH = 14 - pOH = 11.05$

So the pH of a 7.00×10^{-2} M solution of ammonia is 11.05

S35

a. $pH = 6.0$
b. $pH = 3.0$
c. $pH = 8.0$

S36

formula: $pH = -log[H_3O^+]$

$HCl \rightarrow H^+ + Cl^-$

a)

$[H^+] = [H_3O^+] = 0.000014\ M$

$pH = -log(0.000014) = 4.85$

So the indicator will be blue.

b)

$pH = -log(.00014) = 3.85$

So the indicator will remain red.

c)

$pH = -log(.0014) = 2.85$

So the indicator will remain red.

d)

$pH = -log(.0000014) = 5.85$

So the indicator will turn blue

S37

An indicator is a weak acid and can therefore alter the pH of the mixture if added in excess amounts

S38

A point at which both acid and base have been consumed and neither in excess

S39

Whether strong or weak acid, the necessary amount of HA is the same regardless of a strong or weak acid, but the pH at the equivalence point is not the same

S40

$4.04 = pK_a + log \frac{0.020}{0.205}$

$pK_a = 5.05$

$K_a = 8.90\times 10^{-6}$

S41

$(150ml)(0.100M)=15\ mmol\ NaOH$

$20-15= \frac{5}{50+200}=0.02M= [NaOH]$

$pOH=-log (0.02) = 1.70$

$pH=14-1.70= 12.3$

S42

Since sodium ions do not affect the pH of a solution, only the iodine ion needs to be considered. The iodide ion is an anion, meaning in an aqueous solution it will take an H^+ ion from water to form HI and OH^-. Therefore, the solution will be basic.

S43

$pH = 7$

S44

$pK_a + pK_b = 14$

$pK_a +4.8=14$

$pK_a = 9.2$

$pH = pK_a + log\frac{[conjugate\ base]}{[acid]}$

$pH = 9.2 + log(1)$

$pHa = 9.2$

S45

A solution can be prepared with equal conjugate base and weak acid but would not form the given pH because the buffer ratio would be 1.00 making the $pH = pK_a = 9.26$.

S46

$HC_2H_3O_{2(aq)}+H_2O_{(l)}\rightleftharpoons C_2H_3O_{2(aq)}^- + H_3O^+_{(aq)}$

ICE Table	$HC_2H_3O_{2(aq)}$	$H_2O_{(l)}$	$C_2H_3O^-_{2(aq)}$	$H_3O^+_{(aq)}$
Initial	0.50M	-	0.50M	0
Change	-x	-	+x	+x
Equilibrium	0.50-x	-	0.50+x	x

$1.8 \times 10^{-5}= \frac{(0.50+x)x}{0.50-x}$

$x=1.8 \times 10^{-5}$

$pH=-\log(1.8 \times 10^{-5})= 4.74$

S47

$NaOH$ is a base, so an acid must be used to neutralize the solution, therefore HBr is the correct substance to use. $NaOH$ and HBr are both strong acids/bases, so the stoichiometric ratio between them is 1:1.

$Moles\ NaOH = (0.325 L)(0.452 M) = 0.147\ moles\ NaOH$

$Volume\ HBr = \frac{0.147\ mol}{5.00\ M} = 0.0280\ L = 28.0\ mL$

S48

$HCN_{(aq)} + H_2O_{(l)}\rightleftharpoons H_3O^+_{(aq)}+ CN^-_{(aq)}$

$6.2\times 10^{-10} = \frac{x^2}{(0.025 – x)}$

assume x<<<0.025, then solve for x

$x = 3.94\times 10^{-6}\ M = [H_3O^+]$

$pH = - log (3.94\times 10^{-6})$

S49

The effects of a solution that consists the same ions.

$PbCl_{2(s)} \rightleftharpoons Pb^+_{2(aq)} + 2Cl^-_{(aq)}$

S50

0.375 M

S51

Using Ice Table

S52

2< pH < 5

S53

Away from equilibrium

S54

Acidic Solution

S55

Yes

S56

12 L

S57

The molarity is the same therefore the amount does not matter.

S58

Start from bottom and the direction of the curve is going up. The equivalence point is around a pH of 7.

S59

0.05 L

S60

Acidic

S61

The ratio of the concentration of one reaction with another reaction of the same substance.

S62

Strong Base

S63

Strong Base

S64

7.5 < pH < 9.0

S65

$CH_3COOH_{(aq)}+H_2O_{(l)}\rightleftharpoons H_3O^+_{(aq)}+CH_3COO^-_{(aq)}$

	CH_3COOH	H_2O	H_3O^+	CH_3COO^-
I	0.153	--	0	0
C	-x	--	+x	+x
E	0.153-x	--	x	x

$K_a=\frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$

$1.8\times 10^{-5}=\frac{x^2}{(0.153-x)}$

$1.8\times 10^{-5}\times (0.153-x)=x^2$

$0=-x^2 -1.8\times 10^{-5}x+2.754\times 10^{-6}$

$x=[H_3O^+]=1.65\times 10^{-3}$

S66

	$HNO_{2(aq)}$	$H_2O_{(l)}$	$H_3O^+_{(aq)}$	$NO^{2-}_{(aq)}$
I	0.548	--	0	0
C	-x	--	+x	+x
E	0.548-x	--	x	x

$K_a=\frac{[H_3O^+][Cl^-]}{[HCl]}= 7.2\times 10^{-4}=\frac{x^2}{0.548-x}$

$x=[H_3O^+]=0.0198$

$pH=-log[H_3O^+]$

$pH=-log(0.0198)$

$ pH=1.70$

S67

	$HClO_{2(aq)}$	$H_2O_{(l)}$	$H_3O^+_{(aq)}$	$ClO_{2(aq)}$
I	0.10	--	0	0
C	-x	--	+x	+x
E	0.10-x	--	x	x

$K_a=\frac{[H_3O^+] [ClO_2^-]}{[HClO_2]}$

$1.1\times 10^{-2}=\frac{x^2}{(0.1-x)}$

$x=[H_3O^+] =0.028119$

$pH=-log[H_3O^+]$

$pH=1.55$

$pH=pK_a=-log(1.1\times 10^{-2})=1.95$

$\Delta pH=1.95-1.55=0.4$

S68

$pH=pK_a+log\frac{[ClO_2^-]}{[HClO_2]}$

$pH=-log (1.1\times 10^{-2})+log\frac{(\frac{6mmol}{55mL})}{(\frac{0.45mmol}{55mL})}$

$pH=1.95+log(13.33)$

$pH=3.08$

S69

$[H_3PO_4]= \frac{(50\ mL\times 0.3\ M)}{75\ mL}= 0.2\ M$

$[H_2PO_4^-]= \frac{(25mL\times 0.3M)}{(25mL+50mL)}=0.1\ M$

$pH=pK_a+log\frac{[H_2PO_4^-]}{[H_3PO^4]}$

$1.82=pK_a+log(\frac{0.1}{0.2})$

$pK_a=2.12 $

$K_a=10^{-2.12}$

$K_a=7.58\times 10^{-3}$

S70

K_{Hln} of 2×10^{-7}. A titration between a strong acid and strong base reach an

endpoint around a neutral pH, therefore the K_{Hln} should be similar to the K_a, 1×10^{-7}. An indicator with a larger K_{Hln}, will be used in acidic solutions. An incidator with a smaller K_{Hln} will be used in basic solutions. Remember, a large K_{Hln} will correspond with a smaller pH, which means more acidic.

S71

Strong acid-weak base. The endpoint is reached in a low pH, because the acid fully dissociated, but the base does not.

S72

$KOH: \frac{(0.05moles)}{L}\times (0.02L)=0.001moles$

$HI: \frac{(0.05mole)}{L}\times (0.01L)=0.0005moles$

$0.001-0.0005=0.0005moles\ KOH\ in\ excess$

$Strong\ acid,\ fully\ dissociates,\ so\ 0.0005\ moles\ of\ OH^-:$

$[OH^-]=\frac{0.0005}{(.01L+0.02L)}=0.0166M$

$pOH=-log(0.0166)=1.77$

$pH=14-1.77=12.22$

$pK_{Hln}=pH\ (at\ endpoint)= 12.22$

This does not give the accurate endpoint of titration, because that should occur around a pH of 7, since it is strong acid-strong base titration.

S73

$HCl: \frac{(0.007moles)}{L}\times (0.200L)=0.0014moles$

$NaOH: \frac{(0.02moles)}{L}\times (0.300L)=0.006moles$

This is strong acid, strong base titration, so they both fully dissociate into ions.

$ 0.006-0.0014=0.0046\ moles\ of\ OH^-\ unreacted$

$[OH^-]= \frac{0.0046}{(0.2L+0.3L)}=0.0092M$

$pOH=-log(0.0092)=2.036$

S74

The volumes will be the same, because they have the same molarity. However, the pH will be different because HF is a stronger acid than CH_3COOH, so its pH will be around or lower 7, while CH_3COOH is a weak acid, so the pH will be greater than 7 (since the strong base, LiOH, has fully dissociated, but the CH_3COOH has not.)

S75

The greater the K_b, the closer the pH will be to 7 at the equivalence point (inflection point of graph), and the steeper the slope will be. However, because it is a weak base, the pH will be lower than 7 no matter what. The lower the K_b, the less steep the slope will be and the lower the pH at the equivalence point will be.

S76

$CsOH: \frac{(0.02\ moles)}{L}\times (0.04L)=0.0008\ moles\ CsOH$

$\frac{0.05mole}{L}\times (0.005L)=0.00025\ moles\ HBr$

$moles\ OH^- = 0.0008-0.00025=0.00055$

$[OH^-]=\frac{0.00055}{(0.04L+0.005L)}=0.0122$

$pOH=1.91$

$pH=14-1.91=12.09$

$\frac{(0.05mole)}{L}\times (0.01L)=0.0005\ moles\ HBr$

$moles\ OH^- = 0.0008-0.0005 = 0.0003$

$[OH^-]=\frac{0.0003}{(0.01L+0.04L)}=0.006\ M$

$pOH=2.22$

$pH=14-2.22=11.79$

$\frac{(0.05mole)}{L}\times (0.025L)=0.0013\ moles\ HBr$

Now more moles of acid than base so,

$moles\ H^+= 0.0013-0.0008 = 0.0005$

$[H^+] = \frac{0.0005}{(.025+0.04)}=0.0076$

$pH=-log(0.0076)=2.13$

S77

It will be basic. The copper is just a spectator ion. The sulfide ion will form a bond with the H from the water. However, because HS is a weak acid, the concentration of OH^- will be greater than H^+ concentration of the dissociation of HS.

S78

Sulfuric acid can be reduced with sodium carbonate, and then further reduced with $NaOH$.

$H_2SO_4 + CO_3^{2-} \rightarrow HSO_4^- +HCO_3^-$

$HSO_4^- + OH^- \rightarrow SO_4^{2-} + H_2O$

S79

$CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^-$

	CH_3COOH	H_2O	H_3O^+	CH_3COO^-
I	0.1	--	0	0
C	-x	--	+x	+x
E	0.1-x	--	x	x

$K_a= \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.8\times 10^{-5}$

$Assume\ x \ll 0.1\ \rightarrow 1.3\times 10^{-3} = [H_3O^+]$

$pH= -log [H_3O^+] = -log[1.3\times 10^{e-3}]$

$pH= 2.89$

S80

$for\ H_2PO_4^-\ :\ pH = \frac{1}{2}(pK_{a1}+pK_{a2}) = \frac{1}{2}(2.15+7.20) = 4.68$

$for HPO_4^{2-}\ :\ pH = \frac{1}{2}(pK_{a2}+pK_{a3}) = \frac{1}{2}(7.20+12.38) = 9.79$

S81

Simultaneously means ionization as an acid and ionization as a base.

a. yes
b. no
c. yes

S82

$K_{a1} = 7.1\times 10^{-3}$, $K_{a2} = 6.3\times 10^{-8}$, $K_{a3} = 4.2\times 10^{-13}$

$pH= pK_a + log \frac{[base]}{[acid]}$

$pK_a= -log(6.3\times 10^{-8})$

$x= 0.3852\ moles\ NaOH\ needed$

$mL = \frac{0.3852\ moles\ NaOH}{6.0M\ NaOH} \times 1000 = 64.2\ mL$

S83

As a result with buffers;

$pH = -log [H_3O^+] =-log[K_a]$

$pH = -log[6.3\times 10^{-5}]= 4.2$

$buffer\ range = 4.2\pm 1= [3.2, 5.2]$

S84

The common ion effect is the suppression of the ionization of a weak electrolyte caused by adding more of an ion that is a product of this ionization.

Add $H_3O^+ : CH_3COOH + H_2O \rightleftharpoons H_3O^+ + CH_3COO^- K_a = 1.8 \times 10^{-5}$

Equilibrium shifts to form more CH_3COOH

Add $OH^- : NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- K_b = 1.8 \times 10^{-5}$

Equilibrium shifts to form more NH_3

S85

$(0.3\ L)(0.56\ M) = 0.168\ mol\ CH_3COO^-$

$(0.3\ L)(0.25\ M) = 0.075\ mol\ CH_3COOH$

	CH_3COO^-	$H3O^+$	H_2O	CH_3COOH
I	0.168 mol	0.006 mol	--	0.075 mol
C	-0.006 mol	-0.006 mol	--	+0.006 mol
E	0.162 mol	0	--	0.018 mol

$[CH_3COO^-] = \frac{0.162\ mol}{0.3\ L} = 0.54\ M$

$[CH_3COOH] = \frac{0.081\ mol}{0.3\ L} = 0.27\ M$

$pH= pK_a + log \frac{[base]}{[acid]}$

$pH= 4.74 + log \frac{[0.54M]}{[0.27M]} = 5.04$

S86

$HCl_{(aq)} + H_2O_{(l)}\rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$

$H_2O_{(l)} + H_2O_{(l)}\rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

	H_2O	H_2O	H_3O^+	OH^-
I	--	--	10^{-8}	0
C	--	--	+x	+x
E	--	--	10^{-8}+x	x

$K_w = (1\times 10^{-8}+x)(x) = 1.0\times 10^{-14}$

Use the quadratic equation to solve for x;

$x= 1.1\times 10^{-7}$

$[H_3O^+]= (1\times 10^{-8}+x)= (1\times 10^{-8}+ 1.1 \times 10^{-7}) = 1.2\times 10^{-7}$

$pH = -log[H_3O^+] = -log[1.2\times 10^{-7}] = 6.92$

S87

Ionization energy is the energy required to completely remove an electron from a gaseous atom or ion. $NaCl$, $NaHCOO$, and $NaHCO_3$ will decrease the ionization of acidic acid.

S88

a)

	HOCl	H_2O	H_3O^+	OCl^-
I	0.45	-	0.001	0
C	- x	-	+ x	+ x
E	0.45 - x	-	0.001 + x	x

$2.9\times 10^{-8} = \frac{(0.001+x)(x)}{(0.45-x)}$

$x= 1.28\times 10^{-5}$

$[H_3O^+] = 0.001+ x = 0.0010128M$

b)

$[H_3O^+][OH^-]=K_w$

$[OH^-]= 9.9\times10^{-14}\ M$

c)

$[OCl^-]=x= 1.28\times 10^{-5}\ M$

d)

$[Cl^-]=[H_3O^+]= 0.0010128M$

S89

a)

	CH_5N	H_2O	OH^-	CH_5NH^+
I	0.3	-	0	0.015
C	-x	-	+x	+x
E	0.3-x	-	x	0.015+x

$1.5\times 10^{-9}= \frac{(0.015+x)(x)}{(0.3-x)}$

$x= 3.0\times 10^{-8}$

$[OH^-]= x=3.0 \times 10^{-8}\ M$

b)

$[C_5H_5NH^+]=0.015M$

c)

$[I^-]=0.015M$

d)

$[H_3O^+][OH^-]= K_w$

$[H_3O^+](1.3\times 10^{-7})=1.0\times 10^{-14}$

$[H_3O^+]= 7.69\times 10^{-8}$

S90

a)

$pH=pK_a+log \frac{[B]}{[A]}$

$pH= 9.21+ log\frac{0.004}{0.025}$

$pH= 7.414$

$pOH=14-pH=6.586$

b)

$pH= pK_a+ log \frac{(0.004+0.025)}{(0.25-0.025)}$

$pH= 8.32$

c)

$(3)(0.076)= 0.228\ moles\ HNO_3$

$\frac{(0.228-0.004)}{(0.25+0.076)}=0.687\ M$

$pH= -log(0.687)= 0.163$

S91

a)

$(0.32L)(0.406M)= 0.13\ moles\ B$

$(0.7L)(0.644M)= 0.451\ moles\ A$

$pH= pK_a+ log\frac{[B]}{[A]}$

$pH= 7.54+log\frac{0.13}{0.451}$

$pH= 7$

b)

Yes, even if you add 1L of H_2O, the mole ratio will not change

c)

$(0.15L)(0.8M)= 0.12\ moles\ NaOH$

$pH= 7.54+ log \frac{(0.13+0.12)}{(0.45-0.12)}$

$pH=7.54$

d)

$7.7=7.54+log\frac{(0.13+x)}{(0.451-x)}$

$x=0.2135\ moles\ of\ NaOH$

$(0.8M)(V)=0.2135\ mol$

$V= 0.2669L\ of\ 0.8M\ NaOH$

S92

a) blue

b) yellow

c) blue

d) yellow

e) blue

f) blue

S93

$(0.73M)(0.025L)= 0.01825\ moles\ HI$

$2[(0.29M)(0.01L)]= 2(0.0029)=0.0058\ moles\ Ca(OH)_2$

$0.01825-0.0058= 0.01245\ mol$

$[H_3O^+]=\frac{0.01245}{0.035}= 0.3557M$

$pH=-log (0.3557)=0.4489$

$14-pH=pOH=13.55$

S94

It doesn't matter because the strength of the base doesn't change the number of moles of that base. At the equivalence point, the amount of moles of acid equals the moles of base regardless of strength. The pOH will not be the same however because when titrating a weak base at the equivalence point, there will be a concentration of its conjugate acid which acts acidic in solution.

S95

It will act basic because PO_4^{3-} is a base in solution because it is the conjugate base to HPO_4^{2-}

S96

Equation	CH_3CH_2COOH	+	$H_2O_{(l)}$	↔	$CH_3CH_2COO^-$	+	H_3O^+
Initial	0.330 M		---		0		0.01M
Change	-x		---		+x		+x
Equilibrium	(0.330 M – x) M		---		x		(0.01 + x) M

$K_a = \frac{x(0.01 + x)}{(0.330\ M – x)} \approx \frac{0.01x}{0.330}$

$[CH_3CH_2COO^-] = 4.3\times 10^{-4}\ M$

$[H_3O^+] = [I^-] = 0.0104\ M$

$[OH^-] = \frac{K_w}{[H_3O^+]}$

$[OH^-] = 9.6\times 10^{-13}\ M$

S97

Equation	NH_3	+	$H_2O(l)$	↔	NH_4^+	+	OH^-
Initial	0.654 M		---		0.202 M		0
Change	-x		---		+x		+x
Equilibrium	(0.654 M – x) M		---		(0.202+ x) M		x

$K_b = \frac{x(0.202 + x)}{(0.654\ M – x)} \approx \frac{0.202x}{0.654}$

$[OH^-] = 2.9\times 10^{-5}\ M$

$[NH_4^+] = [Cl^-] = 0.202\ M$

$[OH^-] = \frac{K_w}{[H_3O^+]}$

$[H_3O^+] = 3.4\times 10^{-10}\ M$

S98

Equation	HNO_2	+	$H_2O_{(l)}$	↔	HNO_2^-	+	H_3O^+
Initial	0.200 M		---		0		0
Change	-x		---		+x		+x

Equilibrium	(0.200 M – x)		---		x		x

$K_a = 7.2\times 10^{-4}$

$K_a = \frac{x^2}{0.200M-x}\approx \frac{x^2}{0.200M}$

$x = 0.012\ M$

$pH = -log[H_3O^+]$

$pH = 1.9$

$pH = pK_a – (-log[M])$

$\Delta pH = 1.22$

b) Since there is no net change in equilibrium just a shift in ionic strength thus the pH of the second solution will be different.

S99

a)

$pH = pK_a +log\frac{[base]}{[acid]}$

$pH = 3.44$

b) $OH^- = 0.1\ mol\ Ba(OH)_2 \times \frac{2\ mol\ OH^-}{1\ mol\ Ba(OH)_2} = 0.2\ mol\ OH^-$

Equation	$HCHO_2$	+	OH^-	↔	CHO_2^-	+	$H_2O_{(l)}$
Initial	0.5 mols		0.2 mols		0.250 mols		---
Change	-.2 mols		-0.2 mols		+0.2 mols		---
Equilibrium	00.3 mols		0		0.45 mols		---

$pH = pK_a +log\frac{[base]}{[acid]}$

$pH = 3.92$

c) $H_3O^+ = 0.01L \times 12M\ HCl = 0.12\ mol\ H_3O^+$

Equation	CHO_2^-	+	H_3O^+	↔	$HCHO_2$	+	$H_2O_{(l)}$
Initial	0.25 mols		0.37 mols		0.50 mols		---
Change	-0.25 mols		-0.25 mols		+0.25 mols		---
Equilibrium	0		0.12 mols		0.75 mols		---

$[H_3O^+] = \frac{0.12\ mol}{1L + 0.01\ L} = 0.119\ M$

$pH = 0.92$

S100

a)

$pH = pK_a +log\frac{[base]}{[acid]}$

$[NH_3] = 0.035\ L \times \frac{0.220\ M\ NH_3}{0.1\ L} = 0.077\ M$

$[NH_4^+] = 0.065 \times \frac{0.250\ M}{0.1\ L} = 0.163\ M$

$pH = 9.00$

b) The pH would change since adding more solution will dilute the concentration and thus reduce the pH to something near $pH = 7$.

c) pH should still be 9 since the amount of acid is so insignificant

S101

a)

$pOH = 14 – 10 = 4$

$[OH^-] = \frac{10^{-4}\ moles}{1L}\times \frac{1\ mol\ Ba(OH)_2}{2\ mol\ OH^-}=0 .00005\ M$

b)

$pH = pK_a + log\frac{[base]}{[acid]}$

$pK_a = 4.74$

$4.5 – 4.74 = log\frac{[base]}{[acid]}$

$10^{-0.24} = \frac{[0.300M]}{[acid]} = 0.52\ M$

S102

$K_b = \frac{(0.500 M)(1.0\times 10^{-7}\ M)}{x\ M} = 1.8\times 10^{-5}$

$[NH_3] = 2.8\times 10^{-3}\ M$

$V = 1L \times (2.8\times 10^{-3}\ M\ NH_3) \times \frac{1\ L}{10.00\ mol\ NH_3} \times \frac{1\ drop}{0.05\ mL} = 3\ drops$

S103

a)

$0.2M\ NH_3 = 0.0072\ mol\ NH_3$

$0.2M\ NH_4Cl = 0.0128\ mol\ NH_4Cl$

$Moles\ of\ base\ in\ excess = 0.0128 - 0.0072 = 0.0056\ mol$

$Total\ volume = 36\ mL + 64\ mL = 100\ mL\ (1L)$

$[OH^-] = \frac{0.0056\ mol}{1\ L} = 0.0056\ M$

$pOH = -log(0.0056) = 2.25$

$12 – 2.25 \rightarrow pH =9.75$

b.)

$[OH^-] = \frac{0.0056\ mol}{200\ L} = 2.8\times 10^{-5}$

No, the pH would be different because you are changing the concentration of the base, and thereby change the pH and pOH values.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Buffers

S104

1.) The pH starts low. 2.) The pH increases slowly until right before the equivalence point. 3.) At equivalence point, the pH spikes by approximately 6 units for only a couple drops of base. 4.) After equivalence point, the pH rises slowly again. 5.) Any acid-base indicator with color change in a pH range of 4-10 will work.

http://chemwiki.ucdavis.edu/Analytical_Chemistry/Quantitative_Analysis/Titration/Titration_Of_A_Strong_Acid_With_A_Strong_Base

S105

$25g\ malonic\ acid\ in\ 0.300\ L = 0.801\ M\ malonic\ acid$

$H_2M + H_2O \rightleftharpoons HM^- + H_3O^+$

$K_{a1} = \frac{[HM^-] [H_3O^+]}{[H_2M]}$

$[H_3O^+] = -log(2.68) = .0021$

$4.5\times 10^{-12} = \frac{[HM^-](0.0021)}{0.801}$

$[HM^-] = 1.72\times 10^{-9}$

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Aqueous_Solutions/The_pH_Scale/Determining_and_Calculating_pH

S106

$K_a = \frac{[H^+][A^-]}{[HA]}$

$log(K_a) = log(\frac{[H^+][A^-]}{[HA]})$

$logK_a = log[H^+] + log(\frac{[A^-]}{[HA]})$

$-pK_a = -pH + log(\frac{[A^-]}{[HA]})$

$pH = pK_a + log(\frac{[A^-]}{[HA]})$

http://chemwiki.ucdavis.edu/Physical_Chemistry/Acids_and_Bases/Buffers/Henderson-Hasselbalch_Approximation

---

This page titled 15.20: Exercises is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

16: Solubility and Complex-Ion Equilibria

16.1: Solubility Product Constant, Ksp
16.2: Relationship Between Solubility and Ksp
16.3: Common-Ion Effect in Solubility Equilibria
16.4: Limitations of the K_{sp} Concept
16.5: Criteria for Precipitation and its Completeness
16.6: Fractional Precipitation
16.7: Solubility and pH
16.8: Equilibria Involving Complex Ions
16.9: Qualitative Cation Analysis
16.10: Exercises

This page titled 16: Solubility and Complex-Ion Equilibria is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

16.1: Solubility Product Constant, Ksp

Learning Objectives

- To calculate the solubility of an ionic compound from its K_{sp}

We begin our discussion of solubility and complexation equilibria—those associated with the formation of complex ions—by developing quantitative methods for describing dissolution and precipitation reactions of ionic compounds in aqueous solution. Just as with acid–base equilibria, we can describe the concentrations of ions in equilibrium with an ionic solid using an equilibrium constant expression.

The Solubility Product

When a slightly soluble ionic compound is added to water, some of it dissolves to form a solution, establishing an equilibrium between the pure solid and a solution of its ions. For the dissolution of calcium phosphate, one of the two main components of kidney stones, the equilibrium can be written as follows, with the solid salt on the left:

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3\,Ca^{2+}(aq) + 2\,PO_4^{3-}(aq) \tag{16.1.1}$$

In Section 16.3 and in more advanced chemistry courses, it was noted that basic anions, such as S^{2-}, PO_4^{3-}, and CO_3^{2-}, react with water to produce OH^- and the corresponding protonated anion. Consequently, their calculated molarities, assuming no protonation in aqueous solution, are only approximate.

The equilibrium constant for the dissolution of a sparingly soluble salt is the **solubility product (K_{sp})** of the salt. Because the concentration of a pure solid such as $Ca_3(PO_4)_2$ is a constant, it does not appear explicitly in the equilibrium constant expression. The equilibrium constant expression for the dissolution of calcium phosphate is therefore

$$K = \frac{[Ca^{2+}]^3[PO_4^{3-}]^2}{[Ca_3(PO_4)_2]} \tag{16.1.2}$$

$$[Ca_3(PO_4)_2]K = K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2 \tag{16.1.3}$$

At 25°C and pH 7.00, Ksp for calcium phosphate is 2.07×10^{-33}, indicating that the concentrations of Ca^{2+} and PO_4^{3-} ions in solution that are in equilibrium with solid calcium phosphate are very low. The values of K_{sp} for some common salts are listed in Table 16.1.1, which shows that the magnitude of K_{sp} varies dramatically for different compounds. Although K_{sp} is not a function of pH in Equations 16.1.2 and 16.1.3, changes in pH can affect the solubility of a compound as discussed later.

Note

As with K, the concentration of a pure solid does not appear explicitly in K_{sp}.

Table 16.1.1: Solubility Products for Selected Ionic Substances at 25°C

Solid	Color	K_{sp}		Solid	Color	K_{sp}
Acetates				**Iodides**		
$Ca(O_2CCH_3)_2{\cdot}3H_2O$	white	4×10^{-3}		Hg_2I_2*	yellow	5.2×10^{-29}
Bromides				PbI_2	yellow	9.8×10^{-9}
AgBr	off-white	5.35×10^{-13}		**Oxalates**		
Hg_2Br_2*	yellow	6.40×10^{-23}		$Ag_2C_2O_4$	white	5.40×10^{-12}
Carbonates				$MgC_2O_4{\cdot}2H_2O$	white	4.83×10^{-6}
$CaCO_3$	white	3.36×10^{-9}		PbC_2O_4	white	4.8×10^{-10}

*These contain the Hg_2^{2+} ion.

Solid	Color	K_{sp}		Solid	Color	K_{sp}
$PbCO_3$	white	7.40×10^{-14}		**Phosphates**		
Chlorides				Ag_3PO_4	white	8.89×10^{-17}
AgCl	white	1.77×10^{-10}		$Sr_3(PO_4)_2$	white	4.0×10^{-28}
Hg_2Cl_2*	white	1.43×10^{-18}		$FePO_4 \cdot 2H_2O$	pink	9.91×10^{-16}
$PbCl_2$	white	1.70×10^{-5}		**Sulfates**		
Chromates				Ag_2SO_4	white	1.20×10^{-5}
$CaCrO_4$	yellow	7.1×10^{-4}		$BaSO_4$	white	1.08×10^{-10}
$PbCrO_4$	yellow	2.8×10^{-13}		$PbSO_4$	white	2.53×10^{-8}
Fluorides				**Sulfides**		
BaF_2	white	1.84×10^{-7}		Ag_2S	black	6.3×10^{-50}
PbF_2	white	3.3×10^{-8}		CdS	yellow	8.0×10^{-27}
Hydroxides				PbS	black	8.0×10^{-28}
$Ca(OH)_2$	white	5.02×10^{-6}		ZnS	white	1.6×10^{-24}
$Cu(OH)_2$	pale blue	1×10^{-14}				
$Mn(OH)_2$	light pink	1.9×10^{-13}				
$Cr(OH)_3$	gray-green	6.3×10^{-31}				
$Fe(OH)_3$	rust red	2.79×10^{-39}				

*These contain the Hg_2^{2+} ion.

Solubility products are determined experimentally by directly measuring either the concentration of one of the component ions or the solubility of the compound in a given amount of water. However, whereas solubility is usually expressed in terms of mass of solute per 100 mL of solvent, K_{sp}, like K, is defined in terms of the molar concentrations of the component ions.

Example 16.1.1

Calcium oxalate monohydrate [$Ca(O_2CCO_2) \cdot H_2O$, also written as $CaC_2O_4 \cdot H_2O$] is a sparingly soluble salt that is the other major component of kidney stones [along with $Ca_3(PO_4)_2$]. Its solubility in water at 25°C is 7.36×10^{-4} g/100 mL. Calculate its K_{sp}.



A color photograph of a kidney stone, 8 mm in length. Kidney stones form from sparingly soluble calcium salts and are largely composed of $Ca(O_2CCO_2) \cdot H_2O$ and $Ca_3(PO_4)_2$. (Public Doman; Robert R. Wal via Wikipedia).

Given: solubility in g/100 mL

Asked for: K_{sp}

Strategy:

A. Write the balanced dissolution equilibrium and the corresponding solubility product expression.
B. Convert the solubility of the salt to moles per liter. From the balanced dissolution equilibrium, determine the equilibrium concentrations of the dissolved solute ions. Substitute these values into the solubility product expression to calculate K_{sp}.

Solution

A We need to write the solubility product expression in terms of the concentrations of the component ions. For calcium oxalate monohydrate, the balanced dissolution equilibrium and the solubility product expression (abbreviating oxalate as ox^{2-}) are as follows:

$$\mathrm{Ca(O_2CCO_2)\cdot H_2O(s) \rightleftharpoons Ca^{2+}(aq) + {^-O_2CCO_2^-}(aq) + H_2O(l)} \quad K_{sp} = [\mathrm{Ca^{2+}}][\mathrm{ox^{2-}}]$$

Neither solid calcium oxalate monohydrate nor water appears in the solubility product expression because their concentrations are essentially constant.

B Next we need to determine $[Ca^{2+}]$ and $[ox^{2-}]$ at equilibrium. We can use the mass of calcium oxalate monohydrate that dissolves in 100 mL of water to calculate the number of moles that dissolve in 100 mL of water. From this we can determine the number of moles that dissolve in 1.00 L of water. For dilute solutions, the density of the solution is nearly the same as that of water, so dissolving the salt in 1.00 L of water gives essentially 1.00 L of solution. Because each 1 mol of dissolved calcium oxalate monohydrate dissociates to produce 1 mol of calcium ions and 1 mol of oxalate ions, we can obtain the equilibrium concentrations that must be inserted into the solubility product expression. The number of moles of calcium oxalate monohydrate that dissolve in 100 mL of water is as follows:

$$\frac{7.36 \times 10^{-4}\ \mathrm{g}}{146.1\ \mathrm{g/mol}} = 5.04 \times 10^{-6}\ \mathrm{mol\ Ca(O_2CCO_2)\cdot H_2O}$$

The number of moles of calcium oxalate monohydrate that dissolve in 1.00 L of the saturated solution is as follows:

$$\left(\frac{5.04 \times 10^{-6}\ \mathrm{mol\ Ca(O_2CCO_2\cdot)H_2O}}{100\ \mathrm{mL}}\right)\left(\frac{1000\ \mathrm{mL}}{1.00\ \mathrm{L}}\right) = 5.04 \times 10^{-5}\ \mathrm{mol/L} = 5.04 \times 10^{-5}\ \mathrm{M}$$

Because of the stoichiometry of the reaction, the concentration of Ca^{2+} and ox^{2-} ions are both 5.04×10^{-5} M. Inserting these values into the solubility product expression,

$$\begin{aligned} K_{sp} &= [Ca^{2+}][ox^{2-}] \\ &= (5.04 \times 10^{-5})(5.04 \times 10^{-5}) \\ &= 2.54 \times 10^{-9} \end{aligned}$$

In our calculation, we have ignored the reaction of the weakly basic anion with water, which tends to make the actual solubility of many salts greater than the calculated value.

? Exercise 16.1.1

One crystalline form of calcium carbonate ($\mathrm{CaCO_3}$) is the mineral sold as “calcite” in mineral and gem shops. The solubility of calcite in water is 0.67 mg/100 mL. Calculate its K_{sp}.

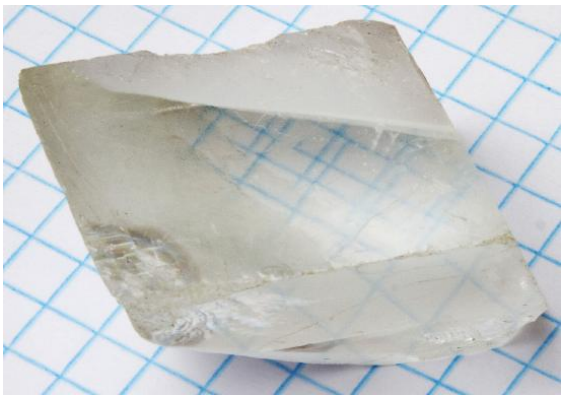
Answer

4.5×10^{-5}

> *The reaction of weakly basic anions with H_2O tends to make the actual solubility of many salts higher than predicted.*

Calcite, a structural material for many organisms, is found in the teeth of sea urchins. The urchins create depressions in limestone that they can settle in by grinding the rock with their teeth. Limestone, however, also consists of calcite, so how can the urchins grind the rock without also grinding their teeth? Researchers have discovered that the teeth are shaped like needles and plates and contain magnesium. The concentration of magnesium increases toward the tip, which contributes to the hardness. Moreover, each

tooth is composed of two blocks of the polycrystalline calcite matrix that are interleaved near the tip. This creates a corrugated surface that presumably increases grinding efficiency. Toolmakers are particularly interested in this approach to grinding.



A crystal of calcite ($CaCO_3$), illustrating the phenomenon of double refraction. When a transparent crystal of calcite is placed over a page, we see two images of the letters. (Cc BY-SA 3.0 Unported; APN MJM via Wikipedia).

Tabulated values of K_{sp} can also be used to estimate the solubility of a salt with a procedure that is essentially the reverse of the one used in Example 16.1.1. In this case, we treat the problem as a typical equilibrium problem and set up a table of initial concentrations, changes in concentration, and final concentrations (ICE Tables), remembering that the concentration of the pure solid is essentially constant.

Example 16.1.2

We saw that the K_{sp} for $Ca_3(PO_4)_2$ is 2.07×10^{-33} at 25°C. Calculate the aqueous solubility of $Ca_3(PO_4)_2$ in terms of the following:

a. the molarity of ions produced in solution
b. the mass of salt that dissolves in 100 mL of water at 25°C

Given: K_{sp}

Asked for: molar concentration and mass of salt that dissolves in 100 mL of water

Strategy:

A. Write the balanced equilibrium equation for the dissolution reaction and construct a table showing the concentrations of the species produced in solution. Insert the appropriate values into the solubility product expression and calculate the molar solubility at 25°C.
B. Calculate the mass of solute in 100 mL of solution from the molar solubility of the salt. Assume that the volume of the solution is the same as the volume of the solvent.

Solution:

A. The dissolution equilibrium for $Ca_3(PO_4)_2$ (Equation 16.1.2) is shown in the following ICE table. Because we are starting with distilled water, the initial concentration of both calcium and phosphate ions is zero. For every 1 mol of $Ca_3(PO_4)_2$ that dissolves, 3 mol of Ca^{2+} and 2 mol of PO_4^{3-} ions are produced in solution. If we let x equal the solubility of $Ca_3(PO_4)_2$ in moles per liter, then the change in $[Ca^{2+}]$ will be $+3x$, and the change in $[PO_4^{3-}]$ will be $+2x$. We can insert these values into the table.

$$\mathbf{Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)}$$

	$Ca_3(PO_4)_2$	$[Ca^{2+}]$	$[PO_4^{3-}]$
initial	pure solid	0	0
change	—	$+3x$	$+2x$
final	pure solid	$3x$	$2x$

Although the amount of solid $Ca_3(PO_4)_2$ changes as some of it dissolves, its molar concentration does not change. We now insert the expressions for the equilibrium concentrations of the ions into the solubility product expression:

$$\begin{aligned} K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2 &= (3x)^3(2x)^2 \\ 2.07 \times 10^{-33} &= 108x^5 \\ 1.92 \times 10^{-35} &= x^5 \\ 1.14 \times 10^{-7}\ M &= x \end{aligned}$$

This is the molar solubility of calcium phosphate at 25°C. However, the molarity of the ions is 2*x* and 3*x*, which means that $[PO_4^{3-}] = 2.28 \times 10^{-7}$ and $[Ca^{2+}] = 3.42 \times 10^{-7}$.

b. To find the mass of solute in 100 mL of solution, we assume that the density of this dilute solution is the same as the density of water because of the low solubility of the salt, so that 100 mL of water gives 100 mL of solution. We can then determine the amount of salt that dissolves in 100 mL of water:

$$\left(\frac{1.14 \times 10^{-7}\ \text{mol}}{1\ \text{L}}\right) 100\ \text{mL} \left(\frac{1\ \text{L}}{1000\ \text{mL}}\right) \left(\frac{310.18\ \text{g}\ Ca_3(PO_4)_2}{1\ \text{mol}}\right) = 3.54 \times 10^{-6}\ \text{g}\ Ca_3(PO_4)_2$$

? Exercise 16.1.2

The solubility product of silver carbonate (Ag_2CO_3) is 8.46×10^{-12} at 25°C. Calculate the following:

a. the molarity of a saturated solution
b. the mass of silver carbonate that will dissolve in 100 mL of water at this temperature

Answer

a. 1.28×10^{-4} M
b. 3.54 mg

The Ion Product

The **ion product (*Q*)** of a salt is the product of the concentrations of the ions in solution raised to the same powers as in the solubility product expression. It is analogous to the reaction quotient (*Q*) discussed for gaseous equilibria. Whereas K_{sp} describes equilibrium concentrations, the ion product describes concentrations that are not necessarily equilibrium concentrations.

> *The ion product Q is analogous to the reaction quotient Q for gaseous equilibria.*

As summarized in Figure 16.1.1, there are three possible conditions for an aqueous solution of an ionic solid:

- $Q < K_{sp}$. The solution is unsaturated, and more of the ionic solid, if available, will dissolve.
- $Q = K_{sp}$. The solution is saturated and at equilibrium.
- $Q > K_{sp}$. The solution is supersaturated, and ionic solid will precipitate.

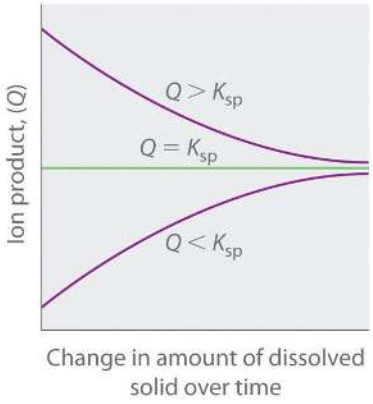


Figure 16.1.1: The Relationship between Q and K_{sp}. If Q is less than K_{sp}, the solution is unsaturated and more solid will dissolve until the system reaches equilibrium (Q = K_{sp}). If Q is greater than K_{sp}, the solution is supersaturated and solid will precipitate until Q = K_{sp}. If Q = K_{sp}, the rate of dissolution is equal to the rate of precipitation; the solution is saturated, and no net change in the amount of dissolved solid will occur.

The process of calculating the value of the ion product and comparing it with the magnitude of the solubility product is a straightforward way to determine whether a solution is unsaturated, saturated, or supersaturated. More important, the ion product tells chemists whether a precipitate will form when solutions of two soluble salts are mixed.

Example 16.1.3

We mentioned that barium sulfate is used in medical imaging of the gastrointestinal tract. Its solubility product is 1.08×10^{-10} at 25°C, so it is ideally suited for this purpose because of its low solubility when a "barium milkshake" is consumed by a patient. The pathway of the sparingly soluble salt can be easily monitored by x-rays. Will barium sulfate precipitate if 10.0 mL of 0.0020 M Na_2SO_4 is added to 100 mL of 3.2×10^{-4} M $BaCl_2$? Recall that NaCl is highly soluble in water.

Given: K_{sp} and volumes and concentrations of reactants

Asked for: whether precipitate will form

Strategy:

A. Write the balanced equilibrium equation for the precipitation reaction and the expression for K_{sp}.
B. Determine the concentrations of all ions in solution when the solutions are mixed and use them to calculate the ion product (Q).
C. Compare the values of Q and K_{sp} to decide whether a precipitate will form.

Solution

A The only slightly soluble salt that can be formed when these two solutions are mixed is $BaSO_4$ because NaCl is highly soluble. The equation for the precipitation of $BaSO_4$ is as follows:

$$BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)}$$

The solubility product expression is as follows:

$$K_{sp} = [Ba^{2+}][SO_4{}^{2-}] = 1.08\times10^{-10}$$

B To solve this problem, we must first calculate the ion product—$Q = [Ba^{2+}][SO_4{}^{2-}]$—using the concentrations of the ions that are present after the solutions are mixed and before any reaction occurs. The concentration of Ba^{2+} when the solutions are mixed is the total number of moles of Ba^{2+} in the original 100 mL of $BaCl_2$ solution divided by the final volume (100 mL + 10.0 mL = 110 mL):

$$\text{moles } Ba^{2+} = 100 \text{ mL}\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right)\left(\frac{3.2 \times 10^{-4} \text{ mol}}{1 \text{ L}}\right) = 3.2 \times 10^{-5} \text{ mol } Ba^{2+}$$

$$[Ba^{2+}] = \left(\frac{3.2 \times 10^{-5} \text{ mol } Ba^{2+}}{110 \text{ mL}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 2.9 \times 10^{-4} \text{ M } Ba^{2+}$$

Similarly, the concentration of SO_4^{2-} after mixing is the total number of moles of SO_4^{2-} in the original 10.0 mL of Na_2SO_4 solution divided by the final volume (110 mL):

$$\text{moles } SO_4^{2-} = 10.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.0020 \text{ mol}}{1 \text{ L}} \right) = 2.0 \times 10^{-5} \text{ mol } SO_4^{2-}$$

$$[SO_4^{2-}] = \left(\frac{2.0 \times 10^{-5} \text{ mol } SO_4^{2-}}{110 \text{ mL}} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 1.8 \times 10^{-4} \text{ M } SO_4^{2-}$$

We can now calculate Q:

$$Q = [Ba^{2+}][SO_4^{2-}] = (2.9\times10^{-4})(1.8\times10^{-4}) = 5.2\times10^{-8}$$

C We now compare Q with the K_{sp}. If $Q > K_{sp}$, then $BaSO_4$ will precipitate, but if $Q < K_{sp}$, it will not. Because $Q > K_{sp}$, we predict that $BaSO_4$ will precipitate when the two solutions are mixed. In fact, $BaSO_4$ will continue to precipitate until the system reaches equilibrium, which occurs when $[Ba^{2+}][SO_4^{2-}] = K_{sp} = 1.08 \times 10^{-10}$.

Exercise 16.1.3

The solubility product of calcium fluoride (CaF_2) is 3.45×10^{-11}. If 2.0 mL of a 0.10 M solution of NaF is added to 128 mL of a 2.0×10^{-5}M solution of $Ca(NO_3)_2$, will CaF_2 precipitate?

Answer

yes ($Q = 4.7 \times 10^{-11} > K_{sp}$)

Summary

The solubility product (K_{sp}) is used to calculate equilibrium concentrations of the ions in solution, whereas the ion product (Q) describes concentrations that are not necessarily at equilibrium. The equilibrium constant for a dissolution reaction, called the solubility product (K_{sp}), is a measure of the solubility of a compound. Whereas solubility is usually expressed in terms of mass of solute per 100 mL of solvent, K_{sp} is defined in terms of the molar concentrations of the component ions. In contrast, the ion product (Q) describes concentrations that are not necessarily equilibrium concentrations. Comparing Q and K_{sp} enables us to determine whether a precipitate will form when solutions of two soluble salts are mixed.

This page titled 16.1: Solubility Product Constant, Ksp is shared under a CC BY-NC-SA 3.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **18.1: Solubility Product Constant, Ksp** is licensed CC BY-NC-SA 3.0.

16.2: Relationship Between Solubility and Ksp

Learning Objectives

- Quantitatively related K_{sp} to solubility

Considering the relation between solubility and K_{sp} is important when describing the solubility of slightly ionic compounds. However, this article discusses ionic compounds that are difficult to dissolve; they are considered "slightly soluble" or "almost insoluble." Solubility product constants (K_{sq}) are given to those solutes, and these constants can be used to find the molar solubility of the compounds that make the solute. This relationship also facilitates finding the K_{sq} of a slightly soluble solute from its solubility.

Introduction

Recall that the definition of *solubility* is the maximum possible concentration of a solute in a solution at a given temperature and pressure. We can determine the solubility product of a slightly soluble solid from that measure of its solubility at a given temperature and pressure, provided that the only significant reaction that occurs when the solid dissolves is its dissociation into solvated ions, that is, the only equilibrium involved is:

$$\mathrm{M}_p\mathrm{X}_q(s) \rightleftharpoons p\mathrm{M}^{\mathrm{m}+}(aq) + q\mathrm{X}^{\mathrm{n}-}(aq) \tag{16.2.1}$$

In this case, we calculate the solubility product by taking the solid's solubility expressed in units of moles per liter (mol/L), known as its **molar solubility**.

Calculation of K_{sp} from Equilibrium Concentrations

We began the chapter with an informal discussion of how the mineral fluorite is formed. Fluorite, $\mathrm{CaF_2}$, is a slightly soluble solid that dissolves according to the equation:

$$\mathrm{CaF_2}(s) \rightleftharpoons \mathrm{Ca^{2+}}(aq) + 2\,\mathrm{F^-}(aq)$$

The concentration of Ca^{2+} in a saturated solution of CaF_2 is 2.1×10^{-4} *M*; therefore, that of F^- is 4.2×10^{-4} *M*, that is, twice the concentration of $\mathrm{Ca^{2+}}$. What is the solubility product of fluorite?

Solution

First, write out the K_{sp} expression, then substitute in concentrations and solve for K_{sp}:

$$\mathrm{CaF_2(s)} \rightleftharpoons \mathrm{Ca^{2+}(aq)} + 2\,\mathrm{F^-(aq)}$$

A saturated solution is a solution at equilibrium with the solid. Thus:

$$\begin{aligned} K_{\mathrm{sp}} &= [\mathrm{Ca^{2+}}][\mathrm{F^-}]^2 \\ &= (2.1 \times 10^{-4})(4.2 \times 10^{-4})^2 \\ &= 3.7 \times 10^{-11} \end{aligned}$$

As with other equilibrium constants, we do not include units with K_{sp}.

Exercise 16.2.1

In a saturated solution that is in contact with solid $Mg(OH)_2$, the concentration of Mg^{2+} is 3.7×10^{-5} *M*. What is the solubility product for $Mg(OH)_2$?

$$\mathrm{Mg(OH)_2}(s) \rightleftharpoons \mathrm{Mg^{2+}}(aq) + 2\,\mathrm{OH^-}(aq)$$

Answer

2.0×10^{-13}

Determination of Molar Solubility from K_{sp}

The K_{sp} of copper(I) bromide, CuBr, is 6.3×10^{-9}. Calculate the molar solubility of copper bromide.

Solution

The solubility product constant of copper(I) bromide is 6.3×10^{-9}.

The reaction is:

$$\mathrm{CuBr}(s) \rightleftharpoons \mathrm{Cu}^{+}(aq) + \mathrm{Br}^{-}(aq)$$

First, write out the solubility product equilibrium constant expression:

$$K_{sp} = [\mathrm{Cu}^{+}][\mathrm{Br}^{-}]$$

Create an ICE table (as introduced in the chapter on fundamental equilibrium concepts), leaving the CuBr column empty as it is a solid and does not contribute to the K_{sp}:

	CuBr ⇌	Cu^+	+ Br^-
Initial concentration (*M*)		0	0
Change (*M*)		*x*	*x*
Equilibrium concentration (*M*)		0 + *x* = *x*	0 + *x* = *x*

At equilibrium:

$$\begin{aligned} K_{sp} &= [\mathrm{Cu}^{+}][\mathrm{Br}^{-}] \\ 6.3 \times 10^{-9} &= (x)(x) = x^2 \\ x &= \sqrt{(6.3 \times 10^{-9})} = 7.9 \times 10^{-5} \end{aligned}$$

Therefore, the molar solubility of CuBr is 7.9×10^{-5} *M*.



Finding the Solubility of a Salt: https://youtu.be/98BuldrICXM

Summary

Solubility is defined as the maximum amount of solute that can be dissolved in a solvent at equilibrium. Equilibrium is the state at which the concentrations of products and reactant are constant after the reaction has taken place. The solubility product constant (K_{sp}) describes the equilibrium between a solid and its constituent ions in a solution. The value of the constant identifies the degree to which the compound can dissociate in water. The higher the K_{sp}, the more soluble the compound is. K_{sq} is defined in terms of activity rather than concentration because it is a measure of a concentration that depends on certain conditions such as temperature, pressure, and composition. It is influenced by surroundings. K_{sp} is used to describe the saturated solution of ionic compounds. (A saturated solution is in a state of equilibrium between the dissolved, dissociated, undissolved solid, and the ionic compound.)

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 16.2: Relationship Between Solubility and Ksp is shared under a CC BY license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **18.2: Relationship Between Solubility and Ksp** by OpenStax is licensed CC BY 4.0.

16.3: Common-Ion Effect in Solubility Equilibria

Learning Objectives

- Recognize common ions from various salts, acids, and bases.
- Calculate concentrations involving common ions.
- Calculate ion concentrations involving chemical equilibrium.

The common-ion effect is used to describe the effect on an equilibrium involving a substance that adds an ion that is a part of the equilibrium.

Introduction

The solubility products K_{sp}'s are equilibrium constants in hetergeneous equilibria (i.e., between two different phases). If several salts are present in a system, they all ionize in the solution. If the salts contain a common cation or anion, these salts contribute to the concentration of the common ion. Contributions from all salts must be included in the calculation of concentration of the common ion. For example, a solution containing sodium chloride and potassium chloride will have the following relationship:

$$[\mathrm{Na^+}] + [\mathrm{K^+}] = [\mathrm{Cl^-}] \tag{16.3.1}$$

Consideration of *charge balance* or *mass balance* or both leads to the same conclusion. The solubility product expression tells us that the equilibrium concentrations of the cation and the anion are inversely related. That is, as the concentration of the anion increases, the maximum concentration of the cation needed for precipitation to occur decreases—and vice versa—so that K_{sp} is constant. Consequently, the solubility of an ionic compound depends on the concentrations of other salts that contain the same ions. Adding a common cation or anion shifts a solubility equilibrium in the direction predicted by Le Chatelier's principle. As a result, the solubility of any sparingly soluble salt is almost always decreased by the presence of a soluble salt that contains a common ion. The exceptions generally involve the formation of complex ions, which is discussed later.

Common Ions

When NaCl and KCl are dissolved in the same solution, the $\mathrm{Cl^-}$ ions are *common* to both salts. In a system containing NaCl and KCl, the $\mathrm{Cl^-}$ ions are common ions.

$$\mathrm{NaCl} \rightleftharpoons \mathrm{Na^+} + \mathrm{Cl^-}$$
$$\mathrm{KCl} \rightleftharpoons \mathrm{K^+} + \mathrm{Cl^-}$$
$$\mathrm{CaCl_2} \rightleftharpoons \mathrm{Ca^{2+}} + 2\mathrm{Cl^-}$$
$$\mathrm{AlCl_3} \rightleftharpoons \mathrm{Al^{3+}} + 3\mathrm{Cl^-}$$
$$\mathrm{AgCl} \rightleftharpoons \mathrm{Ag^+} + \mathrm{Cl^-}$$

For example, when AgCl is dissolved into a solution already containing NaCl (actually $\mathrm{Na^+}$ and $\mathrm{Cl^-}$ ions), the $\mathrm{Cl^-}$ ions come from the ionization of both AgCl and NaCl. Thus, $[\mathrm{Cl^-}]$ differs from $[\mathrm{Ag^+}]$. The following examples show how the concentration of the common ion is calculated.

Example 16.3.1

What are $[\mathrm{Na^+}]$, $[\mathrm{Cl^-}]$, $[\mathrm{Ca^{2+}}]$, and $[\mathrm{H^+}]$ in a solution containing 0.10 M each of NaCl, $\mathrm{CaCl_2}$, and HCl?

Solution

Due to the conservation of ions, we have

$$[\mathrm{Na^+}] = [\mathrm{Ca^{2+}}] = [\mathrm{H^+}] = 0.10\ \mathrm{M}.$$

but

$$\begin{aligned}[\mathrm{Cl^-}] = &\ 0.10\ (\text{due to NaCl}) \\ &+ 0.20\ (\text{due to } \mathrm{CaCl_2}) \\ &+ 0.10\ (\text{due to HCl}) \\ = &\ 0.40\ \mathrm{M}\end{aligned} \tag{16.3.2}$$

Exercise 16.3.1

John poured 10.0 mL of 0.10 M NaCl, 10.0 mL of 0.10 M KOH, and 5.0 mL of 0.20 M HCl solutions together and then he made the total volume to be 100.0 mL. What is $[Cl^-]$ in the final solution?

$$[Cl^-] = \frac{0.1\ M \times 10\ mL + 0.2\ M \times 5.0\ mL}{100.0\ mL} = 0.020\ M$$

Le Châtelier's Principle states that if an equilibrium becomes unbalanced, the reaction will shift to restore the balance. If a common ion is added to a weak acid or weak base equilibrium, then the equilibrium will shift towards the reactants, in this case the weak acid or base.

Example 16.3.2

Consider the lead(II) ion concentration in this *saturated* solution of $PbCl_2$. The balanced reaction is

$$PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$

Defining s as the concentration of dissolved lead(II) chloride, then:

$$[Pb^{2+}] = s$$

$$[Cl^-] = 2s$$

These values can be substituted into the solubility product expression, which can be solved for s:

$$\begin{aligned} K_{sp} &= [Pb^{2+}][Cl^-]^2 \\ &= s \times (2s)^2 \\ 1.7 \times 10^{-5} &= 4s^3 \\ s^3 &= \frac{1.7 \times 10^{-5}}{4} \\ &= 4.25 \times 10^{-6} \\ s &= \sqrt[3]{4.25 \times 10^{-6}} \\ &= 1.62 \times 10^{-2}\ mol\ dm^{-3} \end{aligned}$$

The concentration of lead(II) ions in the solution is 1.62 x 10^{-2} M. Consider what happens if sodium chloride is added to this saturated solution. Sodium chloride shares an ion with lead(II) chloride. The chloride ion is **common** to both of them; this is the origin of the term "common ion effect".

Look at the original equilibrium expression again:

$$PbCl_2\ (s) \rightleftharpoons Pb^{2+}\ (aq) + 2Cl^-\ (aq)$$

What happens to that equilibrium if extra chloride ions are added? According to *Le* Châtelier, the position of equilibrium will shift to counter the change, in this case, by removing the chloride ions by making extra solid lead(II) chloride.

Of course, the concentration of lead(II) ions in the solution is so small that only a tiny proportion of the extra chloride ions can be converted into solid lead(II) chloride. The lead(II) chloride becomes even *less soluble*, and the concentration of lead(II) ions in the solution *decreases*. This type of response occurs with any sparingly soluble substance: it is less soluble in a solution which contains any ion which it has in common. This is the common ion effect.

A Simple Example

If an attempt is made to dissolve some lead(II) chloride in some 0.100 M sodium chloride solution instead of in water, what is the equilibrium concentration of the lead(II) ions this time? As before, define s to be the concentration of the lead(II) ions.

$$[Pb^{2+}] = s \tag{16.3.3}$$

The calculations are different from before. This time the concentration of the chloride ions is governed by the concentration of the sodium chloride solution. The number of ions coming from the lead(II) chloride is going to be tiny compared with the 0.100 M coming from the sodium chloride solution. In calculations like this, it can be assumed that the concentration of the common ion is entirely due to the other solution. This simplifies the calculation.

So we assume:

$$[Cl^-] = 0.100\ M \tag{16.3.4}$$

The rest of the mathematics looks like this:

$$\begin{aligned} K_{sp} &= [Pb^{2+}][Cl^-]^2 \\ &= s \times (0.100)^2 \\ 1.7 \times 10^{-5} &= s \times 0.00100 \end{aligned} \tag{16.3.5}$$

therefore:

$$\begin{aligned} s &= \frac{1.7 \times 10^{-5}}{0.0100} \\ &= 1.7 \times 10^{-3}\ \text{M} \end{aligned} \tag{16.3.6}$$

Finally, compare that value with the simple saturated solution:

Original solution:

$$[Pb^{2+}] = 0.0162\ M \tag{16.3.7}$$

Solution in 0.100 M NaCl solution:

$$[Pb^{2+}] = 0.0017\ M \tag{16.3.8}$$

The concentration of the lead(II) ions has decreased by a factor of about 10. If more concentrated solutions of sodium chloride are used, the solubility decreases further.

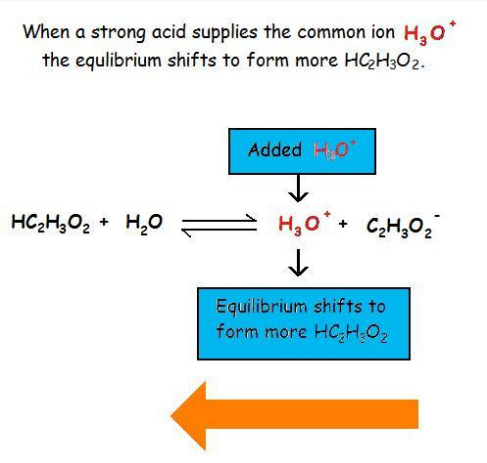
> Adding a common ion to a system at equilibrium affects the equilibrium composition, but **not** the ionization constant.

Common Ion Effect with Weak Acids and Bases

Adding a common ion prevents the weak acid or weak base from ionizing as much as it would without the added common ion. The common ion effect suppresses the ionization of a weak acid by adding more of an ion that is a product of this equilibrium.

Example 16.3.3

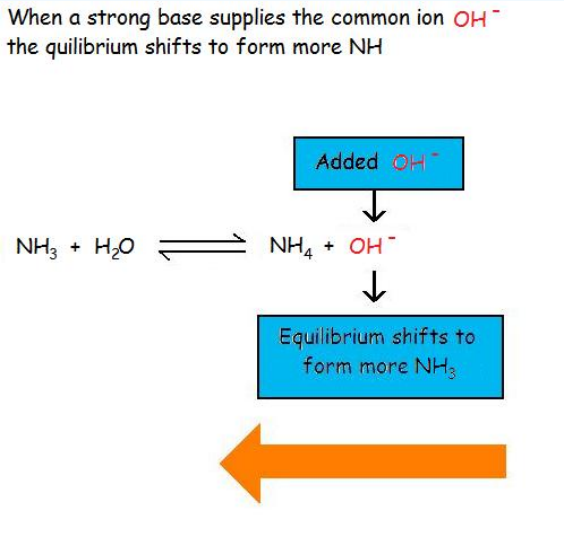
The common ion effect of H_3O^+ on the ionization of acetic acid



The common ion effect suppresses the ionization of a weak base by adding more of an ion that is a product of this equilibrium.

Example 16.3.4

Consider the common ion effect of OH^- on the ionization of ammonia



Adding the common ion of hydroxide shifts the reaction towards the left to decrease the stress (in accordance with Le Châtelier's Principle), forming more reactants. This decreases the reaction quotient, because the reaction is being pushed towards the left to reach equilibrium. The equilibrium constant, $K_b = 1.8 \times 10^{-5}$, does not change. The reaction is put out of balance, or equilibrium.

$$Q_a = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

At first, when more hydroxide is added, the quotient is greater than the equilibrium constant. The reaction then shifts right, causing the denominator to increase, decreasing the reaction quotient and pulling towards equilibrium and causing Q to decrease towards K.

Common Ion Effect on Solubility

Consider, for example, the effect of adding a soluble salt, such as $CaCl_2$, to a saturated solution of calcium phosphate [$Ca_3(PO_4)_2$].

$$\mathrm{Ca_3(PO_4)_2(s) \rightleftharpoons 3\,Ca^{2+}(aq) + 2\,PO_4^{3-}(aq)} \tag{16.3.9}$$

We have seen that the solubility of $Ca_3(PO_4)_2$ in water at 25°C is 1.14×10^{-7} M ($K_{sp} = 2.07 \times 10^{-33}$). Thus a saturated solution of $Ca_3(PO_4)_2$ in water contains

$$3 \times (1.14 \times 10^{-7}\ M) = 3.42 \times 10^{-7}\ M\ \mathrm{Ca^{2+}} \tag{16.3.10}$$

and

$$2 \times (1.14 \times 10^{-7}\ M) = 2.28 \times 10^{-7}\ M\ \mathrm{PO_4^{3-}} \tag{16.3.11}$$

according to the stoichiometry shown in Equation 16.3.9 (neglecting hydrolysis to form HPO_4^{2-}). If $CaCl_2$ is added to a saturated solution of $Ca_3(PO_4)_2$, the Ca^{2+} ion concentration will increase such that [Ca^{2+}] > 3.42×10^{-7} M, making $Q > K_{sp}$. The only way the system can return to equilibrium is for the reaction in Equation 16.3.9 to proceed to the left, resulting in precipitation of $\mathrm{Ca_3(PO_4)_2}$. This will decrease the concentration of both Ca^{2+} and PO_4^{3-} until $Q = K_{sp}$.

> *Adding a common ion decreases solubility, as the reaction shifts toward the left to relieve the stress of the excess product. Adding a common ion to a dissociation reaction causes*

> *the equilibrium to shift left, toward the reactants, causing precipitation.*

Example 16.3.5

Consider the reaction:

$$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$$

What happens to the solubility of $PbCl_2$(s) when 0.1 M NaCl is added?

Solution

$$K_{sp} = 1.7 \times 10^{-5}$$

$$Q_{sp} = 1.8 \times 10^{-5}$$

Identify the common ion: Cl^-

Notice: $Q_{sp} > K_{sp}$ The addition of NaCl has caused the reaction to shift out of equilibrium because there are more dissociated ions. Typically, solving for the molarities requires the assumption that the solubility of $PbCl_2$ is equivalent to the concentration of Pb^{2+} produced because they are in a 1:1 ratio.

Because K_{sp} for the reaction is 1.7×10^{-5}, the overall reaction would be $(s)(2s)^2 = 1.7\times10^{-5}$. Solving the equation for s gives $s = 1.62\times10^{-2}$ M. The coefficient on Cl^- is 2, so it is assumed that twice as much Cl^- is produced as Pb^{2+}, hence the '2s.' The solubility equilibrium constant can be used to solve for the molarities of the ions at equilibrium.

The molarity of Cl^- added would be 0.1 M because Na^+ and Cl^- are in a 1:1 ration in the ionic salt, NaCl. Therefore, the overall molarity of Cl^- would be 2s + 0.1, with 2s referring to the contribution of the chloride ion from the dissociation of lead chloride.

$$\begin{aligned} Q_{sp} &= [Pb^{2+}][Cl^-]^2 \\ 1.8 \times 10^{-5} &= (s)(2s + 0.1)^2 \qquad (16.3.12) \\ s &= [Pb^{2+}] \\ &= 1.8 \times 10^{-3} M \\ 2s &= [Cl^-] \\ &\approx 0.1 M \qquad (16.3.13) \end{aligned}$$

Notice that the molarity of Pb^{2+} is lower when NaCl is added. The equilibrium constant remains the same because of the increased concentration of the chloride ion. To simplify the reaction, it can be assumed that $[Cl^-]$ is approximately 0.1M since the formation of the chloride ion from the dissociation of lead chloride is so small. The reaction quotient for $PbCl_2$ is greater than the equilibrium constant because of the added Cl^-. This therefore shift the reaction left towards equilibrium, causing precipitation and lowering the current solubility of the reaction. Overall, the solubility of the reaction decreases with the added sodium chloride.

> The common ion effect usually decreases the solubility of a sparingly soluble salt.

Example 16.3.6

Calculate the solubility of calcium phosphate [$Ca_3(PO_4)_2$] in 0.20 M $CaCl_2$.

Given: concentration of $CaCl_2$ solution

Asked for: solubility of $Ca_3(PO_4)_2$ in $CaCl_2$ solution

Strategy:

A. Write the balanced equilibrium equation for the dissolution of $Ca_3(PO_4)_2$. Tabulate the concentrations of all species produced in solution.
B. Substitute the appropriate values into the expression for the solubility product and calculate the solubility of $Ca_3(PO_4)_2$.

Solution

A The balanced equilibrium equation is given in the following table. If we let x equal the solubility of $Ca_3(PO_4)_2$ in moles per liter, then the change in $[Ca^{2+}]$ is once again +3x, and the change in $[PO_4^{3-}]$ is +2x. We can insert these values into the ICE table.

$$Ca_3(PO_4)_{2(s)} \rightleftharpoons 3Ca^{2+}_{(aq)} + 2PO^{3-}_{4(aq)} \tag{16.3.14}$$

	$Ca_3(PO_4)_2$	$[Ca^{2+}]$	$[PO_4^{3-}]$
initial	pure solid	0.20	0
change	—	+3x	+2x
final	pure solid	0.20 + 3x	2x

B The K_{sp} expression is as follows:

$$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2 = (0.20 + 3x)^3(2x)^2 = 2.07\times10^{-33}$$

Because $Ca_3(PO_4)_2$ is a sparingly soluble salt, we can reasonably expect that x << 0.20. Thus (0.20 + 3x) M is approximately 0.20 M, which simplifies the K_{sp} expression as follows:

$$\begin{aligned} K_{sp} = (0.20)^3(2x)^2 &= 2.07 \times 10^{-33} \\ x^2 &= 6.5 \times 10^{-32} \\ x &= 2.5 \times 10^{-16}\ \text{M} \end{aligned}$$

This value is the solubility of $Ca_3(PO_4)_2$ in 0.20 M $CaCl_2$ at 25°C. It is approximately nine orders of magnitude less than its solubility in pure water, as we would expect based on Le Chatelier's principle. With one exception, this example is identical to Example 16.3.2—here the initial $[Ca^{2+}]$ was 0.20 M rather than 0.

? Exercise 16.3.4

Calculate the solubility of silver carbonate in a 0.25 M solution of sodium carbonate. The solubility of silver carbonate in pure water is 8.45×10^{-12} at 25°C.

Answer

2.9×10^{-6} M (versus 1.3×10^{-4} M in pure water)



The Common Ion Effect in Solubility Products: https://youtu.be/_P3wozLs0Tc

Summary

Adding a common cation or common anion to a solution of a sparingly soluble salt shifts the solubility equilibrium in the direction predicted by Le Chatelier's principle. The solubility of the salt is almost always **decreased** by the presence of a common ion.

Contributors and Attributions

- Emmellin Tung and Mahtab Danai (UCD)
- Jim Clark (Chemguide.co.uk)
- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)

This page titled 16.3: Common-Ion Effect in Solubility Equilibria is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **18.3: Common-Ion Effect in Solubility Equilibria** is licensed CC BY-NC-SA 4.0.

16.4: Limitations of the K_{sp} Concept

Solubility equilibria defined by a single, simple K_s expression are probably the exception rather than the rule. Such equilibria are often in competition with other reactions with such species as H^+or OH^-, complexing agents, oxidation-reduction, formation of other sparingly soluble species or, in the case of carbonates and sulfites, of gaseous products. The exact treatments of these systems can be extremely complicated, involving the solution of large sets of simultaneous equations. For most practical purposes it is sufficient to recognize the general trends, and to carry out approximate calculations.

The solubility product of an ionic compound describes the concentrations of ions in equilibrium with a solid, but what happens if some of the cations become associated with anions rather than being completely surrounded by solvent? Then predictions of the total solubility of the compound based on the assumption that the solute exists solely as discrete ions would differ substantially from the actual solubility, as would predictions of ionic concentrations. In general, four situations explain why the solubility of a compound may be other than expected: ion pair formation, the incomplete dissociation of molecular solutes, the formation of complex ions, and changes in pH.

The Diverse Ion Effect (The Salt Effect)

A sparingly-soluble salt will be more soluble in a solution that contains non-participating ions. This is just the opposite of the **common ion effect**, and it might at first seem rather counter-intuitive: why would adding more ions of any kind make a salt more likely to dissolve?

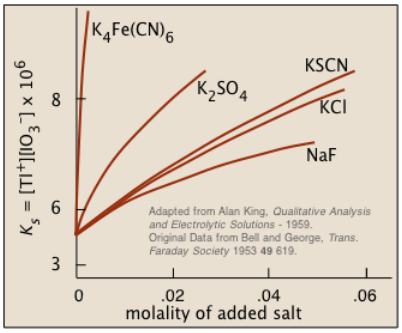


Figure 16.4.1: Solubility of thallium iodate in solutions containing dissolved salts

Although the concentrations of ions in equilibrium with a sparingly soluble solid are so low that they are essentially the same as the activities, the presence of other ions at concentrations of about 0.001 M or greater can materially reduce the activities of the dissolution products, permitting the solubilities to be greater than what simple equilibrium calculations would predict.

> *The Diverse Ion Effect typically has the opposite effect on solubility than the common ion effect.*

The higher the charge of the foreign ion, the more pronounced is the salt effect. This tells us that inter-ionic (and thus electrostatic) interactions must play a role. The details are rather complicated, but the general idea is that all ions in solution, besides possessing tightly-held waters of hydration, tend to attract oppositely-charged ions ("*counter-ions*") around them. This "atmosphere" of counterions is always rather diffuse, but much less so (and more tightly bound) when one or both kinds of ions have greater charges. From a distance, these ion-counterion pairs (or ion pairs) bodies appear to be almost electrically neutral, which keeps them from interacting with each other (as to form a precipitate).

Example 16.4.1

How does the solubility of $AgCl$ change in absence and presence of $NaNO_3$, given that the molar solubility product of $AgCl$ is 1.76×10^{-10}.

Solution

The Uncommon Ion Effect is a "real" effect and requires the activity definition of K_{sp} to address. The exact definition of the solubility constant is:

$$K_{sp} = a\{Ag^+\} \cdot a\{Cl^-\} = [Ag+]\gamma(Ag^+)[Cl^-]\gamma(Cl^-)$$

Where $a\{Ag^+\}$ and $a\{Cl^-\}$ are the activities of these ions in the presence of the $NaNO_3$. The γ values are the are the corresponding activity coefficients for the ions.

- In dilute solutions (i.e. absence of $NaNO_3$\)

$$\gamma(Ag^+) = \gamma(Cl^-) = 1$$

So, the solubility of $AgCl$ is then

$$[Ag^+] = \sqrt{K_{sp}} = 1.33 \times 10^{-5} \tag{16.4.1}$$

- In the presence of $NaNO_3$, the activity coefficients deviate from unity. So

$$[Ag^+] = \frac{\sqrt{K_{sp}}}{\sqrt{\gamma(Ag^+)\gamma(Cl^-)}} \tag{16.4.2}$$

. The denominator has to evaluated from an activity table and is 0.769 so \([Ag^+] = \dfrac{\sqrt{K_{sp}}}{0.769} = 1.73 \times 10^{-5}\].

Hence, in the presence of the $NaNO_3$ soluble salt, the molar solubility of the slightly soluble $AgCl$ salt *increases*. An opposite effect is expected if there were a common ions between the two salts.

Incomplete Dissociation of Solutes into Ions

An **ion pair** consists of a cation and an anion that are in intimate contact in solution, rather than separated by solvent (Figure 16.4.1). The ions in an ion pair are held together by the same attractive electrostatic force in ionic solids. As a result, the ions in an ion pair migrate as a single unit, whose net charge is the sum of the charges on the ions. In many ways, we can view an ion pair as a species intermediate between the ionic solid (in which each ion participates in many cation–anion interactions that hold the ions in a rigid array) and the completely dissociated ions in solution (where each is fully surrounded by water molecules and free to migrate independently).

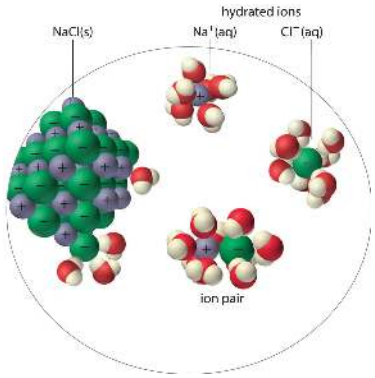


Figure 16.4.2: Ion-Pair Formation. In an ion pair, the cation and the anion are in intimate contact in solution and migrate as a single unit. They are not completely dissociated and individually surrounded by solvent molecules, as are the hydrated ions, which are free to migrate independently.

The overall effect is to reduce the concentrations of the less-shielded ions that are available to combine to form a precipitate. We say that the *thermodynamically-effective concentrations* of these ions are less than their "analytical" concentrations. Chemists refer to these effective concentrations as ionic *activities*, and they denote them by curly brackets $\{Ag^+\}$ as opposed to square brackets $[Ag^+]$ which refer to the nominal or analytical concentrations.

As illustrated for calcium sulfate in the following equation, a second equilibrium must be included to describe the solubility of salts that form ion pairs:

$$\underset{\text{ion pair}}{CaSO_4(s) \rightleftharpoons Ca^{2+} \cdot SO_4^{2-}(aq)} \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq) \tag{16.4.3}$$

The ion pair is represented by the symbols of the individual ions separated by a dot, which indicates that they are associated in solution. The formation of an ion pair is a dynamic process, just like any other equilibrium, so a particular ion pair may exist only briefly before dissociating into the free ions, each of which may later associate briefly with other ions.

Ion-pair formation can have a major effect on the measured solubility of a salt. For example, the measured K_{sp} for calcium sulfate is 4.93×10^{-5} at 25°C. The solubility of $CaSO_4$ should be 7.02×10^{-3} M if the only equilibrium involved were as follows:

$$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO^{2-}_{4(aq)} \tag{16.4.4}$$

In fact, the experimentally measured solubility of calcium sulfate at 25°C is 1.6×10^{-2} M, almost twice the value predicted from its K_{sp}. The reason for the discrepancy is that the concentration of ion pairs in a saturated $CaSO_4$ solution is almost as high as the concentration of the hydrated ions. Recall that the magnitude of attractive electrostatic interactions is greatest for small, highly charged ions. Hence ion pair formation is most important for salts that contain M^{2+} and M^{3+} ions, such as Ca^{2+} and La^{3+}, and is relatively unimportant for salts that contain monopositive cations, except for the smallest, Li^+. We therefore expect a saturated solution of $CaSO_4$ to contain a high concentration of ion pairs and its solubility to be greater than predicted from its K_{sp}.

Simultaneous Equilibria (Side Reactions)

A molecular solute may also be more soluble than predicted by the measured concentrations of ions in solution due to incomplete dissociation. This is particularly common with weak organic acids. Although strong acids (HA) dissociate completely into their constituent ions (H^+ and A^-) in water, weak acids such as carboxylic acids do not ($K_a = 1.5 \times 10^{-5}$). However, the molecular (undissociated) form of a weak acid (HA) is often quite soluble in water; for example, acetic acid (CH_3CO_2H) is completely miscible with water. Many carboxylic acids, however, have only limited solubility in water, such as benzoic acid ($C_6H_5CO_2H$), with $K_a = 6.25 \times 10^{-5}$. Just as with calcium sulfate, we need to include an additional equilibrium to describe the solubility of benzoic acid:

$$C_6H_5CO_2H_{(s)} \rightleftharpoons C_6H_5CO_2H_{(aq)} \rightleftharpoons C_6H_5CO^-_{2(aq)} + H^+_{(aq)} \tag{16.4.5}$$

In a case like this, measuring only the concentration of the ions grossly underestimates the total concentration of the organic acid in solution. In the case of benzoic acid, for example, the pH of a saturated solution at 25°C is 2.85, corresponding to $[H^+] = [C_6H_5CO_2^-] = 1.4 \times 10^{-3}$ M. The total concentration of benzoic acid in the solution, however, is 2.8×10^{-2} M. Thus approximately 95% of the benzoic acid in solution is in the form of hydrated neutral molecules—$C_6H_5CO_2H_{(aq)}$—and only about 5% is present as the dissociated ions (Figure 16.4.3).

> *Incomplete dissociation of a molecular solute that is miscible with water can increase the solubility of the solute.*

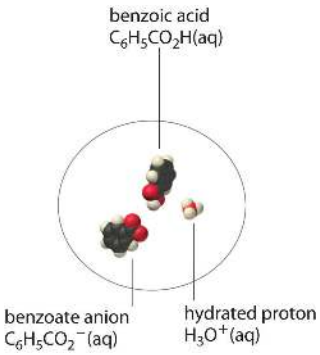


Figure 16.4.3: Incomplete Dissociation of a Molecular Solute. In a saturated solution of benzoic acid in water at 25°C, only about 5% of the dissolved benzoic acid molecules are dissociated to form benzoate anions and hydrated protons. The remaining 95% exists in solution in the form of hydrated neutral molecules. (H_2O molecules are omitted for clarity.)

Although ion pairs, such as $Ca^{2+}{\bullet}SO_4^{2-}$, and undissociated electrolytes, such as $C_6H_5CO_2H$, are both electrically neutral, there is a major difference in the forces responsible for their formation. Simple electrostatic attractive forces between the cation and the anion hold the ion pair together, whereas a polar covalent O−H bond holds together the undissociated electrolyte.

✔ Example 16.4.2: CdI_2 Solubility

The dissolution of cadmium iodide is water is commonly represented as

$$CdI_{2(s)} \rightleftharpoons Cd^{2+} + 2I^- \tag{16.4.6}$$

What you were likely taught about the dissociation of salts in water is wrong! To most students (and to most of their teachers!), this would imply that a 0.1*M* solution of this salt would contain 0.1*M* of $Cd^{2+}(aq)$ — and this would be seriously in error because it fails to take into account that the two ions react with each other.

Firstly, they combine to form neutral, largely-covalent molecular species:

$$Cd^{2+}_{(aq)} + 2I^{-}_{(aq)} \rightarrow CdI_{2(aq)} \tag{16.4.7}$$

This non-ionic form accounts for 78% of the Cd present in the solution! In addition, they form a molecular ion $CdI^{-}_{(aq)}$ according to the following scheme:

Table 16.4.1: CdI2 species in solution

$CdI_{2(s)} \rightleftharpoons Cd^{2+} + 2I^{-}$	$K_1 = 10^{-3.9}$
$Cd^{2+} + I^{-} \rightleftharpoons CdI^{+}$	$K_2 = 10^{+2.3}$
$CdI2_{(s)} \rightleftharpoons CdI^{+} + I^{-}$	$K = 10^{-1.6} = 0.023$

As a consequence, the concentration of "free" Cd^{2+}(aq) in an aqueous cadmium iodide solution is only about 2% of the value you would calculate by taking K_1 as the solubility product. The principal component of such as solution is actually [covalently-bound] CdI_2(aq). It turns out that

Many salts, especially those of metals beyond Group 2, are similarly only partially ionized in aqueous solution:

Table 16.4.2: Incomplete Dissociation of Select Salts

salt	molarity	% cation	other species
KCl	0.52	95	KCl(aq) 5%
$MgSO_4$	0.04	58	$MgSO_4$(aq) 42%
$CaCl_2$	0.44	70	$CaCl^{+}$(aq) 30%
$CuSO_4$	0.045	56	$CuSO_4$(aq) 44%
CdI_2	0.50	2	CdI_2(aq) 76%, CdI^{-}(aq) 22%
$FeCl_3$	0.1	10	$FeCl^{2+}$(aq) 42%, $FeCl_2$(aq) 40%, $FeOH^{2+}$(aq) 6%, $Fe(OH)_2^{+}$(aq) 2%

The data shown Tables 16.4.1 and 16.4.2 are taken from the article *Salts are Mostly NOT Ionized* by Stephen Hawkes: 1996 *J Chem Educ.* 73(5) 421-423. This fact was stated by Arrhenius in 1887, but has been largely ignored and is rarely mentioned in standard textbooks.

Summary

The solubility constant, as with all equilibrium constants, are properly defined in terms of *effective concentrations* (activities) and the use of true concentrations to approximate activities in equilibrium constants can often fail. As with all approximations, it is important to understand the conditions that it will fail. three examples are discussed. The salt effect refers to the fact that the presence of a salt which has no ion in common with the solute, has an effect on the ionic strength of the solution and hence on activity coefficients, so that the equilibrium constant, expressed as a concentration quotient, changes.

Contributors and Attributions

- Wikipedia
- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Yomen Atassi (via StackExchange)

This page titled 16.4: Limitations of the K_{sp} Concept is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **18.4: Limitations of the K_{sp} Concept** is licensed CC BY-NC-SA 4.0.

16.5: Criteria for Precipitation and its Completeness

The equation that describes the equilibrium between solid calcium carbonate and its solvated ions is:

$$\mathrm{CaCO_3}(s) \rightleftharpoons \mathrm{Ca^{2+}}(aq) + \mathrm{CO_3^{2-}}(aq) \tag{16.5.1}$$

We can establish this equilibrium either by adding solid calcium carbonate to water or by mixing a solution that contains calcium ions with a solution that contains carbonate ions. If we add calcium carbonate to water, the solid will dissolve until the concentrations are such that the value of the reaction quotient ($Q = [\mathrm{Ca^{2+}}][\mathrm{CO_3^{2-}}]$) is equal to the solubility product (K_{sp} = 4.8 × 10^{-9}). If we mix a solution of calcium nitrate, which contains Ca^{2+} ions, with a solution of sodium carbonate, which contains $\mathrm{CO_3^{2-}}$ ions, the slightly soluble ionic solid $CaCO_3$ will precipitate, provided that the concentrations of Ca^{2+} and $\mathrm{CO_3^{2-}}$ ions are such that Q is greater than K_{sp} for the mixture. The reaction shifts to the left and the concentrations of the ions are reduced by formation of the solid until the value of Q equals K_{sp}. A saturated solution in equilibrium with the undissolved solid will result. If the concentrations are such that Q is less than K_{sp}, then the solution is not saturated and no precipitate will form.

We can compare numerical values of Q with K_{sp} to predict whether precipitation will occur, as Example 16.5.7 shows. (Note: Since all forms of equilibrium constants are temperature dependent, we will assume a room temperature environment going forward in this chapter unless a different temperature value is explicitly specified.)

The first step in the preparation of magnesium metal is the precipitation of $Mg(OH)_2$ from sea water by the addition of lime, $Ca(OH)_2$, a readily available inexpensive source of OH^- ion:

$$\mathrm{Mg(OH)_2}(s) \rightleftharpoons \mathrm{Mg^{2+}}(aq) + 2\,\mathrm{OH^-}(aq) \quad K_{sp} = 2.1 \times 10^{-13} \tag{16.5.2}$$

The concentration of $Mg^{2+}(aq)$ in sea water is 0.0537 *M*. Will $Mg(OH)_2$ precipitate when enough $Ca(OH)_2$ is added to give a $[OH^-]$ of 0.0010 *M*?

Solution

This problem asks whether the reaction:

$$\mathrm{Mg(OH)_2}(s) \rightleftharpoons \mathrm{Mg^{2+}}(aq) + 2\,\mathrm{OH^-}(aq) \tag{16.5.3}$$

shifts to the left and forms solid $Mg(OH)_2$ when $[Mg^{2+}]$ = 0.0537 *M* and $[OH^-]$ = 0.0010 *M*. The reaction shifts to the left if Q is greater than K_{sp}. Calculation of the reaction quotient under these conditions is shown here:

$$Q = [\mathrm{Mg^{2+}}][\mathrm{OH^-}]^2 = (0.0537)(0.0010)^2 = 5.4 \times 10^{-8} \tag{16.5.4}$$

Because Q is greater than K_{sp} (Q = 5.4 × 10^{-8} is larger than K_{sp} = 2.1 × 10^{-13}), we can expect the reaction to shift to the left and form solid magnesium hydroxide. $Mg(OH)_2(s)$ forms until the concentrations of magnesium ion and hydroxide ion are reduced sufficiently so that the value of Q is equal to K_{sp}.

? Exercise 16.5.1

Use the solubility products in Appendix J to determine whether $CaHPO_4$ will precipitate from a solution with $[Ca^{2+}]$ = 0.0001 *M* and $[\mathrm{HPO_4^{2-}}]$ = 0.001 *M*.

Answer

No precipitation of $CaHPO_4$; Q = 1 × 10^{-7}, which is less than K_{sp}

✔ Precipitation of AgCl upon Mixing Solutions

Does silver chloride precipitate when equal volumes of a 2.0 × 10^{-4}-*M* solution of $AgNO_3$ and a 2.0 × 10^{-4}-*M* solution of NaCl are mixed?

(Note: The solution also contains Na^+ and $\mathrm{NO_3^-}$ ions, but when referring to solubility rules, one can see that sodium nitrate is very soluble and cannot form a precipitate.)

Solution

The equation for the equilibrium between solid silver chloride, silver ion, and chloride ion is:

$$\mathrm{AgCl}(s) \rightleftharpoons \mathrm{Ag}^+(aq) + \mathrm{Cl}^-(aq)$$

The solubility product is 1.8×10^{-10} (see Appendix J).

AgCl will precipitate if the reaction quotient calculated from the concentrations in the mixture of $AgNO_3$ and NaCl is greater than K_{sp}. The volume doubles when we mix equal volumes of $AgNO_3$ and NaCl solutions, so each concentration is reduced to half its initial value. Consequently, immediately upon mixing, $[Ag^+]$ and $[Cl^-]$ are both equal to:

$$\frac{1}{2}(2.0 \times 10^{-4})\ M = 1.0 \times 10^{-4}\ M$$

The reaction quotient, Q, is *momentarily* greater than K_{sp} for AgCl, so a supersaturated solution is formed:

$$Q = [\mathrm{Ag}^+][\mathrm{Cl}^-] = (1.0 \times 10^{-4})(1.0 \times 10^{-4}) = 1.0 \times 10^{-8} > K_{sp}$$

Since supersaturated solutions are unstable, AgCl will precipitate from the mixture until the solution returns to equilibrium, with Q equal to K_{sp}.

Exercise 16.5.2

Will $KClO_4$ precipitate when 20 mL of a 0.050-*M* solution of K^+ is added to 80 mL of a 0.50-*M* solution of ClO_4^-? (Remember to calculate the new concentration of each ion after mixing the solutions before plugging into the reaction quotient expression.)

Answer

No, $Q = 4.0 \times 10^{-3}$, which is less than $K_{sp} = 1.07 \times 10^{-2}$

In the previous two examples, we have seen that $Mg(OH)_2$ or AgCl precipitate when Q is greater than K_{sp}. In general, when a solution of a soluble salt of the M^{m+} ion is mixed with a solution of a soluble salt of the X^{n-} ion, the solid, M_pX_q precipitates if the value of Q for the mixture of M^{m+} and X^{n-} is greater than K_{sp} for M_pX_q. Thus, if we know the concentration of one of the ions of a slightly soluble ionic solid and the value for the solubility product of the solid, then we can calculate the concentration that the other ion must exceed for precipitation to begin. To simplify the calculation, we will assume that precipitation begins when the reaction quotient becomes equal to the solubility product constant.

Precipitation of Calcium Oxalate

Blood will not clot if calcium ions are removed from its plasma. Some blood collection tubes contain salts of the oxalate ion, $C_2O_4^{2-}$, for this purpose (Figure 16.5.4). At sufficiently high concentrations, the calcium and oxalate ions form solid, $CaC_2O_4{\cdot}H_2O$ (which also contains water bound in the solid). The concentration of Ca^{2+} in a sample of blood serum is 2.2×10^{-3} *M*. What concentration of $C_2O_4^{2-}$ ion must be established before $CaC_2O_4{\cdot}H_2O$ begins to precipitate?



Figure 16.5.4: Anticoagulants can be added to blood that will combine with the Ca^{2+} ions in blood serum and prevent the blood from clotting. (credit: modification of work by Neeta Lind)

Solution

The equilibrium expression is:

$$CaC_2O_4(s) \rightleftharpoons Ca^{2+}(aq) + C_2O_4^{2-}(aq) \tag{16.5.5}$$

For this reaction:

$$K_{sp} = [Ca^{2+}][C_2O_4^{2-}] = 2.27 \times 10^{-9} \tag{16.5.6}$$

(see Appendix J)

CaC_2O_4 does not appear in this expression because it is a solid. Water does not appear because it is the solvent.

Solid CaC_2O_4 does not begin to form until Q equals K_{sp}. Because we know K_{sp} and $[Ca^{2+}]$, we can solve for the concentration of $C_2O_4^{2-}$ that is necessary to produce the first trace of solid:

$$Q = K_{sp} = [Ca^{2+}][C_2O_4^{2-}] = 2.27 \times 10^{-9}$$
$$(2.2 \times 10^{-3})[C_2O_4^{2-}] = 2.27 \times 10^{-9}$$
$$[C_2O_4^{2-}] = \frac{2.27 \times 10^{-9}}{2.2 \times 10^{-3}} = 1.0 \times 10^{-6}$$

A concentration of $[C_2O_4^{2-}]$ = 1.0 × 10^{-6} M is necessary to initiate the precipitation of CaC_2O_4 under these conditions.

? Exercise 16.5.3

If a solution contains 0.0020 mol of CrO_4^{2-} per liter, what concentration of Ag^+ ion must be reached by adding solid $AgNO_3$ before Ag_2CrO_4 begins to precipitate? Neglect any increase in volume upon adding the solid silver nitrate.

Answer:

7.0 × 10^{-5} M

It is sometimes useful to know the concentration of an ion that remains in solution after precipitation. We can use the solubility product for this calculation too: If we know the value of K_{sp} and the concentration of one ion in solution, we can calculate the concentration of the second ion remaining in solution. The calculation is of the same type as that in Example—calculation of the concentration of a species in an equilibrium mixture from the concentrations of the other species and the equilibrium constant. However, the concentrations are different; we are calculating concentrations after precipitation is complete, rather than at the start of precipitation.

✓ Concentrations Following Precipitation

Clothing washed in water that has a manganese [$Mn^{2+}(aq)$] concentration exceeding 0.1 mg/L (1.8 × 10^{-6} M) may be stained by the manganese upon oxidation, but the amount of Mn^{2+} in the water can be reduced by adding a base. If a person doing laundry wishes to add a buffer to keep the pH high enough to precipitate the manganese as the hydroxide, $Mn(OH)_2$, what pH is required to keep [Mn^{2+}] equal to 1.8 × 10^{-6} M?

Solution The dissolution of $Mn(OH)_2$ is described by the equation:

$$Mn(OH)_2(s) \rightleftharpoons Mn^{2+}(aq) + 2\,OH^-(aq) \quad K_{sp} = 4.5 \times 10^{-14} \tag{16.5.7}$$

We need to calculate the concentration of OH^- when the concentration of Mn^{2+} is 1.8 × 10^{-6} M. From that, we calculate the pH. At equilibrium:

$$K_{sp} = [Mn^{2+}][OH^-]^2 \tag{16.5.8}$$

or

$$(1.8 \times 10^{-6})[OH^-]^2 = 4.5 \times 10^{-14} \tag{16.5.9}$$

so

$$[OH^-] = 1.6 \times 10^{-4}\ M \tag{16.5.10}$$

Now we calculate the pH from the pOH:

$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.6 \times 10^{-4}) = 3.80$$
$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 3.80 = 10.20$$

If the person doing laundry adds a base, such as the sodium silicate (Na_4SiO_4) in some detergents, to the wash water until the pH is raised to 10.20, the manganese ion will be reduced to a concentration of 1.8×10^{-6} *M*; at that concentration or less, the ion will not stain clothing.

? Exercise 16.5.4

The first step in the preparation of magnesium metal is the precipitation of $Mg(OH)_2$ from sea water by the addition of $Ca(OH)_2$. The concentration of $Mg^{2+}(aq)$ in sea water is 5.37×10^{-2} *M*. Calculate the pH at which $[Mg^{2+}]$ is diminished to 1.0×10^{-5} *M* by the addition of $Ca(OH)_2$.

Answer

11.09

Due to their light sensitivity, mixtures of silver halides are used in fiber optics for medical lasers, in photochromic eyeglass lenses (glass lenses that automatically darken when exposed to sunlight), and—before the advent of digital photography—in photographic film. Even though AgCl ($K_{sp} = 1.6 \times 10^{-10}$), AgBr ($K_{sp} = 7.7 \times 10^{-13}$), and AgI ($K_{sp} = 8.3 \times 10^{-17}$) are each quite insoluble, we cannot prepare a homogeneous solid mixture of them by adding Ag^+ to a solution of Cl^-, Br^-, and I^-; essentially all of the AgI will precipitate before any of the other solid halides form because of its smaller value for K_{sp}. However, we can prepare a homogeneous mixture of the solids by slowly adding a solution of Cl^-, Br^-, and I^- to a solution of Ag^+.

When two anions form slightly soluble compounds with the same cation, or when two cations form slightly soluble compounds with the same anion, the less soluble compound (usually, the compound with the smaller K_{sp}) generally precipitates first when we add a precipitating agent to a solution containing both anions (or both cations). When the K_{sp} values of the two compounds differ by two orders of magnitude or more (e.g., 10^{-2} vs. 10^{-4}), almost all of the less soluble compound precipitates before any of the more soluble one does. This is an example of selective precipitation, where a reagent is added to a solution of dissolved ions causing one of the ions to precipitate out before the rest.



Determining if a Precipitate forms (The Ion Product): https://youtu.be/Naf7PoHPz8Y

Summary

A slightly soluble electrolyte begins to precipitate when the magnitude of the reaction quotient for the dissolution reaction exceeds the magnitude of the solubility product. Precipitation continues until the reaction quotient equals the solubility product. A reagent can be added to a solution of ions to allow one ion to selectively precipitate out of solution. The common ion effect can also play a role in precipitation reactions. In the presence of an ion in common with one of the ions in the solution, Le Chatelier's principle applies and more precipitate comes out of solution so that the molar solubility is reduced.

Glossary

common ion effect
effect on equilibrium when a substance with an ion in common with the dissolved species is added to the solution; causes a decrease in the solubility of an ionic species, or a decrease in the ionization of a weak acid or base

molar solubility
solubility of a compound expressed in units of moles per liter (mol/L)

selective precipitation
process in which ions are separated using differences in their solubility with a given precipitating reagent

solubility product (K_{sp})
equilibrium constant for the dissolution of a slightly soluble electrolyte

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 16.5: Criteria for Precipitation and its Completeness is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **18.5: Criteria for Precipitation and its Completeness** is licensed CC BY-NC-SA 4.0.

16.6: Fractional Precipitation

Learning Objectives

- Calculate ion concentrations to maintain a heterogeneous equilibrium.
- Calculate pH required to precipitate a metal hydroxide.
- Design experiments to separate metal ions in a solution of mixtures of metals.

A mixture of metal ions in a solution can be separated by precipitation with anions such as Cl^-, Br^-, SO_4^{2-}, CO_3^{2-}, S^{2-}, $Cr_2O_4^{2-}$, PO_4^{2-}, OH^- etc. When a metal ion or a group of metal ions form insoluble salts with a particular anion, they can be separated from others by precipitation. We can also separate the anions by precipitating them with appropriate metal ions. There are no definite dividing lines between **insoluble salts, sparingly soluble**, and **soluble salts**, but concentrations of their saturated solutions are small, medium, and large. Solubility products are usually listed for insoluble and sparingly soluble salts, but they are not given for soluble salts. Solubility products for soluble salts are very large.

What type of salts are usually soluble, sparingly soluble and insoluble? The following are some general guidelines, but these are not precise laws.

- All nitrates are soluble. The singly charged large NO_3^- ions form salts with high solubilities. So do ClO_4^-, ClO_3^-, NO_2^-, $HCOO^-$, and CH_3COO^-.
- All chlorides, bromides, and iodides are soluble except those of Ag^+, Hg_2^{2+}, and Pb^{2+}. CaF_2, BaF_2, and PbF_2 are also insoluble.
- All sulfates are soluble, except those of Ba^{2+}, Sr^{2+}, and Pb^{2+}. The doubly charged sulfates are usually less soluble than halides and nitrates.
- Most singly charge cations K^+, Na^+, NH_4^+ form soluble salts. However, $K_3Co(NO_2)_6$ and $(NH_4)_3Co(NO_2)_6$ are insoluble.

These are handy rules for us to have if we deal with salts often. On the other hand, solubility is an important physical property of a substance, and these properties are listed in handbooks.

Chemical Separation of Metal Ions

Formation of crystals from a saturated solution is a **heterogeneous equilibrium** phenomenon, and it can be applied to separate various chemicals or ions in a solution. When solubilities of two metal salts are very different, they can be separated by precipitation. The K_{sp} values for various salts are valuable information, and some data are given in Table E3. In the first two examples, we show how barium and strontium can be separated as chromate.

Example 16.6.1

The K_{sp} for strontium chromate is 3.6×10^{-5} and the K_{sp} for barium chromate is 1.2×10^{-10}. What concentration of potassium chromate will precipitate the maximum amount of either the barium or the strontium chromate from an equimolar 0.30 M solution of barium and strontium ions without precipitating the other?

Solution

Since the K_{sp} for barium chromate is smaller, we know that $BaCrO_4$ will form a precipitate first as $[CrO_4^{2-}]$ increases so that Q_{sp} for $BaCrO_4$ also increases from zero to K_{sp} of $BaCrO_4$, at which point, $BaCrO_4$ precipitates. As $[CrO_4^{2-}]$ increases, $[Ba^{2+}]$ decreases. Further increase of $[CrO_4^{2-}]$ till Q_{sp} for $SrCrO_4$ increases to K_{sp} of $SrCrO_4$; it then precipitates.

Let us write the equilibrium equations and data down to help us think. Let x be the concentration of chromate to precipitate Sr^{2+}, and y be that to precipitate Ba^{2+}:

$$SrCrO_4(s) \rightarrow Sr^{2+}(aq) + CrO_4^{2-}(aq)$$

According to the definition of K_{sp} we have we have $K_{sp} = (0.30)(x) = 3.6 \times 10^{-5}$. Solving for x gives

$$x = \frac{3.6 \times 10^{-5}}{0.30} = 1.2 \times 10^{-4} M$$

Further, let y be the concentration of chromate to precipitate precipitate Ba^{2+}:

$$BaCrO_4(s) \rightarrow Ba^{2+}(aq) + CrO_4^{2-}(aq)$$

with $K_{sp} = (0.30)(y) = 1.2 \times 10^{-10}$. Solving for y gives

$$y = \frac{1.2 \times 10^{-10}}{0.30} = 4.0 \times 10^{-10}\ M$$

The K_{sp}'s for the two salts indicate $BaCrO_4$ to be much less soluble, and it will precipitate before any $SrCrO_4$ precipitates. If chromate concentration is maintained less than $1.2 \times 10^{-4} M$, then all Sr^{2+} ions will remain in the solution.

Discussion

In reality, controling the increase of $[CrO_4^{2-}]$ is very difficult.

Example 16.6.2

The K_{sp} for strontium chromate is 3.6×10^{-5} and the K_{sp} for barium chromate is 1.2×10^{-10}. Potassium chromate is added a small amount at a time to first precipitate $BaCrO_4$. Calculate $[Ba^{2+}]$ when the first trace of $SrCrO_4$ precipitate starts to form in a solution that contains 0.30 M each of Ba^{2+} and Sr^{2+} ions.

Solution

From the solution given in Example 16.6.1, $[CrO_4^{2-}] = 3.6 \times 10^{-4}\ M$ when $SrCrO_4$ starts to form. At this concentration, the $[Ba^{2+}]$ is estimated at $3.6 \times 10^{-4} = 1.2 \times 10^{-10}$.

The K_{sp} of $BaCrO_4$.

Thus,

$$[Ba^{2+}] = 3.33 \times 10^{-7}\ M$$

Very small indeed, compared to 0.30. In the fresh precipitate of $SrCrO_4$, the molar ratio of $SrCrO_4$ to $BaCrO_4$ is

$$\frac{0.30}{3.33 \times 10^{-7}} = 9.0 \times 10^{5}.$$

Hence, the amount of Ba^{2+} ion in the solid is only 1×10^{-6} (i.e., 1 ppm) of all metal ions, providing that all the solid was removed when

$$[CrO_4^{2-}] = 3.6 \times 10^{-4} M.$$

Discussion

The calculation shown here indicates that the separation of Sr and Ba is pretty good. In practice, an impurity level of 1 ppm is a very small value.

Example 16.6.3

What reagent should you use to separate silver and lead ions that are present in a solution? What data or information will be required for this task?

Solution

The K_{sp}'s for salts of silver and lead are required. We list the K_{sp}'s for chlorides and sulfates in a table here. These value are found in the Handbook Menu of our website as Salts K_{sp}.

Solutions to Example 17.6.3

Salt	K_{sp}	Salt	K_{sp}
$AgCl$	1.8×10^{-10}	Ag_2SO_4	1.4×10^{-5}
Hg_2Cl_2	1.3×10^{-18}	$BaSO_4$	1.1×10^{-10}
$PbCl_2$	1.7×10^{-5}	$CaSO_4$	2.4×10^{-5}

Salt	K_{sp}	Salt	K_{sp}
		$PbSO_4$	6.3×10^{-7}
		$SrSO_4$	3.2×10^{-7}

Because the K_{sp}'s $AgCl$ and $PbCl_2$ are very different, chloride, Cl^-, apppears a good choice of negative ions for their separation.

The literature also indicates that $PbCl_2$ is rather soluble in warm water, and by heating the solution to 350 K (80°C), you can keep Pb^{2+} ions in solution and precipitate $AgCl$ as a solid. The solubility of $AgCl$ is very small even at high temperatures.

Discussion

Find more detailed information about the solubility of lead chloride as a function of temperature.

Can sulfate be used to separate silver and lead ions? Which one will form a precipitate first as the sulfate ion concentration increases? What is the $[Pb^{2+}]$ when Ag_2SO_4 begins to precipitate in a solution that contains 0.10 M Ag^+?



The Separation of Two Ions by a Difference in Solubility: The Separation of Two Ions by a Difference in Solubility(opens in new window) [youtu.be]

This page titled 16.6: Fractional Precipitation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **17.6: Precipitation and Separation of Ions** by Chung (Peter) Chieh is licensed CC BY-NC-SA 3.0.

16.7: Solubility and pH

Learning Objectives

- To understand why the solubility of many compounds depends on pH.

The solubility of many compounds depends strongly on the pH of the solution. For example, the anion in many sparingly soluble salts is the conjugate base of a weak acid that may become protonated in solution. In addition, the solubility of simple binary compounds such as oxides and sulfides, both strong bases, is often dependent on pH. In this section, we discuss the relationship between the solubility of these classes of compounds and pH.

The Effect of Acid–Base Equilibria the Solubility of Salts

We begin our discussion by examining the effect of pH on the solubility of a representative salt, $\mathrm{M^+A^-}$, where $\mathrm{A^-}$ is the conjugate base of the weak acid HA. When the salt dissolves in water, the following reaction occurs:

$$\mathrm{MA(s) \rightleftharpoons M^+(aq) + A^-(aq)}$$

with

$$K_{sp} = [\mathrm{M^+}][\mathrm{A^-}]$$

The anion can also react with water in a hydrolysis reaction:

$$\mathrm{A^-(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + HA(aq)} \tag{16.7.1}$$

Because of the reaction described in Equation 16.7.1, the predicted solubility of a sparingly soluble salt that has a basic anion such as S^{2-}, PO_4^{3-}, or CO_3^{2-} is increased. If instead a strong acid is added to the solution, the added H^+ will react essentially completely with A^- to form HA. This reaction decreases $[A^-]$, which decreases the magnitude of the ion product

$$Q = [\mathrm{M^+}][\mathrm{A^-}] \tag{16.7.2}$$

According to Le Chatelier's principle, more MA will dissolve until $Q = K_{sp}$. Hence an acidic pH dramatically increases the solubility of virtually all sparingly soluble salts whose anion is the conjugate base of a weak acid. In contrast, pH has little to no effect on the solubility of salts whose anion is the conjugate base of a stronger weak acid or a strong acid, respectively (e.g., chlorides, bromides, iodides, and sulfates). For example, the hydroxide salt $Mg(OH)_2$ is relatively insoluble in water:

$$Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq) \tag{16.7.3}$$

with

$$K_{sp} = 5.61 \times 10^{-12} \tag{16.7.4}$$

When acid is added to a saturated solution that contains excess solid $Mg(OH)_2$, the following reaction occurs, removing OH^- from solution:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \tag{16.7.5}$$

The overall equation for the reaction of $Mg(OH)_2$ with acid is thus

$$Mg(OH)_{2(s)} + 2H^+(aq) \rightleftharpoons Mg^{2+}(aq) + 2H_2O(l) \tag{16.7.6}$$

As more acid is added to a suspension of $Mg(OH)_2$, the equilibrium shown in Equation 16.7.6 is driven to the right, so more $Mg(OH)_2$ dissolves.

Such pH-dependent solubility is not restricted to salts that contain anions derived from water. For example, CaF_2 is a sparingly soluble salt:

$$CaF_{2(s)} \rightleftharpoons Ca^{2+}(aq) + 2F^-(aq) \tag{16.7.7}$$

with

$$K_{sp} = 3.45 \times 10^{-11} \tag{16.7.8}$$

When strong acid is added to a saturated solution of CaF_2, the following reaction occurs:

$$H^+(aq) + F^-(aq) \rightleftharpoons HF(aq) \tag{16.7.9}$$

Because the forward reaction decreases the fluoride ion concentration, more CaF_2 dissolves to relieve the stress on the system. The net reaction of CaF_2 with strong acid is thus

$$CaF_{2(s)} + 2H^+(aq) \rightarrow Ca^{2+}(aq) + 2HF(aq) \tag{16.7.10}$$

Example 16.7.1 shows how to calculate the solubility effect of adding a strong acid to a solution of a sparingly soluble salt.

> *Sparingly soluble salts derived from weak acids tend to be more soluble in an acidic solution.*

Example 16.7.1

Lead oxalate (PbC_2O_4), lead iodide (PbI_2), and lead sulfate ($PbSO_4$) are all rather insoluble, with K_{sp} values of 4.8×10^{-10}, 9.8×10^{-9}, and 2.53×10^{-8}, respectively. What effect does adding a strong acid, such as perchloric acid, have on their relative solubilities?

Given: K_{sp} values for three compounds

Asked for: relative solubilities in acid solution

Strategy:

Write the balanced chemical equation for the dissolution of each salt. Because the strongest conjugate base will be most affected by the addition of strong acid, determine the relative solubilities from the relative basicity of the anions.

Solution

The solubility Equilibria for the three salts are as follows:

$$PbC_2O_{4(s)} \rightleftharpoons Pb^{2+}(aq) + C_2O^{2-}_{4(aq)}$$

$$PbI_{2(s)} \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$$

$$PbSO_{4(s)} \rightleftharpoons Pb^{2+}(aq) + SO^{2-}_{4(aq)}$$

The addition of a strong acid will have the greatest effect on the solubility of a salt that contains the conjugate base of a weak acid as the anion. Because HI is a strong acid, we predict that adding a strong acid to a saturated solution of PbI_2 will not greatly affect its solubility; the acid will simply dissociate to form H^+(aq) and the corresponding anion. In contrast, oxalate is the fully deprotonated form of oxalic acid (HO_2CCO_2H), which is a weak diprotic acid (pK_{a1} = 1.23 and pK_{a2} = 4.19). Consequently, the oxalate ion has a significant affinity for one proton and a lower affinity for a second proton. Adding a strong acid to a saturated solution of lead oxalate will result in the following reactions:

$$C_2O^{2-}_{4(aq)} + H^+(aq) \rightarrow HO_2CCO^-_{2(aq)}$$

$$HO_2CCO^-_{2(aq)} + H^+(aq) \rightarrow HO_2CCO_2H(aq)$$

These reactions will decrease $[C_2O_4^{2-}]$, causing more lead oxalate to dissolve to relieve the stress on the system. The pK_a of HSO_4^- (1.99) is similar in magnitude to the pK_{a1} of oxalic acid, so adding a strong acid to a saturated solution of $PbSO_4$ will result in the following reaction:

$$SO^{2-}_{4(aq)} + H^+(aq) \rightleftharpoons HSO^-_{4(aq)}$$

Because HSO_4^- has a pKa of 1.99, this reaction will lie largely to the left as written. Consequently, we predict that the effect of added strong acid on the solubility of $PbSO_4$ will be significantly less than for PbC_2O_4.

Exercise 16.7.1

Which of the following insoluble salts—AgCl, Ag_2CO_3, Ag_3PO_4, and/or AgBr—will be substantially more soluble in 1.0 M HNO_3 than in pure water?

Answer

Ag_2CO_3 and Ag_3PO_4



Solubility Products and pH: https://youtu.be/XJ0s5SATZgQ

Caves and their associated pinnacles and spires of stone provide one of the most impressive examples of pH-dependent solubility Equilbria(part (a) in Figure 16.7.1:). Perhaps the most familiar caves are formed from limestone, such as Carlsbad Caverns in New Mexico, Mammoth Cave in Kentucky, and Luray Caverns in Virginia. The primary reactions that are responsible for the formation of limestone caves are as follows:

$$\mathrm{CO_2(aq) + H_2O(l) \rightleftharpoons H^+(aq) + HCO_3^-(aq)} \tag{16.7.11}$$

$$\mathrm{HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)} \tag{16.7.12}$$

$$\mathrm{Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)} \tag{16.7.13}$$

Limestone deposits that form caves consist primarily of $CaCO_3$ from the remains of living creatures such as clams and corals, which used it for making structures such as shells. When a saturated solution of $CaCO_3$ in CO_2-rich water rises toward Earth's surface or is otherwise heated, CO_2 gas is released as the water warms. $CaCO_3$ then precipitates from the solution according to the following equation (part (b) in Figure 16.7.1:):

$$Ca^{2+}(aq) + 2HCO_{3(aq)}^- \rightleftharpoons CaCO_{3(s)} + CO_{2(g)} + H_2O(l) \tag{16.7.14}$$

The forward direction is the same reaction that produces the solid called scale in teapots, coffee makers, water heaters, boilers, and other places where hard water is repeatedly heated.



Figure 16.7.1: The Chemistry of Cave Formation. (a) This cave in Campanet, Mallorca, Spain, and its associated formations are examples of pH-dependent solubility equilibriums. (b) A cave forms when groundwater containing atmospheric CO_2, forming an acidic solution, dissolves limestone ($CaCO_3$) in a process that may take tens of thousands of years. As groundwater seeps into a cave, water evaporates from the solution of $CaCO_3$ in CO_2-rich water, producing a supersaturated solution and a shift in equilibrium that causes precipitation of the $CaCO_3$. The deposited limestone eventually forms stalactites and stalagmites.

When groundwater-containing atmospheric CO_2 (Equations 16.7.11 and 16.7.12) finds its way into microscopic cracks in the limestone deposits, $CaCO_3$ dissolves in the acidic solution in the reverse direction of Equation 16.7.14 The cracks gradually enlarge from 10–50 µm to 5–10 mm, a process that can take as long as 10,000 yr. Eventually, after about another 10,000 yr, a cave forms. Groundwater from the surface seeps into the cave and clings to the ceiling, where the water evaporates and causes the equilibrium in Equation 16.7.14 to shift to the right. A circular layer of solid $CaCO_3$ is deposited, which eventually produces a long, hollow spire of limestone called a stalactite that grows down from the ceiling. Below, where the droplets land when they fall from the ceiling, a similar process causes another spire, called a stalagmite, to grow up. The same processes that carve out hollows below ground are also at work above ground, in some cases producing fantastically convoluted landscapes like that of Yunnan Province in China (Figure 16.7.2).



Figure 16.7.2: Solubility Equilbriain the Formation of Karst Landscapes. Landscapes such as the steep limestone pinnacles of the Stone Forest in Yunnan Province, China, are formed from the same process that produces caves and their associated formations.

Acidic, Basic, and Amphoteric Oxides and Hydroxides

One of the earliest classifications of substances was based on their solubility in acidic versus basic solution, which led to the classification of oxides and hydroxides as being either basic or acidic. **Basic oxides** and hydroxides either react with water to produce a basic solution or dissolve readily in aqueous acid. **Acidic oxides** or hydroxides either react with water to produce an acidic solution or are soluble in aqueous base. There is a clear correlation between the acidic or the basic character of an oxide and the position of the element combined with oxygen in the periodic table. Oxides of metallic elements are generally basic oxides, and oxides of nonmetallic elements are acidic oxides. Compare, for example, the reactions of a typical metal oxide, cesium oxide, and a typical nonmetal oxide, sulfur trioxide, with water:

$$Cs_2O(s) + H_2O(l) \rightarrow 2Cs^+(aq) + 2OH^-(aq) \tag{16.7.15}$$

$$SO_{3(g)} + H_2O(l) \rightarrow H_2SO_{4(aq)} \tag{16.7.16}$$

Cesium oxide reacts with water to produce a basic solution of cesium hydroxide, whereas sulfur trioxide reacts with water to produce a solution of sulfuric acid—very different behaviors indeed

> *Metal oxides generally react with water to produce basic solutions, whereas nonmetal oxides produce acidic solutions.*

The difference in reactivity is due to the difference in bonding in the two kinds of oxides. Because of the low electronegativity of the metals at the far left in the periodic table, their oxides are best viewed as containing discrete M^{n+} cations and O^{2-} anions. At the other end of the spectrum are nonmetal oxides; due to their higher electronegativities, nonmetals form oxides with covalent bonds to oxygen. Because of the high electronegativity of oxygen, however, the covalent bond between oxygen and the other atom, E, is usually polarized: $E^{\delta+}$–$O^{\delta-}$. The atom E in these oxides acts as a Lewis acid that reacts with the oxygen atom of water to produce an oxoacid. Oxides of metals in high oxidation states also tend to be acidic oxides for the same reason: they contain covalent bonds to oxygen. An example of an acidic metal oxide is MoO_3, which is insoluble in both water and acid but dissolves in strong base to give solutions of the molybdate ion (MoO_4^{2-}):

$$MoO_{3(s)} + 2OH^-(aq) \rightarrow MoO^{2-}_{4(aq)} + H_2O(l) \tag{16.7.17}$$

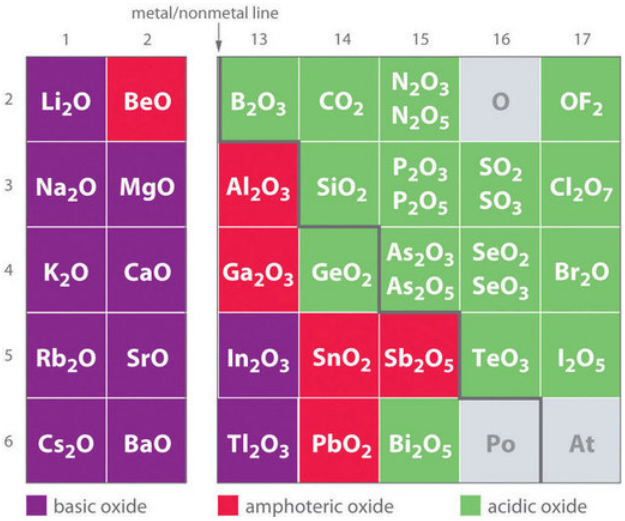


Figure 16.7.3: Classification of the Oxides of the Main Group Elements According to Their Acidic or Basic Character. There is a gradual transition from basic oxides to acidic oxides from the lower left to the upper right in the periodic table. Oxides of metallic elements are generally basic oxides, which either react with water to form a basic solution or dissolve in aqueous acid. In contrast, oxides of nonmetallic elements are acidic oxides, which either react with water to form an acidic solution or are soluble in aqueous base. Oxides of intermediate character, called amphoteric oxides, are located along a diagonal line between the two extremes. Amphoteric oxides either dissolve in acid to produce water or dissolve in base to produce a soluble complex ion. (Radioactive elements are not classified.)

As shown in Figure 16.7.3, there is a gradual transition from basic metal oxides to acidic nonmetal oxides as we go from the lower left to the upper right in the periodic table, with a broad diagonal band of oxides of intermediate character separating the two extremes. Many of the oxides of the elements in this diagonal region of the periodic table are soluble in both acidic and basic solutions; consequently, they are called **amphoteric oxides** (from the Greek ampho, meaning “both,” as in amphiprotic).

Amphoteric oxides either dissolve in acid to produce water or dissolve in base to produce a soluble complex. As shown in Video 16.7.1, for example, mixing the amphoteric oxide $Cr(OH)_3$ (also written as Cr_2O_3•$3H_2O$) with water gives a muddy, purple-brown suspension. Adding acid causes the $Cr(OH)_3$ to dissolve to give a bright violet solution of $Cr^{3+}(aq)$, which contains the $[Cr(H_2O)_6]^{3+}$ ion, whereas adding strong base gives a green solution of the $[Cr(OH)_4]^-$ ion. The chemical equations for the reactions are as follows:

$$\underset{\text{violet}}{Cr(OH)_3(s) + 3H^+(aq) \rightarrow Cr^{3+}(aq)} + 3H_2O(l) \tag{16.7.18}$$

$$Cr(OH)_3(s) + OH^-(aq) \rightarrow \underset{\text{green}}{[Cr(OH)_4]^-(aq)} \tag{16.7.19}$$



Video* 16.7.1*: *Chromium(III) Hydroxide [$Cr(OH)_3$ or Cr_2O_3•$3H_2O$] is an Example of an Amphoteric Oxide. All three beakers originally contained a suspension of brownish purple $Cr(OH)_3(s)$ (center). When concentrated acid (6 M H_2SO_4) was added to the beaker on the left, $Cr(OH)_3$ dissolved to produce violet $[Cr(H_2O)_6]^{3+}$ ions and water. The addition of concentrated base (6 M NaOH) to the beaker on the right caused $Cr(OH)_3$ to dissolve, producing green $[Cr(OH)_4]^-$ions. For a more complete description, see https://www.youtube.com/watch?v=IQNcLH6OZK0*

Example 16.7.2

Aluminum hydroxide, written as either $Al(OH)_3$ or Al_2O_3•$3H_2O$, is amphoteric. Write chemical equations to describe the dissolution of aluminum hydroxide in (a) acid and (b) base.

Given: amphoteric compound

Asked for: dissolution reactions in acid and base

Strategy:

Using Equations 16.7.18and 16.7.19as a guide, write the dissolution reactions in acid and base solutions.

Solution

a. An acid donates protons to hydroxide to give water and the hydrated metal ion, so aluminum hydroxide, which contains three OH^- ions per Al, needs three H^+ ions:

$$Al(OH)_{3(s)} + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$$

In aqueous solution, Al^{3+} forms the complex ion $[Al(H_2O)_6]^{3+}$.

b. In basic solution, OH^- is added to the compound to produce a soluble and stable poly(hydroxo) complex:

$$Al(OH)_{3(s)} + OH^-(aq) \rightarrow [Al(OH)_4]^-(aq)$$

Exercise 16.7.2

Copper(II) hydroxide, written as either $Cu(OH)_2$ or $CuO•H_2O$, is *amphoteric*. Write chemical equations that describe the dissolution of cupric hydroxide both in an acid and in a base.

Answer

$$Cu(OH)_{2(s)} + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2H_2O(l)$$

$$Cu(OH)_{2(s)} + 2OH^-(aq) \rightarrow [Cu(OH)_4]^2_{-(aq)}$$

Selective Precipitation Using pH

Many dissolved metal ions can be separated by the selective precipitation of the cations from solution under specific conditions. In this technique, pH is often used to control the concentration of the anion in solution, which controls which cations precipitate.

> *The concentration of anions in solution can often be controlled by adjusting the pH, thereby allowing the selective precipitation of cations.*

Suppose, for example, we have a solution that contains 1.0 mM Zn^{2+} and 1.0 mM Cd^{2+} and want to separate the two metals by selective precipitation as the insoluble sulfide salts, ZnS and CdS. The relevant solubility equilbria can be written as follows:

$$ZnS(s) \rightleftharpoons Zn^{2+}(aq) + S^{2-}(aq) \tag{16.7.20}$$

with

$$K_{sp} = 1.6 \times 10^{-24} \tag{16.7.21}$$

and

$$CdS(s) \rightleftharpoons Cd^{2+}(aq) + S^{2-}(aq) \tag{16.7.22}$$

with

$$K_{sp} = 8.0 \times 10^{-27} \tag{16.7.23}$$

Because the S^{2-} ion is quite basic and reacts extensively with water to give HS^- and OH^-, the solubility equilbria are more accurately written as $MS(s) \rightleftharpoons M^{2+}(aq) + HS^-(aq) + OH^-$ rather than $MS(s) \rightleftharpoons M^{2+}(aq) + S^{2-}(aq)$. Here we use the simpler form involving S^{2-}, which is justified because we take the reaction of S^{2-} with water into account later in the solution, arriving at the same answer using either equilibrium equation.

The sulfide concentrations needed to cause ZnS and CdS to precipitate are as follows:

$$K_{sp} = [Zn^{2+}][S^{2-}] \tag{16.7.24}$$

$$1.6 \times 10^{-24} = (0.0010\ M)[S^{2-}] \tag{16.7.25}$$

$$1.6 \times 10^{-21}\ M = [S^{2-}] \tag{16.7.26}$$

and

$$K_{sp} = [Cd^{2+}][S^{2-}] \tag{16.7.27}$$

$$8.0 \times 10^{-27} = (0.0010\ M)[S^{2-}] \tag{16.7.28}$$

$$8.0 \times 10^{-24}\ M = [S^{2-}] \tag{16.7.29}$$

Thus sulfide concentrations between 1.6×10^{-21} M and 8.0×10^{-24} M will precipitate CdS from solution but not ZnS. How do we obtain such low concentrations of sulfide? A saturated aqueous solution of H_2S contains 0.10 M H_2S at 20°C. The $pK_{a1}$for H_2S is 6.97, and pK_{a2} corresponding to the formation of $[S^{2-}]$ is 12.90. The equations for these reactions are as follows:

$$H_2S(aq) \rightleftharpoons H^+(aq) + HS^-(aq) \tag{16.7.30}$$

with

$$pK_{a1} = 6.97 \text{ and hence } K_{a1} = 1.1 \times 10^{-7} \tag{16.7.31}$$

$$HS^-(aq) \rightleftharpoons H^+(aq) + S^{2-}(aq) \tag{16.7.32}$$

with

$$pK_{a2} = 12.90 \text{ and hence } K_{a2} = 1.3 \times 10^{-13} \tag{16.7.33}$$

We can show that the concentration of S^{2-} is 1.3×10^{-13} by comparing K_{a1} and K_{a2} and recognizing that the contribution to $[H^+]$ from the dissociation of HS^- is negligible compared with $[H^+]$ from the dissociation of H_2S. Thus substituting 0.10 M in the equation for K_{a1} for the concentration of H_2S, which is essentially constant regardless of the pH, gives the following:

$$K_{a1} = 1.1 \times 10^{-7} = \frac{[H^+][HS^-]}{[H_2S]} = \frac{x^2}{0.10\ M} \tag{16.7.34}$$

$$x = 1.1 \times 10^{-4}\ M = [H^+] = [HS^-]$$

Substituting this value for $[H^+]$ and $[HS^-]$ into the equation for K_{a2},

$$K_{a2} = 1.3 \times 10^{-13} = \frac{[H^+][S^{2-}]}{[HS^-]} = \frac{(1.1 \times 10^{-4}\ M)x}{1.1 \times 10^{-4}\ M} = x = [S^{2-}] \tag{16.7.35}$$

Although $[S^{2-}]$ in an H_2S solution is very low (1.3×10^{-13} M), bubbling H_2S through the solution until it is saturated would precipitate both metal ions because the concentration of S^{2-} would then be much greater than 1.6×10^{-21} M. Thus we must adjust $[S^{2-}]$ to stay within the desired range. The most direct way to do this is to adjust $[H^+]$ by adding acid to the H_2S solution (recall Le Chatelier's principle), thereby driving the equilibrium in Equation 16.7.33 to the left. The overall equation for the dissociation of H_2S is as follows:

$$H_2S(aq) \rightleftharpoons 2H^+(aq) + S^{2-}(aq) \tag{16.7.36}$$

Now we can use the equilibrium constant K for the overall reaction, which is the product of K_{a1} and K_{a2}, and the concentration of H_2S in a saturated solution to calculate the H^+ concentration needed to produce $[S^{2-}]$ of 1.6×10^{-21} M:

$$K = K_{a1}K_{a2} = (1.1 \times 10^{-7})(1.3 \times 10^{-13}) = 1.4 \times 10^{-20} = \frac{[H^+]^2[S^{2-}]}{[H_2S]} \tag{16.7.37}$$

$$[H^+]^2 = \frac{K[H_2S]}{[S^{2-}]} = \frac{(1.4 \times 10^{-20})(0.10\ M)}{1.6 \times 10^{-21}\ M} = 0.88 \tag{16.7.38}$$

$$[H^+] = 0.94 \tag{16.7.39}$$

Thus adding a strong acid such as HCl to make the solution 0.94 M in H^+ will prevent the more soluble ZnS from precipitating while ensuring that the less soluble CdS will precipitate when the solution is saturated with H_2S.

Example 16.7.3

A solution contains 0.010 M Ca^{2+} and 0.010 M La^{3+}. What concentration of HCl is needed to precipitate $La_2(C_2O_4)_3 \cdot 9H_2O$ but not $Ca(C_2O_4) \cdot H_2O$ if the concentration of oxalic acid is 1.0 M? K_{sp} values are 2.32×10^{-9} for $Ca(C_2O_4)$ and 2.5×10^{-27} for $La_2(C_2O_4)_3$; $pK_{a1} = 1.25$ and $pK_{a2} = 3.81$ for oxalic acid.

Given: concentrations of cations, K_{sp} values, and concentration and pK_a values for oxalic acid

Asked for: concentration of HCl needed for selective precipitation of $La_2(C_2O_4)_3$

Strategy:

A. Write each solubility product expression and calculate the oxalate concentration needed for precipitation to occur. Determine the concentration range needed for selective precipitation of $La_2(C_2O_4)_3 \cdot 9H_2O$.

B. Add the equations for the first and second dissociations of oxalic acid to get an overall equation for the dissociation of oxalic acid to oxalate. Substitute the $[ox^{2-}]$ needed to precipitate $La_2(C_2O_4)_3 \cdot 9H_2O$ into the overall equation for the dissociation of oxalic acid to calculate the required $[H^+]$.

Solution

A Because the salts have different stoichiometries, we cannot directly compare the magnitudes of the solubility products. Instead, we must use the equilibrium constant expression for each solubility product to calculate the concentration of oxalate needed for precipitation to occur. Using ox^{2-} for oxalate, we write the solubility product expression for calcium oxalate as follows:

$$K_{sp} = [Ca^{2+}][ox^{2-}] = (0.010)[ox^{2-}] = 2.32 \times 10^{-9}$$

$$[ox^{2-}] = 2.32 \times 10^{-7}\ M$$

The expression for lanthanum oxalate is as follows:

$$K_{sp} = [La^{3+}]^2[ox^{2-}]^3 = (0.010)^2[ox^{2-}]^3 = 2.5 \times 10^{-27}$$

$$[ox^{2-}] = 2.9 \times 10^{-8}\ M$$

Thus lanthanum oxalate is less soluble and will selectively precipitate when the oxalate concentration is between $2.9 \times 10^{-8} M$ and $2.32 \times 10^{-7} M$.

B To prevent Ca^{2+} from precipitating as calcium oxalate, we must add enough H^+ to give a maximum oxalate concentration of 2.32×10^{-7} M. We can calculate the required $[H^+]$ by using the overall equation for the dissociation of oxalic acid to oxalate:

$$HO_2CCO_2H(aq) \rightleftharpoons 2H^+(aq) + C_2O_{4(aq)}^{2-} \tag{16.7.40}$$

$$K = K_{a1}K_{a2} = (10^{-1.25})(10^{-3.81}) = 10^{-5.06} = 8.7\times10^{-6}$$

Substituting the desired oxalate concentration into the equilibrium constant expression,

$$8.7 \times 10^{-6} = \frac{[H^+]^2[ox^{2-}]}{[HO_2CCO_2H]} = \frac{[H^+]^2(2.32 \times 10^{-7})}{1.0} \tag{16.7.41}$$

$$[H^+] = 6.1\ M \tag{16.7.42}$$

Thus adding enough HCl to give $[H^+]$ = 6.1 M will cause only $La_2(C_2O_4)_3{\bullet}9H_2O$ to precipitate from the solution.

? Exercise 16.7.3

A solution contains 0.015 M Fe^{2+} and 0.015 M Pb^{2+}. What concentration of acid is needed to ensure that Pb^{2+} precipitates as PbS in a saturated solution of H_2S, but Fe^{2+} does not precipitate as FeS? K_{sp} values are 6.3×10^{-18} for FeS and 8.0×10^{-28} for PbS.

Answer

0.018 M H^+

Summary

The anion in sparingly soluble salts is often the conjugate base of a weak acid that may become protonated in solution, so the solubility of simple oxides and sulfides, both strong bases, often depends on pH. The anion in many sparingly soluble salts is the conjugate base of a weak acid. At low pH, protonation of the anion can dramatically increase the solubility of the salt. Oxides can be classified as acidic oxides or basic oxides. Acidic oxides either react with water to give an acidic solution or dissolve in strong base; most acidic oxides are nonmetal oxides or oxides of metals in high oxidation states. Basic oxides either react with water to give a basic solution or dissolve in strong acid; most basic oxides are oxides of metallic elements. Oxides or hydroxides that are soluble in both acidic and basic solutions are called amphoteric oxides. Most elements whose oxides exhibit amphoteric behavior are located along the diagonal line separating metals and nonmetals in the periodic table. In solutions that contain mixtures of dissolved metal ions, the pH can be used to control the anion concentration needed to selectively precipitate the desired cation.

Contributors and Attributions

- Anonymous

This page titled 16.7: Solubility and pH is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **18.7: Solubility and pH** is licensed CC BY-NC-SA 4.0.

16.8: Equilibria Involving Complex Ions

Learning Objectives

- To be introduced to complex ions, including ligands.

Previously, you learned that metal ions in aqueous solution are hydrated—that is, surrounded by a shell of usually four or six water molecules. A hydrated ion is one kind of a **complex ion** (or, simply, complex), a species formed between a central metal ion and one or more surrounding **ligands**, molecules or ions that contain at least one lone pair of electrons, such as the $[Al(H_2O)_6]^{3+}$ ion.

A complex ion forms from a metal ion and a ligand because of a Lewis acid–base interaction. The positively charged metal ion acts as a Lewis acid, and the ligand, with one or more lone pairs of electrons, acts as a Lewis base. Small, highly charged metal ions, such as Cu^{2+} or Ru^{3+}, have the greatest tendency to act as Lewis acids, and consequently, they have the greatest tendency to form complex ions.

As an example of the formation of complex ions, consider the addition of ammonia to an aqueous solution of the hydrated Cu^{2+} ion $\{[Cu(H_2O)_6]^{2+}\}$. Because it is a stronger base than H_2O, ammonia replaces the water molecules in the hydrated ion to form the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ ion. Formation of the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex is accompanied by a dramatic color change, as shown in Figure 16.8.1. The solution changes from the light blue of $[Cu(H_2O)_6]^{2+}$ to the blue-violet characteristic of the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ ion.



Figure 16.8.1: The Formation of Complex Ions. An aqueous solution of $CuSO_4$ consists of hydrated Cu^{2+} ions in the form of pale blue $[Cu(H_2O)_6]^{2+}$ (left). The addition of aqueous ammonia to the solution results in the formation of the intensely blue-violet $[Cu(NH_3)_4(H_2O)_2]^{2+}$ ions, usually written as $[Cu(NH_3)_4]^{2+}$ ion (right) because ammonia, a stronger base than H_2O, replaces water molecules from the hydrated Cu^{2+} ion. For a more complete description, see https://www.youtube.com/watch?v=IQNcLH6OZK0.

The Formation Constant

The replacement of water molecules from $[Cu(H_2O)_6]^{2+}$ by ammonia occurs in sequential steps. Omitting the water molecules bound to Cu^{2+} for simplicity, we can write the equilibrium reactions as follows:

$$Cu^{2+}(aq) + NH_{3(aq)} \rightleftharpoons [Cu(NH_3)]^{2+}_{(aq)} \quad K_1 \tag{16.8.1}$$

$$[Cu(NH_3)]^{2+}_{(aq)} + NH_{3(aq)} \rightleftharpoons [Cu(NH_3)_2]^{2+}_{(aq)} \quad K_2 \tag{16.8.2}$$

$$[Cu(NH_3)_2]^{2+}_{(aq)} + NH_{3(aq)} \rightleftharpoons [Cu(NH_3)_3]^{2+}_{(aq)} \quad K_3 \tag{16.8.3}$$

$$[Cu(NH_3)_3]^{2+}_{(aq)} + NH_{3(aq)} \rightleftharpoons [Cu(NH_3)_4]^{2+}_{(aq)} \quad K_4 \tag{16.8.4}$$

The sum of the stepwise reactions is the overall equation for the formation of the complex ion: The hydrated Cu^{2+} ion contains six H_2O ligands, but the complex ion that is produced contains only four NH_3 ligands, not six.

$$Cu^{2+}_{(aq)} + 4NH_{3(aq)} \rightleftharpoons [Cu(NH_3)_4]^{2+}_{(aq)} \tag{16.8.5}$$

The equilibrium constant for the formation of the complex ion from the hydrated ion is called the **formation constant (K_f)**. The equilibrium constant expression for K_f has the same general form as any other equilibrium constant expression. In this case, the expression is as follows:

$$K_f = \frac{\left[[Cu(NH_3)_4]^{2+}\right]}{[Cu^{2+}][NH_3]^4} = 2.1 \times 10^{13} = K_1K_2K_3K_4 \tag{16.8.6}$$

> *The formation constant (K_f) has the same general form as any other equilibrium constant expression.*

Water, a pure liquid, does not appear explicitly in the equilibrium constant expression, and the hydrated $Cu^{2+}(aq)$ ion is represented as Cu^{2+} for simplicity. As for any equilibrium, the larger the value of the equilibrium constant (in this case, K_f), the more stable the product. With $K_f = 2.1 \times 10^{13}$, the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ion is very stable. The formation constants for some common complex ions are listed in Table 16.8.1.

Table 16.8.1: Formation Constants for Selected Complex Ions in Aqueous Solution*

	Complex Ion	Equilibrium Equation	K_f
Ammonia Complexes	$[Ag(NH_3)_2]^+$	$Ag^+ + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]^+$	1.1×10^7
	$[Cu(NH_3)_4]^{2+}$	$Cu^{2+} + 4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{2+}$	2.1×10^{13}
	$[Ni(NH_3)_6]^{2+}$	$Ni^{2+} + 6NH_3 \rightleftharpoons [Ni(NH_3)_6]^{2+}$	5.5×10^8
Cyanide Complexes	$[Ag(CN)_2]^-$	$Ag^+ + 2CN^- \rightleftharpoons [Ag(CN)_2]^-$	1.1×10^{18}
	$[Ni(CN)_4]^{2-}$	$Ni^{2+} + 4CN^- \rightleftharpoons [Ni(CN)_4]^{2-}$	2.2×10^{31}
	$[Fe(CN)_6]^{3-}$	$Fe^{3+} + 6CN^- \rightleftharpoons [Fe(CN)_6]^{3-}$	1×10^{42}
Hydroxide Complexes	$[Zn(OH)_4]^{2-}$	$Zn^{2+} + 4OH^- \rightleftharpoons [Zn(OH)_4]^{2-}$	4.6×10^{17}
	$[Cr(OH)_4]^-$	$Cr^{3+} + 4OH^- \rightleftharpoons [Cr(OH)_4]^-$	8.0×10^{29}
Halide Complexes	$[HgCl_4]^{2-}$	$Hg^{2+} + 4Cl^- \rightleftharpoons [HgCl_4]^{2-}$	1.2×10^{15}
	$[CdI_4]^{2-}$	$Cd^{2+} + 4I \rightleftharpoons [CdI_4]^{2-}$	2.6×10^5
	$[AlF_6]^{3-}$	$Al^{3+} + 6F^- \rightleftharpoons [AlF_6]^{3-}$	6.9×10^{19}
Other Complexes	$[Ag(S_2O_3)_2]^{3-}$	$Ag^+ + 2S_2O_3^{2-} \rightleftharpoons [Ag(S_2O_3)_2]^{3-}$	2.9×10^{13}
	$[Fe(C_2O_4)_3]^{3-}$	$Fe^{3+} + 3C_2O_4^{2-} \rightleftharpoons [Fe(C_2O_4)_3]^{3-}$	2.0×10^{20}

*Reported values are overall formation constants. Source: Data from *Lange's Handbook of Chemistry*, 15th ed. (1999).

Example 16.8.1

If 12.5 g of $Cu(NO_3)_2 \cdot 6H_2O$ is added to 500 mL of 1.00 M aqueous ammonia, what is the equilibrium concentration of $Cu^{2+}(aq)$?

Given: mass of Cu^{2+} salt and volume and concentration of ammonia solution

Asked for: equilibrium concentration of $Cu^{2+}(aq)$

Strategy:

A. Calculate the initial concentration of Cu^{2+} due to the addition of copper(II) nitrate hexahydrate. Use the stoichiometry of the reaction shown in Equation 16.8.5 to construct a table showing the initial concentrations, the changes in concentrations, and the final concentrations of all species in solution.

B. Substitute the final concentrations into the expression for the formation constant (Equation 16.8.6) to calculate the equilibrium concentration of $Cu^{2+}(aq)$.

Solution

Adding an ionic compound that contains Cu^{2+} to an aqueous ammonia solution will result in the formation of $[Cu(NH_3)_4]^{2+}(aq)$, as shown in Equation 16.8.5. We assume that the volume change caused by adding solid copper(II) nitrate to aqueous ammonia is negligible.

A The initial concentration of Cu^{2+} from the amount of added copper nitrate prior to any reaction is as follows:

$$12.5\ \cancel{g}\ Cu(NO_3)_2 \cdot 6H_2O \left(\frac{1\ mol}{295.65\ \cancel{g}}\right)\left(\frac{1}{500\ \cancel{mL}}\right)\left(\frac{1000\ \cancel{mL}}{1\ L}\right) = 0.0846\ M \tag{16.8.7}$$

Because the stoichiometry of the reaction is four NH_3 to one Cu^{2+}, the amount of NH_3 required to react completely with the Cu^{2+} is 4(0.0846) = 0.338 M. The concentration of ammonia after complete reaction is 1.00 M − 0.338 M = 0.66 M. These results are summarized in the first two lines of the following table. Because the equilibrium constant for the reaction is large (2.1×10^{13}), the equilibrium will lie far to the right. Thus we will assume that the formation of $[Cu(NH_3)_4]^{2+}$ in the first step is complete and allow some of it to dissociate into Cu^{2+} and NH_3 until equilibrium has been reached. If we define x as the amount of Cu^{2+} produced by the dissociation reaction, then the stoichiometry of the reaction tells us that the change in the concentration of $[Cu(NH_3)_4]^{2+}$ is $-x$, and the change in the concentration of ammonia is $+4x$, as indicated in the table. The final concentrations of all species (in the bottom row of the table) are the sums of the concentrations after complete reaction and the changes in concentrations.

$Cu^{2+} + 4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{2+}$

	$[Cu^{2+}]$	**$[NH_3]$**	**$[[Cu(NH_3)_4]^{2+}]$**
initial	0.0846	1.00	0
after complete reaction	0	0.66	0.0846
change	$+x$	$+4x$	$-x$
final	x	0.66 + 4x	0.0846 − x

B Substituting the final concentrations into the expression for the formation constant (Equation 16.8.6) and assuming that x << 0.0846, which allows us to remove x from the sum and difference,

$$K_f = \frac{[[Cu(NH_3)_4]^{2+}]}{[Cu^{2+}][NH_3]^4} = \frac{0.0846 - x}{x(0.66 + 4x)^4} \approx \frac{0.0846}{x(0.66)^4} = 2.1 \times 10^{13} \tag{16.8.8}$$

$$x = 2.1 \times 10^{-14} \tag{16.8.9}$$

The value of x indicates that our assumption was justified. The equilibrium concentration of $Cu^{2+}(aq)$ in a 1.00 M ammonia solution is therefore 2.1×10^{-14} M.

? Exercise 16.8.1

The ferrocyanide ion $\{[Fe(CN)_6]^{4-}\}$ is very stable, with a K_f of 1×10^{35}. Calculate the concentration of cyanide ion in equilibrium with a 0.65 M solution of $K_4[Fe(CN)_6]$.

Answer 2×10^{-6} M

The Effect of the Formation of Complex Ions on Solubility

What happens to the solubility of a sparingly soluble salt if a ligand that forms a stable complex ion is added to the solution? One such example occurs in conventional black-and-white photography. Recall that black-and-white photographic film contains light-sensitive microcrystals of AgBr, or mixtures of AgBr and other silver halides. AgBr is a sparingly soluble salt, with a K_{sp} of 5.35×10^{-13} at 25°C. When the shutter of the camera opens, the light from the object being photographed strikes some of the crystals on the film and initiates a photochemical reaction that converts AgBr to black Ag metal. Well-formed, stable negative images appear in tones of gray, corresponding to the number of grains of AgBr converted, with the areas exposed to the most light being darkest. To fix the image and prevent more AgBr crystals from being converted to Ag metal during processing of the film, the unreacted AgBr on the film is removed using a complexation reaction to dissolve the sparingly soluble salt.

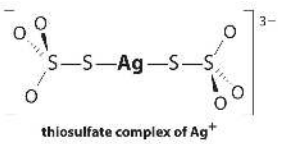
The reaction for the dissolution of silver bromide is as follows:

$$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)} \tag{16.8.10}$$

with

$$K_{sp} = 5.35 \times 10^{-13} \text{ at } 25^\circ \text{C} \tag{16.8.11}$$

The equilibrium lies far to the left, and the equilibrium concentrations of Ag^+ and Br^- ions are very low (7.31×10^{-7} M). As a result, removing unreacted AgBr from even a single roll of film using pure water would require tens of thousands of liters of water and a great deal of time. Le Chatelier's principle tells us, however, that we can drive the reaction to the right by removing one of the products, which will cause more AgBr to dissolve. Bromide ion is difficult to remove chemically, but silver ion forms a variety of stable two-coordinate complexes with neutral ligands, such as ammonia, or with anionic ligands, such as cyanide or thiosulfate ($S_2O_3^{2-}$). In photographic processing, excess AgBr is dissolved using a concentrated solution of sodium thiosulfate.



thiosulfate complex of Ag^+

The reaction of Ag^+ with thiosulfate is as follows:

$$Ag^+_{(aq)} + 2S_2O^{2-}_{3(aq)} \rightleftharpoons [Ag(S_2O_3)_2]^{3-}_{(aq)} \tag{16.8.12}$$

with

$$K_f = 2.9 \times 10^{13} \tag{16.8.13}$$

The magnitude of the equilibrium constant indicates that almost all Ag^+ ions in solution will be immediately complexed by thiosulfate to form $[Ag(S_2O_3)_2]^{3-}$. We can see the effect of thiosulfate on the solubility of AgBr by writing the appropriate reactions and adding them together:

$$\mathrm{AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)} \quad K_{\text{sp}} = 5.35 \times 10^{-13} \tag{16.8.14}$$

$$\mathrm{Ag^+(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Ag(S_2O_3)_2]^{3-}(aq)} \quad K_{\text{f}} = 2.9 \times 10^{13} \tag{16.8.15}$$

$$\mathrm{AgBr(s) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Ag(S_2O_3)_2]^{3-}(aq) + Br^-(aq)} \quad K = K_{\text{sp}}K_{\text{f}} = 15 \tag{16.8.16}$$

Comparing K with K_{sp} shows that the formation of the complex ion increases the solubility of AgBr by approximately 3×10^{13}. The dramatic increase in solubility combined with the low cost and the low toxicity explains why sodium thiosulfate is almost universally used for developing black-and-white film. If desired, the silver can be recovered from the thiosulfate solution using any of several methods and recycled.

> *If a complex ion has a large K_f, the formation of a complex ion can dramatically increase the solubility of sparingly soluble salts.*

Example 16.8.2

Due to the common ion effect, we might expect a salt such as AgCl to be much less soluble in a concentrated solution of KCl than in water. Such an assumption would be incorrect, however, because it ignores the fact that silver ion tends to form a two-coordinate complex with chloride ions ($AgCl_2^-$). Calculate the solubility of AgCl in each situation:

a. in pure water
b. in 1.0 M KCl solution, ignoring the formation of any complex ions
c. the same solution as in part (b) except taking the formation of complex ions into account, assuming that $AgCl_2^-$ is the only Ag^+ complex that forms in significant concentrations

At 25°C, $K_{sp} = 1.77 \times 10^{-10}$ for AgCl and $K_f = 1.1 \times 10^5$ for $AgCl_2^-$.

Given: K_{sp} of AgCl, K_f of $AgCl_2^-$, and KCl concentration

Asked for: solubility of AgCl in water and in KCl solution with and without the formation of complex ions

Strategy:

A. Write the solubility product expression for AgCl and calculate the concentration of Ag^+ and Cl^- in water.
B. Calculate the concentration of Ag^+ in the KCl solution.
C. Write balanced chemical equations for the dissolution of AgCl and for the formation of the $AgCl_2^-$ complex. Add the two equations and calculate the equilibrium constant for the overall equilibrium.

D. Write the equilibrium constant expression for the overall reaction. Solve for the concentration of the complex ion.

Solution

a. **A** If we let x equal the solubility of AgCl, then at equilibrium $[Ag+] = [Cl^-] = x$ M. Substituting this value into the solubility product expression,

$$K_{sp} = [Ag^+][Cl^-] = (x)(x) = x^2 = 1.77\times10^{-10}$$
$$x = 1.33\times10^{-5}$$

Thus the solubility of AgCl in pure water at 25°C is 1.33×10^{-5} M.

b. **B** If x equals the solubility of AgCl in the KCl solution, then at equilibrium $[Ag^+] = x$ M and $[Cl^-] = (1.0 + x)$ M. Substituting these values into the solubility product expression and assuming that $x << 1.0$,

$$K_{sp} = [Ag^+][Cl^-] = (x)(1.0 + x) \approx x(1.0) = 1.77\times10^{-10} = x$$

If the common ion effect were the only important factor, we would predict that AgCl is approximately five orders of magnitude less soluble in a 1.0 M KCl solution than in water.

c. **C** To account for the effects of the formation of complex ions, we must first write the equilibrium equations for both the dissolution and the formation of complex ions. Adding the equations corresponding to K_{sp} and K_f gives us an equation that describes the dissolution of AgCl in a KCl solution. The equilibrium constant for the reaction is therefore the product of K_{sp} and K_f:

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq) \quad K_{sp} = 1.77 \times 10^{-10} \tag{16.8.17}$$
$$Ag^+(aq) + 2Cl^- \rightleftharpoons [AgCl_2]^- \quad K_f = 1.1 \times 10^5 \tag{16.8.18}$$
$$AgCl(s) + Cl^- \rightleftharpoons [AgCl_2]^- \quad K = K_{sp}K_f = 1.9 \times 10^{-5} \tag{16.8.19}$$

D If we let x equal the solubility of AgCl in the KCl solution, then at equilibrium $[AgCl_2^-] = x$ and $[Cl^-] = 1.0 - x$. Substituting these quantities into the equilibrium constant expression for the net reaction and assuming that $x << 1.0$,

$$K = \frac{[AgCl_2^-]}{[Cl^-]} = \frac{x}{1.0 - x} \approx 1.9 \times 10^{-5} = x$$

That is, AgCl dissolves in 1.0 M KCl to produce a 1.9×10^{-5} M solution of the $AgCl_2^-$ complex ion. Thus we predict that AgCl has approximately the same solubility in a 1.0 M KCl solution as it does in pure water, which is 10^5 times greater than that predicted based on the common ion effect. (In fact, the measured solubility of AgCl in 1.0 M KCl is almost a factor of 10 greater than that in pure water, largely due to the formation of other chloride-containing complexes.)

? Exercise 16.8.2

Calculate the solubility of mercury(II) iodide (HgI_2) in each situation:

a. pure water
b. a 3.0 M solution of NaI, assuming $[HgI_4]^{2-}$ is the only Hg-containing species present in significant amounts

$K_{sp} = 2.9 \times 10^{-29}$ for HgI_2 and $K_f = 6.8 \times 10^{29}$ for $[HgI_4]^{2-}$.

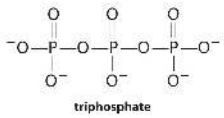
Answer

a. 1.9×10^{-10} M
b. 1.4 M



Solubility of Complex Ions: https://youtu.be/f4pkKDg2XTA

Complexing agents, molecules or ions that increase the solubility of metal salts by forming soluble metal complexes, are common components of laundry detergents. Long-chain carboxylic acids, the major components of soaps, form insoluble salts with Ca^{2+} and Mg^{2+}, which are present in high concentrations in "hard" water. The precipitation of these salts produces a bathtub ring and gives a gray tinge to clothing. Adding a complexing agent such as pyrophosphate ($O_3POPO_3^{4-}$, or $P_2O_7^{4-}$) or triphosphate ($P_3O_{10}^{5-}$) to detergents prevents the magnesium and calcium salts from precipitating because the equilibrium constant for complex-ion formation is large:



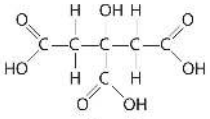
triphosphate

$$Ca^{2+}_{(aq)} + O_3POPO^{4-}_{4(aq)} \rightleftharpoons [Ca(O_3POPO_3)]^{2-}_{(aq)} \tag{16.8.20}$$

with

$$K_f = 4 \times 10^4 \tag{16.8.21}$$

However, phosphates can cause environmental damage by promoting eutrophication, the growth of excessive amounts of algae in a body of water, which can eventually lead to large decreases in levels of dissolved oxygen that kill fish and other aquatic organisms. Consequently, many states in the United States have banned the use of phosphate-containing detergents, and France has banned their use beginning in 2007. "Phosphate-free" detergents contain different kinds of complexing agents, such as derivatives of acetic acid or other carboxylic acids. The development of phosphate substitutes is an area of intense research.



Citric acid,
a hydroxy polycarboxylic acid

Commercial water softeners also use a complexing agent to treat hard water by passing the water over ion-exchange resins, which are complex sodium salts. When water flows over the resin, sodium ion is dissolved, and insoluble salts precipitate onto the resin surface. Water treated in this way has a saltier taste due to the presence of Na^+, but it contains fewer dissolved minerals.

Another application of complexing agents is found in medicine. Unlike x-rays, magnetic resonance imaging (MRI) can give relatively good images of soft tissues such as internal organs. MRI is based on the magnetic properties of the ^1H nucleus of hydrogen atoms in water, which is a major component of soft tissues. Because the properties of water do not depend very much on whether it is inside a cell or in the blood, it is hard to get detailed images of these tissues that have good contrast. To solve this problem, scientists have developed a class of metal complexes known as "MRI contrast agents." Injecting an MRI contrast agent into a patient selectively affects the magnetic properties of water in cells of normal tissues, in tumors, or in blood vessels and allows doctors to "see" each of these separately (Figure 16.8.2). One of the most important metal ions for this application is Gd^{3+}, which with seven unpaired electrons is highly paramagnetic. Because Gd^{3+}(aq) is quite toxic, it must be administered as a very

stable complex that does not dissociate in the body and can be excreted intact by the kidneys. The complexing agents used for gadolinium are ligands such as $DTPA^{5-}$ (diethylene triamine pentaacetic acid), whose fully protonated form is shown here.

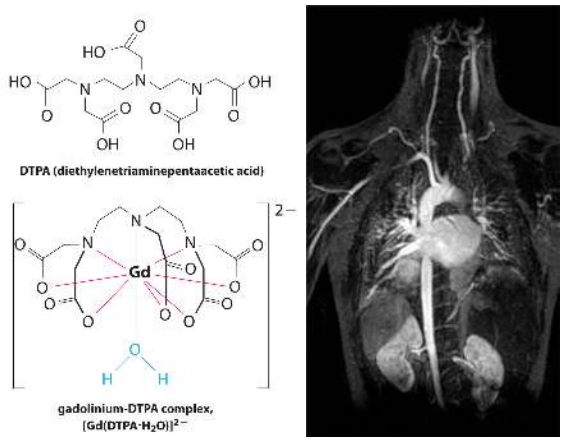


Figure 16.8.2: An MRI Image of the Heart, Arteries, and Veins. When a patient is injected with a paramagnetic metal cation in the form of a stable complex known as an MRI contrast agent, the magnetic properties of water in cells are altered. Because the different environments in different types of cells respond differently, a physician can obtain detailed images of soft tissues.

Summary

The formation of complex ions can substantially increase the solubility of sparingly soluble salts if the complex ion has a large K_f. A complex ion is a species formed between a central metal ion and one or more surrounding ligands, molecules or ions that contain at least one lone pair of electrons. Small, highly charged metal ions have the greatest tendency to act as Lewis acids and form complex ions. The equilibrium constant for the formation of the complex ion is the formation constant (K_f). The formation of a complex ion by adding a complexing agent increases the solubility of a compound.

This page titled 16.8: Equilibria Involving Complex Ions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **18.8: Equilibria Involving Complex Ions** is licensed CC BY-NC-SA 4.0.

16.9: Qualitative Cation Analysis

Learning Objectives

- To know how to separate metal ions by selective precipitation.
- To understand how several common metal cations can be identified in a solution using selective precipitation.

The composition of relatively complex mixtures of metal ions can be determined using **qualitative analysis**, a procedure for discovering the identity of metal ions present in the mixture (rather than quantitative information about their amounts). The procedure used to separate and identify more than 20 common metal cations from a single solution consists of selectively precipitating only a few kinds of metal ions at a time under given sets of conditions. Consecutive precipitation steps become progressively less selective until almost all of the metal ions are precipitated, as illustrated in Figure 16.9.1.

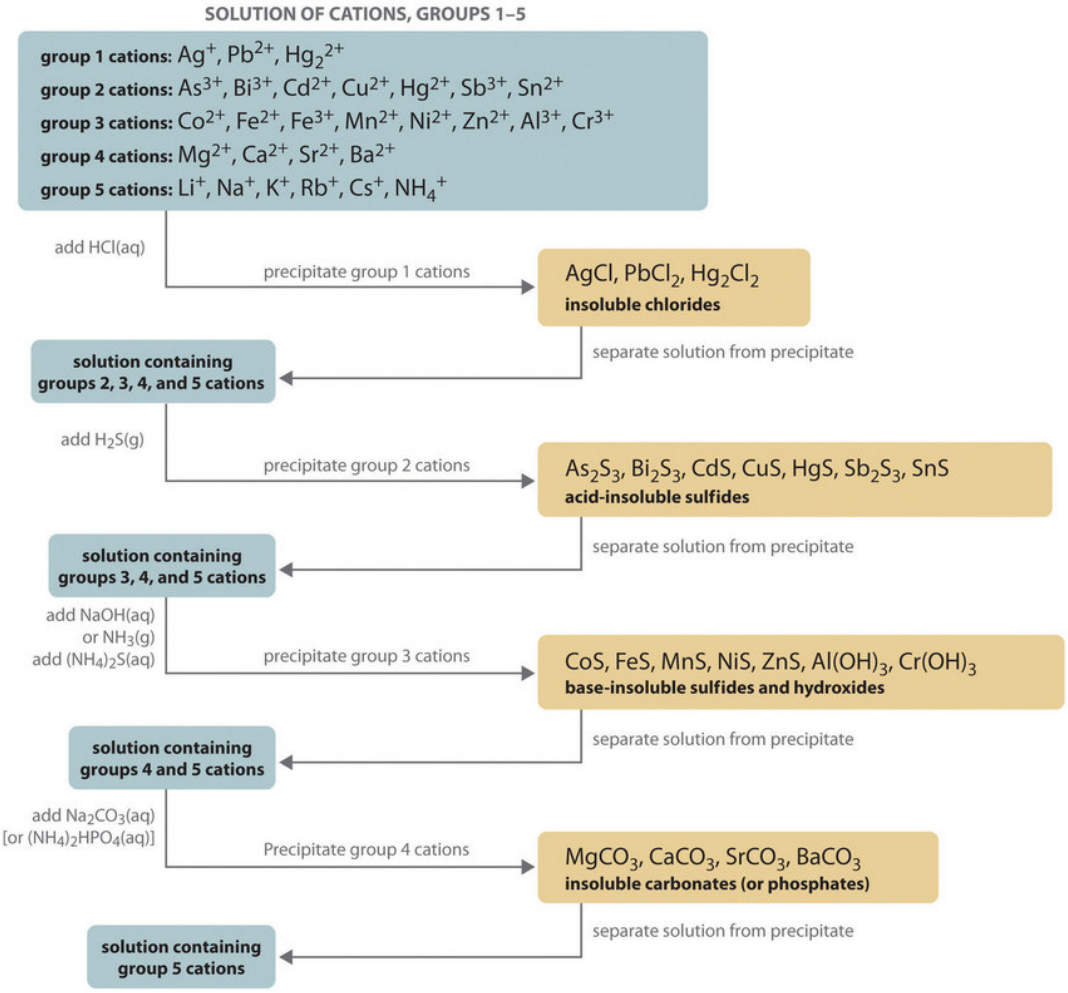


Figure 16.9.1: Steps in a Typical Qualitative Analysis Scheme for a Solution That Contains Several Metal Ions

Group 1: Insoluble Chlorides

Most metal chloride salts are soluble in water; only Ag^+, Pb^{2+}, and Hg_2^{2+} form chlorides that precipitate from water. Thus the first step in a qualitative analysis is to add about 6 M HCl, thereby causing AgCl, $PbCl_2$, and/or Hg_2Cl_2 to precipitate. If no precipitate forms, then these cations are not present in significant amounts. The precipitate can be collected by filtration or centrifugation.

Group 2: Acid-Insoluble Sulfides

Next, the acidic solution is saturated with H_2S gas. Only those metal ions that form very insoluble sulfides, such as As^{3+}, Bi^{3+}, Cd^{2+}, Cu^{2+}, Hg^{2+}, Sb^{3+}, and Sn^{2+}, precipitate as their sulfide salts under these acidic conditions. All others, such as Fe^{2+} and Zn^{2+}, remain in solution. Once again, the precipitates are collected by filtration or centrifugation.

Group 3: Base-Insoluble Sulfides (and Hydroxides)

Ammonia or $NaOH$ is now added to the solution until it is basic, and then $(NH_4)_2S$ is added. This treatment removes any remaining cations that form insoluble hydroxides or sulfides. The divalent metal ions Co^{2+}, Fe^{2+}, Mn^{2+}, Ni^{2+}, and Zn^{2+} precipitate as their sulfides, and the trivalent metal ions Al^{3+} and Cr^{3+} precipitate as their hydroxides: $Al(OH)_3$ and $Cr(OH)_3$. If the mixture contains Fe^{3+}, sulfide reduces the cation to Fe^{2+}, which precipitates as FeS.

Group 4: Insoluble Carbonates or Phosphates

The next metal ions to be removed from solution are those that form insoluble carbonates and phosphates. When Na_2CO_3 is added to the basic solution that remains after the precipitated metal ions are removed, insoluble carbonates precipitate and are collected. Alternatively, adding $(NH_4)_2HPO_4$ causes the same metal ions to precipitate as insoluble phosphates.

Group 5: Alkali Metals

At this point, we have removed all the metal ions that form water-insoluble chlorides, sulfides, carbonates, or phosphates. The only common ions that might remain are any alkali metals (Li^+, Na^+, K^+, Rb^+, and Cs^+) and ammonium (NH_4^+). We now take a second sample from the original solution and add a small amount of $NaOH$ to neutralize the ammonium ion and produce NH_3. (We cannot use the same sample we used for the first four groups because we added ammonium to that sample in earlier steps.) Any ammonia produced can be detected by either its odor or a litmus paper test. A flame test on another original sample is used to detect sodium, which produces a characteristic bright yellow color. The other alkali metal ions also give characteristic colors in flame tests, which allows them to be identified if only one is present.

Metal ions that precipitate together are separated by various additional techniques, such as forming complex ions, changing the pH of the solution, or increasing the temperature to redissolve some of the solids. For example, the precipitated metal chlorides of group 1 cations, containing Ag^+, Pb^{2+}, and Hg_2^{2+}, are all quite insoluble in water. Because $PbCl_2$ is much more soluble in hot water than are the other two chloride salts, however, adding water to the precipitate and heating the resulting slurry will dissolve any $PbCl_2$ present. Isolating the solution and adding a small amount of Na_2CrO_4 solution to it will produce a bright yellow precipitate of $PbCrO_4$ if Pb^{2+} were in the original sample (Figure 16.9.2).



Figure 16.9.2: When a small amount of Na_2CrO_4 solution is added to a sample containing Pb^{2+} ions in water, a bright yellow precipitate of $PbCrO_4$ forms. (CC BY-SA 3.0 Unported; PRHaney via Wikipedia).

As another example, treating the precipitates from group 1 cations with aqueous ammonia will dissolve any $AgCl$ because Ag^+ forms a stable complex with ammonia: $[Ag(NH_3)_2]^+$. In addition, Hg_2Cl_2 disproportionates in ammonia.

$$2\,Hg_2^{2+} \rightarrow Hg + Hg^{2+}$$

to form a black solid that is a mixture of finely divided metallic mercury and an insoluble mercury(II) compound, which is separated from solution:

$$Hg_2Cl_2(s) + 2\,NH_3(aq) \rightarrow Hg(l) + Hg(NH_2)Cl(s) + NH_4^+(aq) + Cl^-(aq)$$

Any silver ion in the solution is then detected by adding HCl, which reverses the reaction and gives a precipitate of white $AgCl$ that slowly darkens when exposed to light:

$$[\mathrm{Ag(NH_3)_2}]^+(\mathrm{aq}) + 2\,\mathrm{H}^+(\mathrm{aq}) + \mathrm{Cl}^-(\mathrm{aq}) \rightarrow \mathrm{AgCl(s)} + 2\,\mathrm{NH_4^+(aq)}$$

Similar but slightly more complex reactions are also used to separate and identify the individual components of the other groups.

Summary

In qualitative analysis, the identity, not the amount, of metal ions present in a mixture is determined. The technique consists of selectively precipitating only a few kinds of metal ions at a time under given sets of conditions. Consecutive precipitation steps become progressively less selective until almost all the metal ions are precipitated. Other additional steps are needed to separate metal ions that precipitate together.

This page titled 16.9: Qualitative Cation Analysis is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.7: Qualitative Analysis for Metallic Elements** by Anonymous is licensed CC BY-NC-SA 3.0.

16.10: Exercises

Q17.1a

Write K_{sp} expression for the following equilibria.

a. $PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-_{(aq)}$
b. $Cu(IO_3)_{2\ (s)} \rightleftharpoons Cu^{2+}(aq) + 2IO^-_{3\ (aq)}$
c. $Ag(OH)_{3\ (s)} \rightleftharpoons Al^{3+}(aq) + 3OH^-_{(aq)}$
d. $CaF_{2\ (s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$

S17.1a

a. $K_{sp} = [Pb^{2+}][I^-]^2$
b. $K_{sp} = [Cu^{2+}][IO^-_3]^2$
c. $K_{sp} = [Al^{3+}][OH^-]^3$
d. $K_{sp} = [Ca^{2+}][F^-]^2$

Q17.1b

Write K_{sp} expression for the following equilibria. For example $MX_{2(s)} \rightleftharpoons M^{2+}_{(aq)} + 2X^-_{(aq)}$, $K_{sp} = [M^{2+}][X^-]^2$

- $Al(OH)_{3(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3OH^-_{(aq)}$
- $CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$
- $AgI_{(s)} \rightleftharpoons Ag^+_{(aq)} + I^-_{(aq)}$
- $BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)}$

S17.1b

a. $K_{sp} = [Al^{3+}][OH^-]^3$
b. $K_{sp} = [Ca^{2+}][F^-]^2$
c. $K_{sp} = [Ag^+][I^-]$
d. $K_{sp} = [Ba^{2+}][SO^{2+}_4]$

Q17.1c

Write K_{sp} expressions for the following equilibrium reactions

a. $AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$
b. $Al(OH)_{3(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3OH^-_{(aq)}$
c. $PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$

S17.1c

a. $K_{sp} = [Ag^+][Br^-]$
b. $K_{sp} = [Al^{3+}][OH^-]^3$
c. $K_{sp} = [Pb^{2+}][Cl^-]^2$

Q17.1d

What would the K_{sp} expression be for the following reactions?

a. $NH_4Br_{(s)} \rightleftharpoons NH^+_{4(aq)} + Br^-_{(aq)}$
b. $BaCrO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CrO^{2-}_{4(aq)}$
c. $Na_2S_2O_{3(s)} \rightleftharpoons 2Na^+_{(aq)} + S_2O^{2-}_{3(aq)}$
d. $Fe_2(Cr_2O_7)_{3(s)} \rightleftharpoons 2Fe^{3+}_{(aq)} + 3Cr_2O^{2-}_{7(aq)}$

S17.1d

- $[NH_4^+][Br^-]$
- $[Ba^{2+}][CrO_4^{2-}]$
- $[Na^+]^2[S_2O_3^{2-}]$
- $[Fe^{3+}]^2[Cr_2O_7^{2-}]^3$

Q17.2

What is the solubility equilibrium and solubility-product constant expression for CaF_2?

S17.2

1. $CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$

2. $K_{sp} = [Ca^{2+}][F^-]^2$

Q17.2

Write the solubility equilibrium equations that are described by the following K_{sp} expressions. For example, $K_{sp} = [Ra^{2+}][IO_3^-]^2$ represents $Ra(IO_3)_2 \rightleftharpoons Ra^{2+}_{(aq)} + 2IO^-_{3(aq)}$

a. $PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$
b. $Mg_3(PO_4)_{2(s)} \rightleftharpoons 3Mg^{2+}_{(aq)} + 2PO^{3-}_{4(aq)}$
c. $SrCO_{3(s)} \rightleftharpoons Sr^{2+}_{(aq)} + CO^{2-}_{(aq)}$
d. $Hg_2Cl_{2(s)} \rightleftharpoons Hg^{2+}_{(aq)} + 2Cl^-_{(aq)}$

S17.2

a. $K_{sp} = [Pb^{2+}][Cl^-]^2$
b. $K_{sp} = [Mg^{2+}]^3[PO_4^{3-}]^2$
c. $K_{sp} = [Sr^{2+}][CO_3^{2-}]$
d. $K_{sp} = [Hg_2^{2+}][Cl^-]^2$

Q17.2

Write the equilibrium reactions for these K_{sp} expressions

a. $K_{sp} = [Ag^+][CrO_4^{2-}]$
b. $K_{sp} = [Ba^{2+}][SO_4^{2-}]$
c. $K_{sp} = [Pb^{2+}][CrO_4^{2-}]$

S17.2

a. $Ag_2CrO_{4(s)} \rightleftharpoons Ag^+_{(aq)} + CrO^{2-}_{4(aq)}$
b. $BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)}$
c. $PbCrO_{4(s)} \rightleftharpoons Pb^{2+}_{(aq)} + CrO^{2-}_{4(aq)}$

Q17.2

Write the solubility equilibrium equations according to the K_{sp} expression provided.

a. $AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$
b. $Ra(IO_3)_{2(s)} \rightleftharpoons Ra^{2+}_{(aq)} + 2IO^-_{3(aq)}$
c. $Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^+_{(aq)} + CrO^{2-}_{4(aq)}$
d. $MgF_{2(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2F^-_{(aq)}$

S17.2

a. $K_{sp} = [Ag^+][Cl^-]$
b. $K_{sp} = [Ra^{2+}][IO_3^-]^2$
c. $K_{sp} = [Ag^+]^2[CrO_4^{2-}]$
d. $K_{sp} = [Mg^{2+}][F^-]^2$

Q17.3a

Write the solubility product expression to which each one applies:

a. $K_{sp}(Hg_2Cl_2) = 1.3 \times 10^{-19}$
b. $K_{sp}(AgI) = 8.5 \times 10^{-17}$
c. $K_{sp}(SrSO_4) = 3.2 \times 10^{-7}$
d. $K_{sp}(Fe(OH)_3) = 4 \times 10^{-38}$

S17.3a

a. $K_{sp} = [Hg^{2+}][Cl^-]^2 = 1.3 \times 10^{-19}$
b. $K_{sp} = [Ag^+][I^-] = 8.5 \times 10^{-17}$
c. $K_{sp} = [Sr^{2+}][SO_4^{2-}] = 3.2 \times 10^{-7}$
d. $K_{sp} = [Fe^{3+}][OH^-]^3 = 4 \times 10^{-38}$

Q17.3b

The following K_{sp} values are found in a table. Write the solubility product expression to which each one applies. For example, $K_{sp}(BaCO_3) = [Ba^{2+}][CO_3^-] = 5.1 \times 10^{-9}$

a. $K_{sp}(CaSO_4) = 9.1 \times 10^{-6}$
b. $K_{sp}(PbI_2) = 7.1 \times 10^{-9}$
c. $K_{sp}(Mg(OH)_2) = 1.8 \times 10^{-11}$
d. $K_{sp}(Ag_2CrO_2) = 1.1 \times 10^{-12}$

S17.3b

a. $[Ca^{2+}][SO_4^{2-}] = 9.1 \times 10^{-6}$
b. $[Pb^{2+}][I^-]^2 = 7.1 \times 10^{-9}$
c. $[Mg^{2+}][OH^-]^2 = 1.8 \times 10^{-11}$
d. $[Ag^+]^2[CrO_4^{2-}] = 1.1 \times 10^{-12}$

Q17.3c

Write the solubility product expression for each K_{sp} value

a. $K_{sp}(BaF_2) = 3.15 \times 10^{-3}$
b. $K_{sp}(Mg(OH)_2) = 1.71 \times 10^{-4}$
c. $K_{sp}(BaCO_3) = 5.1 \times 10^{-9}$

S17.3c

a. $K_{sp} = [Ba^+][F^-]^2 = 3.15 \times 10^{-3}$
b. $K_{sp} = [Mg^{2+}][OH^-]^2 = 1.71 \times 10^{-4}$
c. $K_{sp} = [Ba^{2+}][CO_3^{2-}] = 5.1 \times 10^{-9}$

Q17.3d

Write the solubility product expression for the following.

a. $Ag_2CrO_{4(s)}$
b. $Mg_3(PO_4)_{2(s)}$
c. $PbCl_{2(s)}$

d. $Cr(OH)_3$

S17.3d

a. $[CrO_4^{2-}][Ag^+]^2$
b. $[Mg^{2+}]^3[PO_4^{3-}]^2$
c. $[Pb^{2+}][Cl^-]^2$
d. $[Cr^{3+}][OH^-]^3$

Q17.5

Which of the following would be more soluble in an acidic solution: $Zn(OH)_2$, KBr, $AlCl_3$, PbS?

S17.5

$Zn(OH)_2$ and PbS would be more soluble in an acidic solution.

Q17.25

A particular water sample that is saturated in CaF_2 has a Ca^{2+} content of 130 ppm (i.e., 130 g Ca^{2+} per 10^6 grams of water sample). What is the F^- ion content of the water in ppm?

S17.25

CaF_2 dissociates into Ca^{2+} and $2F^-$. $(CaF_2 \rightleftharpoons Ca^{2+} + 2F^-)$ So the ion content of the F^- ion would be double that of the Ca^{2+} ion content per 10^6 grams of the water sample. $\frac{260\ g\ F^-}{10^6\ g} = 260 ppm$

Q17.25

How much SO_4^{2-} ion content (in ppm) that a sample of water contains when it's saturated in $CaSO_4$ and it has a Ca^{2+} content of 40 ppm?

S17.25

$[SO_4^{2-}] = 2.76\ M = 2.65 \times 10^{-7}\ ppm$

Video Solution

Q17.25a

A sample of water is saturated with $Mg(OH)_2$. The sample has a Mg^{2+} content of 45 ppm (that is, 45 g Mg^{2+} per 10^6 g of water sample. What is the OH^- ion content of the water in ppm?

S17.25a

$[Mg^{2+}] = \frac{45\ g\ Mg^{2+}}{10^6\ g\ soln} \times \frac{1\ mol\ Mg^{2+}}{24.31\ g\ Mg^{2+}} \times \frac{1000 g\ soln}{1\ L\ soln} = 1.85 \times 10^{-3}$

$[Mg^{2+}][OH^-]^2 = K_{sp} = 1.8 \times 10^{-11} = (1.85 \times 10^{-3})[OH^-]^2$

$[OH^-] = 9.86 \times 10^{-5}$

$ppm\ OH^- = \frac{9.86 \times 10^{-5}}{1\ L\ soln} \times \frac{17\ g\ OH^-}{1\ mol\ OH^-} \times \frac{1\ L\ soln}{1000\ g} \times 10^6\ soln = 1.68 ppm$

Q17.25b

Silver chromate is dissolved in 0.1M $AgNO_3$. What affect does the common ion have on the solubility? $K_{sp} = 9 \times 10^{-12}$ molar solubility =$1.3 \times 10^{-4}\ M$

S17.25b

$$Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^+_{(aq)} + CrO^{2-}_{4(aq)} \qquad (16.10.1)$$

ICE	$Ag_2CrO^+_{4(s)}$	$2Ag^+_{(aq)}$	$CrO^{2-}_{4(aq)}$
Initial	-	0.1 M	0
Change	-	+2x	+x
Equilibrium	-	0.1M+2x	x

$$K_{sp} = [Ag_+]^2[CrO_4^{2-}] = (0.1\ M + 2x)^2(x) = 9 \times 10^{-12} \tag{16.10.2}$$

if we assume $x << (0.1M)^2(x) = 9 \times 10^{-12} \approx 9 \times 10^{-10}$

Q17.31

Will a precipitate of lead (II) occur if three drops of 0.20MKI are added to 100.0mL of 0.010M $Pb(NO_3)_2$? assuming 1 drop= 0.05 mL.

S17.31

$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2I^-_{(aq)}$

$K_{sp} = 7.1 \times 10^{-9}$

$amt I- = (3drops) \times \frac{0.05mL}{1drop} \times \frac{L}{1000mL} \times \frac{0.20mol\ KI}{L} \times \frac{mol\ I^-}{mol\ KI} = 3.0 \times 10^{-5} mol\ I^-$

$total\ volume = (0.1L) + 3drops \frac{0.05mL}{1drop} \frac{L}{1000mL} = 0.10015L$

$[I^-] = \frac{3.0 \times 10^{-5}\ mol}{0.10015L} = 3.0 \times 10^{-4}\ M$

$Q_{sp} = [Pb^{2+}][I^-]^2 = (0.010M)(3.0 \times 10^{-4})^2 = 9.0 \times 10^{-10}$

Since $Q_{sp} < K_{sp}$, therefore, $PbI_{2(s)}$ will NOT precipitate

Q17.33c

Can $Ce(IO_3)_3$ precipitate when adding together 100 mL of $1.0 \times 10^{-2}\ M\ Ce(NO_3)_3$ and 200 mL of $2.0 \times 10^{-3}\ M\ KIO_3$? ($K_{sp} = 1.9 \times 10^{-10}$).

S17.33c

Yes, it will precipitate, since Q is bigger than K.

Q17.33

4.7 mL of 0.35 M KI are added to a 150.0 mL of 0.069 M $Pb(NO_3)_2$. Will a precipitate of lead iodide be formed?

The following reactions has a K_{sp} of 7.1×10^{-9}:

$$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2I^-_{(aq)} \tag{16.10.3}$$

S17.33

Two salt solutions are combined so four different salts may precipitate out. However, we are only interseted in the $Pb(NO_3)_2$. So this is a simple compare problem involving calculating Q_{sp} and then comparing to K_{sp}.

$$Q_{sp} = [Pb^{2+}][I^-]^2 \tag{16.10.4}$$

Both $[Pb^{2+}]$ and $[I^-]$ are needed to be known to calculate Q_{sp} for the final solution. This is a changing concentration problem too.

$$(4.7\ \cancel{mL}) \left(\frac{1\ L}{1000\ \cancel{mL}} \right) \left(\frac{0.35\ \cancel{mol\ KI}}{1\ L} \right) \left(\frac{1\ mol\ I^-}{1\ \cancel{mol\ KI}} \right) = 1.65 \times 10^{-3} mol I^- \tag{16.10.5}$$

Convert to concentration

$$[I^-] = \frac{1.65 \times 10^{-3}\ mol\ I^-}{0.150\ L} = 1.07 \times 10^{-2}\ M \tag{16.10.6}$$

$$Q_{sp} = [Pb^{2+}][I^-]^2 = (0.069)(1.07 \times 10^{-2})^2 = 8.3 \times 10^{-6} > K_{sp} = 7.1 \times 10^{-9} \tag{16.10.7}$$

Therefore, $PbI_{2(s)}$ should precipitate after mixing these two solutions.

Q17.35a

When 310 ml of 0.460 M CaS are added to 310.0 ml of 0.112M $NaNO_3$, what percentage of the Na^+ is left unprecipitated?

S17.35a

$$NaNO_3 + CaS \rightarrow NaS + CaNO_3 \tag{16.10.8}$$

$$(310\ mL)(0.460\ M\ CaS) = 142.6\ mmol\ CaS \tag{16.10.9}$$

$$(310\ mL)(0.112\ M\ NaNO_3) = 34.72\ mmol\ NaNO_3 \tag{16.10.10}$$

There are no unprecipitated sodium ions in the solution because there is an excess of CaS.

Q17.35b

How much Mg^{2+} ions will remain in solution after adding $NaOH$ to a solution that's composed of 200.0 mL of 0.010M Mg^{2+} and 0.005M Ca^{2+}?

S17.35b

The Mg^{2+} ions already get used 100%.

Q17.35c

When 70.0 mL sample of 0.372 M $Na_2SO_{4(aq)}$ is added to 100 mL of 0.138 M $Ca(NO_3)_{2(aq)}$, what percentage of Ca^{2+} remains unprecipitated?

S17.35c

$[SO_4^{2-}] = 0.372\ M \times \frac{70\ mL}{(70mL+100mL)} = 0.153\ M$

$[Ca^{2+}] = 0.138\ M \times \frac{100mL}{(70mL+100mL} = 0.0776M$

We determine the value of the ion product and compare it to the solubility product constant value.

$Q_{sp} = [Ca^{2+}][NO_3^-]^2 = (0.0776)(0.153)^2 = 1.82 \times 10^{-3} > 9.1 \times 10^{-6} = K_{sp}\ for\ Ca(NO_3)_2$

Since $Q_{sp} > K_{sp}$, $Ca(NO_3)_2$ should precipitate.

Equation: $Ca(NO_3)_{2(s)} \rightleftharpoons \quad Ca^{2+}_{(aq)} + \quad 2NO^-_{3(aq)}$

Orig. soln: ____ 0.0776 M 0.153M

Form solid: ____ -0.0776 M -0.00219M

Not @ Eq: ____ 0 M 0.151 M

Changes: ____ + X M + 2X M

Equil: ____ X M (0.151 + 2X) M

$K_{sp} = [Ca^{2+}][NO_3^-]^2 = 9.1 \times 10^{-6} = (X)(0.151 + 2X)^2$

$X = 2.08 \times 10^{-7}$

$[Ca^{2+}] = X = 2.08 \times 10^{-7}$

$\%\ Ca^{2+}\ unprecipitated = \frac{2.08 \times 10^{-7}\ M\ final}{0.0776\ M\ initial} \times 100\% = 2.68 \times 10^{-4}\ \%\ unprecipitated$

Q17.35d

Silver ion and bromine ion precipitate into $AgBr_{(s)}$. The Ag^+ concentration is about 0.065M If the Br^- concentration remains at $1.0 \times 10^{17}\ M$, what will be the remaining of $[Ag^+]$ in solution when precipitation stops ($K_{sp} = 7.7 \times 10^{13}$)?

S17.35d

$$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)} \tag{16.10.11}$$

$Q_{sp} = [Ag^+][Br^-] = (0.065)(1.0 \times 10^{15})$

$Q_{sp} > K_{sp}$ therefore, precipitation occurs

$[Ag^+][Br^-] = K_{sp}$

$[Ag^+] \times (1.0 \times 10^{15}) = 7.7 \times 10^{13}$

$[Ag^+]\ remaining = 7.7 \times 10^{-4}\ M$

Q17.37a

$[OH^-]$ is maintained at $3.0 \times 10^{-2}\ M$ in a solution and $[Mg^{2+}] = 0.036\ M$. How much Mg^{2+} remain after $Mg(OH)_2$ precipitate? ($K_{sp} = 1.8 \times 10^{-11}$)

S17.37a

$[Mg^{2+}] = 2.0 \times 10^{-8}\ M$ remains.

Q17.37b

If a constant $[F^-] = 0.003\ M$ is maintained in a solution in which the initial $[Ca^{2+}] = 0.038\ M$, what percentage of Ca^{2+} will remain in solution after $CaF_{2(s)}$ precipitates? What $[F^-]$ should be maintained to ensure only 3.4% of Ca^{2+} remains unprecipitated?

S17.37b

We first use the solubility product constant expression to determine $[Ca^{2+}]$ in a solution with 0.003 M F^-

$K_{sp} = [Ca^{2+}][F^-]^2$

$5.3 \times 10^{-9} = [Ca^{2+}](0.003)^2$

$[Ca^{2+}] = \frac{5.3 \times 10^{-9}}{(0.003)^2} = 5.89 \times 10^{-4}$

$\%\ unprecipitated = \frac{5.89 \times 10^{-4}\ M}{0.038\ M} \times 100\% = 1.5\%$

Now, we want to find $[F^-]$ must be maintained to keep $[Ca^{2+}]_{final}$ =3.4%

$[Ca^{2+}]_{initial} = 0.034 \times 0.038 M = 1.29 \times 10^{-3}\ M$

$K_{sp} = [Ca^{2+}][F^-]^2 = 5.3 \times 10^{-9} = (1.29 \times 10^{-3})[F^-]^2$

$[F^-] = \sqrt{\frac{5.3 \times 10^{-9}}{1.29 \times 10^{-3}}} = 0.0021\ M$

Q17.37c

If K_{sp} of $Zn(CN)_{2(s)}$ is 3×10^{-16}, what percentage of Zn^{2+} is precipitated when we know that

- $[Zn^{2+}]_o = 0.36\ M$
- $[CN^-]_o = 5 \times 10^{-3}\ M$
- $K_{sp} = 3 \times 10^{-16}$ for $Zn(CN)_{2(s)}$.

S17.37c

$$Zn(CN)_{2(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2CN^-_{(aq)} \tag{16.10.12}$$

$[Zn^{2+}][CN]^2 = 3 \times 10^{-16}$

$[Zn^{2+}] = \frac{3 \times 10^{-16}}{(5 \times 10^{-3})^2} = 1.2 \times 10^{-11}$

Due to the precipitation reaction, $[Zn]$ was reduced to 3.33×10^{-9}, signifying that it has essentially fully precipitated out of solution.

Q17.39a

There are 320 g Na^+ per kg of water

a. should NaOH precipitate when $[OH^-] = 1.57 \times 10^{-3}\ M$?
b. is it possible for Na and Cl to be separated in the water?

S17.39a

a. no
b. no

Q17.39b

A sample of seawater contains 0.25 M Mg^{2+} and 0.0033 M Ca^{2+}. Which will precipitate first when $NaOH$ is added? What's the concentration of X^{2+} ion remaining after $X(OH)_2$ precipitate?

S17.39b

$Mg(OH)_2$ will precipitate first, and it gets used almost completely.

Q17.39c

0.01M of Cu^+ and 0.01M of Ag^+ are in solution. A piece of solid NaI is put in inside. What is the concentration of Ag^+ as soon as CuI begins to precipitate?

S17.39c

$Cu^+ + I^- \rightarrow CuI$ and K_{sp} is 5.1×10^{-12}

$Ag^+ + I^- \rightarrow AgI$ and K_{sp} is 8.3×10^{-17}

$[Ag^+] = \frac{K_{sp}\ of\ AgI}{[I^-]\ in\ CuI}$

$[Ag^+] = \frac{8.3 \times 10^{-17}}{5.1 \times 10^{-10}} = 1.627 \times 10^{-7}$

Q17.41

$KI_{(aq)}$ is slowly added to a solution with $[Pb^{2+}] = [Ag^+] = 0.30\ M$. For PbI_2, $K_{sp} = 9.6 \times 10^{-9}$; for AgI, $K_{sp} = 6.2 \times 10^{-18}$.

a. Which precipitate should form first, PbI_2 or AgI?
b. What is fractional precipitation?
c. Can Pb^{2+} and Ag^+ be effectively separated by fractional precipitation of their iodides?

S17.41

a. AgI will precipitate first because it has the smaller Ksp value.
b. Fractional precipitation is the technique of separating two or more ions from a solution by adding a reactant that precipitates first one ion, then another, and so on.
c. Yes because the ratio of the Ksp values compared to the molar concentration of .3 M for [Pb^{2+}] and [Ag^+] shows that both ions will completely dissociate.

http://chemwiki.ucdavis.edu/Inorganic_Chemistry/Reactions_in_Aqueous_Solutions/Precipitation_Reactions

Q17.41

If Cl^- is added to a solution that contains $0.025\ M$ of Ag^+ ions and $0.00448\ M$ of Pb^{2+} ions. What concentration of Cl^- is required to precipitate each ion? $K_{sp}(AgCl) = 1.8 \times 10^{-10}$ and $K_{sp}(PbCl_2) = 1.6 \times 10^{-5}$.

S17.41

- After adding $7.2 \times 10^{-9}\ M$ of $[Cl^-]$, silver will precipitate.
- After adding $3.57 \times 10^{-3}\ M$ of $[Cl^-]$, lead will precipitate.

Q17.41

0.23 M Na^+ is added to a solution of 0.45 M Cl^- and 0.48 M OH^-.

a. Should $NaCl$ or $NaOH$ precipitate first?
b. What concentration of $[OH^-]$ is required for the precipitation?
c. Can these ions be effectively precipitated by fractional precipitation?

S17.41

a. NaCl
b. 0.049 M
c. yes

Q17.43a

A concentration solution of potassium chromate is added to a solution that is 0.1 M Ba^{2+} and 0.1 M Sr^{2+} Which ion will precipitate first? What is the percentage of this ion remaining when the other ion begins to precipitate?

S17.43a

$Q_c = K_{sp}$

a)

$BaCrO_{4(s)} \rightleftharpoons Ba^{2+} + CrO_4^{2-}$

$K_{sp} = [Ba^{2+}][CrO_4^{2-}][CrO_4^{2-}] = 1.2 \times 10^{-9}\ M$

$1.2 \times 10^{-10} = (0.1\ M)[CrO_4^{2-}]$

$SrCrO_{4(s)} \rightleftharpoons Sr^{2+} + CrO_4^{-2}$

$K_{sp} = [Sr^{2+}][CrO_4^{2-}][CrO_4^{2-}] = 3.5 \times 10^{-4}\ M$

$3.5 \times 10^{-5} = (0.1\ M)[CrO_4^{2-}]$

Therefore, $BaCrO_4$ PRECIPITATES FIRST

b)

$BaCrO_{4(s)} \rightleftharpoons Ba^{2+} + CrO_4^{2-}$

$K_{sp} = [Ba^{2+}][CrO_4^{2-}]$

$1.2 \times 10^{-10} = [Ba^{2+}][3.5 \times 10^{-4}]$

$[Ba^{2+}] = 3.4 \times 10^{-7}\ M$

$\%\ Ba^{2+}\ remaining = \frac{3.4 \times 10^{-7}\ M}{0.1} \times 100\% = 0.000343\%$

Q17.43b

0.4 M Cu^{2+} is added to a solution of 0.039 M CO_2^3 and 0.039 M Cl^-.

a. Which precipitates first?
b. When the second ion starts to precipitate, what is the concentration of the first ion?
c. Can these be effectively separated by fractional precipitation?

S17.43b

a. $CuCO_3$
b. 0.028 M
c. yes

Q17.43c

An aqueous solution that 2.00 M in $Cu(CO_3)$ is slowly added from a buret to a aqueous solution that is 0.020 M in Cl^- and 0.30 M in Br^-:

a. Which ion will precipitate first?
b. What is the remaining concentration of the first ion when the second ion begins to precipitate?
c. Is the separation of F^- and Br^- feasible by fractional precipitation in this solution?

S17.43c

From the table of K_{sp} values (Table E3):

$$K_{sp}\ (CuCl) = 1.7 \times 10^{-7} \tag{16.10.13}$$

$$K_{sp}\ (CuBr) = 6.3 \times 10^{-9} \tag{16.10.14}$$

$CuBr$ will be the first to precipitate due to its smaller K_{sp} value.

$CuCl$ will precipitate when $Q_{sp} > K_{sp}$. First find the Cu concentration when $CuCl$ first precipitates.

$$1.7 \times 10^{-7} = 0.02[Cu^+] \tag{16.10.15}$$

$$[Cu^+] = 8.5 \times 10^{-7} \tag{16.10.16}$$

$$6.3 \times 10^{-9} = 8.5 \times 10^{-7}[Br^-] \tag{16.10.17}$$

$$0.007412 = [Br^-] \tag{16.10.18}$$

$[Br^-]$ drop from 0.3 M to 0.007412 M before $CuCl$ precipitate begins to form. This means that \frac{0.007412}{0.3}= 0.0247= 2.47\% \). This small percentage of Br^- ions that remain unreacted in the solution implies that fractional precipitation is feasible.

Q17.43d

If a solution contains $5.2 \times 10^{-3}\ M$ of Pb^{2+} and $4.8 \times 10^{-2}\ M$ of Cu^+. What concentration of I^- is needed to be added to start a precipitation reaction?

S17.43d

$1.10 \times 10^{-10}\ M$ of I^- is needed to begin precipitation.

Q17.59

a) Can $NaOH$ be separated by precipitating $NaCl$ but not HCl if all concentrations are 0.2 M?

b) Will the separation be complete?

S17.59

a) Yes

b) No

Q17.59

Which of the following solids are more soluble in an acidic solution than in pure water: $NaCl$, $Fe(OH)_2$, CaF_2, $ZnCO_3$, or $C_5H_4CO_2H$?

S17.59

- $NaCl$: They will have similar solubilities in both solvents because $NaCl$ does not produce any ions that will directly react with an acid.
- $Fe(OH)_2$: This will be more soluble in an acidic solution because the OH^- ions created will form water with H^+ ions present in the acidic solution.
- CaF_2: This will be more soluble in an acidic solution because the F^- ions created will react with the H^+ ions present in the acidic solution to increase the pH.
- $ZnCO_3$: This will be more soluble in an acidic solution because the CO_3 will lose an oxygen to form water with 2 protons already present in the acidic solution.
- $C_5H_4CO_2H$: This will be more soluble in pure water because in this molecule there are H_3O^+ ions that will not react with the acidic solution at all, and will therefore be more soluble in water.

Q17.59

Is $Ce(IO_3)_3$ able to precipitate from a solution that contains 500 mL of 0.10 M KIO_3 and 200 mL of 0.05 M $Ce(NO_3)_3$? ($K_{sp} = 1.9 \times 10^{-10}$)

S17.59

It's going to precipitate.

Q17.61

Can $Pb(OH)_2$ precipitate from a solution that contains $3.2 \times 10^{-3}\ M\ PbCl_2$ and $0.05\ M\ NH_3$?

S17.61

No precipitation would occur.

Q17.61

A saturated solution of $Ca_3(PO_4)_2$ has $[Ca^{2+}] = [PO_4^{3-}] = 2.9 \times 10^{-7}\ M$. Will $Ca_3(PO_4)_2$ precipitate?

S17.61

no

Q17.61

The solubility of $Ba(OH)_2$ in a particular buffer solution is $0.5 \frac{g}{L}$. What is the pH of the buffer solution? $K_{sp} = 5.0 \times 10^{-3}$

S17.61

First find $[Ba^{2+}]$ using $0.5 \frac{g}{L}$

$$\frac{0.5\ grams}{171.327\ \frac{g}{mol}} = 0.00584\ \frac{mol}{L} = [Ba^{2+}] \tag{16.10.19}$$

$$5 \times 10^{-3} = [Ba^{2+}][OH^-]^2 \tag{16.10.20}$$

$$[OH-] = 0.925 \tag{16.10.21}$$

$$pOH = -log(0.925) = 0.034 \tag{16.10.22}$$

$$pH = 13.9663 \tag{16.10.23}$$

Q17.81

The solubility of Na^{2+} in 2 M $CaCl$ is $1.3 \times 10^{-2} \frac{mol}{L}$. If K_{sp} of Na^{2+} is 3.1×10^{-10}, what is the value of K_f for $NaCl$?

S17.81

$K_f = 2.88 \times 10^{-2}$

Q17.81

If a 0.20 mol sample of $Zn(NH_3)_4SO_4$ is dissolved in 1.00L of an aqueous solution of 1.0 M NH_3, what's the concentration of $[Zn^{2+}]$? Hint: Use equilibrium constants in Table E4.

S17.81

Many complex ions such as tetrammine zinc (II), $Zn(NH_3)_4^{+2}$, are loosely aggregated and tend to dissociate in a water solution until an equilibrium is established between the complex ion and its components.

$$Zn(NH_3)_4^{+2} \rightleftharpoons Zn^{2+} + 4NH_3 \tag{16.10.24}$$

$$K = \frac{[Zn^{2+}][NH_3]^4}{[Zn(NH_3)^{+2}]^4} = 7.8 \times 10^8 \tag{16.10.25}$$

Video Solution

Q17.81

The solubility of $AlF_{3(s)}$ in 1.5 M $OH^-_{(aq)}$ is $2.0 \times 10^{-4} \frac{mol}{L}$. Given that K_{sp} of AlF_3 is 3.7×10^{-9}, what is the value of K_f for $Al(OH)_3$?

S17.81

To get the net reaction we combine the solubility product expression for $AlH_{3(s)}$ with the formation expression of $Al(OH)_{3(s)}$.

Solubility: $AlF_{3(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3F^-_{(aq)}$, $K_{sp} = 3.7 \times 10^{-9}$

Formation: $Al^{3+}_{(aq)} + 3OH^-_{(aq)} \rightleftharpoons Al(OH)_{3(aq)}$, $K_f = ?$

Net Reaction: $AlF_{3(s)} + 3OH^-_{(aq)} \rightleftharpoons Al(OH)_{3(aq)} + 3F^-_{(aq)}$, $K_{overall} = K_{sp}K_f$

$K_{overall} = \frac{[Al(OH)_3][F^-]^3}{[OH^-]^3} = \frac{(2.0 \times 10^{-4})^4}{(1.5)^3} = 4.74 \times 10^{-16} = K_f \times (3.7 \times 10^{-9})$

$K_f = \frac{4.74 \times 10^{-16}}{3.7 \times 10^{-9}} = 1.27 \times 10^{-7}$

Q17.83

A mixture of $PbSO_4(s)$ and $PbS_2O_3(s)$ is shaken with pure water until a saturated solution is formed. Both solids remain in excess. What is $[Pb^{2+}]$ in the saturated solution? For $PbSO_4$, $K_{sp} = 3.4 \times 10^{-7}$; for PbS_2O_3, $K_{sp} = 7.6 \times 10^{-4}$.

S17.83

$K_{sp}(PbSO_4) = [Pb^{2+}][SO_4^{2-}]$ with $[Pb^{2+}] = [SO_4^{2-}]$

So $K_{sp}(PbSO_4) = [Pb^{2+}]^2$

$K_{sp}(PbS_2O_3) = [Pb^{2+}][S_2O_3^{2-}]$ with $[Pb^{2+}] = [S_2O_3^{2-}]$

So $K_{sp}(PbS_2O_3) = [Pb^{2+}]^2$

$K_{sp}(PbSO4) \times K_{sp}(PbS_2O_3) = [Pb^{2+}]^4$

$(3.4 \times 10^{-7})(7.6 \times 10^{-4}) = [Pb^{2+}]^4$

$\mathbf{[Pb^{2+}] = .0040M}$

http://chemwiki.ucdavis.edu/Physical_Chemistry/Physical_Properties_of_Matter/Solutions/Solubilty/Relating_Solubility_and_Kspv

Q17.83

A mixture of $AlPO_4(s)$ and $AlPO_3(s)$ is shaken with water until a saturated solution is formed. Both of the solids remain in excess. What is $[Al^{3+}]$ in the saturated solution? For $AlPO_4$, $K_{sp} = 1.3 \times 10^{-6}$ and for $AlPO_3$, $K_{sp} = 3.5 \times 10^{-4}$.

S17.83

There are three relationships you need for this question:

$1.3 \times 10^{-6} = [Al^{3+}][PO_4^{3-}]$

$3.5 \times 10^{-4} = [Al^{3+}][PO_3^{3-}]$

The third is the electro neutrality equation, which states that the total positive charge concentration must equal the total negative charge concentration.

$[Al^{3+}] = [PO_4^{3-}] + [PO_3^{3-}]$

Use the first two expressions to solve for the concentration of each anion, then substitute them into the electroneutrality expression and solve for $[Al^{3+}]$.

$[PO_4^{3-}] = \frac{1.3 \times 10^{-6}}{[Al^{3+}]}$

$[PO_3^{3-}] = \frac{3.5 \times 10^{-4}}{[Al^{3+}]}$

$[Al^{3+}] = \frac{1.3 \times 10^{-6}}{[Al^{3+}]} + \frac{3.5 \times 10^{-4}}{[Al^{3+}]}$

$[Al^{3+}]^2 = 1.3 \times 10^{-6} + 3.5 \times 10^{-4} = 3.5 \times 10^{-4}$

$[Al^{3+}] = \sqrt{(3.5 \times 10^{-4}} = 0.019\ M$

Q17.83

A mixture of $CaSO_4$ and $AgSO_4$ is shaken with water until the solution is saturated. What is $[Ag^{2+}]$?

S17.83

0.0314 M

Q17.83 - (Answer is not given)

At what pH does $Cr(OH)_3$ just starts to precipitate from an 0.10 M Cr^{3+} solution? From this calculation justify why it is impossible to prepare an 0.10 M Cr^{3+} solution at pH=7. The K_{sp} of $Cr(OH)_3$ is 6.3×10^{-31}.

S17.83

This page titled 16.10: Exercises is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

17: Thermodynamics

This chapter explores the principles of thermochemistry, a branch of chemistry focused on the study of heat involved in chemical reactions and changes of state. Understanding thermochemical concepts is essential for grasping how energy is absorbed, released, and transformed during chemical processes. Previous chapters in this text have described various applications of thermochemistry, an important aspect of thermodynamics concerned with the heat flow accompanying chemical reactions and phase transitions. This chapter will introduce additional thermodynamic concepts, including those that enable the prediction of any chemical or physical changes under a given set of conditions.

As you delve into the concepts of thermochemistry in this chapter, consider how these principles are applied in cutting-edge research to address global energy challenges. Dr. Vonzyy's work illustrates the crucial role of thermochemistry in the innovation and development of sustainable energy solutions.

Faculty Feature

Dr. Alex Vonzyy's research focuses on thermochemistry, emphasizing the study of heat in chemical reactions and its application in energy solutions. His team develops advanced materials for Li-ion batteries, hydrogen storage, CO_2 capture, and photovoltaics, aligning with climate change mitigation and the Paris Agreement goals. Utilizing atomistic simulations, machine learning, and automated synthesis, they rapidly create and test materials to accelerate renewable energy adoption. Dr. Vonzyy's work highlights the critical role of thermochemical principles in designing innovative energy solutions, showcasing the connection between chemical energy transformations and global sustainability efforts.

Read more about Dr. Alex Voznyy and his research here: http://cleanenergy.utoronto.ca/



17.1: Introduction
17.2: Enthalpy
17.3: Spontaneity
17.4: Entropy
17.5: The Second and Third Laws of Thermodynamics
17.6: Free Energy
17.7: Key Terms
17.8: Key Equations
17.9: Summary
17.10: Exercises

This page titled 17: Thermodynamics is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

17.1: Introduction

Among the many capabilities of chemistry is its ability to predict if a process will occur under specified conditions. Thermodynamics, the study of relationships between the energy and work associated with chemical and physical processes, provides this predictive ability. Previous chapters in this text have described various applications of thermochemistry, an important aspect of thermodynamics concerned with the heat flow accompanying chemical reactions and phase transitions. This chapter will introduce additional thermodynamic concepts, including those that enable the prediction of any chemical or physical changes under a given set of conditions.



Figure 17.1.1: Geysers are a dramatic display of thermodynamic principles in nature. Water deep within the underground channels of the geyser is under high pressure and heated to high temperature by magma. When a pocket of water near the surface reaches boiling point and is expelled, the resulting drop in pressure causes larger volumes of water to flash boil, forcefully ejecting steam and water in an impressive eruption. (credit: modification of work by Yellowstone National Park)

This page titled 17.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.0: Introduction** is licensed CC BY 4.0.

17.2: Enthalpy

Learning Objectives

By the end of this section, you will be able to:

- State the first law of thermodynamics
- Define enthalpy and explain its classification as a state function
- Write and balance thermochemical equations
- Calculate enthalpy changes for various chemical reactions
- Explain Hess's law and use it to compute reaction enthalpies

Thermochemistry is a branch of chemical thermodynamics, the science that deals with the relationships between heat, work, and other forms of energy in the context of chemical and physical processes. As we concentrate on thermochemistry in this chapter, we need to consider some widely used concepts of thermodynamics.

Substances act as reservoirs of energy, meaning that energy can be added to them or removed from them. Energy is stored in a substance when the kinetic energy of its atoms or molecules is raised. The greater kinetic energy may be in the form of increased translations (travel or straight-line motions), vibrations, or rotations of the atoms or molecules. When thermal energy is lost, the intensities of these motions decrease and the kinetic energy falls. The total of all possible kinds of energy present in a substance is called the internal energy (U), sometimes symbolized as E.

As a system undergoes a change, its internal energy can change, and energy can be transferred from the system to the surroundings, or from the surroundings to the system. Energy is transferred into a system when it absorbs heat (q) from the surroundings or when the surroundings do work (w) on the system. For example, energy is transferred into room-temperature metal wire if it is immersed in hot water (the wire absorbs heat from the water), or if you rapidly bend the wire back and forth (the wire becomes warmer because of the work done on it). Both processes increase the internal energy of the wire, which is reflected in an increase in the wire's temperature. Conversely, energy is transferred out of a system when heat is lost from the system, or when the system does work on the surroundings.

The relationship between internal energy, heat, and work can be represented by the equation:

$$\Delta U = q + w$$

as shown in Figure 17.2.1. This is one version of the first law of thermodynamics, and it shows that the internal energy of a system changes through heat flow into or out of the system (positive q is heat flow in; negative q is heat flow out) or work done on or by the system. The work, w, is positive if it is done on the system and negative if it is done by the system.

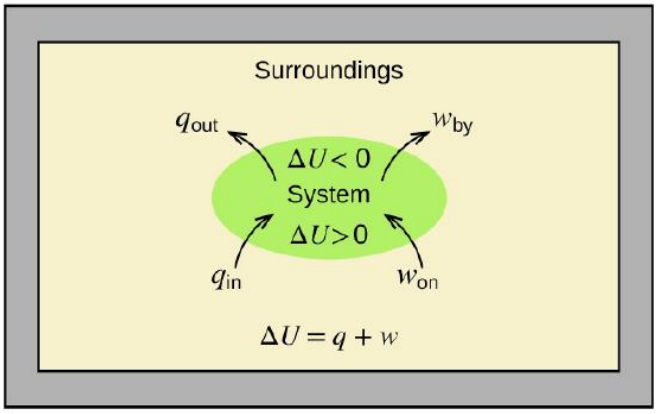


Figure 17.2.1: The internal energy, U, of a system can be changed by heat flow and work. If heat flows into the system, q_{in}, or work is done on the system, w_{on}, its internal energy increases, $\Delta U > 0$. If heat flows out of the system, q_{out}, or work is done by the system, w_{by}, its internal energy decreases, $\Delta U < 0$.

A type of work called expansion work (or pressure-volume work) occurs when a system pushes back the surroundings against a restraining pressure, or when the surroundings compress the system. An example of this occurs during the operation of an internal

combustion engine. The reaction of gasoline and oxygen is exothermic. Some of this energy is given off as heat, and some does work pushing the piston in the cylinder. The substances involved in the reaction are the system, and the engine and the rest of the universe are the surroundings. The system loses energy by both heating and doing work on the surroundings, and its internal energy decreases. (The engine is able to keep the car moving because this process is repeated many times per second while the engine is running.) We will consider how to determine the amount of work involved in a chemical or physical change in the chapter on thermodynamics.

Link to Learning

This view of an internal combustion engine illustrates the conversion of energy produced by the exothermic combustion reaction of a fuel such as gasoline into energy of motion.

As discussed, the relationship between internal energy, heat, and work can be represented as $\Delta U = q + w$. Internal energy is an example of a state function (or state variable), whereas heat and work are not state functions. The value of a state function depends only on the state that a system is in, and not on how that state is reached. If a quantity is not a state function, then its value *does* depend on how the state is reached. An example of a state function is altitude or elevation. If you stand on the summit of Mt. Kilimanjaro, you are at an altitude of 5895 m, and it does not matter whether you hiked there or parachuted there. The distance you traveled to the top of Kilimanjaro, however, is not a state function. You could climb to the summit by a direct route or by a more roundabout, circuitous path (Figure 17.2.2). The distances traveled would differ (distance is not a state function) but the elevation reached would be the same (altitude is a state function).

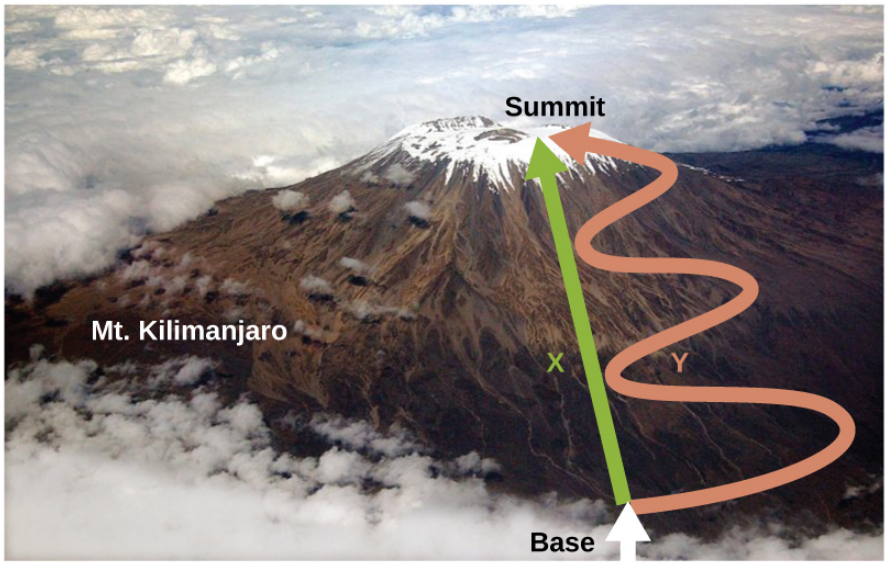


Figure 17.2.2: Paths X and Y represent two different routes to the summit of Mt. Kilimanjaro. Both have the same change in elevation (altitude or elevation on a mountain is a state function; it does not depend on path), but they have very different distances traveled (distance walked is not a state function; it depends on the path). (credit: modification of work by Paul Shaffner)

Chemists ordinarily use a property known as enthalpy (H) to describe the thermodynamics of chemical and physical processes. Enthalpy is defined as the sum of a system's internal energy (U) and the mathematical product of its pressure (P) and volume (V):

$$H = U + PV$$

Enthalpy is also a state function. Enthalpy values for specific substances cannot be measured directly; only enthalpy *changes* for chemical or physical processes can be determined. For processes that take place at constant pressure (a common condition for many chemical and physical changes), the enthalpy change (ΔH) is:

$$\Delta H = \Delta U + P\Delta V$$

The mathematical product $P\Delta V$ represents work (w), namely, expansion or pressure-volume work as noted. By their definitions, the arithmetic signs of ΔV and w will **always** be opposite:

$$P\Delta V = -w$$

Substituting this equation and the definition of internal energy into the enthalpy-change equation yields:

$$\begin{aligned}\Delta H &= \Delta U + P\Delta V \\ &= q_p + w - w \\ &= q_p\end{aligned}$$

where q_p is the **heat of reaction** under conditions of constant pressure.

And so, if a chemical or physical process is carried out at constant pressure with the only work done caused by expansion or contraction, then the heat flow (q_p) and enthalpy change (ΔH) for the process are equal.

The heat given off when you operate a Bunsen burner is equal to the enthalpy change of the methane combustion reaction that takes place, since it occurs at the essentially constant pressure of the atmosphere. On the other hand, the heat produced by a reaction measured in a bomb calorimeter is not equal to ΔH because the closed, constant-volume metal container prevents the pressure from remaining constant (it may increase or decrease if the reaction yields increased or decreased amounts of gaseous species). Chemists usually perform experiments under normal atmospheric conditions, at constant external pressure with $q = \Delta H$, which makes enthalpy the most convenient choice for determining heat changes for chemical reactions.

The following conventions apply when using ΔH:

- A negative value of an enthalpy change, $\Delta H < 0$, indicates an exothermic reaction; a positive value, $\Delta H > 0$, indicates an endothermic reaction. If the direction of a chemical equation is reversed, the arithmetic sign of its ΔH is changed (a process that is endothermic in one direction is exothermic in the opposite direction).
- Chemists use a thermochemical equation to represent the changes in both matter and energy. In a thermochemical equation, the enthalpy change of a reaction is shown as a ΔH value following the equation for the reaction. This ΔH value indicates the amount of heat associated with the reaction involving the number of moles of reactants and products *as shown in the chemical equation*. For example, consider this equation:

  $$\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)} \quad \Delta H = -286 kJ$$

  This equation indicates that when 1 mole of hydrogen gas and mole of oxygen gas at some temperature and pressure change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released to the surroundings. If the coefficients of the chemical equation are multiplied by some factor, the enthalpy change must be multiplied by that same factor (ΔH is an extensive property):

  - two-fold increase in amounts

    $$\mathrm{2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)} \quad \Delta H = 2 \times (-286 kJ) = -572 kJ$$

  - two-fold decrease in amounts

    $$\mathrm{\tfrac{1}{2}H_2(g) + \tfrac{1}{4}O_2(g) \longrightarrow \tfrac{1}{2}H_2O(l)} \quad \Delta H = \frac{1}{2} \times (-286 kJ) = -143 kJ$$

- The enthalpy change of a reaction depends on the physical states of the reactants and products, so these must be shown. For example, when 1 mole of hydrogen gas and mole of oxygen gas change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released. If gaseous water forms, only 242 kJ of heat are released.

  $$\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(g)} \quad \Delta H = -242 kJ$$

Example 17.2.1: Writing Thermochemical Equations

When 0.0500 mol of HCl(*aq*) reacts with 0.0500 mol of NaOH(*aq*) to form 0.0500 mol of NaCl(*aq*), 2.9 kJ of heat are produced. Write a balanced thermochemical equation for the reaction of one mole of HCl.

$$\mathrm{HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)}$$

Solution

For the reaction of 0.0500 mol acid (HCl), q = −2.9 kJ. The reactants are provided in stoichiometric amounts (same molar ratio as in the balanced equation), and so the amount of acid may be used to calculate a molar enthalpy change. Since ΔH is an extensive property, it is proportional to the amount of acid neutralized:

$$\Delta H = 1 \text{ mol HCl} \times \frac{-2.9kJ}{0.0500 \text{ mol HCl}} = -58kJ$$

The thermochemical equation is then

$$\text{HCl(aq)} + \text{NaOH(aq)} \longrightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \quad \Delta H = -58kJ$$

? Exercise 17.2.1

When 1.34 g Zn(*s*) reacts with 60.0 mL of 0.750 M HCl(*aq*), 3.14 kJ of heat are produced. Determine the enthalpy change per mole of zinc reacting for the reaction:

$$\text{Zn(s)} + 2\,\text{HCl(aq)} \longrightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{(g)}$$

Answer

ΔH = −153 kJ

Be sure to take both stoichiometry and limiting reactants into account when determining the ΔH for a chemical reaction.

✓ Example 17.2.2: Writing Thermochemical Equations

A gummy bear contains 2.67 g sucrose, $C_{12}H_{22}O_{11}$. When it reacts with 7.19 g potassium chlorate, $KClO_3$, 43.7 kJ of heat are produced. Write a thermochemical equation for the reaction of one mole of sucrose:

$$\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(aq)} + 8\,\text{KClO}_3\text{(aq)} \longrightarrow 12\,\text{CO}_2\text{(g)} + 11\,\text{H}_2\text{O(l)} + 8\,\text{KCl(aq)}$$

Solution

Unlike the previous example exercise, this one does not involve the reaction of stoichiometric amounts of reactants, and so the *limiting reactant* must be identified (it limits the yield of the reaction and the amount of thermal energy produced or consumed).

The provided amounts of the two reactants are

$$(2.67\,\text{g})\left(\frac{1\,\text{mol}}{342.3\,\text{g}}\right) = 0.00780\,\text{mol}\,\text{C}_{12}\text{H}_{22}\text{O}_{11}$$

$$(7.19\,\text{g})\left(\frac{1\,\text{mol}}{122.5\,\text{g}}\right) = 0.0587\,\text{mol}\,\text{KCIO}$$

The provided molar ratio of perchlorate-to-sucrose is then

$$\frac{0.0587\,\text{mol}\,\text{KClO}_3}{0.00780\,\text{mol}\,\text{C}_{12}\text{H}_{22}\text{O}_{11}} = 7.52$$

The balanced equation indicates 8 mol $KClO_3$ are required for reaction with 1 mol $C_{12}H_{22}O_{11}$. Since the provided amount of $KClO_3$ is less than the stoichiometric amount, it is the limiting reactant and may be used to compute the enthalpy change:

$$\Delta H = \frac{-43.7\,\text{kJ}}{0.0587\,\text{molKClO}_3} = \frac{744\,\text{kJ}}{\text{mol}\,\text{KClO}_3}$$

Because the equation, as written, represents the reaction of 8 mol $KClO_3$, the enthalpy change is

$$\left(\frac{744\,\text{kJ}}{1\ \text{mol}\,\text{KClO}_3}\right)\left(8\ \text{mol}\,\text{KClO}_3\right) = 5960\,\text{kJ}$$

The enthalpy change for this reaction is −5960 kJ, and the thermochemical equation is:

$$\text{C}_{12}\text{H}_{22}\text{O}_{11} + 8\,\text{KClO}_3 \longrightarrow 12\,\text{CO}_2 + 11\,\text{H}_2\text{O} + 8\,\text{KCl} \quad \Delta H = -5960\,\text{kJ}$$

Exercise 17.2.1

When 1.42 g of iron reacts with 1.80 g of chlorine, 3.22 g of $\mathrm{FeCl_2(s)}$ and 8.60 kJ of heat is produced. What is the enthalpy change for the reaction when 1 mole of $\mathrm{FeCl_2(s)}$ is produced?

Answer

$\Delta H = -338$ kJ

Enthalpy changes are typically tabulated for reactions in which both the reactants and products are at the same conditions. A standard state is a commonly accepted set of conditions used as a reference point for the determination of properties under other different conditions. For chemists, the IUPAC standard state refers to materials under a pressure of 1 bar and solutions at 1 M, and does not specify a temperature. Many thermochemical tables list values with a standard state of 1 atm. Because the ΔH of a reaction changes very little with such small changes in pressure (1 bar = 0.987 atm), ΔH values (except for the most precisely measured values) are essentially the same under both sets of standard conditions. We will include a superscripted "o" in the enthalpy change symbol to designate standard state. Since the usual (but not technically standard) temperature is 298.15 K, this temperature will be assumed unless some other temperature is specified. Thus, the symbol is used to indicate an enthalpy change for a process occurring under these conditions. (The symbol ΔH is used to indicate an enthalpy change for a reaction occurring under nonstandard conditions.)

The enthalpy changes for many types of chemical and physical processes are available in the reference literature, including those for combustion reactions, phase transitions, and formation reactions. As we discuss these quantities, it is important to pay attention to the *extensive* nature of enthalpy and enthalpy changes. Since the enthalpy change for a given reaction is proportional to the amounts of substances involved, it may be reported on that basis (i.e., as the ΔH for specific amounts of reactants). However, we often find it more useful to divide one extensive property (ΔH) by another (amount of substance), and report a per-amount *intensive* value of ΔH, often "normalized" to a per-mole basis. (Note that this is similar to determining the intensive property specific heat from the extensive property heat capacity, as seen previously.)

Standard Enthalpy of Combustion

Standard enthalpy of combustion is the enthalpy change when 1 mole of a substance burns (combines vigorously with oxygen) under standard state conditions; it is sometimes called "heat of combustion." For example, the enthalpy of combustion of ethanol, −1366.8 kJ/mol, is the amount of heat produced when one mole of ethanol undergoes complete combustion at 25 °C and 1 atmosphere pressure, yielding products also at 25 °C and 1 atm.

$$\mathrm{C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2 + 3\,H_2O(l)} \quad \Delta H^\circ = -1366.8\,\mathrm{kJ}$$

Enthalpies of combustion for many substances have been measured; a few of these are listed in Table 17.2.1. Many readily available substances with large enthalpies of combustion are used as fuels, including hydrogen, carbon (as coal or charcoal), and hydrocarbons (compounds containing only hydrogen and carbon), such as methane, propane, and the major components of gasoline.

Table 17.2.1: Standard Molar Enthalpies of Combustion

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH_c° (kJ/mol) at 25 °C
carbon	$C(s) + O_2(g) \longrightarrow CO_2(g)$	−393.5
hydrogen	$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$	−285.8
magnesium	$Mg(s) + \frac{1}{2}O_2(g) \longrightarrow MgO(s)$	−601.6
sulfur	$S(s) + O_2(g) \longrightarrow SO_2(g)$	−296.8
carbon monoxide	$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$	−283.0
methane	$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$	−890.8
acetylene	$C_2H_2(g) + \frac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(l)$	−1301.1

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH_c° (kJ/mol) at 25 °C
ethanol	$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$	−1366.8
methanol	$CH_3OH(l) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$	−726.1
isooctane	$C_8H_{18}(l) + \frac{25}{2}O_2(g) \longrightarrow 8CO_2(g) + 9H_2O(l)$	−5461

Example 17.2.3: Using Enthalpy of Combustion

As Figure 17.2.3 suggests, the combustion of gasoline is a highly exothermic process. Let us determine the approximate amount of heat produced by burning 1.00 L of gasoline, assuming the enthalpy of combustion of gasoline is the same as that of isooctane, a common component of gasoline. The density of isooctane is 0.692 g/mL.



Figure 17.2.3: The combustion of gasoline is very exothermic. (credit: modification of work by “AlexEagle”/Flickr)

Solution

Starting with a known amount (1.00 L of isooctane), we can perform conversions between units until we arrive at the desired amount of heat or energy. The enthalpy of combustion of isooctane provides one of the necessary conversions. Table 17.2.1 gives this value as −5460 kJ per 1 mole of isooctane (C_8H_{18}).

Using these data,

$$1.00\ \text{L}\,\cancel{\text{C}_8\text{H}_{18}} \times \left(\frac{1000\ \text{mL}\,\cancel{\text{C}_8\text{H}_{18}}}{1\ \text{L}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \left(\frac{0.692\ \text{g}\,\cancel{\text{C}_8\text{H}_{18}}}{1\ \text{mL}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \left(\frac{1\ \text{mol}\,\cancel{\text{C}_8\text{H}_{18}}}{114\ \text{g}\,\cancel{\text{C}_8\text{H}_{18}}}\right) \times \frac{-5460\ \text{kJ}}{1\ \text{mol}\,\cancel{\text{C}_8\text{H}_{18}}} = -3.31 \times 10^4\ \text{kJ}$$

The combustion of 1.00 L of isooctane produces 33,100 kJ of heat. (This amount of energy is enough to melt 99.2 kg, or about 218 lbs, of ice.)

Note: If you do this calculation one step at a time, you would find:

$$1.00\,\text{L}\,\text{C}_8\text{H}_{18} \longrightarrow 1.00 \times 10^3\ \text{mL}\text{C}_8\text{H}_{18}$$
$$1.00 \times 10^3\ \text{mL}\,\text{C}_8\text{H}_{18} \longrightarrow 692\,\text{g}\,\text{C}_8\text{H}_{18}$$
$$692\,\text{g}\,\text{C}_8\text{H}_{18} \longrightarrow 6.07\,\text{mol}\,\text{C}_8\text{H}_{18}$$
$$6.07\,\text{mol}\,\text{C}_8\text{H}_{18} \longrightarrow -3.31 \times 10^4 kJ$$

Exercise 17.2.3

How much heat is produced by the combustion of 125 g of acetylene?

Answer

6.25 10^3 kJ

https://chem.libretexts.org/@go/page/478968

Chemistry in Everyday Life: Emerging Algae-Based Energy Technologies (Biofuels)

As reserves of fossil fuels diminish and become more costly to extract, the search is ongoing for replacement fuel sources for the future. Among the most promising biofuels are those derived from algae (Figure 5.22). The species of algae used are nontoxic, biodegradable, and among the world's fastest growing organisms. About 50% of algal weight is oil, which can be readily converted into fuel such as biodiesel. Algae can yield 26,000 gallons of biofuel per hectare—much more energy per acre than other crops. Some strains of algae can flourish in brackish water that is not usable for growing other crops. Algae can produce biodiesel, biogasoline, ethanol, butanol, methane, and even jet fuel.



(a) (b) (c)

Figure 5.22 (a) Tiny algal organisms can be (b) grown in large quantities and eventually (c) turned into a useful fuel such as biodiesel. (credit a: modification of work by Micah Sittig; credit b: modification of work by Robert Kerton; credit c: modification of work by John F. Williams)

According to the US Department of Energy, only 39,000 square kilometers (about 0.4% of the land mass of the US or less than of the area used to grow corn) can produce enough algal fuel to replace all the petroleum-based fuel used in the US. The cost of algal fuels is becoming more competitive—for instance, the US Air Force is producing jet fuel from algae at a total cost of under $5 per gallon.[3] The process used to produce algal fuel is as follows: grow the algae (which use sunlight as their energy source and CO_2 as a raw material); harvest the algae; extract the fuel compounds (or precursor compounds); process as necessary (e.g., perform a transesterification reaction to make biodiesel); purify; and distribute (Figure 17.2.4).



Figure 17.2.4: Algae convert sunlight and carbon dioxide into oil that is harvested, extracted, purified, and transformed into a variety of renewable fuels.

Link to Learning

Click here to learn more about the process of creating algae biofuel.

Standard Enthalpy of Formation

A standard enthalpy of formation is an enthalpy change for a reaction in which exactly 1 mole of a pure substance is formed from free elements in their most stable states under standard state conditions. These values are especially useful for computing or predicting enthalpy changes for chemical reactions that are impractical or dangerous to carry out, or for processes for which it is difficult to make measurements. If we have values for the appropriate standard enthalpies of formation, we can determine the enthalpy change for any reaction, which we will practice in the next section on Hess's law.

The standard enthalpy of formation of $CO_2(g)$ is −393.5 kJ/mol. This is the enthalpy change for the exothermic reaction:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H_f^\circ = \Delta H^\circ = -393.5\ kJ}$$

starting with the reactants at a pressure of 1 atm and 25 °C (with the carbon present as graphite, the most stable form of carbon under these conditions) and ending with one mole of CO_2, also at 1 atm and 25 °C. For nitrogen dioxide, $NO_2(g)$, is 33.2 kJ/mol. This is the enthalpy change for the reaction:

$$\tfrac{1}{2}\mathrm{N_2(g)} + \mathrm{O_2(g)} \longrightarrow \mathrm{NO_2(g)} \quad \Delta H_f^\circ = \Delta H^\circ = +33.2\,\mathrm{kJ}$$

A reaction equation with mole of $\mathrm{N_2}$ and 1 mole of $\mathrm{O_2}$ is correct in this case because the standard enthalpy of formation always refers to 1 mole of product, $\mathrm{NO_2(g)}$.

You will find a table of standard enthalpies of formation of many common substances in Appendix G. These values indicate that formation reactions range from highly exothermic (such as −2984 kJ/mol for the formation of $\mathrm{P_4O_{10}}$) to strongly endothermic (such as +226.7 kJ/mol for the formation of acetylene, $\mathrm{C_2H_2}$). By definition, the standard enthalpy of formation of an element in its most stable form is equal to zero under standard conditions, which is 1 atm for gases and 1 M for solutions.

✔ Example 17.2.4: Evaluating an Enthalpy of Formation

Ozone, $\mathrm{O_3(g)}$, forms from oxygen, $\mathrm{O_2(g)}$, by an endothermic process. Ultraviolet radiation is the source of the energy that drives this reaction in the upper atmosphere. Assuming that both the reactants and products of the reaction are in their standard states, determine the standard enthalpy of formation, of ozone from the following information:

$$3\,\mathrm{O_2(g)} \longrightarrow 2\,\mathrm{O_3(g)} \quad \Delta H^\circ = +286\,\mathrm{kJ}$$

Solution

ΔH_f° is the enthalpy change for the formation of one mole of a substance in its standard state from the elements in their standard states. Thus, for $\mathrm{O_3(g)}$ is the enthalpy change for the reaction:

$$\tfrac{3}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{O_3(g)}$$

For the formation of 2 mol of $\mathrm{O_3(g)}$, $\Delta H^\circ = +286\,\mathrm{kJ}$. This ratio, can be used as a conversion factor to find the heat produced when 1 mole of $O_3(g)$ is formed, which is the enthalpy of formation for $\mathrm{O_3(g)}$:

$$\Delta H^\circ \text{ for 1 mole of } \mathrm{O_3}(g) = 1\ \mathrm{mol\ \cancel{O_3}} \times \frac{286\,\mathrm{kJ}}{2\,\mathrm{mol\ O_3}} = 143\,\mathrm{kJ}$$

Therefore,

$$\Delta H_f^\circ\,[O_3(g)] = +143\,\mathrm{kJ/mol}.$$

? Exercise 17.2.4

Hydrogen gas, $\mathrm{H_2}$, reacts explosively with gaseous chlorine, $\mathrm{Cl_2}$, to form hydrogen chloride, $\mathrm{HCl(g)}$. What is the enthalpy change for the reaction of 1 mole of $\mathrm{H_2(g)}$ with 1 mole of $\mathrm{Cl_2(g)}$ if both the reactants and products are at standard state conditions? The standard enthalpy of formation of $\mathrm{HCl(g)}$ is −92.3 kJ/mol.

Answer

For the reaction

$$\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl(g)} \quad \Delta H^\circ = -184.6\,\mathrm{kJ}$$

✔ Example 17.2.5: Writing Reaction Equations for

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OH(l)}$
b. $\mathrm{Ca_3(PO_4)_2(s)}$

Solution

Remembering that reaction equations are for forming 1 mole of the compound from its constituent elements under standard conditions, we have:

a. $$2\,\mathrm{C(s,\,graphite)} + 3\,\mathrm{H_2(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{C_2H_5OH(l)}$$

b. $$3\,\mathrm{Ca(s)} + \tfrac{1}{2}\mathrm{P_4(s)} + 4\,\mathrm{O_2(g)} \longrightarrow \mathrm{Ca_3(PO_4)_2(s)}$$

Note: The standard state of carbon is graphite, and phosphorus exists as $\mathrm{P_4}$.

Exercise 17.2.5

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OC_2H_5(l)}$
b. $\mathrm{Na_2CO_3(s)}$

Answer

a. $$4\,\mathrm{C(s,\,graphite)} + 5\,\mathrm{H_2(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{C_2H_5OC_2H_5(l)}$$

b. $$2\,\mathrm{Na(s)} + \mathrm{C(s,\,graphite)} + \tfrac{3}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{Na_2CO_3(s)}$$

Hess's Law

There are two ways to determine the amount of heat involved in a chemical change: measure it experimentally, or calculate it from other experimentally determined enthalpy changes. Some reactions are difficult, if not impossible, to investigate and make accurate measurements for experimentally. And even when a reaction is not hard to perform or measure, it is convenient to be able to determine the heat involved in a reaction without having to perform an experiment.

This type of calculation usually involves the use of Hess's law, which states: *If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps.* Hess's law is valid because enthalpy is a state function: Enthalpy changes depend only on where a chemical process starts and ends, but not on the path it takes from start to finish. For example, we can think of the reaction of carbon with oxygen to form carbon dioxide as occurring either directly or by a two-step process. The direct process is written:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)} \quad \Delta H^\circ = -394kJ$$

In the two-step process, first carbon monoxide is formed:

$$\mathrm{C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g)} \quad \Delta H^\circ = -111kJ$$

Then, carbon monoxide reacts further to form carbon dioxide:

$$\mathrm{CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g)} \quad \Delta H^\circ = -283kJ$$

The equation describing the overall reaction is the sum of these two chemical changes:

step 1	$\mathrm{C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g)}$
step 2	$\mathrm{CO(g) + \tfrac{1}{2}O_2(g) \longrightarrow CO_2(g)}$
sum	$\mathrm{C(s) + \cancel{CO(g)} + O_2(g) \longrightarrow \cancel{CO(g)} + CO_2(g)}$

Because the $\mathrm{CO(g)}$ produced in Step 1 is consumed in Step 2, the net change is:

$$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)}$$

According to Hess's law, the enthalpy change of the reaction will equal the sum of the enthalpy changes of the steps.

step 1	$\mathrm{C(s) + \frac{1}{2}O_2(g) \longrightarrow \cancel{CO(g)}}$	$\Delta H^\circ = -111\ \mathrm{kJ}$
step 2	$\mathrm{\cancel{CO(g)} + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)}$	$\Delta H^\circ = -283\ \mathrm{kJ}$
sum	$\mathrm{C(s) + O_2(g) \longrightarrow CO_2(g)}$	$\Delta H^\circ = -394\ \mathrm{kJ}$

The result is shown in Figure 17.2.5. We see that ΔH of the overall reaction is the same whether it occurs in one step or two. This finding (overall ΔH for the reaction = sum of ΔH values for reaction "steps" in the overall reaction) is true in general for chemical and physical processes.

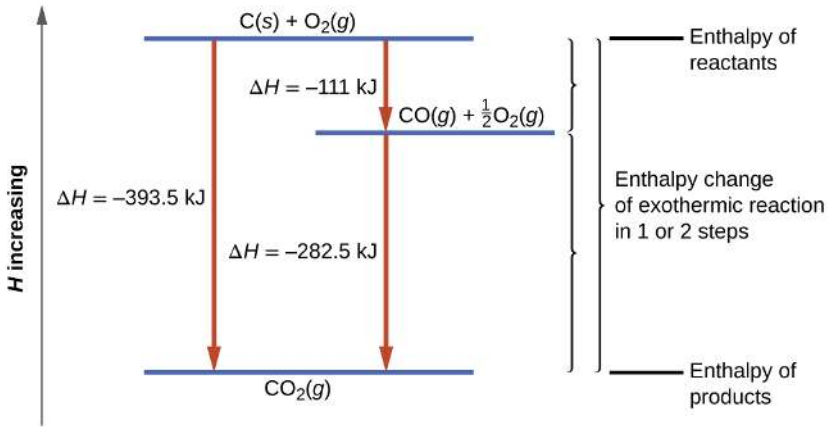


Figure 17.2.5: The formation of $\mathrm{CO_2(g)}$ from its elements can be thought of as occurring in two steps, which sum to the overall reaction, as described by Hess's law. The horizontal blue lines represent enthalpies. For an exothermic process, the products are at lower enthalpy than are the reactants.

Before we further practice using Hess's law, let us recall two important features of ΔH.

1. ΔH is directly proportional to the quantities of reactants or products. For example, the enthalpy change for the reaction forming 1 mole of $NO_2(g)$ is +33.2 kJ:

$$\mathrm{\frac{1}{2}N_2(g) + O_2(g) \longrightarrow NO_2(g)} \quad \Delta H = +33.2\ \mathrm{kJ}$$

When 2 moles of $\mathrm{NO_2}$ (twice as much) are formed, the ΔH will be twice as large:

$$\mathrm{N_2(g) + 2\,O_2(g) \longrightarrow 2\,NO_2(g)} \quad \Delta H = +66.4\ \mathrm{kJ}$$

In general, if we multiply or divide an equation by a number, then the enthalpy change should also be multiplied or divided by the same number.

2. ΔH for a reaction in one direction is equal in magnitude and opposite in sign to ΔH for the reaction in the reverse direction. For example, given that:

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)} \quad \Delta H = -184.6\ \mathrm{kJ}$$

Then, for the "reverse" reaction, the enthalpy change is also "reversed":

Example 17.2.6: Stepwise Calculation of Heat of Formation Using Hess's Law

Determine the enthalpy of formation, of $\mathrm{FeCl_3(s)}$ from the enthalpy changes of the following two-step process that occurs under standard state conditions:

$$\mathrm{Fe(s) + Cl_2(g) \longrightarrow FeCl_2(s)} \quad \Delta H^\circ = -341.8\ \mathrm{kJ}$$

$$\mathrm{FeCl_2(s) + \frac{1}{2}Cl_2(g) \longrightarrow FeCl_3(s)} \quad \Delta H^\circ = -57.7\ \mathrm{kJ}$$

Solution

We are trying to find the standard enthalpy of formation of $\mathrm{FeCl_3(s)}$, which is equal to ΔH° for the reaction:

$$\mathrm{Fe(s) + \frac{3}{2}Cl_2(g) \longrightarrow FeCl_3(s)} \quad \Delta H_f^\circ = ?$$

Looking at the reactions, we see that the reaction for which we want to find ΔH° is the sum of the two reactions with known ΔH values, so we must sum their ΔH values:

step 1	$\mathrm{Fe(s) + Cl_2(g) \longrightarrow \cancel{FeCl_2(s)}}$	$\Delta H^\circ = -341.8\,\mathrm{kJ}$
step 2	$\mathrm{\cancel{FeCl_2(s)} + \frac{1}{2}Cl_2(g) \longrightarrow FeCl_3(s)}$	$\Delta H^\circ = -57.7\,\mathrm{kJ}$
sum	$\mathrm{Fe(s) + \frac{3}{2}Cl_2(g) \longrightarrow FeCl_3(s)}$	$\Delta H^\circ = -399.5\,\mathrm{kJ}$

The enthalpy of formation, ΔH_f° of $\mathrm{FeCl_3(s)}$ is −399.5 kJ/mol.

Exercise 17.2.6

Calculate ΔH for the process:

$$\mathrm{N_2(g) + 2\,O_2(g) \longrightarrow 2\,NO_2(g)}$$

from the following information:

$$\mathrm{N_2(g) + O_2(g) \longrightarrow 2\,NO(g)} \quad \Delta H = 180.5\,\mathrm{kJ}$$

$$\mathrm{NO(g) + \tfrac{1}{2}O_2(g) \longrightarrow NO_2(g)} \quad \Delta H = -57.06\,\mathrm{kJ}$$

Answer

66.4 kJ

Here is a less straightforward example that illustrates the thought process involved in solving many Hess's law problems. It shows how we can find many standard enthalpies of formation (and other values of ΔH) if they are difficult to determine experimentally.

Example 17.2.7: A More Challenging Problem Using Hess's Law

Chlorine monofluoride can react with fluorine to form chlorine trifluoride:

i. $$\mathrm{ClF(g) + F_2(g) \longrightarrow ClF_3(g)} \quad \Delta H^\circ = ?$$

Use the reactions here to determine the ΔH^o for reaction (i):

ii. $$\mathrm{2\,OF_2(g) \longrightarrow O_2(g) + 2\,F_2(g)} \quad \Delta H^\circ_{(ii)} = -49.4\,\mathrm{kJ}$$

iii. $$\mathrm{2\,ClF(g) + O_2(g) \longrightarrow Cl_2O(g) + OF_2(g)} \quad \Delta H^\circ_{(iii)} = +214.0\,\mathrm{kJ}$$

iv. $$\mathrm{ClF_3(g) + O_2(g) \longrightarrow \tfrac{1}{2}Cl_2O(g) + \tfrac{3}{2}OF_2(g)} \quad \Delta H^\circ_{(in)} = +236.2\,\mathrm{kJ}$$

Solution

Our goal is to manipulate and combine reactions (ii), (iii), and (iv) such that they add up to reaction (i). Going from left to right in (i), we first see that ClF(g) is needed as a reactant. This can be obtained by multiplying reaction (iii) by $\frac{1}{2}$ which means that the ΔH^o change is also multiplied by $\frac{1}{2}$:

$$\mathrm{ClF(g) + \tfrac{1}{2}O_2(g) \longrightarrow \tfrac{1}{2}Cl_2O(g) + \tfrac{1}{2}OF_2(g)} \quad \Delta H^\circ = \frac{1}{2}(214.0) = +107.0\,\mathrm{kJ}$$

Next, we see that F_2 is also needed as a reactant. To get this, reverse and halve reaction *(ii)*, which means that the ΔH^o changes sign and is halved:

$$\mathrm{\tfrac{1}{2}O_2(g) + F_2(g) \longrightarrow OF_2(g)} \quad \Delta H^\circ = +24.7\,\mathrm{kJ}$$

To get $\mathrm{ClF_3}$ as a product, reverse (iv), changing the sign of ΔH^o:

$$\mathrm{\tfrac{1}{2}Cl_2O(g) + \tfrac{3}{2}OF_2(g) \longrightarrow ClF_3(g) + O_2(g)} \quad \Delta H^\circ = -236.2kJ$$

Now check to make sure that these reactions add up to the reaction we want:

step 1	$\mathrm{ClF(g)} + \cancel{\frac{1}{2}\mathrm{O_2(g)}} \longrightarrow \cancel{\frac{1}{2}\mathrm{Cl_2O(g)}} + \cancel{\frac{1}{2}\mathrm{OF_2(g)}}$	$\Delta H^\circ = +107.0\,\mathrm{kJ}$
step 2	$\cancel{\frac{1}{2}\mathrm{O_2(g)}} + \mathrm{F_2(g)} \longrightarrow \cancel{\mathrm{OF_2(g)}}$	$\Delta H^\circ = +24.7\,\mathrm{kJ}$
step 3	$\cancel{\frac{1}{2}\mathrm{Cl_2O(g)}} + \cancel{\frac{3}{2}\mathrm{OF_2(g)}} \longrightarrow \mathrm{ClF_3(g)} + \cancel{\mathrm{O_2(g)}}$	$\Delta H^\circ = -236.2\,\mathrm{kJ}$
sum	$\mathrm{ClF(g)} + \mathrm{F_2(g)} \longrightarrow \mathrm{ClF_3(g)}$	$\Delta H^\circ = -104.5\,\mathrm{kJ}$

Reactants $\frac{1}{2}\mathrm{O_2}$ and $\frac{1}{2}\mathrm{O_2}$ cancel out the product $\mathrm{O_2}$; the $\frac{1}{2}\mathrm{Cl_2O}$ product cancels $\frac{1}{2}\mathrm{Cl_2O}$ reactant and $\frac{3}{2}\mathrm{Cl_2O}$ reactant is cancelled by products $\frac{1}{2}\mathrm{OF_2}$ and $\mathrm{OF_2}$. This leaves only reactants $\mathrm{ClF(g)}$ and $\mathrm{F_2(g)}$ and product $\mathrm{ClF_3(g)}$, which are what we want. Since summing these three modified reactions yields the reaction of interest, summing the three modified ΔH^o values will give the desired ΔH^o:

$$\Delta H^\circ = (+107.0kJ) + (24.7kJ) + (-236.2kJ) = -104.5\,\mathrm{kJ}$$

Exercise 17.2.7

Aluminum chloride can be formed from its elements:

i. $$2\,\mathrm{Al(s)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{AlCl_3(s)} \quad \Delta H^\circ = ?$$

Use the reactions here to determine the ΔH^o for reaction *(i)*:

ii. $$\mathrm{HCl(g)} \longrightarrow \mathrm{HCl(aq)} \quad \Delta H^\circ_{(ii)} = -74.8\,\mathrm{kJ}$$

iii. $$\mathrm{H_2(g)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{HCl(g)} \quad \Delta H^\circ_{(iii)} = -185\,\mathrm{kJ}$$

iv. $$\mathrm{AlCl_3(aq)} \longrightarrow \mathrm{AlCl_3(s)} \quad \Delta H^\circ_{(iv)} = +323\,\mathrm{kJ}$$

v. $$2\,\mathrm{Al(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{AlCl_3(aq)} + 3\,\mathrm{H_2(g)} \quad \Delta H^\circ_{(v)} = -1049\,\mathrm{kJ}$$

Answer

−1407 kJ

We also can use Hess's law to determine the enthalpy change of any reaction if the corresponding enthalpies of formation of the reactants and products are available. The stepwise reactions we consider are:

i. decompositions of the reactants into their component elements (for which the enthalpy changes are proportional to the negative of the enthalpies of formation of the reactants), followed by
ii. re-combinations of the elements to give the products (with the enthalpy changes proportional to the enthalpies of formation of the products).

The standard enthalpy change of the overall reaction is therefore equal to: (ii) the sum of the standard enthalpies of formation of all the products plus (i) the sum of the negatives of the standard enthalpies of formation of the reactants. This is usually rearranged slightly to be written as follows, with ∑ representing "the sum of" and *n* standing for the stoichiometric coefficients:

$$\Delta H^\circ_{\text{reaction}} = \sum n \times \Delta H^\circ_f(\text{ products }) - \sum n \times \Delta H^\circ_f(\text{ reactants })$$

The following example shows in detail why this equation is valid, and how to use it to calculate the enthalpy change for a reaction of interest.

Example 17.2.8: Using Hess's Law

What is the standard enthalpy change for the reaction:

$$3\,\mathrm{NO_2(g)} + \mathrm{H_2O(l)} \longrightarrow 2\,\mathrm{HNO_3(aq)} + \mathrm{NO(g)} \quad \Delta H^\circ = ?$$

Solution: Using the Equation

Use the special form of Hess's law given previously, and values from Appendix G:

$$\begin{aligned}\Delta H^\circ_{\text{reaction}} &= \sum n \times \Delta H^\circ_f(\text{ products }) - \sum n \times \Delta H^\circ_f(\text{ reactants }) \\ &= \left[2\ \text{mol}\,\cancel{\text{HNO}_3\text{(aq)}} \times \frac{-207.4\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{HNO}_3\text{(aq)}}} + 1\ \text{mol}\,\cancel{\text{NO(g)}} \times \frac{+90.2\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{NO(g)}}}\right] \\ &\quad - \left[3\ \text{mol}\,\cancel{\text{NO}_2\text{(g)}} \times \frac{+33.2\,\text{kJ}}{1\ \text{mole}\,\cancel{\text{NO}_2\text{(g)}}} + 1\ \text{mol}\,\cancel{\text{H}_2\text{O(l)}} \times \frac{-285.8\,\text{kJ}}{1\ \text{mol}\,\cancel{\text{H}_2\text{O(l)}}}\right] \\ &= [2 \times (-206.64) + 90.25] - [3 \times 33.2 + -(-285.83)] \\ &= -323.03 + 186.23 \\ &= -136.80\,\text{kJ}\end{aligned}$$

Solution: Supporting Why the General Equation is Valid

Alternatively, we can write this reaction as the sum of the decompositions of $3\,\text{NO}_2\text{(g)}$ and $H2O(l)$ into their constituent elements, and the formation of $2\,\text{HNO}_3\text{(aq)}$ and $1\,\text{NO(g)}$ from their constituent elements. Writing out these reactions, and noting their relationships to the ΔH° values for these compounds (from Appendix G), we have:

$$\begin{aligned} 3\,\text{NO}_2\text{(g)} &\longrightarrow \tfrac{3}{2}\text{N}_2\text{(g)} + 3\,\text{O}_2\text{(g)} && \Delta H^\circ_1 = -99.6\,\text{kJ} \\ \text{H}_2\text{O(l)} &\longrightarrow \text{H}_2\text{(g)} + \tfrac{1}{2}\text{O}_2\text{(g)} && \Delta H^\circ_2 = +285.8\,\text{kJ}\left[-1 \times \Delta H^\circ_f(\text{H}_2\text{O})\right] \\ \text{H}_2\text{(g)} + \text{N}_2\text{(g)} + 3\,\text{O}_2\text{(g)} &\longrightarrow 2\,\text{HNO}_3\text{(aq)} && \Delta H^\circ_3 = -414.8\,\text{kJ}\left[2 \times \Delta H^\circ_f(\text{HNO}_3)\right] \\ \tfrac{1}{2}\text{N}_2\text{(g)} + \tfrac{1}{2}\text{O}_2\text{(g)} &\longrightarrow \text{NO(g)} && \Delta H^\circ_4 = +90.2\,\text{kJ}\left[1 \times \Delta H^\circ_f(\text{NO})\right] \end{aligned}$$

Summing these reaction equations gives the reaction we are interested in:

$$3\,\text{NO}_2\text{(g)} + \text{H}_2\text{O(l)} \longrightarrow 2\,\text{HNO}_3\text{(aq)} + \text{NO(g)}$$

Summing their enthalpy changes gives the value we want to determine:

$$\begin{aligned}\Delta H^\circ_{rxn} &= \Delta H^\circ_1 + \Delta H^\circ_2 + \Delta H^\circ_3 + \Delta H^\circ_4 \\ &= (-99.6\,\text{kJ}) + (+285.8\,\text{kJ}) + (-414.8\,\text{kJ}) + (+90.2\,\text{kJ}) \\ &= -138.4\,\text{kJ}\end{aligned}$$

So the standard enthalpy change for this reaction is ΔH° = −138.4 kJ.

Note that this result was obtained by:

a. multiplying the of each product by its stoichiometric coefficient and summing those values,
b. multiplying the of each reactant by its stoichiometric coefficient and summing those values, and then
c. subtracting the result found in (2) from the result found in (1).

This is also the procedure in using the general equation, as shown.

? Exercise 17.2.8

Calculate the heat of combustion of 1 mole of ethanol, $\text{C}_2\text{H}_5\text{OH(l)}$, when $\text{H}_2\text{O(l)}$ and $\text{CO}_2\text{(g)}$ are formed. Use the following enthalpies of formation: $\text{C}_2\text{H}_5\text{OH(l)}$, −278 kJ/mol; $\text{H}_2\text{O(l)}$, −286 kJ/mol; and $\text{CO}_2\text{(g)}$, −394 kJ/mol.

Answer

−1368 kJ/mol

Footnotes

- 3For more on algal fuel, see http://www.theguardian.com/environme...n-fuel-problem.

This page titled 17.2: Enthalpy is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **5.3: Enthalpy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

17.3: Spontaneity

Learning Objectives

By the end of this section, you will be able to:

- Distinguish between spontaneous and nonspontaneous processes
- Describe the dispersal of matter and energy that accompanies certain spontaneous processes

Processes have a natural tendency to occur in one direction under a given set of conditions. Water will naturally flow downhill, but uphill flow requires outside intervention such as the use of a pump. Iron exposed to the earth's atmosphere will corrode, but rust is not converted to iron without intentional chemical treatment. A **spontaneous process** is one that occurs naturally under certain conditions. A **nonspontaneous process**, on the other hand, will not take place unless it is "driven" by the continual input of energy from an external source. A process that is spontaneous in one direction under a particular set of conditions is nonspontaneous in the reverse direction. At room temperature and typical atmospheric pressure, for example, ice will spontaneously melt, but water will not spontaneously freeze.

The spontaneity of a process is *not* correlated to the speed of the process. A spontaneous change may be so rapid that it is essentially instantaneous or so slow that it cannot be observed over any practical period of time. To illustrate this concept, consider the decay of radioactive isotopes, a topic more thoroughly treated in the chapter on nuclear chemistry. Radioactive decay is by definition a spontaneous process in which the nuclei of unstable isotopes emit radiation as they are converted to more stable nuclei. All the decay processes occur spontaneously, but the rates at which different isotopes decay vary widely. Technetium-99m is a popular radioisotope for medical imaging studies that undergoes relatively rapid decay and exhibits a half-life of about six hours. Uranium-238 is the most abundant isotope of uranium, and its decay occurs much more slowly, exhibiting a half-life of more than four billion years (Figure 17.3.1).

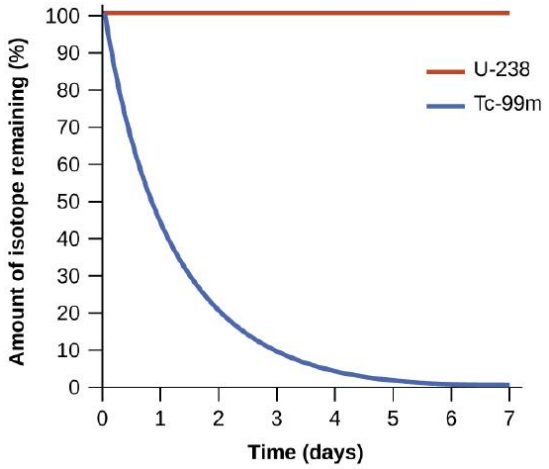


Figure 17.3.1: Both U-238 and Tc-99m undergo spontaneous radioactive decay, but at drastically different rates. Over the course of one week, essentially all of a Tc-99m sample and none of a U-238 sample will have decayed.

As another example, consider the conversion of diamond into graphite (Figure 17.3.2).

$$C(s, \text{diamond}) \longrightarrow C(s, \text{graphite})$$

The phase diagram for carbon indicates that graphite is the stable form of this element under ambient atmospheric pressure, while diamond is the stable allotrope at very high pressures, such as those present during its geologic formation. Thermodynamic calculations of the sort described in the last section of this chapter indicate that the conversion of diamond to graphite at ambient pressure occurs spontaneously, yet diamonds are observed to exist, and persist, under these conditions. Though the process is spontaneous under typical ambient conditions, its rate is extremely slow; so, for all practical purposes diamonds are indeed "forever." Situations such as these emphasize the important distinction between the thermodynamic and the kinetic aspects of a process. In this particular case, diamonds are said to be *thermodynamically unstable* but *kinetically stable* under ambient conditions.

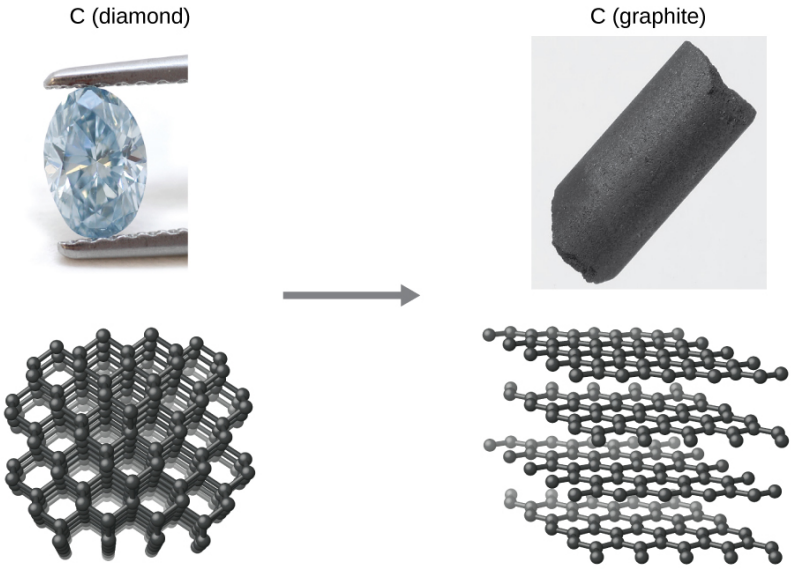


Figure 17.3.2: The conversion of carbon from the diamond allotrope to the graphite allotrope is spontaneous at ambient pressure, but its rate is immeasurably slow at low to moderate temperatures. This process is known as *graphitization*, and its rate can be increased to easily measurable values at temperatures in the 1000–2000 K range. (credit "diamond" photo: modification of work by "Fancy Diamonds"/Flickr; credit "graphite" photo: modification of work by images-of-elements.com/carbon.php)

Dispersal of Matter and Energy

Extending the discussion of thermodynamic concepts toward the objective of predicting spontaneity, consider now an isolated system consisting of two flasks connected with a closed valve. Initially there is an ideal gas in one flask and the other flask is empty ($P = 0$) (Figure 17.3.3). When the valve is opened, the gas spontaneously expands to fill both flasks equally. Recalling the definition of pressure-volume work from the chapter on thermochemistry, note that no work has been done because the pressure in a vacuum is zero.

$$w = -P\Delta V = 0 \qquad (P = 0 \text{ in a vacuum}) \tag{17.3.1}$$

Note as well that since the system is isolated, no heat has been exchanged with the surroundings (q = 0). The *first law of thermodynamics* confirms that there has been no change in the system's internal energy as a result of this process.

$$\Delta U = q + w = 0 + 0 = 0$$

The spontaneity of this process is therefore not a consequence of any change in energy that accompanies the process. Instead, the driving force appears to be related to the *greater, more uniform dispersal of matter* that results when the gas is allowed to expand. Initially, the system was comprised of one flask containing matter and another flask containing nothing. After the spontaneous expansion took place, the matter was distributed both more widely (occupying twice its original volume) and more uniformly (present in equal amounts in each flask).

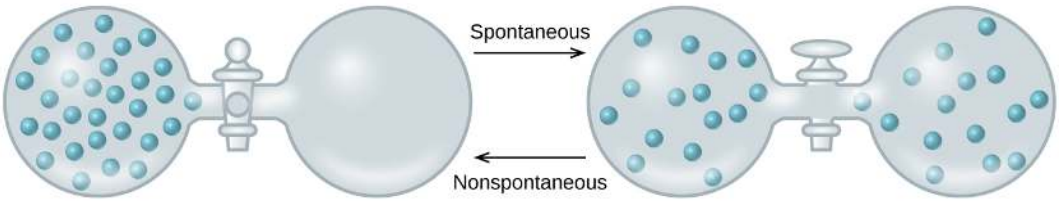


Figure 17.3.3: An isolated system consists of an ideal gas in one flask that is connected by a closed valve to a second flask containing a vacuum. Once the valve is opened, the gas spontaneously becomes evenly distributed between the flasks.

Now consider two objects at different temperatures: object X at temperature T_X and object Y at temperature T_Y, with $T_X > T_Y$ (Figure 17.3.4). When these objects come into contact, heat spontaneously flows from the hotter object (X) to the colder one (Y). This corresponds to a loss of thermal energy by X and a gain of thermal energy by Y.

$$q_X < 0 \quad \text{and} \quad q_Y = -q_X > 0$$

From the perspective of this two-object system, there was no net gain or loss of thermal energy, rather the available thermal energy was redistributed among the two objects. This spontaneous process resulted in a *more uniform dispersal of energy*.



Figure 17.3.4: When two objects at different temperatures come in contact, heat spontaneously flows from the hotter to the colder object.

As illustrated by the two processes described, an important factor in determining the spontaneity of a process is the extent to which it changes the dispersal or distribution of matter and/or energy. In each case, a spontaneous process took place that resulted in a more uniform distribution of matter or energy.

Example 17.3.1: Redistribution of Matter during a Spontaneous Process

Describe how matter is redistributed when the following spontaneous processes take place:

a. A solid sublimes.
b. A gas condenses.
c. A drop of food coloring added to a glass of water forms a solution with uniform color.

Solution



Figure 17.3.5: (credit a: modification of work by Jenny Downing; credit b: modification of work by "Fuzzy Gerdes"/Flickr; credit c: modification of work by Paul A. Flowers)

a. Sublimation is the conversion of a solid (relatively high density) to a gas (much lesser density). This process yields a much greater dispersal of matter, since the molecules will occupy a much greater volume after the solid-to-gas transition.
b. Condensation is the conversion of a gas (relatively low density) to a liquid (much greater density). This process yields a much lesser dispersal of matter, since the molecules will occupy a much lesser volume after the gas-to-liquid transition.
c. The process in question is diffusion. This process yields a more uniform dispersal of matter, since the initial state of the system involves two regions of different dye concentrations (high in the drop of dye, zero in the water), and the final state of the system contains a single dye concentration throughout.

Exercise 17.3.1

Describe how energy is redistributed when a spoon at room temperature is placed in a cup of hot coffee.

Answer

Heat will spontaneously flow from the hotter object (coffee) to the colder object (spoon), resulting in a more uniform distribution of thermal energy as the spoon warms and the coffee cools.

This page titled 17.3: Spontaneity is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.1: Spontaneity** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

17.4: Entropy

Learning Objectives

By the end of this section, you will be able to:

- Define entropy
- Explain the relationship between entropy and the number of microstates
- Predict the sign of the entropy change for chemical and physical processes

In 1824, at the age of 28, Nicolas Léonard Sadi Carnot (Figure 17.4.1) published the results of an extensive study regarding the efficiency of steam heat engines. A later review of Carnot's findings by Rudolf Clausius introduced a new thermodynamic property that relates the spontaneous heat flow accompanying a process to the temperature at which the process takes place. This new property was expressed as the ratio of the *reversible* heat (q_{rev}) and the kelvin temperature (T). In thermodynamics, a **reversible process** is one that takes place at such a slow rate that it is always at equilibrium and its direction can be changed (it can be "reversed") by an infinitesimally small change in some condition. Note that the idea of a reversible process is a formalism required to support the development of various thermodynamic concepts; no real processes are truly reversible, rather they are classified as *irreversible*.

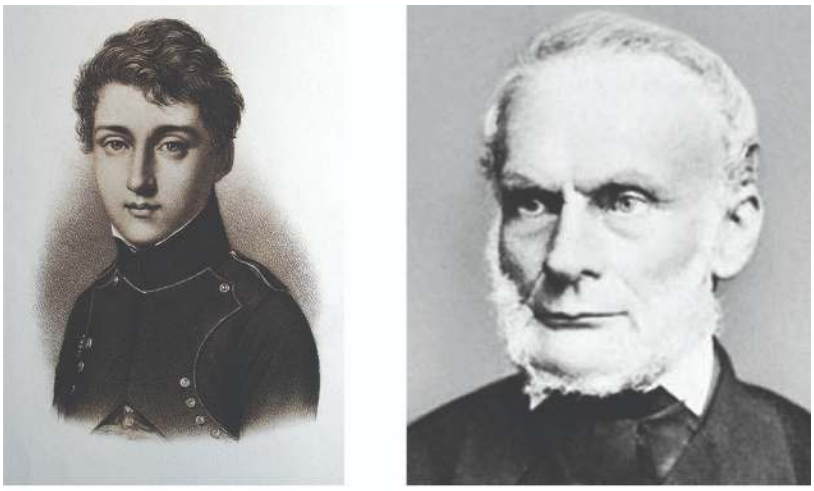


Figure 17.4.1: (a) Nicholas Léonard Sadi Carnot's research into steam-powered machinery and (b) Rudolf Clausius's later study of those findings led to groundbreaking discoveries about spontaneous heat flow processes.

Similar to other thermodynamic properties, this new quantity is a state function, so its change depends only upon the initial and final states of a system. In 1865, Clausius named this property entropy (S) and defined its change for any process as the following:

$$\Delta S = \frac{q_{rev}}{T}$$

The entropy change for a real, irreversible process is then equal to that for the theoretical reversible process that involves the same initial and final states.

Entropy and Microstates

Following the work of Carnot and Clausius, Ludwig Boltzmann developed a molecular-scale statistical model that related the entropy of a system to the *number of microstates* (W) possible for the system. A **microstate** is a specific configuration of all the locations and energies of the atoms or molecules that make up a system. The relation between a system's entropy and the number of possible microstates is

$$S = k \ln W$$

where k is the Boltzmann constant, 1.38 10^{-23} J/K.

As for other state functions, the change in entropy for a process is the difference between its final (S_f) and initial (S_i) values:

$$\Delta S = S_f - S_i = k \ln W_f - k \ln W_i = k \ln \frac{W_f}{W_i}$$

For processes involving an increase in the number of microstates, $W_f > W_i$, the entropy of the system increases and $\Delta S > 0$. Conversely, processes that reduce the number of microstates, $W_f < W_i$, yield a decrease in system entropy, $\Delta S < 0$. This molecular-scale interpretation of entropy provides a link to the probability that a process will occur as illustrated in the next paragraphs.

Consider the general case of a system comprised of N particles distributed among n boxes. The number of microstates possible for such a system is n^N. For example, distributing four particles among two boxes will result in $2^4 = 16$ different microstates as illustrated in Figure 17.4.2 Microstates with equivalent particle arrangements (not considering individual particle identities) are grouped together and are called *distributions*. The probability that a system will exist with its components in a given distribution is proportional to the number of microstates within the distribution. Since entropy increases logarithmically with the number of microstates, *the most probable distribution is therefore the one of greatest entropy*.

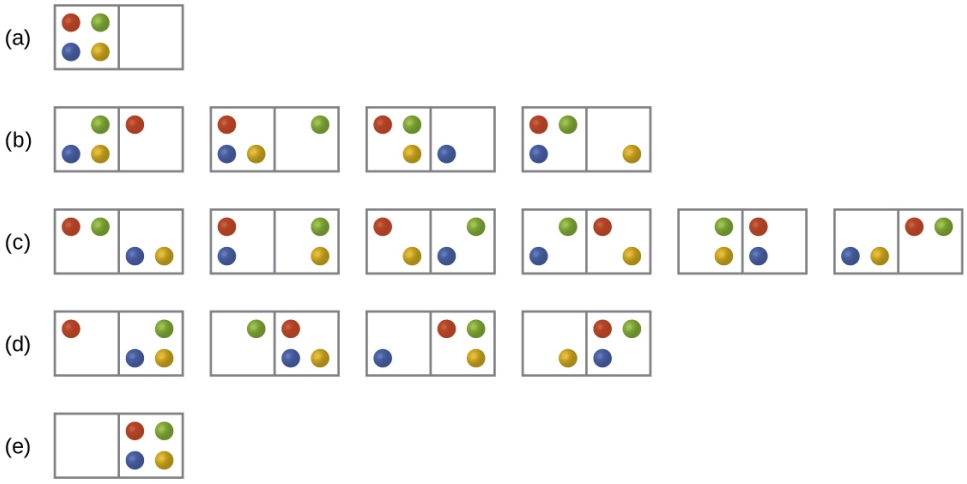


Figure 17.4.2: The sixteen microstates associated with placing four particles in two boxes are shown. The microstates are collected into five distributions—(a), (b), (c), (d), and (e)—based on the numbers of particles in each box.

For this system, the most probable configuration is one of the six microstates associated with distribution (c) where the particles are evenly distributed between the boxes, that is, a configuration of two particles in each box. The probability of finding the system in this configuration is $\frac{6}{16}$ or $\frac{3}{8}$ The least probable configuration of the system is one in which all four particles are in one box, corresponding to distributions (a) and (e), each with a probability of The probability of finding all particles in only one box (either the left box or right box) is then or

$$\left(\frac{1}{16} + \frac{1}{16}\right) = \frac{2}{16} \text{ or } \frac{1}{8}$$

As you add more particles to the system, the number of possible microstates increases exponentially (2^N). A macroscopic (laboratory-sized) system would typically consist of moles of particles ($N \sim 10^{23}$), and the corresponding number of microstates would be staggeringly huge. Regardless of the number of particles in the system, however, the distributions in which roughly equal numbers of particles are found in each box are always the most probable configurations.

This matter dispersal model of entropy is often described qualitatively in terms of the *disorder* of the system. By this description, microstates in which all the particles are in a single box are the most ordered, thus possessing the least entropy. Microstates in which the particles are more evenly distributed among the boxes are more disordered, possessing greater entropy.

The previous description of an ideal gas expanding into a vacuum (Figure 16.4) is a macroscopic example of this particle-in-a-box model. For this system, the most probable distribution is confirmed to be the one in which the matter is most uniformly dispersed or distributed between the two flasks. Initially, the gas molecules are confined to just one of the two flasks. Opening the valve between the flasks increases the volume available to the gas molecules and, correspondingly, the number of microstates possible for the system. Since $W_f > W_i$, the expansion process involves an increase in entropy ($\Delta S > 0$) and is spontaneous.

A similar approach may be used to describe the spontaneous flow of heat. Consider a system consisting of two objects, each containing two particles, and two units of thermal energy (represented as "*") in Figure 17.4.3 The hot object is comprised of

particles **A** and **B** and initially contains both energy units. The cold object is comprised of particles **C** and **D**, which initially has no energy units. Distribution (a) shows the three microstates possible for the initial state of the system, with both units of energy contained within the hot object. If one of the two energy units is transferred, the result is distribution (b) consisting of four microstates. If both energy units are transferred, the result is distribution (c) consisting of three microstates. Thus, we may describe this system by a total of ten microstates. The probability that the heat does not flow when the two objects are brought into contact, that is, that the system remains in distribution (a), is $\frac{3}{10}$. More likely is the flow of heat to yield one of the other two distribution, the combined probability being $\frac{7}{10}$. The most likely result is the flow of heat to yield the uniform dispersal of energy represented by distribution (b), the probability of this configuration being $\frac{4}{10}$. This supports the common observation that placing hot and cold objects in contact results in spontaneous heat flow that ultimately equalizes the objects' temperatures. And, again, this spontaneous process is also characterized by an increase in system entropy.

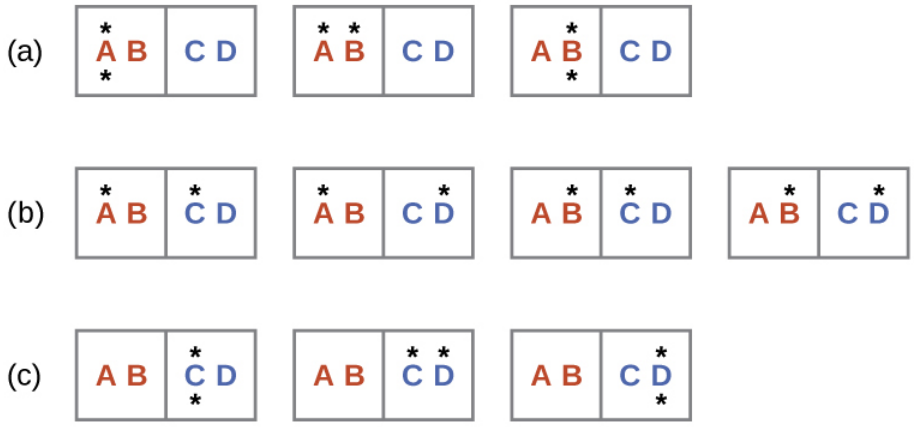
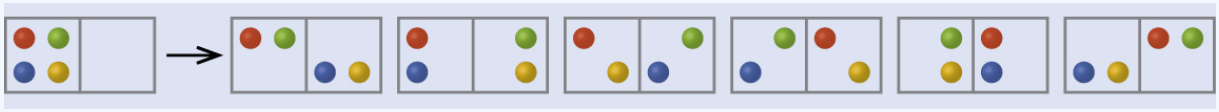


Figure 17.4.3: This shows a microstate model describing the flow of heat from a hot object to a cold object. (a) Before the heat flow occurs, the object comprised of particles **A** and **B** contains both units of energy and as represented by a distribution of three microstates. (b) If the heat flow results in an even dispersal of energy (one energy unit transferred), a distribution of four microstates results. (c) If both energy units are transferred, the resulting distribution has three microstates.

Example 17.4.1: Determination of ΔS

Calculate the change in entropy for the process depicted below.



Solution

The initial number of microstates is one, the final six:

$$\Delta S = k \ln \frac{W_c}{W_a} = 1.38 \times 10^{-23} J/K \times \ln \frac{6}{1} = 2.47 \times 10^{-23} J/K$$

The sign of this result is consistent with expectation; since there are more microstates possible for the final state than for the initial state, the change in entropy should be positive.

Exercise 17.4.1

Consider the system shown in Figure 17.4.3. What is the change in entropy for the process where *all* the energy is transferred from the hot object (**AB**) to the cold object (**CD**)?

Answer

0 J/K

Predicting the Sign of ΔS

The relationships between entropy, microstates, and matter/energy dispersal described previously allow us to make generalizations regarding the relative entropies of substances and to predict the sign of entropy changes for chemical and physical processes. Consider the phase changes illustrated in Figure 17.4.4. In the solid phase, the atoms or molecules are restricted to nearly fixed positions with respect to each other and are capable of only modest oscillations about these positions. With essentially fixed locations for the system's component particles, the number of microstates is relatively small. In the liquid phase, the atoms or molecules are free to move over and around each other, though they remain in relatively close proximity to one another. This increased freedom of motion results in a greater variation in possible particle locations, so the number of microstates is correspondingly greater than for the solid. As a result, $S_{liquid} > S_{solid}$ and the process of converting a substance from solid to liquid (melting) is characterized by an increase in entropy, $\Delta S > 0$. By the same logic, the reciprocal process (freezing) exhibits a decrease in entropy, $\Delta S < 0$.

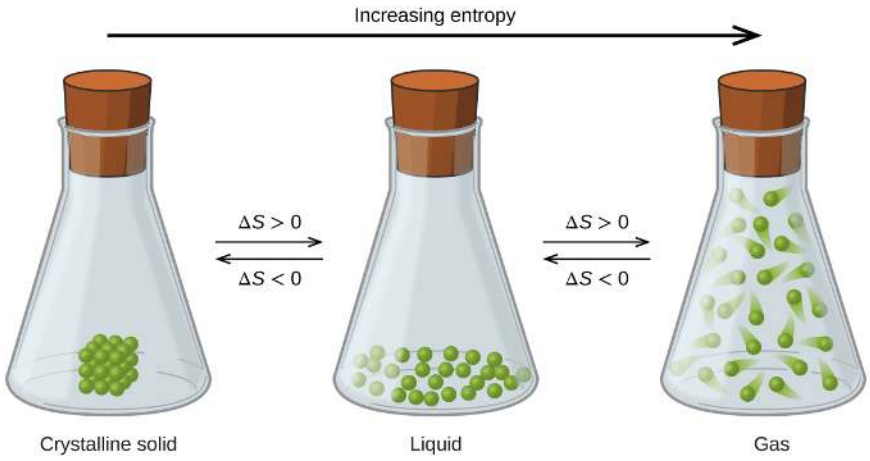


Figure 17.4.4: The entropy of a substance increases ($\Delta S > 0$) as it transforms from a relatively ordered solid, to a less-ordered liquid, and then to a still less-ordered gas. The entropy decreases ($\Delta S < 0$) as the substance transforms from a gas to a liquid and then to a solid.

Now consider the gaseous phase, in which a given number of atoms or molecules occupy a *much* greater volume than in the liquid phase. Each atom or molecule can be found in many more locations, corresponding to a much greater number of microstates. Consequently, for any substance, $S_{gas} > S_{liquid} > S_{solid}$, and the processes of vaporization and sublimation likewise involve increases in entropy, $\Delta S > 0$. Likewise, the reciprocal phase transitions, condensation and deposition, involve decreases in entropy, $\Delta S < 0$.

According to kinetic-molecular theory, the temperature of a substance is proportional to the average kinetic energy of its particles. Raising the temperature of a substance will result in more extensive vibrations of the particles in solids and more rapid translations of the particles in liquids and gases. At higher temperatures, the distribution of kinetic energies among the atoms or molecules of the substance is also broader (more dispersed) than at lower temperatures. Thus, the entropy for any substance increases with temperature (Figure 17.4.5).

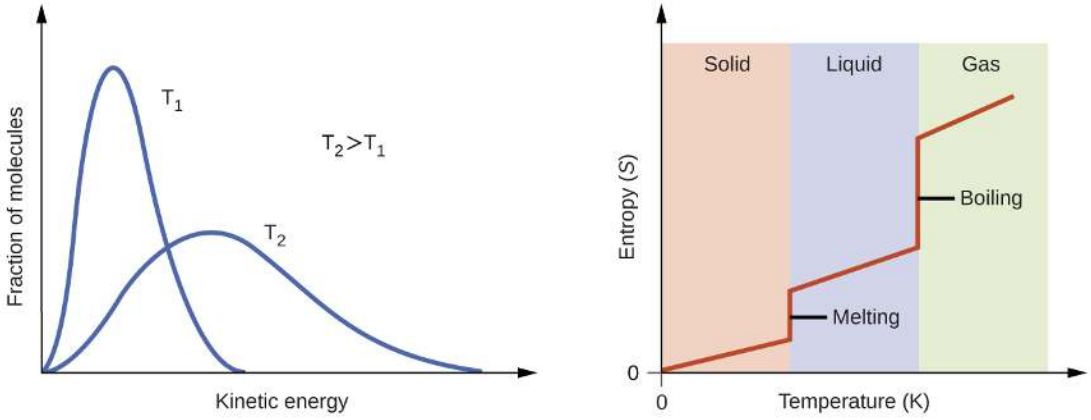


Figure 17.4.5: Entropy increases as the temperature of a substance is raised, which corresponds to the greater spread of kinetic energies. When a substance undergoes a phase transition, its entropy changes significantly.

Link to Learning

Try this simulator with interactive visualization of the dependence of particle location and freedom of motion on physical state and temperature.

The entropy of a substance is influenced by the structure of the particles (atoms or molecules) that comprise the substance. With regard to atomic substances, heavier atoms possess greater entropy at a given temperature than lighter atoms, which is a consequence of the relation between a particle's mass and the spacing of quantized translational energy levels (a topic beyond the scope of this text). For molecules, greater numbers of atoms increase the number of ways in which the molecules can vibrate and thus the number of possible microstates and the entropy of the system.

Finally, variations in the types of particles affects the entropy of a system. Compared to a pure substance, in which all particles are identical, the entropy of a mixture of two or more different particle types is greater. This is because of the additional orientations and interactions that are possible in a system comprised of nonidentical components. For example, when a solid dissolves in a liquid, the particles of the solid experience both a greater freedom of motion and additional interactions with the solvent particles. This corresponds to a more uniform dispersal of matter and energy and a greater number of microstates. The process of dissolution therefore involves an increase in entropy, $\Delta S > 0$.

Considering the various factors that affect entropy allows us to make informed predictions of the sign of ΔS for various chemical and physical processes as illustrated in Example 17.4.2

Example 17.4.2: Predicting the Sign of ΔS

Predict the sign of the entropy change for the following processes. Indicate the reason for each of your predictions.

a. One mole liquid water at room temperature one mole liquid water at 50 °C
b. $Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$
c. \(C_6 H_6(l)+\frac{15}{2} O_2(g) \longrightarrow 6 CO_2(g)+3 H_2O(l))
d. $NH_3(s) \longrightarrow NH_3(l)$

Solution

a. positive, temperature increases
b. negative, reduction in the number of ions (particles) in solution, decreased dispersal of matter
c. negative, net decrease in the amount of gaseous species
d. positive, phase transition from solid to liquid, net increase in dispersal of matter

Exercise 17.4.1

Predict the sign of the entropy change for the following processes. Give a reason for your prediction.

a. $NaNO_3(s) \longrightarrow Na^+(aq) + NO_3{}^-(aq)$
b. the freezing of liquid water
c. $CO_2(s) \longrightarrow CO_2(g)$
d. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

Answer

(a) Positive; The solid dissolves to give an increase of mobile ions in solution.

(b) Negative; The liquid becomes a more ordered solid.

(c) Positive; The relatively ordered solid becomes a gas.

(d) Positive; There is a net increase in the amount of gaseous species.

This page titled 17.4: Entropy is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.2: Entropy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

17.5: The Second and Third Laws of Thermodynamics

Learning Objectives

By the end of this section, you will be able to:

- State and explain the second and third laws of thermodynamics
- Calculate entropy changes for phase transitions and chemical reactions under standard conditions

The Second Law of Thermodynamics

In the quest to identify a property that may reliably predict the spontaneity of a process, a promising candidate has been identified: entropy. Processes that involve an increase in entropy *of the system* ($\Delta S > 0$) are very often spontaneous; however, examples to the contrary are plentiful. By expanding consideration of entropy changes to include *the surroundings*, we may reach a significant conclusion regarding the relation between this property and spontaneity. In thermodynamic models, the system and surroundings comprise everything, that is, the universe, and so the following is true:

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$$

To illustrate this relation, consider again the process of heat flow between two objects, one identified as the system and the other as the surroundings. There are three possibilities for such a process:

1. The objects are at different temperatures, and heat flows from the hotter to the cooler object. *This is always observed to occur spontaneously.* Designating the hotter object as the system and invoking the definition of entropy yields the following:

   $$\Delta S_{sys} = \frac{-q_{rev}}{T_{sys}} \quad \text{and} \quad \Delta S_{surr} = \frac{q_{rev}}{T_{surr}}$$

   The magnitudes of $-q_{\text{rev}}$ and q_{rev} are equal, their opposite arithmetic signs denoting loss of heat by the system and gain of heat by the surroundings. Since $T_{\text{sys}} > T_{\text{surr}}$ in this scenario, the entropy *decrease* of the system will be less than the entropy *increase* of the surroundings, and so *the entropy of the universe will increase*:

   $$|\Delta S_{\text{sys}}| < |\Delta S_{\text{surr}}|$$
   $$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0$$

2. The objects are at different temperatures, and heat flows from the cooler to the hotter object. *This is never observed to occur spontaneously.* Again designating the hotter object as the system and invoking the definition of entropy yields the following:

   $$\Delta S_{sys} = \frac{q_{rev}}{T_{sys}} \quad \text{and} \quad \Delta S_{surr} = \frac{-q_{rev}}{T_{surr}}$$

   The arithmetic signs of q_{rev} denote the gain of heat by the system and the loss of heat by the surroundings. The magnitude of the entropy change for the surroundings will again be greater than that for the system, but in this case, the signs of the heat changes (that is, *the direction of the heat flow*) will yield a negative value for ΔS_{univ}. *This process involves a decrease in the entropy of the universe.*
3. The objects are at essentially the same temperature, $T_{\text{sys}} \approx T_{\text{surr}}$, and so the magnitudes of the entropy changes are essentially the same for both the system and the surroundings. In this case, the entropy change of the universe is zero, and the system is *at equilibrium*.

   $$|\Delta S_{\text{sys}}| \approx |\Delta S_{\text{surr}}|$$
   $$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 0$$

These results lead to a profound statement regarding the relation between entropy and spontaneity known as the second law of thermodynamics: *all spontaneous changes cause an increase in the entropy of the universe.* A summary of these three relations is provided in Table 17.5.1.

Table 17.5.1: The Second Law of Thermodynamics

$\Delta S_{\text{univ}} > 0$	spontaneous

$\Delta S_{univ} < 0$	nonspontaneous (spontaneous in opposite direction)
$\Delta S_{univ} = 0$	at equilibrium

For many realistic applications, the surroundings are vast in comparison to the system. In such cases, the heat gained or lost by the surroundings as a result of some process represents a very small, nearly infinitesimal, fraction of its total thermal energy. For example, combustion of a fuel in air involves transfer of heat from a system (the fuel and oxygen molecules undergoing reaction) to surroundings that are infinitely more massive (the earth's atmosphere). As a result, q_{surr} is a good approximation of q_{rev}, and the second law may be stated as the following:

$$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr} = \Delta S_{sys} + \frac{q_{surr}}{T}$$

We may use this equation to predict the spontaneity of a process as illustrated in Example 17.5.1.

Example 17.5.1: Will Ice Spontaneously Melt?

The entropy change for the process

$$H_2O(s) \longrightarrow H_2O(l)$$

is 22.1 J/K and requires that the surroundings transfer 6.00 kJ of heat to the system. Is the process spontaneous at −10.00 °C? Is it spontaneous at +10.00 °C?

Solution

We can assess the spontaneity of the process by calculating the entropy change of the universe. If ΔS_{univ} is positive, then the process is spontaneous. At both temperatures, ΔS_{sys} = 22.1 J/K and q_{surr} = −6.00 kJ.

At −10.00 °C (263.15 K), the following is true:

$$\begin{aligned}\Delta S_{\text{univ}} &= \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = \Delta S_{\text{sys}} + \frac{q_{\text{surr}}}{T} \\ &= 22.1 J/K + \frac{-6.00 \times 10^3 J}{263.15 K} = -0.7 J/K\end{aligned}$$

$S_{univ} < 0$, so melting is nonspontaneous (*not* spontaneous) at −10.0 °C.

At 10.00 °C (283.15 K), the following is true:

$$\begin{aligned}\Delta S_{\text{univ}} &= \Delta S_{\text{sys}} + \frac{q_{\text{surr}}}{T} \\ &= 22.1 J/K + \frac{-6.00 \times 10^3 J}{283.15 K} = +0.9 J/K\end{aligned}$$

$S_{univ} > 0$, so melting *is* spontaneous at 10.00 °C.

Exercise 17.5.1

Using this information, determine if liquid water will spontaneously freeze at the same temperatures. What can you say about the values of S_{univ}?

Answer

Entropy is a state function, so $\Delta S_{freezing} = -\Delta S_{melting}$ = −22.1 J/K and q_{surr} = +6.00 kJ. At −10.00 °C spontaneous, +0.7 J/K; at +10.00 °C nonspontaneous, −0.9 J/K.

The Third Law of Thermodynamics

The previous section described the various contributions of matter and energy dispersal that contribute to the entropy of a system. With these contributions in mind, consider the entropy of a pure, perfectly crystalline solid possessing no kinetic energy (that is, at

a temperature of absolute zero, 0 K). This system may be described by a single microstate, as its purity, perfect crystallinity and complete lack of motion means there is but one possible location for each identical atom or molecule comprising the crystal (W = 1). According to the Boltzmann equation, the entropy of this system is zero.

$$S = k \ln W = k \ln(1) = 0$$

This limiting condition for a system's entropy represents the third law of thermodynamics: *the entropy of a pure, perfect crystalline substance at 0 K is zero.*

Careful calorimetric measurements can be made to determine the temperature dependence of a substance's entropy and to derive absolute entropy values under specific conditions. **Standard entropies** ($S°$) are for one mole of substance under standard conditions (a pressure of 1 bar and a temperature of 298.15 K; see details regarding standard conditions in the thermochemistry chapter of this text). The **standard entropy change** ($\Delta S°$) for a reaction may be computed using standard entropies as shown below:

$$\Delta S^\circ = \sum \nu S^\circ(\text{ products }) - \sum \nu S^\circ(\text{ reactants })$$

where ν represents stoichiometric coefficients in the balanced equation representing the process. For example, $\Delta S°$ for the following reaction at room temperature

$$mA + nB \longrightarrow xC + yD$$

is computed as:

$$= [xS^\circ(C) + yS^\circ(D)] - [mS^\circ(A) + nS^\circ(B)]$$

A partial listing of standard entropies is provided in Table 17.5.2, and additional values are provided in Appendix G. The example exercises that follow demonstrate the use of $S°$ values in calculating standard entropy changes for physical and chemical processes.

Table 17.5.2: Standard entropies for selected substances measured at 1 atm and 298.15 K. (Values are approximately equal to those measured at 1 bar, the currently accepted standard state pressure.)

Substance	(J mol^{-1} K^{-1})
carbon	
C(*s*, graphite)	5.740
C(*s*, diamond)	2.38
CO(*g*)	197.7
CO_2(*g*)	213.8
CH_4(*g*)	186.3
C_2H_4(*g*)	219.5
C_2H_6(*g*)	229.5
CH_3OH(*l*)	126.8
C_2H_5OH(*l*)	160.7
hydrogen	
H_2(*g*)	130.57
H(*g*)	114.6
H_2O(*g*)	188.71
H_2O(*l*)	69.91
HCI(*g*)	186.8
H_2S(*g*)	205.7

Substance	($J\ mol^{-1}\ K^{-1}$)
oxygen	
$O_2(g)$	205.03

Example 17.5.2: Determination of ΔS°

Calculate the standard entropy change for the following process:

$$H_2O(g) \longrightarrow H_2O(l)$$

Solution

Calculate the entropy change using standard entropies as shown above:

$$\Delta S^\circ = (1mol)\left(70.0Jmol^{-1}K^{-1}\right) - (1mol)\left(188.8Jmol^{-1}K^{-1}\right) = -118.8J/K$$

The value for Δ*S*° is negative, as expected for this phase transition (condensation), which the previous section discussed.

Exercise 17.5.2

Calculate the standard entropy change for the following process:

$$H_2(g) + C_2H_4(g) \longrightarrow C_2H_6(g)$$

Answer

$-120.6\ J\ K^{-1}\ mol^{-1}$

Example 17.5.3: Determination of ΔS°

Calculate the standard entropy change for the combustion of methanol, CH_3OH:

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l)$$

Solution

Calculate the entropy change using standard entropies as shown above:

$$\Delta S^\circ = \sum \nu S^\circ(\text{ products }) - \sum \nu S^\circ(\text{ reactants })$$

$$[2mol \times S^\circ(CO_2(g)) + 4mol \times S^\circ(H_2O(l))] - [2mol \times S^\circ(CH_3OH(l)) + 3mol \times S^\circ(O_2(g))]$$

$$= \{[2(213.8) + 4 \times 70.0] - [2(126.8) + 3(205.2)]\} = -161.6J/K$$

Exercise 17.5.3

Calculate the standard entropy change for the following reaction:

$$Ca(OH)_2(s) \longrightarrow CaO(s) + H_2O(l)$$

Answer

24.7 J/K

This page titled 17.5: The Second and Third Laws of Thermodynamics is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.3: The Second and Third Laws of Thermodynamics** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

17.6: Free Energy

Learning Objectives

By the end of this section, you will be able to:

- Define Gibbs free energy, and describe its relation to spontaneity
- Calculate free energy change for a process using free energies of formation for its reactants and products
- Calculate free energy change for a process using enthalpies of formation and the entropies for its reactants and products
- Explain how temperature affects the spontaneity of some processes
- Relate standard free energy changes to equilibrium constants

One of the challenges of using the second law of thermodynamics to determine if a process is spontaneous is that it requires measurements of the entropy change for the system *and* the entropy change for the surroundings. An alternative approach involving a new thermodynamic property defined in terms of system properties only was introduced in the late nineteenth century by American mathematician Josiah Willard Gibbs. This new property is called the **Gibbs free energy** (G) (or simply the *free energy*), and it is defined in terms of a system's enthalpy and entropy as the following:

$$G = H - TS$$

Free energy is a state function, and at constant temperature and pressure, the free energy change (ΔG) may be expressed as the following:

$$\Delta G = \Delta H - T\Delta S$$

(For simplicity's sake, the subscript "sys" will be omitted henceforth.)

The relationship between this system property and the spontaneity of a process may be understood by recalling the previously derived second law expression:

$$\Delta S_{\text{univ}} = \Delta S + \frac{q_{\text{surr}}}{T}$$

The first law requires that $q_{surr} = -q_{sys}$, and at constant pressure $q_{sys} = \Delta H$, so this expression may be rewritten as:

$$\Delta S_{\text{univ}} = \Delta S - \frac{\Delta H}{T}$$

Multiplying both sides of this equation by $-T$, and rearranging yields the following:

$$-T\Delta S_{\text{univ}} = \Delta H - T\Delta S$$

Comparing this equation to the previous one for free energy change shows the following relation:

$$\Delta G = -T\Delta S_{\text{univ}}$$

The free energy change is therefore a reliable indicator of the spontaneity of a process, being directly related to the previously identified spontaneity indicator, ΔS_{univ}. Table 17.6.1 summarizes the relation between the spontaneity of a process and the arithmetic signs of these indicators.

Table 17.6.1: Relation between Process Spontaneity and Signs of Thermodynamic Properties

$\Delta S_{univ} > 0$	$\Delta G < 0$	moves spontaneously in the forward direction, as written, to reach equilibrium
$\Delta S_{univ} < 0$	$\Delta G > 0$	nonspontaneous in the forward direction, as written, but moves spontaneously in the reverse direction, as written, to reach equilibrium
$\Delta S_{univ} = 0$	$\Delta G = 0$	reversible (at equilibrium)

What's "Free" about ΔG?

In addition to indicating spontaneity, the free energy change also provides information regarding the amount of useful work (w) that may be accomplished by a spontaneous process. Although a rigorous treatment of this subject is beyond the scope of an introductory chemistry text, a brief discussion is helpful for gaining a better perspective on this important thermodynamic property.

For this purpose, consider a spontaneous, exothermic process that involves a decrease in entropy. The free energy, as defined by

$$\Delta G = \Delta H - T\Delta S$$

may be interpreted as representing the difference between the energy produced by the process, Δ*H*, and the energy lost to the surroundings, *T*Δ*S*. The difference between the energy produced and the energy lost is the energy available (or "free") to do useful work by the process, Δ*G*. If the process somehow could be made to take place under conditions of thermodynamic reversibility, the amount of work that could be done would be maximal:

$$\Delta G = w_{max}$$

where w_{max} refers to all types of work except expansion (pressure-volume) work.

However, as noted previously in this chapter, such conditions are not realistic. In addition, the technologies used to extract work from a spontaneous process (e.g., batteries) are never 100% efficient, and so the work done by these processes is always less than the theoretical maximum. Similar reasoning may be applied to a nonspontaneous process, for which the free energy change represents the *minimum* amount of work that must be done *on* the system to carry out the process.

Calculating Free Energy Change

Free energy is a state function, so its value depends only on the conditions of the initial and final states of the system. A convenient and common approach to the calculation of free energy changes for physical and chemical reactions is by use of widely available compilations of standard state thermodynamic data. One method involves the use of standard enthalpies and entropies to compute **standard free energy changes**, ΔG^o, according to the following relation.

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

It is important to understand that for phase changes, ΔG° tells you if the phase change is spontaneous or not; will it happen, or not happen. For chemical reactions, ΔG° tells you the **extent** of a reaction. In other words, ΔG° for a reaction tells you how much product will be present at equilibrium. A reaction with $\Delta G^\circ < 0$ is considered **product-favored at equilibrium**; there will be more products than reactants when the reaction reaches equilibrium. A reaction with $\Delta G^\circ > 0$ is considered **reactant-favored at equilibrium**; there will be more reactants than products when the reaction reaches equilibrium.

✔ Example 17.6.1: Using Standard Enthalpy and Entropy Changes to Calculate ΔG°

Use standard enthalpy and entropy data from Appendix G to calculate the standard free energy change for the vaporization of water at room temperature (298 K). What does the computed value for ΔG^o say about the spontaneity of this process?

Solution

The process of interest is the following:

$$\mathrm{H_2O(l) \longrightarrow H_2O(g)}$$

The standard change in free energy may be calculated using the following equation:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

From Appendix G:

Substance	$\Delta H_f^\circ (\mathrm{kJ/mol})$	$S^\circ (\mathrm{kJ/K\ mol})$
$H_2O(l)$	−285.83	70.0
$H_2O(g)$	−241.82	188.8

Using the appendix data to calculate the standard enthalpy and entropy changes yields:

$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ\left(\mathrm{H_2O(g)}\right) - \Delta H_f^\circ\left(\mathrm{H_2O(l)}\right)\\ &= \left[-241.82\ \mathrm{kJ/mol} - (-285.83)\right]\ \mathrm{kJ/mol} = 44.01\ \mathrm{kJ/mol}\\ \Delta S^\circ &= 1\ \mathrm{mol} \times S^\circ\left(\mathrm{H_2O(g)}\right) - 1\ \mathrm{mol} \times S^\circ\left(\mathrm{H_2O(l)}\right)\\ &= (1\ \mathrm{mol})\ 188.8 J/mol \cdot K - (1\ \mathrm{mol})\ 70.0\ \mathrm{J/mol\ K} = 118.8\ \mathrm{J\ /\ K}\\ \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ\end{aligned}$$

Substitution into the standard free energy equation yields:

$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ\\ &= 44.01\ \mathrm{kJ} - (298K \times 118.8\ \mathrm{J/K}) \times \frac{1\ \mathrm{kJ}}{1000\ \mathrm{J}}\\ &= 44.01\ \mathrm{kJ} - 35.4\ \mathrm{kJ}\\ &= 8.6\ \mathrm{kJ}\end{aligned}$$

At 298 K (25 °C) so boiling is nonspontaneous (*not* spontaneous).

? Exercise 17.6.1

Use standard enthalpy and entropy data from Appendix G to calculate the standard free energy change for the reaction shown here (298 K). What does the computed value for $\Delta G°$ say about the spontaneity of this process?

$$\mathrm{C_2H_6(g) \longrightarrow H_2(g) + C_2H_4(g)}$$

Answer

the reaction is nonspontaneous (*not* spontaneous) at 25 °C.

The standard free energy change for a reaction may also be calculated from **standard free energy of formation** ΔG_f^o values of the reactants and products involved in the reaction. The standard free energy of formation is the free energy change that accompanies the formation of one mole of a substance from its elements in their standard states. Similar to the standard enthalpy of formation, ΔG_f^o is by definition zero for elemental substances in their standard states. The approach used to calculate ΔG^o for a reaction from ΔG_f^o values is the same as that demonstrated previously for enthalpy and entropy changes. For the reaction

$$mA + nB \longrightarrow xC + yD$$

the standard free energy change at room temperature may be calculated as

$$\begin{aligned}\Delta G^\circ &= \sum \nu \Delta G^\circ(\text{ products }) - \sum \nu \Delta G^\circ(\text{ reactants })\\ &= \left[x\Delta G_f^\circ(C) + y\Delta G_f^\circ(D)\right] - \left[m\Delta G_f^\circ(A) + n\Delta G_f^\circ(B)\right].\end{aligned}$$

✓ Example 17.6.2: Using Standard Free Energies of Formation to Calculate ΔG°

Consider the decomposition of yellow mercury(II) oxide.

$$\mathrm{HgO(s, \backslash text\{yellow\}) \longrightarrow Hg(l) + \tfrac{1}{2}O_2(g)}$$

Calculate the standard free energy change at room temperature, using (a) standard free energies of formation and (b) standard enthalpies of formation and standard entropies. Do the results indicate the reaction to be spontaneous or nonspontaneous under standard conditions?

Solution

The required data are available in Appendix G and are shown here.

Compound	ΔG_f° (kJ / mol)	ΔH_f° (kJ / mol)	S° (kJ / K mol)
HgO (*s*, yellow)	−58.43	−90.46	71.13
Hg(*l*)	0	0	75.9
$O_2(g)$	0	0	205.2

(a) Using free energies of formation:

$$\begin{aligned}\Delta G^\circ &= \sum \nu G_f^\circ(\text{ products }) - \sum \nu \Delta G_f^\circ(\text{ reactants }) \\ &= \left[1\Delta G_f^\circ\ \mathrm{Hg(l)} + \frac{1}{2}\Delta G_f^\circ\ \mathrm{O_2(g)}\right] - 1\Delta G_f^\circ\ \mathrm{HgO(s, yellow)} \\ &= \left[1\ \mathrm{mol(0\ kJ\ /\ mol)} + \frac{1}{2}\ \mathrm{mol(0\ kJ\ /\ mol)}\right] - 1\ \mathrm{mol(-58.43\ kJ\ /\ mol)} = 58.43\ \mathrm{kJ\ /\ mol}\end{aligned}$$

(b) Using enthalpies and entropies of formation:

$$\begin{aligned}\Delta H^\circ &= \sum \nu \Delta H_f^\circ(\text{ products }) - \sum \nu \Delta H_f^\circ(\text{ reactants }) \\ &= \left[1\Delta H_f^\circ\ \mathrm{Hg(l)} + \frac{1}{2}\Delta H_f^\circ\ \mathrm{O_2(g)}\right] - 1\Delta H_f^\circ\ \mathrm{HgO(s, yellow)} \\ &= \left[1\ \mathrm{mol(0\ kJ\ /\ mol)} + \frac{1}{2}\ \mathrm{mol(0\ kJ\ /\ mol)}\right] - 1\ \mathrm{mol(-90.46\ kJ\ /\ mol)} = 90.46\ \mathrm{kJ\ /\ mol}\end{aligned}$$

$$\begin{aligned}\Delta S^\circ &= \sum \nu \Delta S^\circ(\text{ products }) - \sum \nu \Delta S^\circ(\text{ reactants }) \\ &= \left[1\ \Delta S^\circ\ \mathrm{Hg(l)} + \frac{1}{2}\Delta S^\circ\ \mathrm{O_2(g)}\right] - 1\Delta S^\circ\ \mathrm{HgO(s, yellow)} \\ &= \left[1\ \mathrm{mol(75.9\ J\ /\ mol\ K)} + \frac{1}{2}\ \mathrm{mol(205.2\ J\ /\ mol\ K)}\right] - 1\mathrm{mol(71.13\ J\ /\ mol\ K)} = 107.4\ \ \mathrm{J\ /\ mol\ K}\end{aligned}$$

$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ = 90.46\ \mathrm{kJ} - 298.15\ \mathrm{K} \times 107.4\ \mathrm{J\ /\ K \cdot mol} \times \frac{1\ \mathrm{kJ}}{1000\ \mathrm{J}} \\ &= (90.46 - 32.01)\ \mathrm{kJ\ /\ mol} = 58.45\ \mathrm{kJ\ /\ mol}\end{aligned}$$

Both ways to calculate the standard free energy change at 25 °C give the same numerical value (to three significant figures), and both predict that the process is nonspontaneous (*not* spontaneous) at room temperature.

? Exercise 17.6.1

Calculate $\Delta G°$ using (a) free energies of formation and (b) enthalpies of formation and entropies (Appendix G). Do the results indicate the reaction to be spontaneous or nonspontaneous at 25 °C?

$$\mathrm{C_2H_4(g) \longrightarrow H_2(g) + C_2H_2(g)}$$

Answer

a. 140.8 kJ/mol, nonspontaneous
b. 141.5 kJ/mol, nonspontaneous

Free Energy Changes for Coupled Reactions

The use of free energies of formation to compute free energy changes for reactions as described above is possible because ΔG is a state function, and the approach is analogous to the use of Hess' Law in computing enthalpy changes (see the chapter on thermochemistry). Consider the vaporization of water as an example:

$$\mathrm{H_2O(l) \longrightarrow H_2O(g)}$$

An equation representing this process may be derived by adding the formation reactions for the two phases of water (necessarily reversing the reaction for the liquid phase). The free energy change for the sum reaction is the sum of free energy changes for the two added reactions:

$$\begin{array}{ll} \mathrm{H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)} & \Delta G_f^\circ \text{ gas} \\ \mathrm{H_2O(l) \longrightarrow H_2(g) + \frac{1}{2}O_2(g)} & -\Delta G_f^o \text{ liquid} \\ \hline \mathrm{H_2O(l) \longrightarrow H_2O(g)} & \Delta G^\circ = \Delta G_f^\circ \text{ gas} - \Delta G_f^\circ \text{ liquid} \end{array}$$

This approach may also be used in cases where a nonspontaneous reaction is enabled by coupling it to a spontaneous reaction. For example, the production of elemental zinc from zinc sulfide is thermodynamically unfavorable, as indicated by a positive value for ΔG^o:

$$\mathrm{ZnS(s) \rightarrow Zn(s) + S(s)} \quad \Delta G_1^\circ = 201.3kJ$$

The industrial process for production of zinc from sulfidic ores involves coupling this decomposition reaction to the thermodynamically favorable oxidation of sulfur:

$$\mathrm{S(s) + O_2(g) \rightarrow SO_2(g)} \quad \Delta G_2^\circ = -300.1kJ$$

The coupled reaction exhibits a negative free energy change and is spontaneous:

$$\mathrm{ZnS(s) + O_2(g) \rightarrow Zn(s) + SO_2(g)}$$

$$\Delta G^\circ = 201.3 \text{ kJ} + -300.1 \text{ kJ} = -98.8 \text{ kJ}$$

This process is typically carried out at elevated temperatures, so this result obtained using standard free energy values is just an estimate. The gist of the calculation, however, holds true.

Example 17.6.3: Calculating Free Energy Change for a Coupled Reaction

Is a reaction coupling the decomposition of ZnS to the formation of $\mathrm{H_2S}$ expected to be spontaneous under standard conditions?

Solution

Following the approach outlined above and using free energy values from Appendix G:

Decomposition of zinc sulfide:	$\mathrm{ZnS(s) \rightarrow Zn(s) + S(s)}$	$\Delta G_1^\circ = 201.3$ kJ
Formation of hydrogen sulfide:	$\mathrm{S(s) + H_2(g) \rightarrow H_2S(g)}$	$\Delta G_1^\circ = -33.4$ kJ
Coupled reaction:	$\mathrm{ZnS(s) + H_2(g) \rightarrow Zn(s) + H_2S(g)}$	$\Delta G^\circ = 201.3 \text{ kJ} + -33.4 \text{ kJ} = 167.9 \text{ kJ}$

The coupled reaction exhibits a positive free energy change and is thus nonspontaneous.

Exercise 17.6.3

What is the standard free energy change for the reaction below? Is the reaction expected to be spontaneous under standard conditions?

$$\mathrm{FeS(s) + O_2(g) \rightarrow Fe(s) + SO_2(g)}$$

Answer

−199.7 kJ; spontaneous

Temperature Dependence of Spontaneity and Extent of Reaction

As was previously demonstrated in this chapter's section on entropy, the spontaneity of a process may depend upon the temperature of the system. Phase transitions, for example, will proceed spontaneously in one direction or the other depending upon the temperature of the substance in question. In a similar, but not identical fashion, some chemical reactions can switch from being product-favored at equilibrium, to being reactant-favored at equilibrium, depending on the temperature.

> **Note**
>
> The numerical value of ΔG° is always dependent on the temperature. In this section we are determining whether or not the **sign** of ΔG° is dependent on the temperature.

To illustrate this concept, the equation relating free energy change to the enthalpy and entropy changes for the process is considered:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

The extent of a process, as reflected in the arithmetic sign of its standard free energy change, is then determined by the signs of the enthalpy and entropy changes and, in some cases, the absolute temperature. Since T is the absolute (Kelvin) temperature, it can only have positive values. Four possibilities therefore exist with regard to the signs of the enthalpy and entropy changes:

1. **Both ΔH^o and ΔS^o are positive.** This condition describes an endothermic process that involves an increase in system entropy. In this case, ΔG^o will be negative if the magnitude of the $T\Delta S^o$ term is greater than ΔH^o. If the $T\Delta S^o$ term is less than ΔH^o, the free energy change will be positive. Such a process is *product-favored at equilibrium at high temperatures and reactant-favored at equilibrium at low temperatures.*
2. **Both ΔH^o and ΔS^o are negative.** This condition describes an exothermic process that involves a decrease in system entropy. In this case, ΔG^o will be negative if the magnitude of the $T\Delta S^o$ term is less than ΔH^o. If the $T\Delta S^o$ term's magnitude is greater than ΔH^o, the free energy change will be positive. Such a process is *product-favored at equilibrium at low temperatures* and *reactant-favored at equilibrium at high temperatures.*
3. **ΔH^o is positive and ΔS^o is negative.** This condition describes an endothermic process that involves a decrease in system entropy. In this case, ΔG^o will be positive regardless of the temperature. Such a process is *reactant-favored at equilibrium at all temperatures.*
4. **ΔH^o is negative and ΔS^o is positive.** This condition describes an exothermic process that involves an increase in system entropy. In this case, ΔG^o will be negative regardless of the temperature. Such a process is *product-favored at equilibrium at all temperatures.*

These four scenarios are summarized in Table 17.6.1

Table 17.6.1

Sign of ΔH^o	Sign of ΔS^o	Sign of ΔG^o	Temperature Dependence of ΔG^o
-	+	-	The sign of ΔG^o does not depend on the temperature.The reaction is product-favored at equilibrium at all temperatures.
+	-	+	The sign of ΔG^o does not depend on the temperature.The reaction is reactant-favored at equilibrium at all temperatures.
-	-	- or +	The sign of ΔG^o **does** depend on the temperature. The reaction will be product-favored at equilibrium at lower temperatures.

Sign of ΔH^o	Sign of ΔS^o	Sign of ΔG^o	Temperature Dependence of ΔG^o
+	+	- or +	The sign of ΔG^o **does** depend on the temperature. The reaction will be product-favored at equilibrium at lower temperatures.

Example 17.6.3: Predicting the Temperature Dependence of Spontaneity

The incomplete combustion of carbon is described by the following equation:

$$2\,\mathrm{C}(s) + \mathrm{O}_2(g) \longrightarrow 2\,\mathrm{CO}(g)$$

Does the sign of ΔG° of this process depend upon temperature?

Solution

Combustion processes are exothermic ($\Delta H^\circ < 0$). This particular reaction involves an increase in entropy due to the accompanying increase in the amount of gaseous species (net gain of one mole of gas, $\Delta S^\circ > 0$). The reaction is therefore product-favored at equilibrium ($\Delta G^\circ < 0$) at all temperatures.

Exercise 17.6.3

Popular chemical hand warmers generate heat by the air-oxidation of iron:

$$4\,\mathrm{Fe}(s) + 3\,\mathrm{O}_2(g) \longrightarrow 2\,\mathrm{Fe}_2\mathrm{O}_3(s)$$

Does the sign of ΔG^o of this process depend upon temperature?

Answer

ΔH^o and ΔS^o are both negative; the reaction is product-favored at equilibrium at low temperatures.

When considering the conclusions drawn regarding the temperature dependence of the sign of ΔG^o, it is important to keep in mind what the terms “high” and “low” mean. Since these terms are adjectives, the temperatures in question are deemed high or low relative to some reference temperature. A process that is reactant-favored at equilibrium at one temperature but product-favored at equilibrium at another temperature will necessarily undergo a change in “extent” (as reflected by its ΔG^o) as temperature varies. This is clearly illustrated by a graphical presentation of the free energy change equation, in which ΔG^o is plotted on the y axis versus T on the x axis:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$y = b + mx$$

Such a plot is shown in Figure 17.6.2 A process whose enthalpy and entropy changes are of the same arithmetic sign will exhibit a temperature-dependence for the sign of ΔG^o as depicted by the two yellow lines in the plot. Each line crosses from one spontaneity domain (positive or negative ΔG^o) to the other at a temperature that is characteristic of the process in question. This temperature is represented by the x-intercept of the line, that is, the value of T for which ΔG^o is zero:

$$\Delta G^\circ = 0 = \Delta H^\circ - T\Delta S^\circ$$

$$T = \frac{\Delta H^\circ}{\Delta S^\circ}$$

Thus, saying a process is product-favored at equilibrium at “high” or “low” temperatures means the temperature is above or below, respectively, that temperature at which ΔG^o for the process is zero.

Note

In this discussion, we have used two different descriptions for the meaning of the sign of ΔG^o. You should be aware of the meaning of each description.

a) **Extent of Reaction:** This description is used to predict the ratio of the product and reactant concentrations at equilibrium. In this description, we use the thermodynamic term ΔG^o to tell us the same information as the equilibrium constant, K. When $\Delta G^o < 0$, $K > 1$, and the reaction will be product-favored at equilibrium. When $\Delta G^o > 0$, $K < 1$, and the reaction is reactant-favored at equilibrium. When $\Delta G^o = 0$, $K = 1$, and the reaction will have roughly equal amounts of products and reactants at equilibrium. In all cases, the reaction will form a mixture of products and reactants at equilibrium. We use the sign and magnitude of ΔG^o to tell us how much product will be made if the reaction is allowed to reach equilibrium.

b) **Spontaneity:** This description is much more complicated because it involves two different interpretations of how a reaction at standard state occurs. One interpretation involves the hypothetical process in which the reaction proceeds from a starting point of pure reactants to a finishing point of pure products, with all substances isolated in their own containers under standard state conditions. In the second, more realistic interpretation, the reaction starts with all reactants and all products in their standard state in one container. We then allow this specific mixture to react an infinitesimally small amount so that we can obtain a rate of change in free energy with respect to the extent of reaction when all reactants and products are mixed and (essentially) in their standard states. Although each interpretation describes a different reaction scenario, the value of the difference in free energy and the value of the rate of change in free energy are the same number. If $\Delta G^o < 0$, we say that the reaction is spontaneous, meaning that the reaction would proceed in the forward direction, as written, to form pure products in their standard state. If $\Delta G^o > 0$, we say that the reaction is nonspontaneous, meaning that the reaction would proceed in the reverse direction, as written, to form pure reactants in their standard state. If $\Delta G^o = 0$, we say that the neither the reactants nor the products are favored to be formed.

A detailed treatment of the meaning of ΔG^o can be found in the paper, "Free Energy versus Extent of Reaction" by Richard S. Treptow, Journal of Chemical Education, **1996**, Volume 73 (1), 51-54.

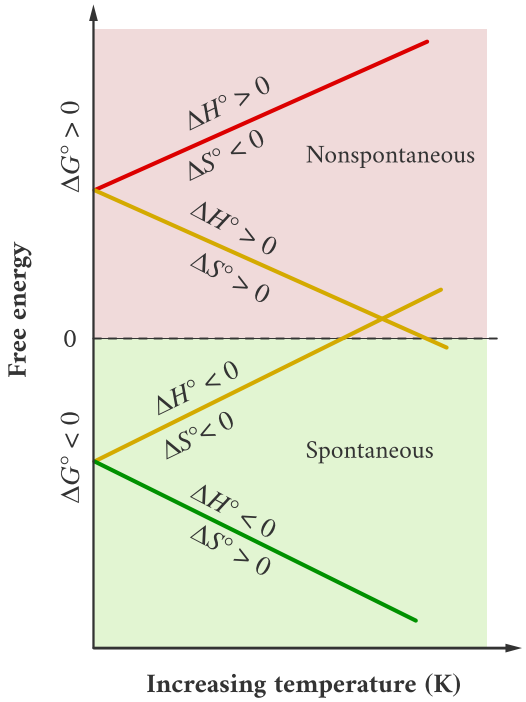


Figure 17.6.2: These plots show the variation in ΔG° with temperature for the four possible combinations of arithmetic sign for ΔH° and ΔS°. Note that in this graph of ΔG°, "spontaneous" is synonymous with "product-favored at equilibrium", and "nonspontaneous" is synonymous with "reactant-favored at equilibrium".

Example 17.6.4: Equilibrium Temperature for a Phase Transition

As defined in the chapter on liquids and solids, the boiling point of a liquid is the temperature at which its solid and liquid phases are in equilibrium (that is, when vaporization and condensation occur at equal rates). Use the information in Tables T1 or T2 to estimate the boiling point of water.

Solution

The process of interest is the following phase change:

$$\mathrm{H_2O}(l) \longrightarrow \mathrm{H_2O}(g)$$

When this process is at equilibrium, $\Delta G = 0$, so the following is true:

$$0 = \Delta H^\circ - T\Delta S^\circ \qquad \text{or} \qquad T = \frac{\Delta H^\circ}{\Delta S^\circ}$$

Using the standard thermodynamic data from Tables T1 or T2,

$$\begin{aligned}\Delta H^\circ &= \Delta H_\mathrm{f}^\circ(\mathrm{H_2O}(g)) - \Delta H_\mathrm{f}^\circ(\mathrm{H_2O}(l)) \\ &= -241.82\ \mathrm{kJ/mol} - (-285.83\ \mathrm{kJ/mol}) = 44.01\ \mathrm{kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta S^\circ &= \Delta S_{298}^\circ(\mathrm{H_2O}(g)) - \Delta S_{298}^\circ(\mathrm{H_2O}(l)) \\ &= 188.8\ \mathrm{J/K\cdot mol} - 70.0\ \mathrm{J/K\cdot mol} = 118.8\ \mathrm{J/K\cdot mol}\end{aligned}$$

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{44.01 \times 10^3\ \mathrm{J/mol}}{118.8\ \mathrm{J/K\cdot mol}} = 370.5\ \mathrm{K} = 97.3\ ^\circ\mathrm{C}$$

The accepted value for water's normal boiling point is 373.2 K (100.0 °C), and so this calculation is in reasonable agreement. Note that the values for enthalpy and entropy changes data used were derived from standard data at 298 K (Tables T1 or T2.). If desired, you could obtain more accurate results by using enthalpy and entropy changes determined at (or at least closer to) the actual boiling point.

? Exercise 17.6.4

Use the information in Tables T1 or T2 to estimate the boiling point of CS_2.

Answer

313 K (accepted value 319 K).

Free Energy and Equilibrium

The free energy change for a process may be viewed as a measure of its driving force. A negative value for ΔG represents a driving force for the process in the forward direction, while a positive value represents a driving force for the process in the reverse direction. When ΔG is zero, the forward and reverse driving forces are equal, and the process occurs in both directions at the same rate (the system is at equilibrium).

In the chapter on equilibrium the *reaction quotient*, Q, was introduced as a convenient measure of the status of an equilibrium system. Recall that Q is the numerical value of the mass action expression for the system, and that you may use its value to identify the direction in which a reaction will proceed in order to achieve equilibrium. When Q is lesser than the equilibrium constant, K, the reaction will proceed in the forward direction until equilibrium is reached and $Q = K$. Conversely, if $Q > K$, the process will proceed in the reverse direction until equilibrium is achieved.

The free energy change for a process taking place with reactants and products present under *nonstandard conditions* (pressures other than 1 bar; concentrations other than 1 M) is related to the standard free energy change according to this equation:

$$\Delta G = \Delta G^\circ + RT \ln Q$$

R is the gas constant (8.314 J/K mol), T is the kelvin or absolute temperature, and Q is the reaction quotient. For gas phase equilibria, the pressure-based reaction quotient, Q_P, is used. The concentration-based reaction quotient, Q_C, is used for condensed phase equilibria. This equation may be used to predict the spontaneity for a process under any given set of conditions as illustrated in Example 17.6.6.

✓ Example 17.6.6: Calculating ΔG under Nonstandard Conditions

What is the free energy change for the process shown here

$$2NH_3(g) \longrightarrow 3H_2(g) + N_2(g)$$

under the specified conditions?

- $T = 25^\circ C$,
- $P_{N_2} = 0.870$ atm
- $P_{H_2} = 0.250$ atm
- $P_{NH_3} = 12.9$ atm

$$\Delta G^\circ = 33.0 kJ/mol$$

Solution

The equation relating free energy change to standard free energy change and reaction quotient may be used directly:

$$\begin{aligned}\Delta G &= \Delta G^\circ + RT \ln Q \\ &= 33.0 \frac{\text{kJ}}{\text{mol}} + \left(8.314 \frac{\text{J}}{\text{mol K}} \times 298 K \times \ln \frac{\left(0.250^3\right) \times 0.870}{12.9^2} \right) \\ &= 9680 \frac{\text{J}}{\text{mol}} \text{ or } 9.68 \text{ kJ/mol}\end{aligned}$$

Since the computed value for ΔG is positive, the reaction is nonspontaneous under these conditions.

Exercise 17.6.6

Calculate the free energy change for this same reaction at 875 °C in a 5.00 L mixture containing 0.100 mol of each gas. Is the reaction spontaneous under these conditions?

Answer

ΔG = –123.5 kJ/mol; yes

For a system at equilibrium, $Q = K$ and $\Delta G = 0$, and the previous equation may be written as

$$0 = \Delta G^\circ + RT \ln K \quad \text{(at equilibrium)}$$

$$\Delta G^\circ = -RT \ln K$$

or

$$K = e^{-\frac{\Delta G}{RT}}$$

This form of the equation provides a useful link between these two essential thermodynamic properties, and it can be used to derive equilibrium constants from standard free energy changes and vice versa. The relations between standard free energy changes and equilibrium constants are summarized in Table 17.6.2

Table 17.6.2: Relations between Standard Free Energy Changes and Equilibrium Constants

K	$\Delta G°$	Composition of an Equilibrium Mixture
> 1	< 0	Products are more abundant
< 1	> 0	Reactants are more abundant
= 1	= 0	Reactants and products are comparably abundant

Example 17.6.7: Calculating an Equilibrium Constant using Standard Free Energy Change

Given that the standard free energies of formation of $Ag^+(aq)$, $Cl^-(aq)$, and $AgCl(s)$ are 77.1 kJ/mol, −131.2 kJ/mol, and −109.8 kJ/mol, respectively, calculate the solubility product, K_{sp}, for AgCl.

Solution

The reaction of interest is the following:

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq) \quad K_{sp} = \left[Ag^+\right]\left[Cl^-\right]$$

The standard free energy change for this reaction is first computed using standard free energies of formation for its reactants and products:

$$\begin{aligned}\Delta G^\circ &= \left[\Delta G_f^\circ\left(\mathrm{Ag^+(aq)}\right) + \Delta G_f^\circ\left(\mathrm{Cl^-(aq)}\right)\right] - \left[\Delta G_f^\circ(\mathrm{AgCl(s)})\right] \\ &= [77.1\ \mathrm{kJ/mol} - 131.2\ \mathrm{kJ/mol}] - [-109.8\ \mathrm{kJ/mol}] = 55.7\ \mathrm{kJ/mol}\end{aligned}$$

The equilibrium constant for the reaction may then be derived from its standard free energy change:

$$K_{sp} = e^{-\frac{\Delta G^\circ}{RT}} = \exp\left(-\frac{\Delta G^\circ}{RT}\right) = \exp\left(-\frac{55.7 \times 10^3 J/mol}{8.314 J/mol \cdot K \times 298.15 K}\right) = \exp(-22.470) = e^{-22.470} = 1.74 \times 10^{-10}$$

This result is in reasonable agreement with the value provided in Appendix J.

Exercise 17.6.7

Use the thermodynamic data provided in Appendix G to calculate the equilibrium constant for the dissociation of dinitrogen tetroxide at 25 °C.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

Answer

$K = 3.1$

To further illustrate the relation between these two essential thermodynamic concepts, consider the observation that reactions spontaneously proceed in a direction that ultimately establishes equilibrium. As may be shown by plotting the free energy versus the extent of the reaction (for example, as reflected in the value of Q), equilibrium is established when the system's free energy is minimized (Figure 17.6.3). If a system consists of reactants and products in nonequilibrium amounts ($Q \neq K$), the reaction will proceed spontaneously in the direction necessary to establish equilibrium.

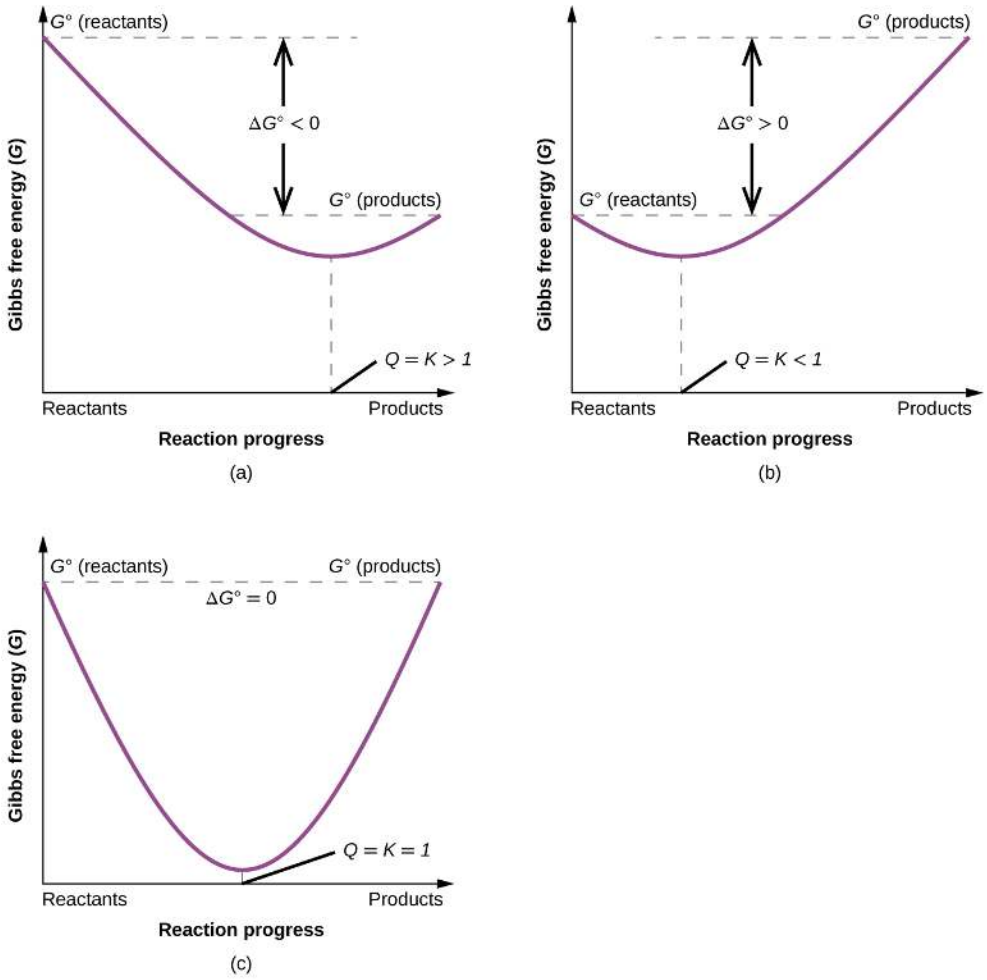


Figure 17.6.3: These plots show the free energy versus reaction progress for systems whose standard free energy changes are (a) negative, (b) positive, and (c) zero. Nonequilibrium systems will proceed spontaneously in whatever direction is necessary to minimize free energy and establish equilibrium.

This page titled 17.6: Free Energy is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.4: Free Energy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

- **16.4: Gibbs Energy** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

17.7: Key Terms

entropy (*S*) | state function that is a measure of the matter and/or energy dispersal within a system, determined by the number of system microstates; often described as a measure of the disorder of the system

Gibbs free energy change (*G*) | thermodynamic property defined in terms of system enthalpy and entropy; all spontaneous processes involve a decrease in *G*

microstate | possible configuration or arrangement of matter and energy within a system

nonspontaneous process | process that requires continual input of energy from an external source

reversible process | process that takes place so slowly as to be capable of reversing direction in response to an infinitesimally small change in conditions; hypothetical construct that can only be approximated by real processes

second law of thermodynamics | all spontaneous processes involve an increase in the entropy of the universe

spontaneous change | process that takes place without a continuous input of energy from an external source

standard entropy (*S*°) | entropy for one mole of a substance at 1 bar pressure; tabulated values are usually determined at 298.15 K

standard entropy change (Δ*S*°) | change in entropy for a reaction calculated using the standard entropies

standard free energy change (Δ*G*°) | change in free energy for a process occurring under standard conditions (1 bar pressure for gases, 1 M concentration for solutions)

standard free energy of formation | change in free energy accompanying the formation of one mole of substance from its elements in their standard states

third law of thermodynamics | entropy of a perfect crystal at absolute zero (0 K) is zero

---

This page titled 17.7: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

- **16.5: Key Terms** is licensed CC BY 4.0.

17.8: Key Equations

$$S = k \ln W$$

$$\Delta S_{univ} = \Delta S_{sys} + \Delta S_{surr}$$

$$\Delta G = \Delta H - T\Delta S$$

This page titled 17.8: Key Equations is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.6: Key Equations** is licensed CC BY 4.0.

17.9: Summary

Chemical and physical processes have a natural tendency to occur in one direction under certain conditions. A spontaneous process occurs without the need for a continual input of energy from some external source, while a nonspontaneous process requires such. Systems undergoing a spontaneous process may or may not experience a gain or loss of energy, but they will experience a change in the way matter and/or energy is distributed within the system.

Entropy (S) is a state function that can be related to the number of microstates for a system (the number of ways the system can be arranged) and to the ratio of reversible heat to kelvin temperature. It may be interpreted as a measure of the dispersal or distribution of matter and/or energy in a system, and it is often described as representing the "disorder" of the system.

For a given substance, entropy depends on phase with $S_{solid} < S_{liquid} < S_{gas}$. For different substances in the same physical state at a given temperature, entropy is typically greater for heavier atoms or more complex molecules. Entropy increases when a system is heated and when solutions form. Using these guidelines, the sign of entropy changes for some chemical reactions and physical changes may be reliably predicted.

The second law of thermodynamics states that a spontaneous process increases the entropy of the universe, $S_{univ} > 0$. If $\Delta S_{univ} < 0$, the process is nonspontaneous, and if $\Delta S_{univ} = 0$, the system is at equilibrium. The third law of thermodynamics establishes the zero for entropy as that of a perfect, pure crystalline solid at 0 K. With only one possible microstate, the entropy is zero. We may compute the standard entropy change for a process by using standard entropy values for the reactants and products involved in the process.

Gibbs free energy (G) is a state function defined with regard to system quantities only and may be used to predict the spontaneity of a process. A negative value for ΔG indicates a spontaneous process; a positive ΔG indicates a nonspontaneous process; and a ΔG of zero indicates that the system is at equilibrium. A number of approaches to the computation of free energy changes are possible.

This page titled 17.9: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **16.7: Summary** is licensed CC BY 4.0.

17.10: Exercises

These are homework exercises to accompany the Textmap created for "General Chemistry: Principles and Modern Applications " by Petrucci et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Q18.1a

Explain how you know if each of the following change in states result in an increase or decrease in entropy.

a. solid to liquid
b. liquid to gas
c. gas to liquid

S18.1a

a. Increase in entropy. Liquids have more entropy than solids do.
b. Increase in entropy. Gas has more entropy than liquids do.
c. Decrease in entropy. Liquids have less entropy than gases.

Q18.1b

Determine whether each of the following is an increase or decrease in entropy.

a. fusion
b. freezing
c. combustion

S18.1b

a. increase
b. decrease
c. increase

Q18.2a

Which of the following changes in states represent the smallest Delta S. Which one has the largest Delta S? How do you know?

a. freezing of water to ice
b. melting of ice to liquid water
c. sublimation of ice to gas

S18.2a

The freezing of water is a decrease in entropy so it has the smallest Delta S because solids have less entropy than liquids. The other two options represent an increase in Entropy. The sublimation of ice is a large increase in Entropy because gas has more Entropy than solids. The sublimation of ice to gas is more of an increase than the melting of ice to water because gas has more entropy than liquid and therefore the change must have a larger Delta S.

Q18.2c

Arrange the compounds in order of increasing entropy (S) and justify your order.

- $CH_3OH_{(s)}$
- $CH_3OH_{(l)}$,
- $CH_3OH_{(g)}$,
- $CH_3CH_2OH_{(g)}$

S18.2c

A- CH_3 OH(s)<CH_3 OH (l)< CH_3 OH(g)< CH_3 CH_2 OH(g)

The bigger the compound, the greater the entropy as there are more degrees of freedom.

Q18.2d

Given the reaction of diamond converting to graphite

$$2C_{(s\ diamond)} \rightarrow 2C_{(s\ (graphite)} \tag{17.10.1}$$

Determine ΔG at 298 K and determine if this reaction is spontaneous or not. What does ΔG say about the rate of this reaction?

- $\Delta H^\circ{}_f(C\ (s)\ diamond) = 1.9\ kJ/mol$
- $S^\circ(C_{(s)}\ diamond = 2.38\ J/(mol\ K)$
- $S^\circ(C_{(s)}\ graphite) = 5.74\ J/(mol\ K)$

S18.2d

$$\Delta G = \Delta H - T\,\Delta\,S \tag{17.10.2}$$

ΔH=(2mol C(s)_(graphite)×ΔH°_(f C(s)graphite))-(2mol C(s)_diamond×ΔH°_(f C(s)diamond))

ΔH=2(0)-2(1.9 kJ/mol)=-3.8kJ

ΔS=(2mol C(s)_(graphite)× S_(C(s)graphite))-(2mol C(s)_diamond× S_C(s)diamond)

ΔS=2(5.74 J/(mol K))- 2(2.38 J/(mol K))=+6.72 J/mol

ΔG=-3.8×10^3 J-298.15(6.72 J/mol K)=-5.51 kJ

This reaction is spontaneous. ΔG says nothing about the kinetics; while this reaction is thermodynamics favored, it occurs on a 100 million year timescale. Diamonds are kinetically meta-stable materials.

Q18.2

Given these reactions reactions, determine whether the ΔS increases or decreases or stays the same at 25 °C;

a. $H_2SO_4(l,\ 1atm) \rightarrow H_2SO_{4(s)}(1\ atm)$
b. $H_2O(l,\ 1atm) \rightarrow H_2O(g,\ 1atm)$
c. $CH_{4(g)} + 2O_{2(g)} \rightleftharpoons CO_{2(g)} + 2H_2O(l)$

S18.2

a. decreases because of a phase change from a liquid to a solid.
b. increases because of a phase change from a liquid to a gas.
c. decreases because there are more moles of gas in the reactants.

Q18.3

“Spontaneous reactions are faster than non spontaneous reactions”. Access this statement and explain why it is true or false.

S18.3

This statement is false because spontaneity does not determine speed; there are spontaneous reactions that range from very slow like melting an ice cube in cold water, to very fast like melting an ice cube in hot water.

Q18.25a

Given the dissolving of $CaCl_{2(s)}$

$$CaCl_{2(s)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} \tag{17.10.3}$$

- the $\Delta G^\circ{}_f$ of $CaCl_2$ (s)= -748.1 kJ/mol,
- the ΔH of $CaCl_2$= -795.8 kJ
- the ΔS^o
  - Ca^{2+}=-53.1 J/Kmol,
  - $CaCl_2$=104.6 J/Kmol,
  - Cl-=56.5

Calculate the temperature of this reaction.

S18.25a

$$\Delta G^\circ = \Delta H - T\Delta S \tag{17.10.4}$$

$$T = \frac{\Delta H - \Delta G^\circ}{S} \tag{17.10.5}$$

T=(0-[-795.8])-(-748.1))[(2(56.5)+(-53.1)]-[104.6]

T=-34.539 K

Q18.25b

At what temperature would the following reaction occur spontaneously? (Delta H°= 1256.4 kJ; Delta S°=587 J/K)

S18.25b

$$2Cr_2O_3 \rightarrow 4Cr + 3O_2 \tag{17.10.6}$$

Delta G°= Delta H°-T DeltaS°=0

1256.4kJ=T (0.587 kJ/K) T=2140.37K

Q18.27a

Calculate ΔG^o for

$$CS_{2(l)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2SO_{2(g)} \tag{17.10.7}$$

using **only** the following ΔG values for the reaction:

- $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ with $\Delta G^o = -394.39\ kJ/mol$
- $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$ with $\Delta G^o = -300.13\ kJ/mol$
- $C_{(s)} + 2S_{(s)} \rightarrow CS_{2(l)}$ with $\Delta G^o = 67.1\ kJ/mol$

S18.27a

A- C(s)+O_2 (g)→CO_2 (g);ΔG=-394.39 kJ/mol

2[S(s)+O_2 (g)→SO_2 (g)];ΔG=2[-300.13] kJ/mol

CS_2 (l)→C(s)+2S(s);ΔG=-(67.1) kJ/mol

ΔG= -394.39 kJ/mol + 2[-300.13] kJ/mol + -(67.1) kJ/mol = -1061.75

Q18.27b

Use the given the standard Gibbs energy changes for these equations:

a. $2Fe2O3(s) \rightarrow 4Fe(s) + 3O_2(g) \quad \Delta G^\circ = -742.2\ kJ/mol$
b. $Fe(OH)_3(s) \rightarrow 3Fe(s) + OH^-(g) \quad \Delta G^\circ = -696.5\ kJ/mol$
c. $Fe_3O_4(s) \rightarrow 3Fe(s) + 2O_2(g) \quad \Delta G^\circ = -1015\ kJ/mol$

to identify the ΔG° for the following reaction

$$6Fe_2O_3(s) \rightarrow O_2(g) + 4Fe_3O_4(s) \tag{17.10.8}$$

S18.27b

6Fe2O3(s)-->O2(g)+4Fe3O4(s)

Fe3O4 (s) -->3Fe(s)+2O2(g) ΔG°=4(-1015) kJ/mol

2Fe2O3 (s)-->4Fe(s)+3O2 (g) -3(-742.2) kJ/mol

---

net: 6Fe2O3(s)-->O2(g)+4Fe3O4(s) ΔG°=(-3(-742.2))+4(-1015)=3837.4 kJ/mol

Q18.27c

Find delta G° at 298.15 K for reaction

$$2CO + O_2 \rightarrow 2CO_2 \tag{17.10.9}$$

DH°= -128.3 kJ DS°=- 159.5 J K^{-1}

S18.27c

DG°= DH° - TDS°

Delta G°= (-128.3 kJ) – [298.15K x (-159.5 JK^{-1} x (1 J/1000kJ))]= -80.75 kJ

Q18.29a

Write a the chemical equation for the complete combustion of methane, CH_4 gas and calculate ΔG° at 298.15 K.

S18.29a

$$5CH_{4(g)} + 5O_2(g) \rightleftharpoons 5CO_{2(g)} + H_2O_{(g)} \tag{17.10.10}$$

From Tables T1 or T2

- ΔG° of $CO_{2(g)}$ is -393.3 kJ/mol
- ΔG° of $H_2O_{(l)}$ is -228.6 kJ/mol

So the standard products minus reactions approach:

ΔG° = (5*[(CO)_2] + [H_2 O]) – (5*[(CH)_4) = ((5*-394.4) + (-228.6)) – (5*-50.72) ΔG° = -1947 kJ

Q18.29

Solve for ΔG° at 298 K for the combustion of $B_2H_{6\ (g)}$ in $B_2O_{3(s)}$ and $H_2O_{(l)}$.

S18.29

First the balanced reaction must be identified

$$B_2H_{6(g)} + 3O_{2(g)} \rightleftharpoons B_2O_{3(s)} + 3H_2O_{(l)} \tag{17.10.11}$$

Using the parameters from Table T1 or T2

$$\Delta G^\circ = [3(-237.1) + (-1194)] - [86.7 + 3(0)] = -11,992\ kJ/mol \tag{17.10.12}$$

The reaction is (strongly) spontaneous since $\Delta G^\circ < 0$ and essentially goes to completion (irreversible reaction).

Q18.31a

Assess the feasibility of the reaction below by evaluating (ΔS^o_{rxn}, ΔH^o_{rxn} and ΔG^o_{rxn}) for this reaction at 25°C. If the reaction is spontaneous, is it enthalpically favored, entropically favored or both?

$$(NH)_4Cl_{(s)} \rightarrow NH_{3(g)} + HCl_{(g)} \tag{17.10.13}$$

S18.31a

$$\Delta S^o = \sum(\Delta S^o)(products) - \sum(\Delta S^o)(reactants) \tag{17.10.14}$$

(192.5 + 186.9) – (94.6) = 284.8 kJ/mol K

$$\Delta H^o_{rxn} = \sum(\Delta H^o)(products) - \sum(\Delta H^o)(reactants) \tag{17.10.15}$$

(-46.11 + -92.31) – (-314.4) = 175.98 kJ/mol

The Gibbs energy can be also tabulated,

$$\Delta G^o_{rxn} = \sum(\Delta G^o)(products) - \sum(\Delta G^o)(reactants) \tag{17.10.16}$$

but can be calculate directly from ΔS^o_{rxn} and ΔH^o_{rxn}

$$\Delta G^o_{rxn} = \Delta H^o_{rxn} - T\,\Delta\,S^o_{rxn} \tag{17.10.17}$$

$$\Delta G^o_{rxn} = 175.98\ kJ/mol - 298\ K(284.8) = -84,737\ kJ/mol \tag{17.10.18}$$

The thermal decomposition of (NH)_4Cl_{(s)} is a strongly spontaneous process that is entropically driven, but not enthalpically driven.

Q18.31a

Given the reaction at 298 K;

$$4NH_{3\ (g)} + 5O_{2\ (g)} \rightarrow 4NO_{(g)} + 6H_2O_{(l)} \tag{17.10.19}$$

Determine the S

S18.31a

ΔS°=ΔS°products-ΔS°reactants

ΔS°=(69.91+210.8)-(111.3)

ΔS°=169.41 J/molK

Q18.31b

Assess the feasibility of the reaction:

$ClNO_2$ (g) + NO (g) --> NO_2 (g) + ClNO (g)

by calculating each of the following for the reaction at 25 °C.

a. ΔS° (standard molar entropy for $ClNO_2$ (g) is 272.23 J/mol•K; entropy for NO (g) is 210.8 J/mol•K; entropy for NO_2 (g) is 240.1 J/mol•K; entropy for ClNO (g) is 261.58 J/mol•K.)
b. ΔH° (ΔH_f for $ClNO_2$ (g) is 12.5 kJ/mol; ΔH_f for NO (g) is 90.25 kJ/mol; ΔH_f for NO_2 (g) is 33.18 kJ/mol; ΔH_f for ClNO (g) is 51.71 kJ/mol.)
c. ΔG°

S18.31b

a) ΔS° = (240.1 + 261.58) – (272.23 + 210.8) = 18.65 J / mol • K

b) ΔH° = (33.18 + 51.71) – (12.5 + 90.25) = -17.86 kJ / mol

c) ΔG° = ΔH° - TΔS° = -17860 J / mol – (298.15 K)(18.65 J / mol • K) = -23420 J / mol = -23.42 kJ / mol

ΔG° < 0... therefore... reaction is spontaneous.

Q18.37

Determine the K_p at 298.15 K for the reaction:

$$2CO_{(g)} + O_{2(g)} \rightleftharpoons 2CO_2 \tag{17.10.20}$$

S18.37

$$\Delta G^\circ = \sum(\Delta G^\circ)_p roducts - \sum(\Delta G^\circ)_r eactants \tag{17.10.21}$$

(2*-137.2) – (2*-394.4) = 514.4 kJ/mol

$$\Delta G^p = -RT \ln K_p \tag{17.10.22}$$

$$\ln K_p = \frac{\Delta G^\circ}{-RT} \tag{17.10.23}$$

- (514.4 kJ/mol)/((8.3145×10)^3*298.15 K)) = -207.5

K_p = e^(-207.5) = (7.65×10)^(-91)

Q18.37

Given the reaction at 298 K and the thermodynamic values in Table T1

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \tag{17.10.24}$$

estiamte K_p.

S18.37

$$\Delta G^\circ = (2(-371.10)) - (2(-300.2)) \tag{17.10.25}$$

$$\Delta G^\circ = -141.8 kJ/mol \tag{17.10.26}$$

$$lnKp = -(-141.8)/((8.3145E-3)(298)) \tag{17.10.27}$$

$$\ln K_p = 57.23 \tag{17.10.28}$$

$$K_p = e^{57.23} = 7.156 \times 10^{24} \tag{17.10.29}$$

Video Solution

Q18.39

Determine values of ΔG° and K at 298.15 K for the following reaction:

$$(2SO)_2(g) + O_2(g) \rightleftharpoons (2SO)_3(g) \tag{17.10.30}$$

S18.39

ΔG° = ∑ (ΔG°)_products - ∑(ΔG°)_reactants

(2*-371.1) – (2* - 300.2) = -141.8 kJ/mol

lnK= (-ΔG)/RT = ((141.8×10)^3 J/mol)/((8.3145 J/mol)(298.15 K)) = 57.2

K= e^57.2 = (6.94×10)^24

Q18.39

Given the reaction at 298 K and the thermodynamic values in Table T1

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(l) \tag{17.10.31}$$

solve for K_p.

S18.39

ΔG°=(6(-237.6)+(4(86.55))-(4(-16.45))

=-1010.6kJ/mol

=-RTlnKp

lnKp=-(-1010.6)/((8.3145E-3)(298))

Kp=$e^{407.87}$

Video Solution

Q18.43

ΔG° for the reaction below is 31.4 kJ/mol at 298.15 K. Use the thermodynamics quantity to decide in which direction the reaction is spontaneous when the concentrations of $H_2SO_{4(aq)}$, $HSO^-_{4(aq)}$, and $H_3O^+_{(aq)}$ are 0.2 M, 0.1 M, and 0.1 M respectively.

$$H_2SO_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO^-_{4(aq)} \tag{17.10.32}$$

S18.43

$$Q_C = \frac{[H_3O^+][HSO_4^-]}{[H_2SO_4]} = \frac{(0.1\ M)^2}{0.2\ M} = 0.05 \tag{17.10.33}$$

$$\Delta G = 31.4kJ + ((8.3145 \times 10)^3\ kJ/(mol * K))(298.15\ K) \ln 0.05 = 23.97kJ \tag{17.10.34}$$

Because ΔG is positive, the reaction is not spontaneous going in the forward direction under these conditions.

Q18.43

Given the ionization of ethanol in water at 298 K;

$$CH_3COH_{(g)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3CO^-_{(aq)} \tag{17.10.35}$$

when the ΔG°=28 kJ/mol and the concentration of CH3CHO(g)=0.1 M, H3O+(aq)= 0.09129 M CH3CO-(aq)=0.09129 M consider in which direction the spontaneous reaction will go.

S18.43

Qc=0.091292(0.1-0.09129) Qc=0.949 ΔG°=28 +(8.314E-3)(298)(ln.949)=+27.87

spontaneous rxn happens in backward direction .

Q18.47

Kp=0.12520.3352 Kp=0.139

ΔG°1000k=-RTlnK =-(8.3145J/molK)(1000K)ln0.139 =1.64 kJ/mol

Qp=(0.032)(0.022)(0.071)(0.095)

Qp=0.104

Because $Q_p < K_p$, the spontaneous reaction will go to the right to form more products.

S18.45a

Kp can be calculated by using the partial pressures of each element in the system. However, since partial pressures are not given, one can use the mole fractions of products over reactants, which in this case is the same as moles of products over reactants. This is because A, B, C, and D are all divided by the same total moles.

$$K_p = \frac{(0.6)(0.6)}{(0.4)(0.4)} = 2.25 \tag{17.10.36}$$

ΔG°500K=-RTlnKp= -3.371 KJ

Qp=(.2X.017)/(.01X.3)= 1.133<2.25, which implies that the reaction will proceed to the right to reach equilibrium.

Q18.45b

For the reaction

$$NH_{3\ (aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4\ (aq)} + OH^-_{(aq)} \tag{17.10.37}$$

which equation is *invalid*.

a. $\Delta G = \Delta G^\circ + RT \ln Q$
b. $K = K_p$
c. $K_p = e^{-\Delta G^\circ / RT}$

S18.45b

C is incorrect because the correct equation is G=-RTlnKp, thus correct format would be Kp=e-GRT

Q18.45c

For reactions involving only gases, which of the following statements is false?

a. ΔG°1/(-RT)=lnk
b. TΔS=ΔH°-ΔG
c. Kc=([products])/([reactants])
d. K=e^(ΔG°1/RT)

Q18.45c

D is incorrect because the exponent,ΔG°1/RT, is supposed to be multiplied by negative 1.

Q18.47

In the reaction at 1000K:

$$NO_{2\ (g)} + H_{2\ (g)} \rightleftharpoons NO_{(g)} + H_2O_{(g)} \tag{17.10.38}$$

calculate Kp when there are 0.335 mol NO2(g), 0.335 mol H2(g), 0.125 mol NO(g), and 0.125 mol H2O(g). Solve for ΔG°at 1000K. Evaluate in what direction the spontaneous reaction will go in if there are 0.071 mol NO2(g), 0.095 molH2(g), 0.032 molNO(g), and 0.022 mol H2O(g).

Q18.47

The imaginary reaction A + B= C + D is at equilibrium at 500 K when there is .4 mol A, .4 mol B, .6 mol C, and .6 mol D.

a. Calculate Kp at 500 K.
b. Calculate ΔG° at 500K
c. What direction would a spontaneous reaction occur if the following were the new number of moles present in the system; .01 mol A, .3 mol B, 0.2 mol C, .017 mol D?

Q18.49a

Calculate ΔG° for each of the following equilibrium reactions at the given temperature. (Kc values not verified for accuracy. Just for practice purposes).

a. 2H2S(g)+ CH4(g)= 4H2(g) +Cs2(g) Kc=5.27X10-8 at 400K
b. SO3(g) + NO(g)→SO2(g) + NO2(g) Kc= 2.4X10-6 at 300 K
c. C3H8(g)+CH4(g)→ 2C2H6(g) Kc= 3.6X10-4 550 K

S18.49a

$$\Delta G^\circ = -RT\ln K_p \tag{17.10.39}$$

a. $\Delta G^\circ = -(8.3145)(400\ K)\ln(5.27\times10^{-8}) = 55.7\ KJ$
b. $\Delta G^\circ = -(8.3145)(300\ K)\ln(2.4\times10^{-6}) = 32.2\ KJ$
c. $\Delta G^\circ = -(8.3145)(550\ K)\ln(3.6\times10^{-4}) = 36.3\ KJ$

Q18.49

Given the following equations solve for the

a. 2SO2(g)+2H2(g)2SO3(g) 125 C and $K_c = 28$
b. 4NH3(g)+5O2(g)4NO(g)+6H_2O(l), 332 C and $K_c = 32$

S18.49

a. ΔG°=-(8.3145E-3 kJ/molK)(125+273K)(ln28)=-11.02 kJ/mol
b. ΔG°=-(8.3145E-3 kJ/molK)(332+273K)(ln32)=-17.43 kJ/mol

Q18.53A

Consider the following imaginary reaction:

$$A(s) + B(g) \rightleftharpoons C(s) \tag{17.10.40}$$

At 30 degrees, ΔG°= 40.00 KJ

Prove that the partial pressure of B(g) is very large.

What conditions could be changed to decrease the B(g) partial pressure.

S18.53

ΔG°=-RTlnKp

40 KJ=-.0083145X303XlnKp

Kp=e-15.8774

Partial pressure of B(g)=1/(e^-15.8774) which is a very large number.

To decrease the partial pressure of B(g), one must make Kp larger which can occur by decreasing the temperature.

Q18.55

Use data from Appendix D to calculate at 298 K for the following reaction:

H2O2(g) + CH4(g) = CH3OH (g) + H2O(g)

ΔS° b) ΔH° c) ΔG° d) K

S18.55

ΔS° = 239.8 J + 188.8 J – 186.3 J +188.8 J = 53.5 J

ΔH° = -200.7-241.8+136.3+ 74.81 = -231.39 KJ

ΔG° = - 162 – 228.6 +105.6 +50.72 = -234.28 KJ

ΔG°=-RTlnKp

-234.28 KJ= -.0083145X298XlnK

K=e^944.555

Q18.57A

What is the temperature of the following reaction if ΔG°=-53.2 kJ, ΔH°=-21.5 kJ, and ΔS°=14.7J/K?

$$2NO(g) + O_2(g) \rightarrow (2NO)_2(g) \tag{17.10.41}$$

S18.57A

In this problem we are asked to calculate the temperature of the given reaction. This can be done after rearranging ΔG°=ΔH°-TΔS° in order to find T.

ΔG°=ΔH°-TΔS°

TΔS°=ΔH°-ΔG°

T=(ΔH°-ΔG°)/(ΔS°)

Find T:

T=(-21.5×(10)^3 J-(-53.2×(10)^3 J))/((14.7 J)/K)=2.6 ×(10)^3 K

Q18.57

What must be the temperature (K) if the following reaction has ΔG°= -38.5 kJ, ΔH°= -14.6 kJ, and ΔS°= 0.152 kJ/ K?

Fe2O3 (s) + 3CO (g) → 2Fe (s) + 3CO2 (g)

S18.57

ΔG°=ΔH°-T ΔS°

(-38.5 kJ) = (-14.6 kJ) - T(0.152 kJ/K)

-157.24 = -T

T=157.24 K

Q18.53

For the decomposition reaction

$$2MgO_{(s)} \rightleftharpoons 2Mg_{(s)} + O_{2\ (g)} \tag{17.10.42}$$

where at 298 K

- $H = -601.7 kJ/mol$ and
- $G = -569.4 kJ/mol$

solve for K_p.

Q18.55

Using the equation FeCO3(s) Fe(s)+CO3(aq), solve for H& S at 298K.

Q18.54

Determine the equilibrium constant at 298 K for the following reaction:

$$Mg(OH)_{2(s)} + 2H^+_{(aq)} \rightleftharpoons Mg^{2+}_{(aq)} + 2H_2O_{(l)} \tag{17.10.43}$$

- ΔG°: [H_2O (l)]= -237.1 kJ/mol,
- ΔG°: [Mg^2+ (aq)]= -454.8 kJ/mol,
- ΔG°: [Mg(OH)2(s)]= -833.5 kJ/mol

S18.54

ΔG°=[2[H_2O(l)]+ [Mg^2+ (aq)]] - [Mg(OH)2 (s)]

ΔG°=[2[-237.1 kJ/mol]+ [-454.8 kJ/mol] - [-833.5 kJ/mol]

ΔG°=-95.5 kJ/ mol = -95.5e3 J/mol

ΔG°= -RTlnK

lnK = -ΔG°/ RT= -(-95.5e3 J/mol)/ (8.3145 J/ mol K* 298K) = 38.5

K = e^(38.5) = 5e16

Q57A

What must the temperature be if the following reaction has ΔG°=-53.2 kJ,ΔH°=-21.5 kJ,and ΔS°=14.7J/K?

$$2NO_{(g)} + O_{2(g)} \rightarrow (2NO)_{2(g)} \tag{17.10.44}$$

S57A

In this problem we are asked to calculate the temperature of the given reaction. This can be done after rearranging ΔG°=ΔH°-TΔS° in order to find T.

ΔG°=ΔH°-TΔS°

TΔS°=ΔH°-ΔG°

T=(ΔH°-ΔG°)/(ΔS°)

Find T:

T=(-21.5×(10)^3 J-(-53.2×(10)^3 J))/((14.7 J)/K)=2.6 ×(10)^3 K

Q59A

The synthesis of water occurs by the reaction H+(aq) + OH-(aq) →H2O(l) at 25°C. Using the following information, and assuming that ΔH° and ΔS° are essentially unchanged, estimate Kp at 25°C.

ΔH°(H^+)=0; ΔH°((OH)^-)=-230.0 kJ/mol; ΔH°(H_2 O)=-285.8 kJ/mol

ΔS°(H^+)=0; ΔS°((OH)^-)=-10.75 J/(mol K); ΔS°(H_2 O)=69.91 J/(mol K)

S59A

First, determine the value of ΔG° at 25°C, from the ΔH° and ΔS° values.

ΔH°=Δ(H°)_f [H_2 O(l)]-Δ(H°)_f [H^+ (aq)]-Δ(H°)_f [(OH)^- (aq)]

=(-285.8 kJ/mol)-(0)-(-230.0 kJ/mol)=-55.8 kJ/mol H^+

ΔS°=S°[H_2 O(l)]-S°[H^+ (aq)]-S°[(OH)^- (aq)]

=(69.91 J/(mol K))-(0)—10.75 J/(mol K)=80.66 J/(mol K) (H)^+

ΔG°=ΔH°-TΔS°=-55.8 kJ/mol-298K×(.08066 kJ/(mol K))=-79.84 kJ/mol=-RT ln(K_p)

ln(K_p=(-ΔG°)/RT=(79.84×(10)^3 J/mol)/(8.3145 J/(mol K)×298 K)=32.2; K_p=e^32.2=9.6×(10)^13)

Q59

The synthesis of water occurs by the reaction H+(aq) + OH-(aq) →H2O(l) at 25°C. Using the following information, and assuming that ΔH° and ΔS° are essentially unchanged, estimate Kp at 25°C.

ΔH°(H^+)=0; ΔH°((OH)^-)=-230.0 kJ/mol; ΔH°(H_2 O)=-285.8 kJ/mol

ΔS°(H^+)=0; ΔS°((OH)^-)=-10.75 J/(mol K); ΔS°(H_2 O)=69.91 J/(mol K)

S59

First, determine the value of ΔG° at 25°C, from the ΔH° and ΔS° values.

ΔH°=Δ(H°)_f [H_2 O(l)]-Δ(H°)_f [H^+ (aq)]-Δ(H°)_f [(OH)^- (aq)]

=(-285.8 kJ/mol)-(0)-(-230.0 kJ/mol)=-55.8 kJ/mol H^+

ΔS°=S°[H_2 O(l)]-S°[H^+ (aq)]-S°[(OH)^- (aq)]

=(69.91 J/(mol K))-(0)—10.75 J/(mol K)=80.66 J/(mol K) (H)^+

ΔG°=ΔH°-TΔS°=-55.8 kJ/mol-298K×(.08066 kJ/(mol K))=-79.84 kJ/mol=-RT ln(K_p)

ln(K_p=(-ΔG°)/RT=(79.84×(10)^3 J/mol)/(8.3145 J/(mol K)×298 K)=32.2; K_p=e^32.2=9.6×(10)^13)

Q61

Determine the temperature at which the K_p=1.0×(10)^5 for the reaction \[N_{2 (g)}+(3H)_{2 (g)} \rightarrow (2NH)_{3 (g)}\). Using the following information:

ΔH°(N_2)=0; 3ΔH°(H_2)=0; 2ΔH°((NH)_3)=-46.11 kJ/mol

ΔS°(N_2)=191.6 J/(mol K); 3ΔS°(H_2)=130.7 J/(mol K); 2ΔS°((NH)_3)=192.5 J/(mol K)

S61

ΔH°=2Δ(H°)_f [(NH)_3 (g)]-Δ(H°)_f [N_2 (g)]-3Δ(H°)_f [H_2 (g)]

=2(-46.11 kJ/mol)-(0)-(0)=-92.22 kJ/mol

ΔS°=2S°[(NH)_3 (g)]-S°[N_2 (g)]-3S°[H_2 (g)]

=2(192.5 J/(mol K))-(191.6 J/(mol K))-3(130.7 J/(mol K))=-198.7 J/(mol K)

ΔG°=ΔH°-TΔS°=-RT lnK

ΔH°-TΔS°-RT lnK

T=(ΔH°)/(ΔS°-R lnK)

T=(-92.22×(10)^3 J/mol)/(-198.7 J/(mol K)-8.3145 J/(mol K) ln(1.0×(10)^5))=313 K

Q61

Determine the temperature at which the $K_p = 1.0 \times 10^5$ for the reaction:

$$N_2(g) + (3H)_2(g) \rightleftharpoons (2NH)_3(g) \tag{17.10.45}$$

Using the following information:

- ΔH°(N_2)=0; 3ΔH°(H_2)=0; 2ΔH°((NH)_3)=-46.11 kJ/mol
- ΔS°(N_2)=191.6 J/(mol K); 3ΔS°(H_2)=130.7 J/(mol K); 2ΔS°((NH)_3)=192.5 J/(mol K)

S61

ΔH°=2Δ(H°)_f [(NH)_3 (g)]-Δ(H°)_f [N_2 (g)]-3Δ(H°)_f [H_2 (g)]

=2(-46.11 kJ/mol)-(0)-(0)=-92.22 kJ/mol

ΔS°=2S°[(NH)_3 (g)]-S°[N_2 (g)]-3S°[H_2 (g)]

=2(192.5 J/(mol K))-(191.6 J/(mol K))-3(130.7 J/(mol K))=-198.7 J/(mol K)

ΔG°=ΔH°-TΔS°=-RT lnK

ΔH°-TΔS°-RT lnK

T=(ΔH°)/(ΔS°-R lnK)

T=(-92.22×(10)^3 J/mol)/(-198.7 J/(mol K)-8.3145 J/(mol K) ln(1.0×(10)^5))=313 K

Q61

Kp = 1e6 for the following reaction:

2SO2 (g) + O2 (g) ↔ 2SO3 (g)

Determine the temperature at which this Kp occurs.

T1= ?, K1= 1e6; T2= 800K, K2= 9.1e2, and ΔH°= -1.8e5 J/mol

S18.61

ln (K2/K1) = ΔH°/R ((1/T1)/(1/t2))

ln(9.1e2/ 1e6)= -1.8e5/ 8.3145((1/T1)/(1/800K))

1/T1= (3.2e-4)+(1.25e-3)= 1.57e-3

T1=1/1.57e-3= 6.37e2 K

Q63A

For the reaction

$$C_6H_{6(g)} \rightleftharpoons (3C)_2H_{2(g)} \tag{17.10.46}$$

,ΔH°=66.3 kJ/mol and K_p=0.214 at 298K.

a. What is Kp at 0°C?
b. At what temperature will Kp =2.50?

S63A

a)

ln K_2/K_1 =(ΔH°)/R (1/T_1 -1/T_2)=(66.3×(10)^3 J/mol)/(8.3145 J/(mol K)) (1/298K-1/273K)=-2.45

K_2/K_1 =e^(-2.45)=0.0863 K_2=0.0863×0.214=0.018 at 273K

b)

ln K_2/K_1 =(ΔH°)/R (1/T_1 -1/T_2)=(66.3×(10)^3 J/mol)/(8.3145 J/(mol K)) (1/T_1 -1/273K)=ln 0.214/2.50=-2.458

(1/T_1 -1/298K)=(-2.458×8.3145)/(66.3×(10)^3) K^(-1)=-3.08×(10)^(-4) K^(-1)

1/T_1 =1/298-3.08×(10)^(-4)=3.05×(10)^(-3)-3.08×(10)^(-4)=2.74×(10)^(-3) K^(-1); T_1=365 K

Q63

At what temperature does this reaction become spontaneous? (ΔH°= 1118.4 kJ, ΔS°= 347.2 J/K)

$$Fe_3O_{4(s)} \rightarrow 3Fe_{(s)} + 2O_{2(g)} \tag{17.10.47}$$

S63

$\Delta G° = \Delta H° - T\,\Delta S° = 0$

$-1118.4\ kJ = -T(0.3472\ kJ/K)$

$T = 3221\ K = 2948°C$

Q71

When referencing a table, what would you predict the normal boiling point of mercury and the vapor pressure of mercury is at 25°C?

After referencing a table with both phases of mercury as gas and liquid, we can conclude that the normal boiling point of liquid mercury (Hg) is 356.73°C = 629.88 K, and vapor pressure of mercury is about 0.002 mmHg at room temperature.

Q74

While mercury is useful in barometers, mercury vapor can be toxic. Given that mercury has Hvap of 56.9 kJ/mol and its normal boiling point is 356.7°C, calculate the vapor pressure in mmHg at room temperature, 25°C.

S18.74

$\Delta Hvap/R * ((1/T1) - (1/T2))\ lnP$

$((56.9e3\ J/mol)/8.314\ J/mol\ K) * ((1/298) - (1/629.7K)) =$

$\ln(12.10) =$

$e^{(12.10)} = 179{,}433\ mmHg$

Q83

Consider the following reaction at 25°C:

$H_2O(g) \leftrightarrow H2O\ (l)$

$\Delta Gf°$ (kJ/ mol): H2O (l)= -237.129, H2O (g) = -228.572

$\Delta Hf°$ (kJ/ mol): H2O (l)= -285.83, H2O (g) = -241.818

S°(J/ mol): H2O (l)=69.91, H2O (g) =188.825

Calculate the Keq for the given reaction.

S18.83

$\Delta H° = [-285.83] - [-241.818] = -44.012\ kJ/mol$

$\Delta S° = [69.91] - [188.825] = -118.915\ J/mol\ k$

$\Delta G° = \Delta H° - T\,\Delta S° = -44012\ J/mol - (298K)(-118.915\ J/mol\ K) = -8575\ J/mol = -8.575\ kJ/mol$

$\Delta G° = -RT\ln Keq$

$\ln K = -\Delta G°/RT = -(-8557J/mol)/(8.3145\ J/mol\ K * 298K) = 38.5$

$K = e^{(3.45)} = 31.6$

Q88

Calculate $\Delta G°$ (in kJ) for the following reaction in which occurs at 25.0° C.

Fe3O4 (s) → 3Fe (s) + 2O2 (g)

ΔHf (kJ/ mol): Fe3O4 (s) = -1118.4

S°(J/ mol): Fe3O4 (s) = 146.4, Fe(s)= 27.8, O2 (g) = 205.1

$\Delta H° = [0] - [-1118.4] = 1118.4\ kJ/mol$

$\Delta S° = [3(27.8) + 2(205.1)] - [146.4] = 347.2\ J/mol\ K$

$\Delta G° = \Delta H° - T\,\Delta S° = 1118.4\ kJ - (298K)(0.3472\ kJ/mol\ K) = 1015\ kJ/mol$

Q104

If ΔG°> 0 for a reaction, it must also be true that: _____

S18.104

Forward reaction is nonspontaneous at standard state and Keq<1.

Q1

Indicate whether each of the following changes represents an increase or decrease in entropy in a system, and explain your reasoning;

a. melting ice
b. burning gasoline
c. freezing liquid bromine.

S1

a. increases entropy, molecular bonds breaking -> greater disorder
b. increases entropy, molecules and compounds released in the burning of anything
c. decreases entropy, frozen molecules are more ordered

Q2

Which of the following processes has the highest ΔS at 25C:

$$CO_{2(s,1atm)} \rightarrow CO_{2(g,10\ mmHg)} \tag{17.10.48}$$

or

$$H_2O_{(s,1atm)} \rightarrow H_2O_{(g,10\ mmHg)} \tag{17.10.49}$$

S2

Converting CO2(s, 1atm)-->CO2(g, 10mmHg) *would have a much higher* ΔS *because it takes much more energy to convert a substance directly from a solid state to a gaseous state.*

Q3

Why is it incorrect to say that the entropy of the world decreases and the total energy of the world fluctuates?

S3

First law of thermodynamics states that energy is neither created nor destroyed, so there is a fixed amount of energy in the universe at a given time. Also, the entropy of the universe increases for all spontaneous, naturally occurring processes, so the entropy of the universe is boundless.

Q4

Calculate the ΔH and ΔS for the following reaction at 25C:

$$4NO_{(g)} \rightarrow 2N_2O_{(g)} + O_{2(g)} \tag{17.10.50}$$

S3

From Table T1, the following properties are identified:

- $NO_{(g)}$: ΔH=90.9 kJ/mol, S=210.76 kJ/mol
- $N_2O_{(g)}$: ΔH=82.05 kJ/mol, S=219.85 kJ/mol
- $O_{2(g)}$: ΔH=0 kJ/mol, S=205.14 kJ/mol

ΔH=[Sum of ΔH (products)]-[Sum of ΔH (reactants)]

ΔS=[Sum of ΔS (products)]-[Sum of ΔS (reactants)]

ΔH=[(2*82.05 kJ/mol)+(0 kJ/mol)]-[(4*90.9 kJ/mol)]=-199.5 kJ/mol

ΔS=[(2*219.85 kJ/mol)+(205.14 kJ/mol)]-[(4*210.76 kJ/mol)]=-198.2 kJ/mol

Q5

Determine ΔG at 265.25 K for the reaction:

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)} \tag{17.10.51}$$

S5

From Table T1, the following properties are identified:

- (ΔH=-114.1 kJ,
- ΔS=-146.5 J/K).

$$\Delta G = \Delta H - T(\Delta S) \tag{17.10.52}$$

$\Delta G = (-114.1\ kJ) - (265.25\ K)(0.1465\ kJ/K)$

$\Delta G = -153.0\ kJ$

Q6

What is the Gibbs energy change ΔG for the following reaction at 25C:

$$N_2 + 3H_2 \rightarrow 2NH_3 \tag{17.10.53}$$

S6

From Table T1, the following properties are identified:

- ΔH=-91.8 kJ and
- ΔS=-198.0 J/K.

$$\Delta G = \Delta H - T(\Delta S) \tag{17.10.54}$$

$$\Delta G = (-91.8\ kJ) - (298\ K)(0.1980\ kJ/K) = -32.8\ kJ \tag{17.10.55}$$

Q7

Which of the following is spontaneous?

a. $\Delta H > 0$ and $T(\Delta S > 0$
b. $\Delta H < 0$ and $T\Delta S > 0$
c. $\Delta H > 0$, and $T(\Delta S < 0$
d. $\Delta H < 0$, and $T\Delta S < 0$

S7

Answer: B

Q8

At what temperatures are the following reaction spontaneous:

$$Br_{2(l)} \rightarrow Br_{2(g)} \tag{17.10.56}$$

S8

- ΔH=30.91 kJ/mol,
- ΔS=93.2 J/molK

For the reaction to be spontaneous:

$$\Delta G = \Delta H - T\Delta S < 0 \tag{17.10.57}$$

(30.91 kJ/mol)-T(0.0932 J/molK <0)

T=331.6 K

Q9

Determine ΔG for the following reaction:

$$C_2H_{4(g)} + H_2O_{(l)} \rightarrow C_2H_5OH_{(l)} \tag{17.10.58}$$

S9

From Table T1, the following properties are identified:

ΔG C2H5OH(l)=-175 kJ/mol

ΔG C2H4(g)=68 kJ/mol

ΔG H2O(l)=-237 kJ/mol

ΔG=(Sum of ΔG products)-(Sum of ΔG reactants)

ΔG=-175 kJ-68 kJ-(-237 kJ)

ΔG=-6 kJ

-6 kJ<0, therefore, spontaneous

Q10

Is the following reaction spontaneous under standard conditions?

$$4KClO_{3(s)} \rightarrow 3KClO_{4(s)} + KCl_{(s)} \tag{17.10.59}$$

S10

From Table T1, the following properties are identified:

- $KClO_3$:
  - $\Delta H_f^o = -397.7\ kJ/mol$
  - $S^o = 143.1\ J/molK$
- $KClO_4$:
  - $\Delta H_f^o = -432.8\ kJ/mol$
  - $S^o = 151.0\ J/molK$
- KCl:
  - $\Delta H_f^o = -436.7\ kJ/mol$,
  - $S^o = 82.6\ J/molK$

Now calculate the enthalpy of the reaction (change of entropy) under standard conditions:

- heat of reaction=3(-432.8 kJ)+(-436.7 kJ)-4(397.7 kJ)=-144 kJ

Now calculate the entropy of the reaction under standard conditions:

- ΔS=3S(KClO4)+S(KCl)-4S(KClO3)
- ΔS=3(151.0 J/K)+(82.6 J/K)-4(143.1\; J/K)=-36.8 J/K

Now combine them together under standard conditions:

$$\Delta G = \Delta H - T\Delta S \tag{17.10.60}$$

- $\Delta G = -144\ kJ - (298\ K)(-38.6\ J/K)\left(\frac{1\ kJ}{1000;\ J}\right) = -133\ kJ$

since $\Delta G < 0$, the reaction is *spontaneous* under standard conditions.

Q11

For what temperatures will this reaction be spontaneous?

$$4KClO_{3(s)} \rightarrow 3KClO_{4(s)} + KCl_{(s)} \tag{17.10.61}$$

S11

The reaction will be spontaneous when $\Delta G < 0$. So all temperatures T such that this inequality is satisfied:

$$\Delta G = \Delta H - T(\Delta S) < 0 \tag{17.10.62}$$

or solving for T

$$\frac{\Delta H}{\Delta S} < T \tag{17.10.63}$$

$$\frac{-144\ kJ}{(-38.6\ J)(1kJ/1000J)} < 3731\ K \tag{17.10.64}$$

Q12

The following reaction occurring at 1 atm is spontaneous at ~331 K:

$$Br_{2(l)} \rightarrow Br_{2(g)} \tag{17.10.65}$$

If the temperature increased, will the reaction?

a. remain spontaneous
b. become non-spontaneous
c. approach equilibrium
d. can not say

S12

Answer: A

Q13

Use data from Appendix D to establish for the following reaction to solve for ΔG:

$$2N_2O_{4(g)} + O_{2(g)} \rightarrow 2N_2O_{5(g)} \tag{17.10.66}$$

S13

ΔG=[(2*115.1 kJ/mol)]-[(2*97.54 kJ/mol)]

ΔG=35.12 kJ/mol

Q14

Using the same reaction and ΔG calculated from the above equation, solve for K_p at 298 K.

S14

$$\Delta G = -RT \ln K_p \tag{17.10.67}$$

$$\ln K_p = \frac{\Delta G}{-RT} \tag{17.10.68}$$

$$\ln K_p = \frac{35120J}{[(-8.314J/mol*K)*298K)} \tag{17.10.69}$$

$$K_p = 6.98 \times 10^{-7} \tag{17.10.70}$$

Q15

The standard Gibbs energy change for the reaction

$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{17.10.71}$$

is 29.05 kJ/mol at 298 K/ use this thermodynamic quantity to decide in which direction the reaction is spontaneous when the concentrations of $NH_{3(aq)}$, $NH^+_{4(aq)}$, and $OH^-_{(aq)}$ are 0.10 M, 1.0*10^-3 M, and 1.0*10^-3 M, respectively.

S15

Qc=[NH4][OH-]/[NH3]=[(1.0*10^-3M)*(1.0*10^-3 M)]/(0.10 M)

Qc=1.0*10^-5

ΔG=ΔG°+RTlnQc

ΔG=(29050 kJ/mol)+(8.314 J/mol*K)(298 K)(ln1.0*10^-5)

ΔG=525.9 kJ/mol

Q16

For a process to occur spontaneously:

a. the entropy of the system must increase
b. the entropy of the surroundings must increase
c. both the entropy of the system and the entropy of the surroundings must increase
d. the net change in entropy of the system and surroundings considered together must be a positive quantity
e. the entropy of the universe must remain constant

S16

Answer: D

Q17

The Gibbs energy change of a reaction can be used to assess

a. how much work the system does on the surroundings
b. the net direction in which the reaction occurs to reach equilibrium
c. how much heat absorbed from surroundings
d. the proportion of the heat evolved in an exothermic reaction that can be converted to various forms of work

S17

Answer: B

Q18.29B

Calculate ΔG for a reaction at 400K if the Keq at that temperature is 200.

ΔG = -400(8.314)ln(200) = -17620 J

Q18.37B

Determine K_p at 298 K for the reaction CO (g) + O_2 (g) àCO_2 (g)

(ΔG_f for CO (g) = -137.2 kJ / mol; ΔG_f (CO_2) = -394.4 kJ / mol)

ΔG° = -RT ln K_p ln K_p = - ΔG° / RT

$\Delta G°_{rxn}$ = -394.4 + 137.2 = -257.2 kJ / mol

ln K_p = 257.2 / (8.3145)(298) K_p = 1.11

Q18.39B

is ΔG positive or negative for this reaction?

$$HCl + NaOH \rightarrow H_2O + NaCl \qquad (17.10.72)$$

S18.39B

ΔH = - 110.27 KJ/mol

ΔG = ΔH (- sign) – T (always positive)(ΔS)

Because we go from aq to solid in this reaction we assume that ΔS is decreasing (increasing in order, so therefore with the given information it is impossible to know)

Q18.47B

Calculate Delta G° for the reaction at 30°C

$$2ZnO \rightarrow 2Zn + O_2 \tag{17.10.73}$$

ZnO Delta H_f°= - 1034 kJ/mol S° = 145 J/molK

Zn Delta H_f° = 0 S°= 68 J/molK

O_2 Delta H_f° = 0 S°= 225 J/molK

S18.47B

Delta H°= (0) –(-1034)= 1034 kJ

Delta S°= [2(68)+1(225)]-(145)=216 J/K

Delta G°= 1034 kJ- (303K)(0.216kJ/K)=968.55 kJ

Q18.49B

Determine K_c or K_p for the following reactions:

a. $Na^+_{(aq)} + Cl^-_{(aq)} \rightleftharpoons NaCl_{(aq)}$ with $\Delta G^\circ = -45.2\ kJ/mol$ @ $267\ K$
b. $CN_{(g)} + H_{2(g)} \rightleftharpoons HCN_{(g)}$ with $\Delta G^\circ = -324\ kJ/mol$ @ $312\ K$
c. $SO_{2(g)} + ½O_{2(g)} \rightarrow SO_{3(g)}$ with $\Delta G^\circ = -215\ kJ/mol$ @ $417\ K$

S18.49B

a) ln K_c = - ΔG° / RT = 45.2 / (8.3145)(267)

K_c = 1.02

b) ln K_p = - ΔG° / RT = 324 / (8.3145)(312)

K_p = 1.13

c) ln K_c = - ΔG° / RT = (215) / (8.3145)(417)

K_c = 1.06

Q18.53B

Calculate the ΔS for the following fake reaction:

$$A + B \rightarrow C + 2D \tag{17.10.74}$$

S°: 160J, 205J, 213.6J, 69.9J

To solve this simply do products minus reactants:

(2(69.9) + 213.6) – (160+205) = ΔS

Δ = -11.6 J

Q55B

Find Delta S^o, Delta H^o, and Delta G^o for the following reaction. Use Appendix D.

$$2NaOH_{(aq)} + H2SO_{4(aq)} \rightarrow Na2SO_{4(aq)} + H_2O_{(l)} \tag{17.10.75}$$

S55B

Delta S^o = 138.1 + 69.9 - (2*48.1 + 20.1) = 91.7 J/mol*K

Delta H^o = -1390 + -285.8 - (2*-470.1 +-909.3) =-1675.8+1849.5 = 173.7 kJ/mol

Delta G^o = -1268 +-237.1 - (2*-419.2+-744.5) = - 1505.1 + 1582.9 = 77.8 kJ/mol

Q18.57B

Using an equation for Gibbs free energy, solve to express T as a function of delta G,H,S

T= (delta H - delta G)/delta S

Q18.59

Find K_{eq} at 25°C for the oxidation of iron (rusting).

$$4Fe + 3O_2 \rightarrow 2Fe_2O_3 \tag{17.10.76}$$

S18.59B

Delta H_f° 0 0 -921kJ/mol

Delta S° 31 210 81 J/mol K

Delta G°= -RT lnK

Delta G°= Delta H°-TDeltaS°

Delta H°= [(-921)2]-[(0)4+(0)3]= -1842 kJ -1.842 x 10^6J

Delta S°=[(81)2]-[(31)4+(210)3]= -592 J/K

Delta G°= Delta H°- T Delta S°

=(-1.842 x 10^6J)-(298K)(-592 J/K)=-1665548. (spontaneous because negative)

K=e^[-DeltaG°/RT]= e^[(-1665548J)/(8.3145)(298)]

Q18.61B

Determine the temperature, in Kelvin, at which K_p = 2.3e-5 for the following reaction, given K_p = 1.0e3 at 345 K, and $\Delta H°_{rxn}$ = 3.2e4 J / mol.

$$SO_{3(g)} + O_{2(g)} \rightarrow SO_{4(g)} \tag{17.10.77}$$

S18.61B

ln (2.3e-5 / 1.0e3) = (3.2e4 / 8.3145)[(1/345) – ($1/T_1$)]

T_1 = 134 K

Q18.63B

For the reaction

$$N_2O_4 \rightarrow 2NO_2 \tag{17.10.78}$$

and $K_p = 0.25$ at 298 K

At what temperature does $K_p = 1.00$?

S18.63B

First solve for Kc.

0.25 = K_c (8.314x298)^(2)

K_c = 4.0x10^-8

Plug in values for K_c and K_p

1 = 4.0x10^-8 (8.314xT)^2

T = .000024 K

Q71B

Using the Clausius-Clapyron equation, calculate the boiling point of benzene for a pressure of 260 mmHg. The normal boiling point of Benzene is 80.1 degC and has a heat of vaporization of 30.72 kJ/mole

S18.71B

The Clausius-Clapeyron Equation is:

$$\text{Ln}(P2/P1) = (\Delta H/R)*(1/T1-1/T2)$$

Substituting:

Ln(260/760) = (30720/8.314)*(1/T1 - 1/(80.1+273.15))

Solving for T1 gives T1 = 310degK -273 = 37degC

Q18.74B

Define a in the equation S = Sinitial - R*ln(a).

a = (the effective concentration of a substance in a system) / (the effective concentration of tha substance in a standard reference state)

Q18.83B

What temperature will the equilibrium constant for formation of $COCl_2$ be $K_p = 1.6 \times 10^4$?

$$CO + Cl_2 \rightarrow COCl_2 \tag{17.10.79}$$

S18.83B

From Table T1 or T2

- $\Delta H^\circ = -95\ kJ\ mol^{-1}$
- $\Delta S^\circ = -256\ Jmol^{-1}K^{-1}$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -RT\ln K_p \tag{17.10.80}$$

$$T = \frac{(-9500\ J\ mol^{-1})}{[-256\ J\ mol^{-1}K^{-1} - (8.3145\ J\ mol^{-1}K^{-1}\ln(1.6\times10^4)} \tag{17.10.81}$$

$$= \frac{-9500}{-336.5} = 28.23K \tag{17.10.82}$$

Q18.88B

Consider the rusting of iron at 298 K:

$$4Fe_{(s)} + 3O_{2(g)} \quad rightleftharpoons 2Fe_2O_{3(s)} \tag{17.10.83}$$

a. calculate ΔG° for this reaction at 298 K
b. determine whether the reaction proceeds spontaneously in the forward or reverse direction.



Heavy rust on the links of a chain. of Wikipedia (Marlith)

S18.88B

a) ΔS°_{rxn} = 2(67.4) – 4(45.3) = -46.4 J / mol • K

ΔG°_{rxn} = -349000 J / mol – (-46.4)(298) = -21.1 kJ / mol

b) reaction is spontaneous in the forward direction.

Q18.104B

if the K_{eq} is zero, does Δ G = 0?

S18.104B

No, Δ G would be zero if K_{eq} =1, the equivalence point.

Q18.3D

What is Boltzmann's equation for entropy? Explain in detail how he derived this equation and how it shows the relationship between entropy and the number of microstates particles can occupy.

S18.3D

S = KlnW.

S is the entropy, K is Boltzmann's constant, and W is the number of microstates. He was able to derive this equation based on the fact that the more microstates a system has, the greater the entropy; so as the W increases, as will S.

http://chemwiki.ucdavis.edu/Physical_Chemistry/Statistical_Mechanics/Boltzmann_Average/The_Boltzmann_constant

Q18.37D

Use data from Appendix D of Petrucci to determine K_p at 298 K for the reaction

$$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g). \quad (17.10.84)$$

S18.37D

$$K_p = K_c(RT)^{\Delta n} \quad (17.10.85)$$

$$\Delta G^\circ = -RT\backslash \ln K_c \quad (17.10.86)$$

-128.6 = (.0083145)(298 K)lnKc (ΔG calculated from difference of values found in Appendix D)

-128.6/2.477 = lnKc

e^-51.90 = Kc

Kc = 2.88 x 10^-23

Kp = (2.88 x 10^-23)(.0821)(298)(-1)

Kp = -7.14 x 10^-20

http://chemwiki.ucdavis.edu/Analytical_Chemistry/Electrochemistry/Electrochemistry_and_Thermodynamics

Q18.49D

For the following equilibrium reactions, calculate ΔG° at the indicated temperature. [Hint: How is each equilibrium constant related to a thermodynamic equilibrium constant, K?]

a. $HCl(g) + O_2(g) \rightleftharpoons H_2O(g) + Cl_2(g)$ with $K_c = 2.3 \times 10^{-12}$ at 25°C
b. $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ with $K_c = 4.5 \times 10^{-7}$ at 25°C
c. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ with $K_c = 1.5 \times 10^{-13}$ at 25°C

S18.49D

Use equation ΔG° = -RTlnK

a.) ΔG° = -(.008314)(298 K)ln(2.3 x 10^-12)

ΔG° = 66.45 kJ/mol

b.) ΔG° = -(.008314)(298 k)ln(4.5 x 10^-7)

ΔG° = 36.207 kJ/mol

c.) ΔG° = (.008314)(298 K)ln(1.5 x 10^-13)

ΔG° = 73.15 kJ/mol

http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Free_Energy/Gibb's_Free_Energy/Gibbs_Free_Energy

Q18.61D

Use the equations ΔG = ΔH –TΔS and ΔG° = -RTlnK, and Appendix D to estimate the temperature at which Kp = 2.6 x 10^3 for the reaction 2SO2(g) + O2(g) ↔ 2SO3(g).

S18.61D

Assume ΔG° = -70.9 (from appendix D, difference of both ΔG° values)

-70.9 = -(.0083145)T(ln(2.6 x 10^3))

T = approximately -10844.4 K

http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Free_Energy/Gibb's_Free_Energy/Gibbs_Free_Energy

Q18.88D

The standard molar entropy of solid hydrazine at its melting point of 1.53°C is 67.15 J/mol*K. The enthalpy of fusion is 12.66 kJ/mol. For N2H4(l) in the interval from 1.53°C to 298.15K, the molar heat capacity at constant pressure is given by the expression Cp = 97.8 + 0.0586(T – 280). Determine the standard molar entropy of $N_2H_{4(l)}$ at 500 K. [Hint: The heat absorbed to produce an infinitesimal change in temperature of a substance is Δq_{rev}= (Cp)(ΔT)].

S18.88D

Cp = 97.8 + .0586(500-280)

Cp = 110.7

Δq_{rev} = (110.7)(23.47 K) = 2598.13

ΔS = q_{rev}/T

ΔS = 2598.13/500 K = **5.196 J/K**

http://chemwiki.ucdavis.edu/Physical_Chemistry/Thermodynamics/State_Functions/Entropy

Q18.1E

Is the entropy increasing or decreasing?

a. Tear open and start using a cold pack.
b. Boiling water.
c. Freezing water.

S18.1E

(a) Increasing (b) Increasing (c) Decreasing

Q18.2E

Put them in order of increasing △S at 25℃.

a. CO_2 (g, 1atm) ↔ CO (s, 1atm)
b. CO_2 (s, 1atm) ↔ CO_2 (l, 1atm)
c. H_2O (s, 1atm) ↔ H_2O (g, 1atm)

S18.2E

Answer: a < b <c.

Q18.27E

Calculate △G° for the reaction

$$COS_{(s)} + H_2O(g) \rightleftharpoons CO_{2\ (g)} + H_2S_{(g)} \tag{17.10.87}$$

given the known thermodynamic properties of the following reactions

a. $CO_{(g)} + H_{2\ (g)} \rightleftharpoons COS_{(g)} + H_{2(g)}$ △G° = 14.7 kJ
b. $CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2\ (g)} + H_{2\ (g)}$ △G° = -28.6 kJ

S18.27E

△G° = -27.2 kJ.

Q18.29E

What's △G° For the reaction

$$2SO_{2\ (g)} + O_{2\ (g)} \rightleftharpoons 2SO_{3\ (g)} \tag{17.10.88}$$

$\triangle G°_f$ of SO_2 is -300kJ/mol, and its -371 kJ/mol of SO_3.

S18.29E

△G° = -142kJ.

Q18.31E

Find △H° if the reaction $Fe_3O_{4(s)} \leftrightarrow 3Fe_{(s)} + 2O_{2(g)}$ happens spontaneously. The temperature is 2948℃,△S° = 347.2 J/K.

S18.31E

△H° = 1118.3 kJ.

Q18.37E

If the Keq for the reaction $CaCO_{3(s)} \leftrightarrow CaO_{(s)} + CO_{2(g)}$ is 8.1×10^{-5}, find △G when T=25℃. Given: $[CaCO_3]$ = 0.003M, [CaO] = 0.0004M, and $[CO_2]$=0.0005M.

S18.37E

△G = -482.5 J.

Q18.39E

What's the △S° of the reaction $2NO_{(g)} + O_{2(g)} \leftrightarrow 2NO_{2(g)}$? If the △S° is 210.8 J/K , 205.1 J/K, and 240.1 J/K respectively.

S18.39E

△S° = -146.6 J/K.

Q18.43E

For the reaction, $N_2O_{(g)}$ + ½ $O_{2(g)} \leftrightarrow 2NO_{2(g)}$, the partial pressure are 0.1atm, 0.5atm, and 0.8atm respectively. Is the reaction spontaneous?

S18.43E

It's nonspontaneous.

Q18.45E

For reaction $3Fe_2O_{3(s)} + H_{2(s)} \leftrightarrow 2Fe_3O_{4(s)} + H_2O_{(g)}$, which of the following is correct and why?

a. △G = △G° + RT lnQ.
b. K=-kp.
c. $Kp = e^{\triangle G°/RT}$

S18.45E

A

Q18.47E

The gas concentrations of the reaction

$$2A_{(g)} + 3B_{(g)} \rightleftharpoons 2C_{(g)} + 4D_{(g)} \tag{17.10.89}$$

are [A] = 1.8x10^{-2}M, [B] = 3.4x10^{-3}M, [C] = 5.6x10^{-1}M, [D]=2.8x10^{-4}M. Assume T=25 ℃, and $K_{eq} = 4.42 \times 10^{-9}$. Find Q, △G°, and △G.

S18.47E

1.51x10^{-4}J, 4.77x10^{4}J, 25.9kJ

Q18.49E

At 298K, $Mg(OH)_{2\,(s)} + 2H^{+}_{(aq)} \leftrightarrow Mg^{2+}_{(aq)} + 2H_2O_{(l)}$ has △G° = -95.5kJ/mol. If the Q is 2.389x10^{-4}, what's △G?

S18.49E

△G = 1.97x10^{6} kJ.53.

Q18.53E

Find the K, K_c, and K_{eq} of the reaction

$$Si(s) + 2Cl_2(g) \rightleftharpoons SiCl4(g) \tag{17.10.90}$$

S18.53E

K= [SiCl]/[Cl_2^-]^2=Kc=Ksp.

Q18.55E

If T has a positive value, △H has a negative value, △S has a negative value, when will △G be positive or negative?

S18.55E

△G will be positive/negative when △H is smaller/greater than T△S

Q18.57E

If T has a positive value, △H has a negative value, △S has a positive value, when will △G be positive or negative?

S18.57E

△G will be positive/negative when △H is greater/smaller than T△S.

Q18.59E

What's the temperature of the reaction

$$H_2(g) + I(g) \rightleftharpoons 2HI\ (g) \tag{17.10.91}$$

if $\Delta H^\circ = -30kJ$, $\Delta S^\circ = 30J/K$, and $K_{eq} = 2.5 \times 10^{-3}$?

S18.59E

Video Solution

T= 855.8K

Q18.61E

Given K_2=2, K_1=3. Temperature increases from 100℃ to 120℃. What's △H°?

S18.61E

△H° = 358.8J.

Q18.63E

For the reaction

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \tag{17.10.92}$$

calculate △G at 25℃ if it's known that △Gf° (SO_3) = -150kJ, △Gf° (SO_2) = -125kJ, Q= 2.0×10^{-3}

S18.63E

△G= 159.6 kJ.

Q18.71E

For problem 61, if the value of Q is not given, instead, it says that SO_2 gas is at 2atm, O_2 is at 0.5atm, and SO_3 gas at 4 atm. Calculate the new △G.

S18.71E

△G = 176.7 kJ.

Q18.74E

For the reaction, $Br_{2(l)} \leftrightarrow Br_{2(g)}$, what's the normal boiling point of it says that △H° = 31.0kJ/mol, △S° = 93.07 J/Kmol?

S18.74E

△H° = 333K.

Q18.83E

Calculate △G° for the oxidation of iron:

$$4Fe(s) + 3O_2(g) \rightleftharpoons 2Fe_2O_3(s) \tag{17.10.93}$$

given:

- △Hf° of $2Fe_2O_3(s)$ is -826 kJ/mol,
- S_m° of $Fe_{(s)}$ is 27 J/mol K and 205 J/mol K for O_2 and 90 J/mol K for $Fe_2O_{3(s)}$.

S18.83E

△G°= 10^{261}J.

Q18.88E

What would happen if increasing the temperature of the reaction

$$H_{2\ (g)} + I_{2\ (g)} \rightleftharpoons 2HI_{(s)} \tag{17.10.94}$$

? Given that $K_p = 620$ at 298 K and $\triangle\ H^\circ = -9.5\ kJ/mol$.

S18.88E

products are favored.

Q18.104E

For the reaction

$$2SO_{2\ (g)} + O_{2\ (g)} \rightleftharpoons 2SO_{3\ (g)} \tag{17.10.95}$$

if it's known that $K_p = 500$ at 25 ℃ and $\triangle\ H^\circ = -7.23\ kJ/mol$, what would happen if we decrease the temperature?

S18.104E

the reactants are favored.

Misc

Q1

Using the following values for entropy determine if a reaction would be spontaneous.

I) ΔSsys = 30 J/K , ΔSsurr = 50 J/K

II) ΔSsys = 60 J/K, ΔSsurr = -85 J/K

A reaction will be spontaneous if the total entropy change is positive. It will not be spontaneous if the total entropy change is negative.

ΔStotal = ΔSsys + ΔSsurr

where

ΔStotal is the total entropy change

ΔSsys is the entropy change of the system

ΔSsurr is the entropy change of the surroundings

I.) ΔStotal = ΔSsys + ΔSsurr

ΔStotal = 30 J/K + 50 J/K

ΔStotal = 80 J/K

The total entropy change is positive, therefore the reaction will be spontaneous.

II.) ΔStotal = ΔSsys + ΔSsurr

ΔStotal = 60 J/K + -85 J/K

ΔStotal = -25 J/K

The total entropy change is negative, therefore the reaction will not be spontaneous.

Q2

Under what conditions are the following reactions spontaneous, and why?

a. 3N2(g) ⇄ 2N3(g)
b. H2O(l) ⇄ H2O(g) where ΔH is negative
c. 2NH4NO2(s) ⇄ 2N2(g) + 4H2O(g) + O2(g) where ΔH is negative

S2

a. Non spontaneous at all temperatures because entropy is decreased
b. Spontaneous when temperature is high, above 100˚C at STP
c. This reaction is spontaneous at all temperatures because of negative ΔH and increasing entropy of the system

Q25

Calculate the change in Gibbs free energy.

HCl+NaOCH3 → NaCl+HOCH3

pKa HCl=-7pKa HOCH3=16 T=298K

Keq=[NaCl][HOCH3][H+]/[HCl][NaOCH3][H+]

Keq=KaHCl/KaHOCH3

Keq=10^7/10^-16= 10^23

ΔG=-RTlnK

ΔG=-8.314x298xln10^23

ΔG=-13.1kJ ← favored towards products

Q35

In the synthesis of gaseous methanol from carbon monoxide gas and hydrogen gas, the following equilibrium concentrations were determined at 444 K: [CO(g)]=0.0977M, [H2O(g)]=0.0799M, and [CH3OH(g)]=0.00799M. Calculate the equilibrium constant and Gibbs energy for this reaction.

S35

CO(g)+2H2(g) → CH3OH(g)

Keq=[CH3OH]/([CO][H2]^2)

Keq=[0.00799]/([0.0977][0.0799]^2)

Keq=12.8

ΔG°=-RTln(Keq)

ΔG°=-(8.314J/Kmol)(444K)ln(12.8)(kJ/1000J)

ΔG°=-9.4kJ/mol

Q37

For the following reaction, what would ΔG° be at 298K ?

Fe3O4(s) → 3 Fe (s) + 2O2(g)

ΔHf° (kj/mol)-1118.4

ΔS° (J/molK) 146.427.8205.1

ΔH° 0 - (-1118.4) = +1118.4 kJ

ΔS° (3(27.8) + 2(205.1)) - 146.4 = 347.2 J/K

ΔG° = ΔH°-TΔS° = 1118.4kJ-(298.15K)(.3472kJ/K)= 1015kJ

Q41

If delta H = 158 kJ and delta S = 411 J/k. At what temperature will this reaction be spontaneous?

S41

deltaG= deltaH-TdeltaS

0> 158000J – T(411 J/K)

T(411 J/K)/ 411 J/K > 158000 J/ 411 J/K

T > 384 K

Q55

Calculate ΔG of this reaction

ΔH=-537.22kJΔS=13.74J/KT=25°C

H2(g)+F2(g) → 2HF

ΔG=ΔH-TΔS

ΔG=(-537.22kJ)-(298K)(13.74J/K)

ΔG=(-537.22kJ)-(298K)(.01374kJ/K)

ΔG=-541.31kJ ← SPONTANEOUS

Q57

What must be the temperature if the following reaction has ΔG°=-52.9 kJ, ΔH°=-34.7kJ, and ΔS°=12.4J/K?

$Fe_2O_3(s)+3CO(g) \rightarrow 2Fe(s)+3CO_2(g)$

$\Delta G^\circ=\Delta H^\circ-T\Delta S^\circ$

(-52.9 kJ)(1000J/kJ)=(-34.7 kJ)(1000J/kJ)-T(12.4J/K)

T=1470K

Q67

Label which reaction is spontaneous or nonspontaneous, and compute the overall reaction, given that it is spontaneous.

$Cu_2O\ (s) \rightarrow 2\ Cu(s) + ½\ O_2\ (g)\ \Delta G^\circ = 125$ kJ

$C(s) + ½\ O_2(g) \rightarrow CO(g)\ \Delta G^\circ = -175$

Reaction 1: nonspontaneous; reaction 2: spontaneous

Net reaction:

$Cu_2O\ (s) + C\ (s) \rightarrow 2\ Cu\ (s) + CO(g)$

ΔG°= 125+ (-175) = -50 kJ

Net reaction is spontaneous.

This page titled 17.10: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

CHAPTER OVERVIEW

18: Electrochemistry

Electrochemistry deals with chemical reactions that produce electricity and the changes associated with the passage of electrical current through matter. The reactions involve electron transfer, and so they are oxidation-reduction (or redox) reactions. Many metals may be purified or electroplated using electrochemical methods. Devices such as automobiles, smartphones, electronic tablets, watches, pacemakers, and many others use batteries for power. Batteries use chemical reactions that produce electricity spontaneously and that can be converted into useful work. All electrochemical systems involve the transfer of electrons in a reacting system. In many systems, the reactions occur in a region known as the cell, where the transfer of electrons occurs at electrodes.

This chapter introduces the fundamental principles of electrochemistry, a field that explores the relationship between electrical energy and chemical reactions. Understanding these principles is crucial for a variety of applications, including the development of advanced technologies in energy storage, corrosion prevention, and medical diagnostics.

Faculty Feature

Dr. Kagan Kerman's groundbreaking research focuses on developing advanced biosensors to accelerate the discovery of multi-target Alzheimer's therapies. These sensors modulate amyloid aggregation and target biomolecules like acetylcholinesterase, using miniaturized, cost-effective, and rapid techniques. Dr. Kerman's projects span protein modifications, neurodegenerative drug screening, and bioconjugate synthesis with nanomaterials, showcasing the practical applications of electrochemical principles. His work highlights the critical role of electrochemistry in creating innovative tools for diagnosing and treating diseases, especially Alzheimer's, by leveraging nanoscale material properties and reactions.

Read more about Dr. Kagan Kerman and his research here: https://bioelectrochemistry.utoronto.ca/



18.1: Introduction
18.2: Review of Redox Chemistry
18.3: Galvanic Cells
18.4: Electrode and Cell Potentials
18.5: Potential, Free Energy, and Equilibrium
18.6: Batteries and Fuel Cells
18.7: Electrolysis
18.8: Key Terms
18.9: Key Equations
18.10: Summary
18.11: Exercises

This page titled 18: Electrochemistry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar via source content that was edited to the style and standards of the LibreTexts platform.

18.1: Introduction

Another chapter in this text introduced the chemistry of reduction-oxidation (redox) reactions. This important reaction class is defined by changes in oxidation states for one or more reactant elements, and it includes a subset of reactions involving the transfer of electrons between reactant species. Around the turn of the nineteenth century, chemists began exploring ways these electrons could be transferred *indirectly* via an external circuit rather than directly via intimate contact of redox reactants. In the two centuries since, the field of *electrochemistry* has evolved to yield significant insights on the fundamental aspects of redox chemistry as well as a wealth of technologies ranging from industrial-scale metallurgical processes to robust, rechargeable batteries for electric vehicles (Figure 18.1.1). In this chapter, the essential concepts of electrochemistry will be addressed.



Figure 18.1.1: Electric vehicles are powered by batteries, devices that harness the energy of spontaneous redox reactions. (credit: modification of work by Robert Couse-Baker)

This page titled 18.1: Introduction is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.0: Introduction** is licensed CC BY 4.0.

18.2: Review of Redox Chemistry

Learning Objectives

By the end of this section, you will be able to:

- Describe defining traits of redox chemistry
- Identify the oxidant and reductant of a redox reaction
- Balance chemical equations for redox reactions using the half-reaction method

Since reactions involving electron transfer are essential to the topic of electrochemistry, a brief review of redox chemistry is provided here that summarizes and extends the content of an earlier text chapter (see chapter on reaction stoichiometry). Readers wishing additional review are referred to the text chapter on reaction stoichiometry.

Oxidation Numbers

By definition, a redox reaction is one that entails changes in *oxidation number* (or *oxidation state*) for one or more of the elements involved. The oxidation number of an element in a compound is essentially an assessment of how the electronic environment of its atoms is different in comparison to atoms of the pure element. By this description, *the oxidation number of an atom in an element is equal to zero*. For an atom in a compound, *the oxidation number is equal to the charge the atom would have in the compound if the compound were ionic*. Consequential to these rules, *the sum of oxidation numbers for all atoms in a molecule is equal to the charge on the molecule*. To illustrate this formalism, examples from the two compound classes, ionic and covalent, will be considered.

Simple ionic compounds present the simplest examples to illustrate this formalism, since by definition the elements' oxidation numbers are numerically equivalent to ionic charges. Sodium chloride, NaCl, is comprised of Na^+ cations and Cl^- anions, and so oxidation numbers for sodium and chlorine are, +1 and −1, respectively. Calcium fluoride, CaF_2, is comprised of Ca^{2+} cations and F^- anions, and so oxidation numbers for calcium and fluorine are, +2 and −1, respectively.

Covalent compounds require a more challenging use of the formalism. Water is a covalent compound whose molecules consist of two H atoms bonded separately to a central O atom via polar covalent O−H bonds. The shared electrons comprising an O−H bond are more strongly attracted to the more electronegative O atom, and so it acquires a partial negative charge in the water molecule (relative to an O atom in elemental oxygen). Consequently, H atoms in a water molecule exhibit partial positive charges compared to H atoms in elemental hydrogen. The sum of the partial negative and partial positive charges for each water molecule is zero, and the water molecule is neutral.

Imagine that the polarization of shared electrons within the O−H bonds of water were 100% complete—the result would be *transfer* of electrons from H to O, and water would be an ionic compound comprised of O^{2-} anions and H^+ cations. And so, the oxidations numbers for oxygen and hydrogen in water are −2 and +1, respectively. Applying this same logic to carbon tetrachloride, CCl_4, yields oxidation numbers of +4 for carbon and −1 for chlorine. In the nitrate ion, NO_3^-, the oxidation number for nitrogen is +5 and that for oxygen is −2, summing to equal the 1− charge on the molecule:

$$(1N\text{ atom})\left(\frac{+5}{N\text{ atom}}\right)+(3O\text{ atoms})\left(\frac{-2}{O\text{ atom}}\right)=+5+-6=-1$$

Balancing Redox Equations

The unbalanced equation below describes the decomposition of molten sodium chloride:

$$\mathrm{NaCl(l)}\longrightarrow\mathrm{Na(l)}+\mathrm{Cl_2(g)}\qquad\text{(unbalanced)}$$

This reaction satisfies the criterion for redox classification, since the oxidation number for Na is decreased from +1 to 0 (it undergoes *reduction*) and that for Cl is increased from −1 to 0 (it undergoes *oxidation*). The equation in this case is easily balanced by inspection, requiring stoichiometric coefficients of 2 for the NaCl and Na:

$$2\,\mathrm{NaCl(l)}\longrightarrow2\,\mathrm{Na(l)}+\mathrm{Cl_2(g)}\qquad\text{(balanced)}$$

Redox reactions that take place in aqueous solutions are commonly encountered in electrochemistry, and many involve water or its characteristic ions, $H^+(aq)$ and $OH^-(aq)$, as reactants or products. In these cases, equations representing the redox reaction can be

very challenging to balance by inspection, and the use of a systematic approach called the *half-reaction method* is helpful.

Balancing REDOX reactions via the half-reaction Method

This approach involves the following steps:

1. Write skeletal equations for the oxidation and reduction half-reactions.
2. Balance each half-reaction for all elements except H and O.
3. Balance each half-reaction for O by adding H_2O.
4. Balance each half-reaction for H by adding H^+.
5. Balance each half-reaction for charge by adding electrons.
6. If necessary, multiply one or both half-reactions so that the number of electrons consumed in one is equal to the number produced in the other.
7. Add the two half-reactions and simplify.
8. If the reaction takes place in a basic medium, add OH^- ions the equation obtained in step 7 to neutralize the H^+ ions (add in equal numbers to both sides of the equation) and simplify.

The examples below demonstrate the application of this method to balancing equations for aqueous redox reactions.

Example 18.2.1: Balancing Equations for Redox Reactions in Acidic Solutions

Write the balanced equation representing reaction between solid copper and nitric acid to yield aqueous copper(II) ions and nitrogen monoxide gas.

Solution

Following the steps of the half-reaction method:

Step 1: Write skeletal equations for the oxidation and reduction half-reactions.

$$\text{oxidation:} \quad \mathrm{Cu(s) \longrightarrow Cu^{2+}(aq)}$$

$$\text{reduction:} \quad \mathrm{HNO_3(aq) \longrightarrow NO(g)}$$

Step 2: Balance each half-reaction for all elements except H and O.

$$\text{oxidation:} \quad \mathrm{Cu(s) \longrightarrow Cu^{2+}(aq)}$$

$$\text{reduction:} \quad \mathrm{HNO_3(aq) \longrightarrow NO(g)}$$

Step 3: Balance each half-reaction for O by adding H_2O.

$$\text{oxidation:} \quad \mathrm{Cu(s) \longrightarrow Cu^{2+}(aq)}$$

$$\text{reduction:} \quad \mathrm{HNO_3(aq) \longrightarrow NO(g) + \mathbf{2\,H_2O(l)}}$$

Step 4: Balance each half-reaction for H by adding H^+.

$$\text{oxidation:} \quad \mathrm{Cu(s) \longrightarrow Cu^{2+}(aq)}$$

$$\text{reduction:} \quad \mathrm{\mathbf{3\,H^+(aq)} + HNO_3(aq) \longrightarrow NO(g) + 2\,H_2O(l)}$$

Step 5: Balance each half-reaction for charge by adding electrons.

$$\text{oxidation:} \quad \mathrm{Cu(s) \longrightarrow Cu^{2+}(aq) + \mathbf{2\,e^-}}$$

$$\text{reduction:} \quad \mathrm{\mathbf{3\,e^-} + 3\,H^+(aq) + HNO_3(aq) \longrightarrow NO(g) + 2\,H_2O(l)}$$

Step 6: If necessary, multiply one or both half-reactions so that the number of electrons consumed in one is equal to the number produced in the other.

$$\text{oxidation } (\times 3): \quad \mathrm{\mathbf{3}\,Cu(s) \longrightarrow \mathbf{3}\,Cu^{2+}(aq) + \mathbf{6}\ \cancel{2}e^-}$$

$$\text{reduction } (\times 2): \quad \mathrm{\mathbf{6}\ \cancel{3}e^- + \mathbf{6}\ \cancel{3}H^+(aq) + \mathbf{2}\,HNO_3(aq) \longrightarrow \mathbf{2}\,NO(g) + \mathbf{4}\ \cancel{2}H_2O(l)}$$

https://chem.libretexts.org/@go/page/475644

Step 7: Add the two half-reactions and simplify.

$$3\,Cu(s) + \cancel{6\,e^-} + 6\,H^+(aq) + 2\,HNO_3(aq) \longrightarrow 3\,Cu^{2+}(aq) + \cancel{6\,e^-} + 2\,NO(g) + 4\,H_2O(l)$$

$$3\,Cu(s) + 6\,H^+(aq) + 2\,HNO_3(aq) \longrightarrow 3\,Cu^{2+}(aq) + 2\,NO(g) + 4\,H_2O(l)$$

Step 8: If the reaction takes place in a basic medium, add OH^- ions the equation obtained in step 7 to neutralize the H^+ ions (add in equal numbers to both sides of the equation) and simplify.

This step not necessary since the solution is stipulated to be acidic.

The balanced equation for the reaction in an acidic solution is then

$$3\,Cu(s) + 6\,H^+(aq) + 2\,HNO_3(aq) \longrightarrow 3\,Cu^{2+}(aq) + 2\,NO(g) + 4\,H_2O(l)$$

Exercise 18.2.1

The reaction above results when using relatively diluted nitric acid. If concentrated nitric acid is used, nitrogen dioxide is produced instead of nitrogen monoxide. Write a balanced equation for this reaction.

Answer

$$Cu(s) + 2\,H^+(aq) + 2\,HNO_3(aq) \longrightarrow Cu^{2+}(aq) + 2\,NO_2(g) + 2\,H_2O(l)$$

Example 18.2.2: Balancing Equations for Redox Reactions in Basic Solutions

Write the balanced equation representing reaction between aqueous permanganate ion, , and solid chromium(III) hydroxide, $Cr(OH)_3$, to yield solid manganese(IV) oxide, MnO_2, and aqueous chromate ion, The reaction takes place in a basic solution.

Solution

Step 1: Write skeletal equations for the oxidation and reduction half-reactions.

$$\text{oxidation:} \quad Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq)$$

$$\text{reduction:} \quad MnO_4^-(aq) \longrightarrow MnO_2(s)$$

Step 2: Balance each half-reaction for all elements except H and O.

$$\text{oxidation:} \quad Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq)$$

$$\text{reduction:} \quad MnO_4^-(aq) \longrightarrow MnO_2(s)$$

Step 3: Balance each half-reaction for O by adding H_2O.

$$\text{oxidation:} \quad \mathbf{H_2O(l)} + Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq)$$

$$\text{reduction:} \quad MnO_4^-(aq) \longrightarrow MnO_2(s) + \mathbf{2\,H_2O(l)}$$

Step 4: Balance each half-reaction for H by adding H^+.

$$\text{oxidation:} \quad H_2O(l) + Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq) + \mathbf{5\,H^+(aq)}$$

$$\text{reduction:} \quad \mathbf{4\,H^+(aq)} + MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\,H_2O(l)$$

Step 5: Balance each half-reaction for charge by adding electrons.

$$\text{oxidation:} \quad H_2O(l) + Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq) + 5\,H^+(aq) + \mathbf{3\,e^-}$$

$$\text{reduction:} \quad \mathbf{3\,e^-} + 4\,H^+(aq) + MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\,H_2O(l)$$

Step 6: If necessary, multiply one or both half-reactions so that the number of electrons consumed in one is equal to the number produced in the other.

This step is not necessary since the number of electrons is already in balance.

Step 7: Add the two half-reactions and simplify.

$$\cancel{H_2O(l)} + Cr(OH)_3(s) \longrightarrow CrO_4^{2-}(aq) + 5\cancel{)}^{1}\, H^+(aq) + \cancel{3e^-}$$

$$+ \quad \cancel{3e^-} + 4\,H^+\cancel{(aq)} + MnO_4^-(aq) \longrightarrow MnO_2(s) + 2\cancel{)}^{1}\, H_2O(l)$$

$$Cr(OH)_3(s) + MnO_4^-(aq) \longrightarrow CrO_4^{2-}(aq) + MnO_2(s) + H^+(aq) + H_2O(l)$$

Step 8: If the reaction takes place in a basic medium, add OH^- ions the equation obtained in step 7 to neutralize the H^+ ions (add in equal numbers to both sides of the equation) and simplify.

$$OH^-(aq) + Cr(OH)_3(s) + MnO_4^-(aq) \longrightarrow CrO_4^{2-}(aq) + MnO_2(s) + \cancel{H^+(aq)} + H_2O(l) + \cancel{OH^-(aq)}$$

after neutralizing the H^+, we get the balanced equation for the reaction in a basic solution

$$OH^-(aq) + Cr(OH)_3(s) + MnO_4^-(aq) \longrightarrow CrO_4^{2-}(aq) + MnO_2(s) + 2\,H_2O(l)$$

? Exercise 18.2.2

Aqueous permanganate ion may also be reduced using aqueous bromide ion, Br^-, the products of this reaction being solid manganese(IV) oxide and aqueous bromate ion, BrO_3^-. Write the balanced equation for this reaction occurring in a basic medium.

Answer

$$H_2O(l) + 2\,MnO_4^-(aq) + Br^-(aq) \longrightarrow 2\,MnO_2(s) + BrO_3^-(aq) + 2\,OH^-(aq)$$

This page titled 18.2: Review of Redox Chemistry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang via source content that was edited to the style and standards of the LibreTexts platform.

- **17.1: Review of Redox Chemistry** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.3: Galvanic Cells

Learning Objectives

By the end of this section, you will be able to:

- Describe the function of a galvanic cell and its components
- Use cell notation to symbolize the composition and construction of galvanic cells

As demonstration of spontaneous chemical change, Figure 18.3.1: shows the result of immersing a coiled wire of copper into an aqueous solution of silver nitrate. A gradual but visually impressive change spontaneously occurs as the initially colorless solution becomes increasingly blue, and the initially smooth copper wire becomes covered with a porous gray solid.



Figure 18.3.1: A copper wire and an aqueous solution of silver nitrate (left) are brought into contact (center) and a spontaneous transfer of electrons occurs, creating blue $Cu^{2+}(aq)$ and gray $Ag(s)$ (right).

These observations are consistent with (i) the oxidation of elemental copper to yield copper(II) ions, $Cu^{2+}(aq)$, which impart a blue color to the solution, and (ii) the reduction of silver(I) ions to yield elemental silver, which deposits as a fluffy solid on the copper wire surface. And so, *the direct transfer of electrons from the copper wire to the aqueous silver ions is spontaneous* under the employed conditions. A summary of this redox system is provided by these equations:

Consider the construction of a device that contains all the reactants and products of a redox system like the one here, but prevents physical contact between the reactants. Direct transfer of electrons is, therefore, prevented; transfer, instead, takes place indirectly through an external circuit that contacts the separated reactants. Devices of this sort are generally referred to as *electrochemical cells*, and those in which a spontaneous redox reaction takes place are called galvanic cells (or voltaic cells).

A galvanic cell based on the spontaneous reaction between copper and silver(I) is depicted in Figure 18.3.2 The cell is comprised of two half-cells, each containing the redox conjugate pair (“couple”) of a single reactant. The half-cell shown at the left contains the Cu(0)/Cu(II) couple in the form of a solid copper foil and an aqueous solution of copper nitrate. The right half-cell contains the Ag(I)/Ag(0) couple as solid silver foil and an aqueous silver nitrate solution. An external circuit is connected to each half-cell at its solid foil, meaning the Cu and Ag foil each function as an *electrode*. By definition, the anode of an electrochemical cell is the electrode at which oxidation occurs (in this case, the Cu foil) and the cathode is the electrode where reduction occurs (the Ag foil). The redox reactions in a galvanic cell occur only at the interface between each half-cell’s reaction mixture and its electrode. To keep the reactants separate while maintaining charge-balance, the two half-cell solutions are connected by a tube filled with inert electrolyte solution called a salt bridge. The spontaneous reaction in this cell produces Cu^{2+} cations in the anode half-cell and consumes Ag^{+} ions in the cathode half-cell, resulting in a compensatory flow of inert ions from the salt bridge that maintains charge balance. Increasing concentrations of Cu^{2+} in the anode half-cell are balanced by an influx of NO_3^{-} from the salt bridge, while a flow of Na^{+} into the cathode half-cell compensates for the decreasing Ag^{+} concentration.

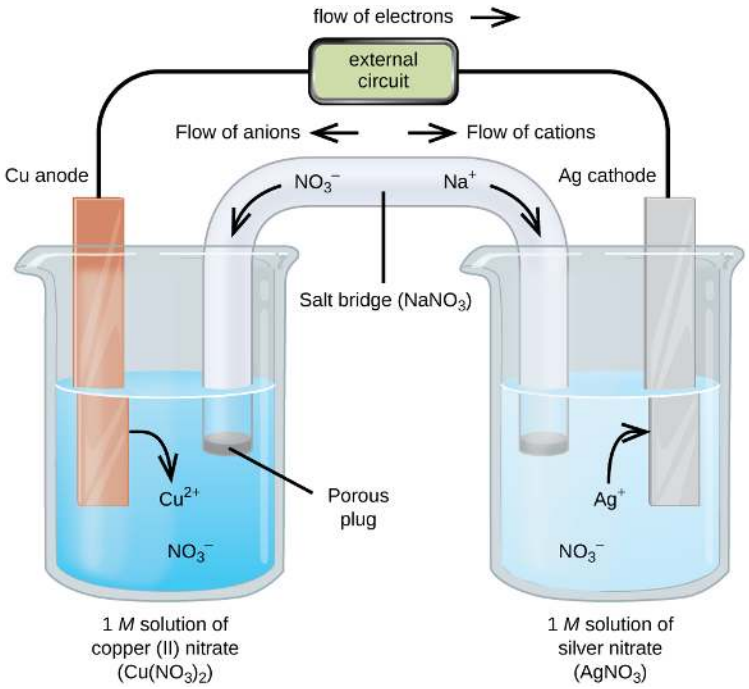


Figure 18.3.2: A galvanic cell based on the spontaneous reaction between copper and silver(I) ions.

Cell Notation

Abbreviated symbolism is commonly used to represent a galvanic cell by providing essential information on its composition and structure. These symbolic representations are called cell notations or cell schematics, and they are written following a few guidelines:

- The relevant components of each half-cell are represented by their chemical formulas or element symbols
- All interfaces between component phases are represented by vertical parallel lines; if two or more components are present in the same phase, their formulas are separated by commas
- By convention, the schematic begins with the anode and proceeds left-to-right identifying phases and interfaces encountered within the cell, ending with the cathode

A verbal description of the cell as viewed from anode-to-cathode is often a useful first-step in writing its schematic. For example, the galvanic cell shown in Figure 18.3.2 consists of a solid copper anode immersed in an aqueous solution of copper(II) nitrate that is connected via a salt bridge to an aqueous silver(I) nitrate solution, immersed in which is a solid silver cathode. Converting this statement to symbolism following the above guidelines results in the cell schematic:

Consider a different galvanic cell (see Figure 18.3.3) based on the spontaneous reaction between solid magnesium and aqueous iron(III) ions:

In this cell, a solid magnesium anode is immersed in an aqueous solution of magnesium chloride that is connected via a salt bridge to an aqueous solution containing a mixture of iron(III) chloride and iron(II) chloride, immersed in which is a platinum cathode. The cell schematic is then written as

Notice the cathode half-cell is different from the others considered thus far in that its electrode is comprised of a substance (Pt) that is neither a reactant nor a product of the cell reaction. This is required when neither member of the half-cell's redox couple can reasonably function as an electrode, which must be electrically conductive and in a phase separate from the half-cell solution. In this case, both members of the redox couple are solute species, and so Pt is used as an inert electrode that can simply provide or accept electrons to redox species in solution. Electrodes constructed from a member of the redox couple, such as the Mg anode in this cell, are called active electrodes.

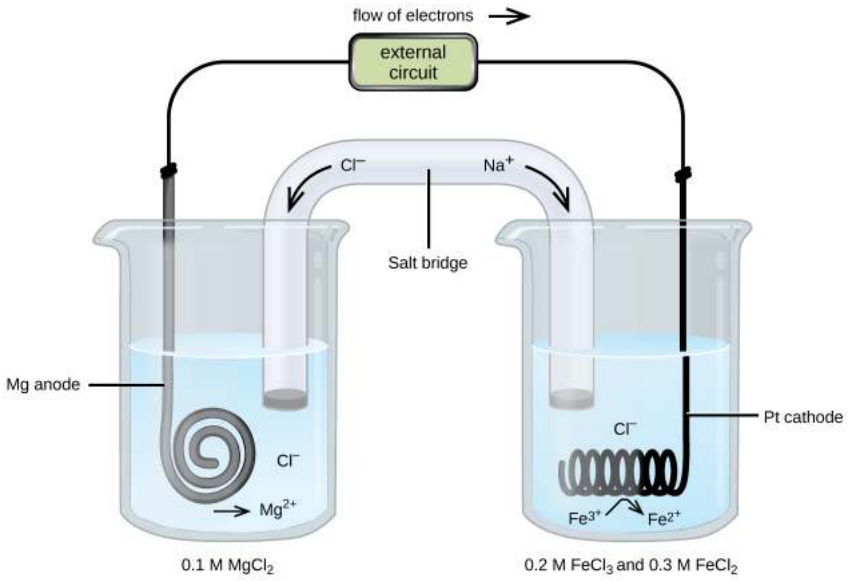


Figure 18.3.3: A galvanic cell based on the spontaneous reaction between magnesium and iron(III) ions.

Example 18.3.1: Writing Galvanic Cell Schematics

A galvanic cell is fabricated by connecting two half-cells with a salt bridge, one in which a chromium wire is immersed in a 1 M $CrCl_3$ solution and another in which a copper wire is immersed in 1 M $CuCl_2$. Assuming the chromium wire functions as an anode, write the schematic for this cell along with equations for the anode half-reaction, the cathode half-reaction, and the overall cell reaction.

Solution

Since the chromium wire is stipulated to be the anode, the schematic begins with it and proceeds left-to-right, symbolizing the other cell components until ending with the copper wire cathode:

The half-reactions for this cell are

Multiplying to make the number of electrons lost by Cr and gained by Cu^{2+} equal yields

Adding the half-reaction equations and simplifying yields an equation for the cell reaction:

Exercise 18.3.1

Omitting solute concentrations and spectator ion identities, write the schematic for a galvanic cell whose net cell reaction is shown below.

Answer

This page titled 18.3: Galvanic Cells is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang via source content that was edited to the style and standards of the LibreTexts platform.

- **17.2: Galvanic Cells** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.4: Electrode and Cell Potentials

Learning Objectives

By the end of this section, you will be able to:

- Describe and relate the definitions of electrode and cell potentials
- Interpret electrode potentials in terms of relative oxidant and reductant strengths
- Calculate cell potentials and predict redox spontaneity using standard electrode potentials

Unlike the spontaneous oxidation of copper by aqueous silver(I) ions described previously, immersing a copper wire in an aqueous solution of lead(II) ions yields no reaction. The two species, $Ag^{+}(aq)$ and $Pb^{2+}(aq)$, thus show a distinct difference in their redox activity towards copper: the silver ion spontaneously oxidized copper, but the lead ion did not. Electrochemical cells permit this relative redox activity to be quantified by an easily measured property, *potential*. This property is more commonly called *voltage* when referenced in regard to electrical applications, and it is a measure of energy accompanying the transfer of charge. Potentials are measured in the volt unit, defined as one joule of energy per one coulomb of charge, V = J/C.

When measured for purposes of electrochemistry, a potential reflects the driving force for a specific type of charge transfer process, namely, the transfer of electrons between redox reactants. Considering the nature of potential in this context, it is clear that the potential of a single half-cell or a single electrode can't be measured; "transfer" of electrons requires both a donor and recipient, in this case a reductant and an oxidant, respectively. Instead, a half-cell potential may only be assessed relative to that of another half-cell. It is only the *difference in potential* between two half-cells that may be measured, and these measured potentials are called cell potentials, E_{cell}, defined as

$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

where $E_{cathode}$ and E_{anode} are the potentials of two different half-cells functioning as specified in the subscripts. As for other thermodynamic quantities, the standard cell potential, $E°_{cell}$, is a cell potential measured when both half-cells are under standard-state conditions (1 *M* concentrations, 1 bar pressures, 298 K):

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

To simplify the collection and sharing of potential data for half-reactions, the scientific community has designated one particular half-cell to serve as a universal reference for cell potential measurements, assigning it a potential of exactly 0 V. This half-cell is the **standard hydrogen electrode** (SHE) and it is based on half-reaction below:

$$2\,\mathrm{H^+(aq)} + 2\,\mathrm{e^-} \longrightarrow \mathrm{H_2(g)}$$

A typical SHE contains an inert platinum electrode immersed in precisely 1 *M* aqueous H^+ and a stream of bubbling H_2 gas at 1 bar pressure, all maintained at a temperature of 298 K (see Figure 18.4.1).

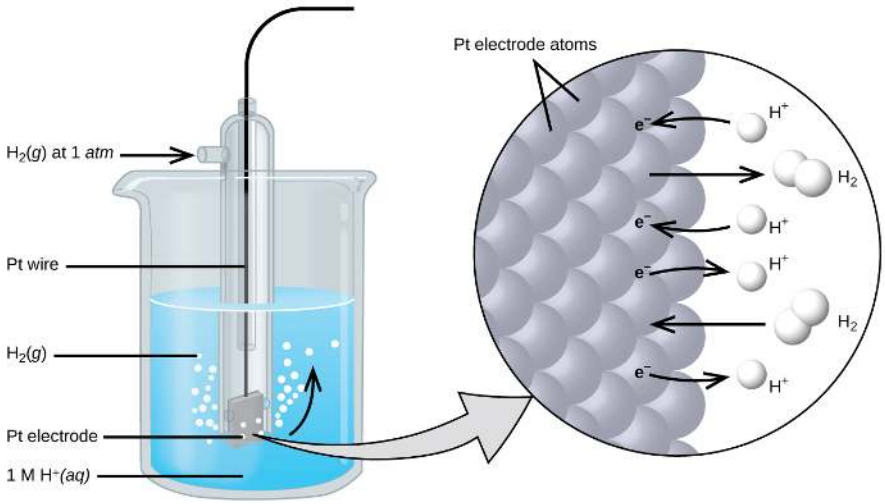


Figure 18.4.1: A standard hydrogen electrode (SHE).

The assigned potential of the SHE permits the definition of a conveniently measured potential for a single half-cell. The electrode potential (E_X) for a half-cell *X* is defined as *the potential measured for a cell comprised of X acting as cathode and the SHE acting as anode*:

$$E_{\text{cell}} = E_X - E_{SHE}$$
$$E_{SHE} = 0V \text{ (defined)}$$
$$E_{\text{cell}} = E_X$$

When the half-cell *X* is under standard-state conditions, its potential is the standard electrode potential, $E°_X$. Since the definition of cell potential requires the half-cells function as cathodes, these potentials are sometimes called *standard reduction potentials*.

This approach to measuring electrode potentials is illustrated in Figure 18.4.2 , which depicts a cell comprised of an SHE connected to a copper(II)/copper(0) half-cell under standard-state conditions. A voltmeter in the external circuit allows measurement of the potential difference between the two half-cells. Since the Cu half-cell is designated as the cathode in the definition of cell potential, it is connected to the red (positive) input of the voltmeter, while the designated SHE anode is connected to the black (negative) input. These connections insure that the sign of the measured potential will be consistent with the sign conventions of electrochemistry per the various definitions discussed above. A cell potential of +0.337 V is measured, and so

$$E°_{\text{cell}} = E°_{\text{Cu}} = +0.337V$$

Tabulations of E° values for other half-cells measured in a similar fashion are available as reference literature to permit calculations of cell potentials and the prediction of the spontaneity of redox processes.

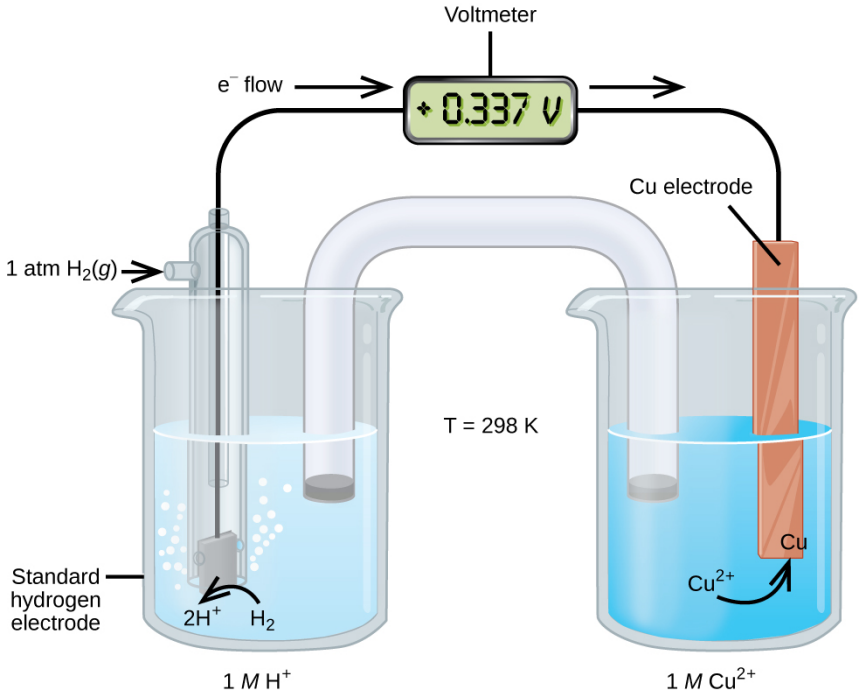


Figure 18.4.2: A cell permitting experimental measurement of the standard electrode potential for the half-reaction

Table 18.4.1 provides a listing of standard electrode potentials for a selection of half-reactions in numerical order, and a more extensive alphabetical listing is given in Appendix L.

Table 18.4.1: Selected Standard Reduction Potentials at 25 °C

Half-Reaction	*E*° (V)
$\mathrm{F_2(g) + 2\,e^- \longrightarrow 2\,F^-(aq)}$	+2.866
$\mathrm{PbO_2(s) + SO_4^{2-}(aq) + 4\,H^+(aq) + 2\,e^- \longrightarrow PbSO_4(s) + 2\,H_2O(l)}$	+1.69
$\mathrm{MnO_4^-(aq) + 8\,H^+(aq) + 5\,e^- \longrightarrow Mn^{2+}(aq) + 4\,H_2O(l)}$	+1.507

Half-Reaction	$E°$ (V)
$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$	+1.498
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.35827
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.229
$Pt^{2+}(aq) + 2e^- \longrightarrow Pt(s)$	+1.20
$Br_2(aq) + 2e^- \longrightarrow 2Br^-(aq)$	+1.0873
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.7996
$Hg_2^{2+}(aq) + 2e^- \longrightarrow 2Hg(l)$	+0.7973
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.771
$MnO_4^-(aq) + 2H_2O(l) + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$	+0.558
$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.5355
$NiO_2(s) + 2H_2O(l) + 2e^- \longrightarrow Ni(OH)_2(s) + 2OH^-(aq)$	+0.49
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
$Hg_2Cl_2(s) + 2e^- \longrightarrow 2Hg(l) + 2Cl^-(aq)$	+0.26808
$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$	+0.22233
$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$	+0.151
$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	-0.1262
$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	-0.1375
$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$	-0.257
$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$	-0.28
$PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$	-0.3505
$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$	-0.4030
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	-0.447
$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$	-0.744
$Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$	-1.185
$Zn(OH)_2(s) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$	-1.245
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	-0.7618
$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	-1.662
$Mg_2(aq) + 2e^- \longrightarrow Mg(s)$	-2.372
$Na^+(aq) + e^- \longrightarrow Na(s)$	-2.71
$Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$	-2.868
$Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$	-2.912
$K^+(aq) + e^- \longrightarrow K(s)$	-2.931
$K^+(aq) + e^- \longrightarrow K(s)$	-3.04

Example 18.4.1: Calculating Standard Cell Potentials

What is the standard potential of the galvanic cell shown in Figure 17.3?

Solution

The cell in Figure 17.3 is galvanic, the spontaneous cell reaction involving oxidation of its copper anode and reduction of silver(I) ions at its silver cathode:

$$\begin{array}{lll} \text{anode:} & \mathrm{Cu(s) \longrightarrow Cu^{2+} + 2\,e^-} & \\ \text{cathode:} & \mathrm{2\,Ag^+(aq) + 2\,e^- \longrightarrow 2\,Ag(s)} & \\ \hline \text{cell:} & \mathrm{Cu(s) + 2\,Ag^+(aq) \longrightarrow Cu^{2+} + 2\,Ag(s)} & E^o_{\text{cell}} = 0\,\text{V} \end{array}$$

The standard cell potential computed as

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= E^\circ_{Ag} - E^\circ_{Cu} \\ &= 0.7996V - 0.34V \\ &= +0.46V \end{aligned}$$

Exercise

What is the standard cell potential expected if the silver cathode half-cell in Figure 17.3 is replaced with a lead half-cell:

$$\mathrm{Pb^{2+}(aq) + 2\,e^- \longrightarrow Pb(s)}$$

Answer

−0. 47 V

Intrepreting Electrode and Cell Potentials

Thinking carefully about the definitions of cell and electrode potentials and the observations of spontaneous redox change presented thus far, a significant relation is noted. The previous section described the spontaneous oxidation of copper by aqueous silver(I) ions, but no observed reaction with aqueous lead(II) ions. Results of the calculations in Example 18.4.1 have just shown *the spontaneous process is described by a positive cell potential* while *the nonspontaneous process exhibits a negative cell potential*. And so, with regard to the relative effectiveness ("strength") with which aqueous Ag^+ and Pb^{2+} ions oxidize Cu under standard conditions, *the stronger oxidant is the one exhibiting the greater standard electrode potential, E°*. Since by convention electrode potentials are for reduction processes, an increased value of E^o corresponds to an increased driving force behind the reduction of the species (hence increased effectiveness of its action as an *oxidizing agent* on some other species). Negative values for electrode potentials are simply a consequence of assigning a value of 0 V to the SHE, indicating the reactant of the half-reaction is a weaker oxidant than aqueous hydrogen ions.

Applying this logic to the numerically ordered listing of standard electrode potentials in Table 18.4.1 shows this listing to be likewise in order of the oxidizing strength of the half-reaction's reactant species, decreasing from strongest oxidant (most positive *E*°) to weakest oxidant (most negative *E*°). Predictions regarding the spontaneity of redox reactions under standard state conditions can then be easily made by simply comparing the relative positions of their table entries. By definition, $E°_{\text{cell}}$ is positive when $E^o_{\text{cathode}} > E^o_{\text{anode}}$, and so any redox reaction in which the oxidant's entry is above the reductant's entry is predicted to be spontaneous.

Reconsideration of the two redox reactions in Example 18.4.1 provides support for this fact. The entry for the silver(I)/silver(0) half-reaction is above that for the copper(II)/copper(0) half-reaction, and so the oxidation of Cu by Ag^+ is predicted to be spontaneous ($E°_{\text{cathode}} > E°_{\text{anode}}$ and so $E°_{\text{cell}} > 0$). Conversely, the entry for the lead(II)/lead(0) half-cell is beneath that for copper(II)/copper(0), and the oxidation of Cu by Pb^{2+} is nonspontaneous ($E°_{\text{cathode}} < E°_{\text{anode}}$ and so $E°_{\text{cell}} < 0$).

Recalling the chapter on thermodynamics, the spontaneities of the forward and reverse reactions of a reversible process show a reciprocal relationship: if a process is spontaneous in one direction, it is non-spontaneous in the opposite direction. As an indicator of spontaneity for redox reactions, the potential of a cell reaction shows a consequential relationship in its arithmetic sign. The spontaneous oxidation of copper by lead(II) ions is *not* observed,

$$\mathrm{Cu(s) + Pb^{2+}(aq) \longrightarrow Cu^{2+}(aq) + Pb(s)} \quad E^\circ_{\text{forward}} = -0.47\ \mathrm{V (negative, non-spontaneous)}$$

and so the reverse reaction, the oxidation of lead by copper(II) ions, is predicted to occur spontaneously:

$$\mathrm{Pb(s) + Cu^{2+}(aq) \longrightarrow Pb^{2+}(aq) + Cu(s)} \quad E^\circ_{\text{forward}} = +0.47\ \mathrm{V (positive, spontaneous)}$$

Note that reversing the direction of a redox reaction effectively interchanges the identities of the cathode and anode half-reactions, and so the cell potential is calculated from electrode potentials in the reverse subtraction order than that for the forward reaction. In practice, a voltmeter would report a potential of −0.47 V with its red and black inputs connected to the Pb and Cu electrodes, respectively. If the inputs were swapped, the reported voltage would be +0.47 V.

Example 18.4.2: Predicting Redox Spontaneity

Are aqueous iron(II) ions predicted to spontaneously oxidize elemental chromium under standard state conditions? Assume the half-reactions to be those available in Table 18.4.1.

Solution

Referring to the tabulated half-reactions, the redox reaction in question can be represented by the equations below:

$$\mathrm{Cr(s) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe(s)}$$

The entry for the putative oxidant, Fe^{2+}, appears *above* the entry for the reductant, Cr, and so a spontaneous reaction is predicted per the quick approach described above. Supporting this predication by calculating the standard cell potential for this reaction gives

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= E^\circ_{\text{Fe(II)}} - E^\circ_{\text{Cr}} \\ &= -0.447V - -0.744V = +0.297V \end{aligned}$$

The positive value for the standard cell potential indicates the process is spontaneous under standard state conditions.

Exercise

Use the data in Table 18.4.1to predict the spontaneity of the oxidation of bromide ion by molecular iodine under standard state conditions, supporting the prediction by calculating the standard cell potential for the reaction. Repeat for the oxidation of iodide ion by molecular bromine.

Answer

$$\begin{aligned} \mathrm{I_2(s) + 2\,Br^-(aq)} &\longrightarrow \mathrm{2\,I^-(aq) + Br_2(l)} \qquad E_{\text{cell}} = -0.5518\ \mathrm{V\ (nonspontaneous)} \\ \mathrm{Br_2(s) + 2\,I^-(aq)} &\longrightarrow \mathrm{2\,Br^-(aq) + I_2(l)} \qquad E^\circ_{\text{cell}} = +0.5518\ \mathrm{V\ (spontaneous)} \end{aligned}$$

This page titled 18.4: Electrode and Cell Potentials is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar via source content that was edited to the style and standards of the LibreTexts platform.

- **17.3: Electrode and Cell Potentials** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.5: Potential, Free Energy, and Equilibrium

Learning Objectives

By the end of this section, you will be able to:

- Explain the relations between potential, free energy change, and equilibrium constants
- Perform calculations involving the relations between cell potentials, free energy changes, and equilibrium
- Use the Nernst equation to determine cell potentials under nonstandard conditions

So far in this chapter, the relationship between the cell potential and reaction *spontaneity* has been described, suggesting a link to the free energy change for the reaction (see chapter on thermodynamics). The interpretation of potentials as measures of oxidant *strength* was presented, bringing to mind similar measures of acid-base strength as reflected in equilibrium constants (see the chapter on acid-base equilibria). This section provides a summary of the relationships between potential and the related thermodynamic properties ΔG and K.

E° and ΔG°

The standard free energy change of a process, ΔG°, was defined in a previous chapter as the maximum work that could be performed by a system, w_{max}. In the case of a redox reaction taking place within a galvanic cell under standard state conditions, essentially all the work is associated with transferring the electrons from reductant-to-oxidant, w_{elec}:

$$\Delta G^\circ = w_{\text{max}} = w_{\text{elec}}$$

The work associated with transferring electrons is determined by the total amount of charge (coulombs) transferred and the cell potential:

$$\Delta G^\circ = w_{\text{elec}} = -nFE^\circ_{\text{cell}}$$

where n is the number of moles of electrons transferred, F is **Faraday's constant**, and E°_{cell} is the standard cell potential. The relation between free energy change and standard cell potential confirms the sign conventions and spontaneity criteria previously discussed for both of these properties: spontaneous redox reactions exhibit positive potentials and negative free energy changes.

E° and K

Combining a previously derived relation between ΔG° and K (see the chapter on thermodynamics) and the equation above relating ΔG° and $E°_{\text{cell}}$ yields the following:

$$\Delta G^\circ = -RT\ln K = -nFE^\circ_{\text{cell}} \tag{18.5.1}$$

with

$$E^\circ_{\text{coll}} = \left(\frac{RT}{nF}\right)\ln K$$

This equation indicates redox reactions with large (positive) standard cell potentials will proceed far towards completion, reaching equilibrium when the majority of reactant has been converted to product. A summary of the relations between $E°$, $\Delta G°$ and K is depicted in Figure 18.5.1, and a table correlating reaction spontaneity to values of these properties is provided in Table 18.5.1.

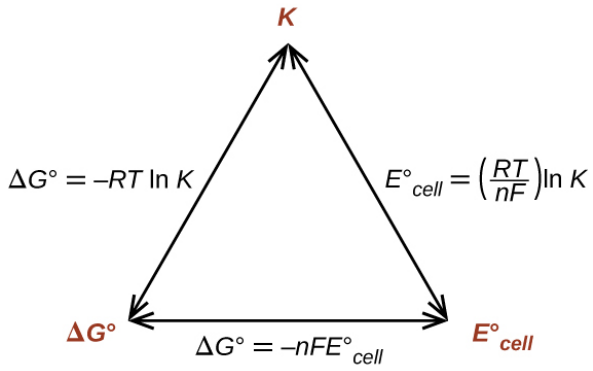


Figure 18.5.1: Graphic depicting the relation between three important thermodynamic properties.

Table 18.5.1

K	$\Delta G°$	$E°_{\text{cell}}$	Comment
> 1	< 0	> 0	Reaction is spontaneous under standard conditions Products more abundant at equilibrium
< 1	> 0	< 0	Reaction is non-spontaneous under standard conditions Reactants more abundant at equilibrium
= 1	= 0	= 0	Reaction is at equilibrium under standard conditions Reactants and products equally abundant

Example 18.5.1: Equilibrium Constants, Standard Cell Potentials, and Standard Free Energy Changes

Use data from Appendix L to calculate the standard cell potential, standard free energy change, and equilibrium constant for the following reaction at 25 °C. Comment on the spontaneity of the forward reaction and the composition of an equilibrium mixture of reactants and products.

$$2\,\text{Ag}^+(\text{aq}) + \text{Fe}(\text{s}) \rightleftharpoons 2\,\text{Ag}(\text{s}) + \text{Fe}^{2+}(\text{aq})$$

Solution

The reaction involves an oxidation-reduction reaction, so the standard cell potential can be calculated using the data in Appendix L.

anode:	$\text{Fe}(\text{s}) \longrightarrow \text{Fe}^{2+}(\text{aq}) + 2\,\text{e}^-$	$E^o_{\text{Fe}^{2+}/\text{Fe}} = -0.447\ \text{V}$
cathode:	$2 \times (\text{Ag}^+(\text{aq}) + \text{e}^- \longrightarrow \text{Ag}(\text{s}))$	$E^o_{\text{Ag}^+/\text{Ag}} = +0.7996\ \text{V}$
cell:	$\text{Fe}(\text{s}) + 2\,\text{Ag}^+(\text{aq}) \longrightarrow \text{Fe}^{2+}(\text{aq}) + 2\,\text{Ag}(\text{s})$	$E^o_{\text{cell}} = +1.2247\ \text{V}$

with

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = E^o_{\text{Fe}^{2+}/\text{Fe}} - E^o_{\text{Ag}^+/\text{Ag}} = +1.2247\ \text{V}$$

With $n = 2$, the equilibrium constant is then

$$E^\circ_{\text{cell}} = \frac{0.0592V}{n}\log K$$

$$K = 10^{n\times E^\circ_{\text{cal}}/0.0592V}$$

$$= 10^{2\times 1.247V/0.0592V}$$

$$= 10^{42.128}$$

$$= 1.3\times 10^{42}$$

The standard free energy is then Equation 18.5.1

$$\Delta G^\circ = -nFE^\circ_{\text{cell}}$$

$$= -2\times 96,485\frac{C}{mol}\times 1.247\frac{J}{C} = -240.6\frac{kJ}{mol}$$

The reaction is spontaneous, as indicated by a negative free energy change and a positive cell potential. The *K* value is very large, indicating the reaction proceeds to near completion to yield an equilibrium mixture containing mostly products.

? Exercise 18.5.1

What is the standard free energy change and the equilibrium constant for the following reaction at room temperature? Is the reaction spontaneous?

$$\text{Sn(s)} + 2\,\text{Cu}^{2+}\text{(aq)} \rightleftharpoons \text{Sn}^{2+}\text{(aq)} + 2\,\text{Cu}^{+}\text{(aq)}$$

Answer

Spontaneous

$n = 2$

$E^\circ_{\text{cell}} = +0.291V$

$[\Delta G^\circ = -56.2\frac{kJ}{mol}$

K = 6.8 10^9.

Potentials at Nonstandard Conditions: The Nernst Equation

Most of the redox processes that interest science and society do not occur under standard state conditions, and so the potentials of these systems under nonstandard conditions are a property worthy of attention. Having established the relationship between potential and free energy change in this section, the previously discussed relation between free energy change and reaction mixture composition can be used for this purpose.

$$\Delta G = \Delta G^\circ + RT\ln Q$$

Notice the reaction quotient, *Q*, appears in this equation, making the free energy change dependent upon the composition of the reaction mixture. Substituting the equation relating free energy change to cell potential yields the **Nernst equation**:

$$-nFE_{\text{cell}} = -nFE^\circ_{\text{cell}} + RT\ln Q$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF}\ln Q$$

This equation describes how the potential of a redox system (such as a galvanic cell) varies from its standard state value, specifically, showing it to be a function of the number of electrons transferred, *n*, the temperature, *T*, and the reaction mixture composition as reflected in *Q*. A convenient form of the Nernst equation for most work is one in which values for the fundamental constants (R and F) and standard temperature (298) K), along with a factor converting from natural to base-10 logarithms, have been included:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592V}{n}\log Q \tag{18.5.2}$$

Example 18.5.2: Predicting Redox Spontaneity Under Nonstandard Conditions

Use the Nernst equation to predict the spontaneity of the redox reaction shown below at room temperature.

$$\mathrm{Co(s) + Fe^{2+}(aq,\ 1.94\ M) \longrightarrow Co^{2+}(aq,\ 0.15\ M) + Fe(s)}$$

Solution

Collecting information from Appendix L and the problem,

anode:	$\mathrm{Co(s) \longrightarrow Co^{2+}(aq) + 2\,e^-}$	$E^o_{\mathrm{Co^{2+}/Co}} = -0.28\ \mathrm{V}$
cathode:	$\mathrm{Fe^{2+}(aq) + 2\,e^- \longrightarrow Fe(s)}$	$E^o_{\mathrm{Fe^{+}/Fe}} = -0.447\ \mathrm{V}$
cell:	$\mathrm{Co(s) + Fe^{2+}(Aq) \longrightarrow Fe(s) + Co^{2+}(aq)}$	$E^o_{\mathrm{cell}} = -0.17\ \mathrm{V}$

with

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.447V - (-0.28V) = -0.17V$$

Notice the negative value of the standard cell potential indicates the process is not spontaneous under standard conditions. Substitution of the Nernst equation terms for the nonstandard conditions requires calculating the reaction quotient Q

$$Q = \frac{\left[\mathrm{Co^{2+}}\right]}{\left[\mathrm{Fe^{2+}}\right]} = \frac{0.15M}{1.94M} = 0.077$$

them Q and n are substituted into Equation \red{Nernst at room temperature}

$$\begin{aligned} E_{\text{cell}} &= E^\circ_{\text{cell}} - \frac{0.0592\ \mathrm{V}}{n}\log Q \\ &= -0.17\ \mathrm{V} - \frac{0.0592\ \mathrm{V}}{2}\log 0.077 \\ &= -0.17\ \mathrm{V} + 0.033\ \mathrm{V} \\ &= -0.14\ \mathrm{V} \end{aligned}$$

The cell potential remains negative (slightly) under the specified conditions, and so the reaction remains nonspontaneous.

Exercise 18.5.2

For the cell schematic below at room temperature, identify values for *n* and *Q*, and calculate the cell potential, E_{cell}.

$$\mathrm{Al(s)|Al^{3+}(aq,\ 0.15\ M)|Cu^{2+}(aq,\ 0.025\ M)|Cu(s)}$$

Answer

n = 6; *Q* = 1440; E_{cell} = +1.97 V, spontaneous.

A **concentration cell** is constructed by connecting two nearly identical half-cells, each based on the same half-reaction and using the same electrode, varying only in the concentration of one redox species. The potential of a concentration cell, therefore, is determined only by the difference in concentration of the chosen redox species. The example problem below illustrates the use of the Nernst equation in calculations involving concentration cells.

Example 18.5.3: Concentration Cells

What is the cell potential of the concentration cell operating at room temperature described by

$$\mathrm{Zn(s)|Zn^{2+}(aq,\ 0.10\ M)|Zn^{2+}(aq,\ 0.50\ M)|Zn(s)}$$

Solution

From the information given:

anode:	$\mathrm{Zn(s) \longrightarrow Zn^{2+}(aq,\ 0.10\ M) + 2\,e^-}$	$E^o_{\mathrm{Co^{2+}/Co}} = -0.7618\ \mathrm{V}$
cathode:	$\mathrm{Zn^{2+}(aq,\ 0.50\ M) + 2\,e^- \longrightarrow Zn(s)}$	$E^o_{\mathrm{Fe^{+}/Fe}} = -0,7618\ \mathrm{V}$
cell:	$\mathrm{\cancel{Zn(s)} + Zn^{2+}(aq,\ 0.50\ M) \longrightarrow Zn^{2+}(aq,\ 0.10\ M) + \cancel{Zn(s)}}$	$E^o_{\mathrm{cell}} = 0\ \mathrm{V}$

Substituting into the Nernst equation (Equation 18.5.2),

$$E_{\mathrm{cell}} = 0.000\ \mathrm{V} - \frac{0.0592\ \mathrm{V}}{2}\log\frac{0.10}{0.50} = +0.021\ \mathrm{V}$$

The positive value for cell potential indicates the overall cell reaction (see above) is spontaneous. This spontaneous reaction is one in which the zinc ion concentration in the cathode falls (it is reduced to elemental zinc) while that in the anode rises (it is produced by oxidation of the zinc anode). A greater driving force for zinc reduction is present in the cathode, where the zinc(II) ion concentration is greater ($E_{\mathrm{cathode}} > E_{\mathrm{anode}}$).

? Exercise 18.5.1

The concentration cell above was allowed to operate until the cell reaction reached equilibrium at room temperature. What are the cell potential and the concentrations of zinc(II) in each half-cell for the cell now?

Answer

E_{cell} = 0.000 V; $[Zn^{2+}]_{\mathrm{cathode}} = [Zn^{2+}]_{\mathrm{anode}}$ = 0.30 *M*

This page titled 18.5: Potential, Free Energy, and Equilibrium is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang via source content that was edited to the style and standards of the LibreTexts platform.

- **17.4: Potential, Free Energy, and Equilibrium** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.6: Batteries and Fuel Cells

Learning Objectives

By the end of this section, you will be able to:

- Describe the electrochemistry associated with several common batteries
- Distinguish the operation of a fuel cell from that of a battery

There are many technological products associated with the past two centuries of electrochemistry research, none more immediately obvious than the battery. A **battery** is a galvanic cell that has been specially designed and constructed in a way that best suits its intended use a source of electrical power for specific applications. Among the first successful batteries was the *Daniell cell*, which relied on the spontaneous oxidation of zinc by copper(II) ions (Figure 18.6.1):

$$\mathrm{Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)}$$

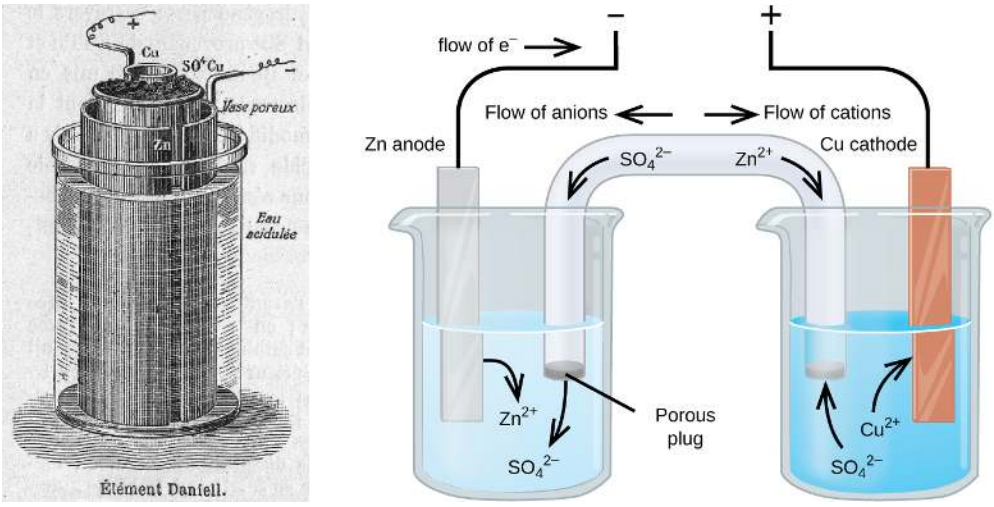


Figure 18.6.1: Illustration of a Daniell cell taken from a 1904 journal publication (left) along with a simplified illustration depicting the electrochemistry of the cell (right). The 1904 design used a porous clay pot to both contain one of the half-cell's content and to serve as a salt bridge to the other half-cell.

Modern batteries exist in a multitude of forms to accommodate various applications, from tiny button batteries that provide the modest power needs of a wristwatch to the very large batteries used to supply backup energy to municipal power grids. Some batteries are designed for single-use applications and cannot be recharged (primary cells), while others are based on conveniently reversible cell reactions that allow recharging by an external power source (secondary cells). This section will provide a summary of the basic electrochemical aspects of several batteries familiar to most consumers, and will introduce a related electrochemical device called a *fuel cell* that can offer improved performance in certain applications.

Link to Learning

Visit this site to learn more about batteries.

Single-Use Batteries

A common primary battery is the dry cell, which uses a zinc can as both container and anode ("–" terminal) and a graphite rod as the cathode ("+" terminal). The Zn can is filled with an electrolyte paste containing manganese(IV) oxide, zinc(II) chloride, ammonium chloride, and water. A graphite rod is immersed in the electrolyte paste to complete the cell. The spontaneous cell reaction involves the oxidation of zinc:

$$\mathrm{Zn(s) \longrightarrow Zn^{2+}(aq) + 2\,e^-} \qquad \text{(anode)}$$

and the reduction of manganese(IV)

$$2\,MnO_2(s) + 2\,NH_4Cl(aq) + 2\,e^- \longrightarrow Mn_2O_3(s) + 2\,NH_3(aq) + H_2O(l) + 2\,Cl^- \qquad \text{(cathode)}$$

which together yield the cell reaction:

$$2\,MnO_2(s) + 2\,NH_4Cl(aq) + Zn(s) \longrightarrow Zn^{2+}(aq) + Mn_2O_3(s) + 2\,NH_3(aq) + H_2O(l) + 2\,Cl^- \qquad \text{(cell)}$$

The voltage (*cell potential*) of a dry cell is approximately 1.5 V. Dry cells are available in various sizes (e.g., D, C, AA, AAA). All sizes of dry cells comprise the same components, and so they exhibit the same voltage, but larger cells contain greater amounts of the redox reactants and therefore are capable of transferring correspondingly greater amounts of charge. Like other galvanic cells, dry cells may be connected in series to yield batteries with greater voltage outputs, if needed.

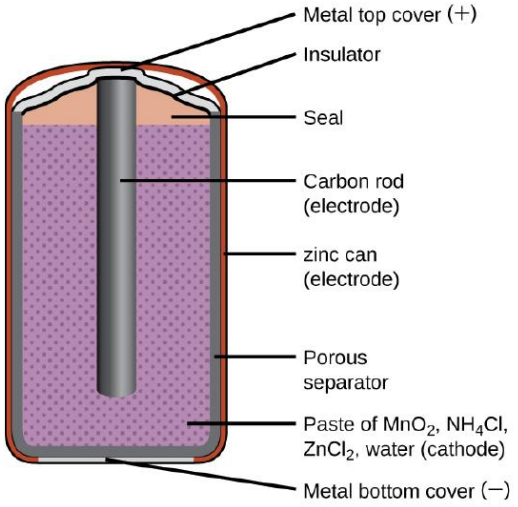


Figure 18.6.2: A schematic diagram shows a typical dry cell.

Link to Learning

Visit this site to learn more about zinc-carbon batteries.

Alkaline batteries (Figure 18.6.3) were developed in the 1950s to improve on the performance of the dry cell, and they were designed around the same redox couples. As their name suggests, these types of batteries use alkaline electrolytes, often potassium hydroxide. The reactions are

anode:	$Zn(s) + 2\,OH^-(aq) \longrightarrow ZnO(s) + H_2O(l) + 2\,e^-$	
cathode:	$2\,MnO_2(s) + H_2O(l) + 2\,e^- \longrightarrow Mn_2O_3(s) + 2\,OH^-(aq)$	
cell:	$Zn(s) + 2\,MnO_2(s) \longrightarrow ZnO(s) + Mn_2O_3(s)$	$E_{cell} = +1.43\ V$

An alkaline battery can deliver about three to five times the energy of a zinc-carbon dry cell of similar size. Alkaline batteries are prone to leaking potassium hydroxide, so they should be removed from devices for long-term storage. While some alkaline batteries are rechargeable, most are not. Attempts to recharge an alkaline battery that is not rechargeable often leads to rupture of the battery and leakage of the potassium hydroxide electrolyte.

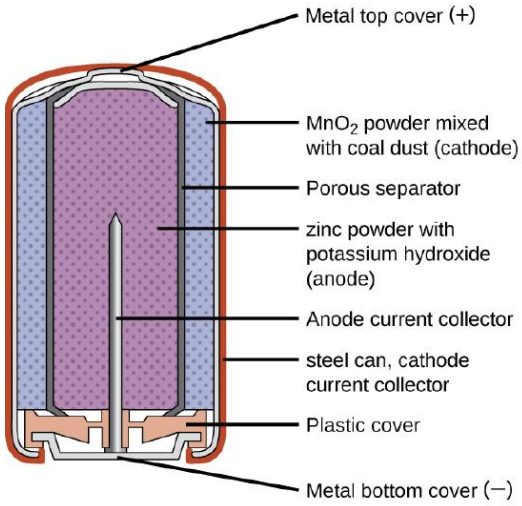


Figure 18.6.3: Alkaline batteries were designed as improved replacements for zinc-carbon (dry cell) batteries.

Link to Learning

Visit this site to learn more about alkaline batteries.

Rechargeable (Secondary) Batteries

Nickel-cadmium, or NiCd, batteries (Figure 18.6.4) consist of a nickel-plated cathode, cadmium-plated anode, and a potassium hydroxide electrode. The positive and negative plates, which are prevented from shorting by the separator, are rolled together and put into the case. This is a "jelly-roll" design and allows the NiCd cell to deliver much more current than a similar-sized alkaline battery. The reactions are

$$\begin{array}{lll} \text{anode:} & \mathrm{Cd(s) + 2\,OH^-(aq) \longrightarrow Cd(OH)_2(s) + 2\,e^-} & \\ \text{cathode:} & \mathrm{NiO_2(s) + 2\,H_2O(l) + 2\,e^- \longrightarrow Ni(OH)_2(s) + 2\,OH^-(aq)} & \\ \hline \text{cell:} & \mathrm{Cd(s) + NiO_2(s) + 2\,H_2O(l) \longrightarrow Cd(OH)_2(s) + Ni(OH)_2(s)} & E_{\text{cell}} \approx +1.2\,\text{V} \end{array}$$

When properly treated, a NiCd battery can be recharged about 1000 times. Cadmium is a toxic heavy metal so NiCd batteries should never be ruptured or incinerated, and they should be disposed of in accordance with relevant toxic waste guidelines.

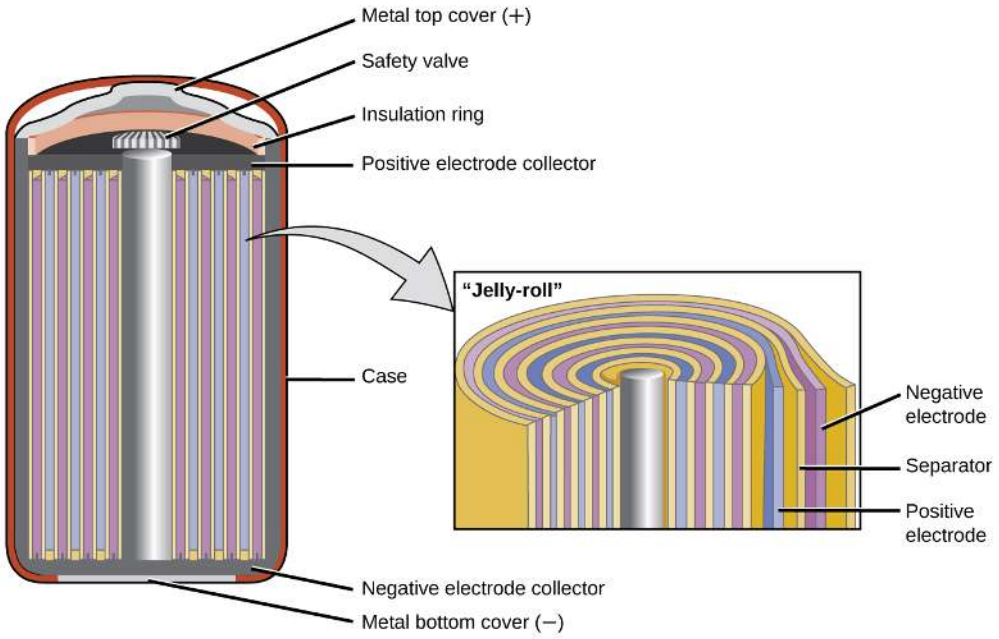


Figure 18.6.4: NiCd batteries use a "jelly-roll" design that significantly increases the amount of current the battery can deliver as compared to a similar-sized alkaline battery.

Link to Learning

Visit this site for more information about nickel cadmium rechargeable batteries.

Lithium ion batteries (Figure 18.6.5) are among the most popular rechargeable batteries and are used in many portable electronic devices. The reactions are

$$\begin{array}{lll} \text{anode:} & \mathrm{LiCoO_2 \longrightarrow Li_{1-x}CoO_2 + x\,Li^+ + x\,e^-} & \\ \text{cathode:} & \mathrm{x\,Li^+ + x\,e^- + x\,C_6 \longrightarrow x\,LiC_6} & \\ \hline \text{cell:} & \mathrm{LiCoO_2 + x\,C_6 \longrightarrow Li_{1-x}CoO_2 + x\,LiC_6} & E_{\text{cell}} \approx +3.7\ \text{V} \end{array}$$

The variable stoichiometry of the cell reaction leads to variation in cell voltages, but for typical conditions, *x* is usually no more than 0.5 and the cell voltage is approximately 3.7 V. Lithium batteries are popular because they can provide a large amount current, are lighter than comparable batteries of other types, produce a nearly constant voltage as they discharge, and only slowly lose their charge when stored.

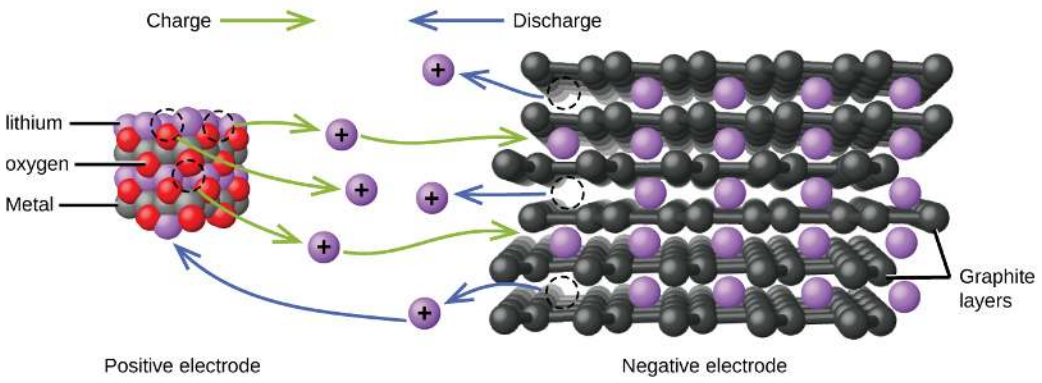


Figure 18.6.5: In a lithium ion battery, charge flows as the lithium ions are transferred between the anode and cathode.

Link to Learning

Visit this site for more information about lithium ion batteries.

The lead acid battery (Figure 18.6.6) is the type of secondary battery commonly used in automobiles. It is inexpensive and capable of producing the high current required by automobile starter motors. The reactions for a lead acid battery are

$$\begin{array}{ll} \text{anode:} & \mathrm{Pb(s) + HSO_4^-(aq) \longrightarrow PbSO_4(s) + H^+(aq) + 2\,e^-} \\ \text{cathode:} & \mathrm{PbO_2(s) + HSO_4^-(aq) + 3\,H^+(aq) + 2\,e^- \longrightarrow PbSO_4(s) + 2\,H_2O(l)} \\ \hline \text{cell:} & \mathrm{Pb(s) + PbO_2(s) + 2\,H_2SO_4(aq) \longrightarrow 2\,PbSO_4(s) + 2\,H_2O(l)} \qquad E_{\text{cell}} \approx +2\text{ V} \end{array}$$

Each cell produces 2 V, so six cells are connected in series to produce a 12-V car battery. Lead acid batteries are heavy and contain a caustic liquid electrolyte, $H_2SO_4(aq)$, but are often still the battery of choice because of their high current density. Since these batteries contain a significant amount of lead, they must always be disposed of properly.

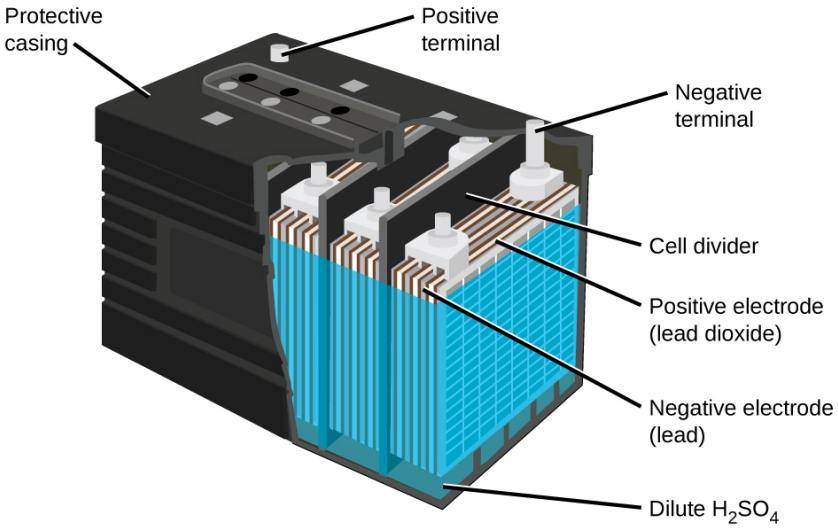


Figure 18.6.6: The lead acid battery in your automobile consists of six cells connected in series to give 12 V.

Link to Learning

Visit this site for more information about lead acid batteries.

Fuel Cells

A fuel cell is a galvanic cell that uses traditional combustive fuels, most often hydrogen or methane, that are continuously fed into the cell along with an oxidant. (An alternative, but not very popular, name for a fuel cell is a *flow battery*.) Within the cell, fuel and oxidant undergo the same redox chemistry as when they are combusted, but via a catalyzed electrochemical that is significantly more efficient. For example, a typical hydrogen fuel cell uses graphite electrodes embedded with platinum-based catalysts to accelerate the two half-cell reactions:

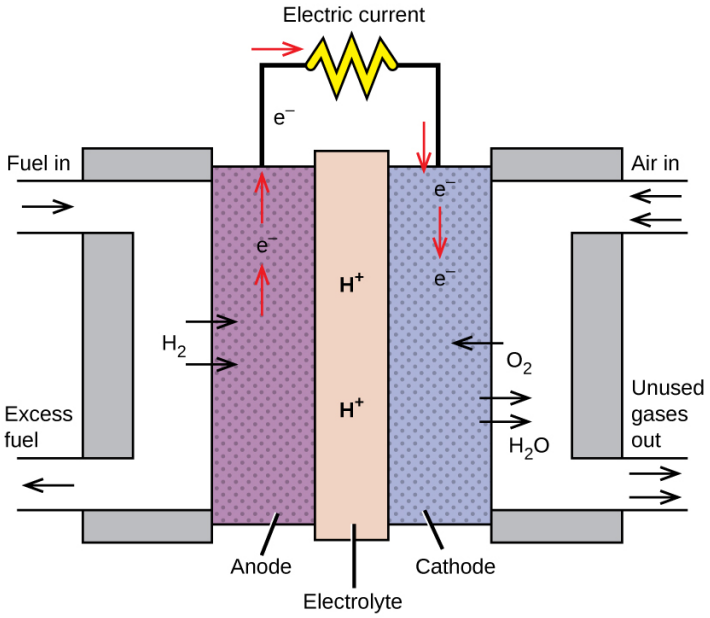


Figure 18.6.7: In this hydrogen fuel cell, oxygen from the air reacts with hydrogen, producing water and electricity.

$$\begin{array}{lll} \text{anode:} & 2\,H_2(g) \longrightarrow 4\,H^+(aq) + 4\,e^- & \\ \text{cathode:} & O_2(g) + 4\,H^+(aq) + 4\,e^- \longrightarrow 2\,H_2O(g) & \\ \hline \text{cell:} & 2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) & E_{\text{cell}} \approx +1.2\,\text{V} \end{array}$$

These types of fuel cells generally produce voltages of approximately 1.2 V. Compared to an internal combustion engine, the energy efficiency of a fuel cell using the same redox reaction is typically more than double (~20%–25% for an engine versus ~50%–75% for a fuel cell). Hydrogen fuel cells are commonly used on extended space missions, and prototypes for personal vehicles have been developed, though the technology remains relatively immature.

Link to Learning

Check out this link to learn more about fuel cells.

This page titled 18.6: Batteries and Fuel Cells is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar via source content that was edited to the style and standards of the LibreTexts platform.

- **17.5: Batteries and Fuel Cells** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.7: Electrolysis

Learning Objectives

By the end of this section, you will be able to:

- Describe the process of electrolysis
- Compare the operation of electrolytic cells with that of galvanic cells
- Perform stoichiometric calculations for electrolytic processes

Electrochemical cells in which spontaneous redox reactions take place (*galvanic cells*) have been the topic of discussion so far in this chapter. In these cells, *electrical work is done by a redox system on its surroundings* as electrons produced by the redox reaction are transferred through an external circuit. This final section of the chapter will address an alternative scenario in which *an external circuit does work on a redox system* by imposing a voltage sufficient to drive an otherwise nonspontaneous reaction, a process known as **electrolysis**. A familiar example of electrolysis is recharging a battery, which involves use of an external power source to drive the spontaneous (discharge) cell reaction in the reverse direction, restoring to some extent the composition of the half-cells and the voltage of the battery. Perhaps less familiar is the use of electrolysis in the refinement of metallic ores, the manufacture of commodity chemicals, and the *electroplating* of metallic coatings on various products (e.g., jewelry, utensils, auto parts). To illustrate the essential concepts of electrolysis, a few specific processes will be considered.

The Electrolysis of Molten Sodium Chloride

Metallic sodium, Na, and chlorine gas, Cl_2, are used in numerous applications, and their industrial production relies on the large-scale electrolysis of molten sodium chloride, NaCl(*l*). The industrial process typically uses a *Downs cell* similar to the simplified illustration shown in Figure 18.7.1. The reactions associated with this process are:

$$\begin{array}{lll} \text{oxidation:} & 2\,\mathrm{Cl^-(l)} & \longrightarrow \mathrm{Cl_2(g} + 2e^- \\ + \quad \text{reduction:} & \mathrm{Na^+(l)} + \mathrm{e^-} & \longrightarrow \mathrm{Na(l)} \\ \hline \text{cell:} & 2\,\mathrm{Cl^-(l)} + 2\,\mathrm{Na^+(l)} & \longrightarrow 2\,\mathrm{Na(l)} + \mathrm{Cl_2(g} + 2e^- \end{array}$$

The cell potential for the above process is negative, indicating the reaction as written (decomposition of liquid NaCl) is not spontaneous. To force this reaction, *a positive potential of magnitude greater than the negative cell potential* must be applied to the cell.

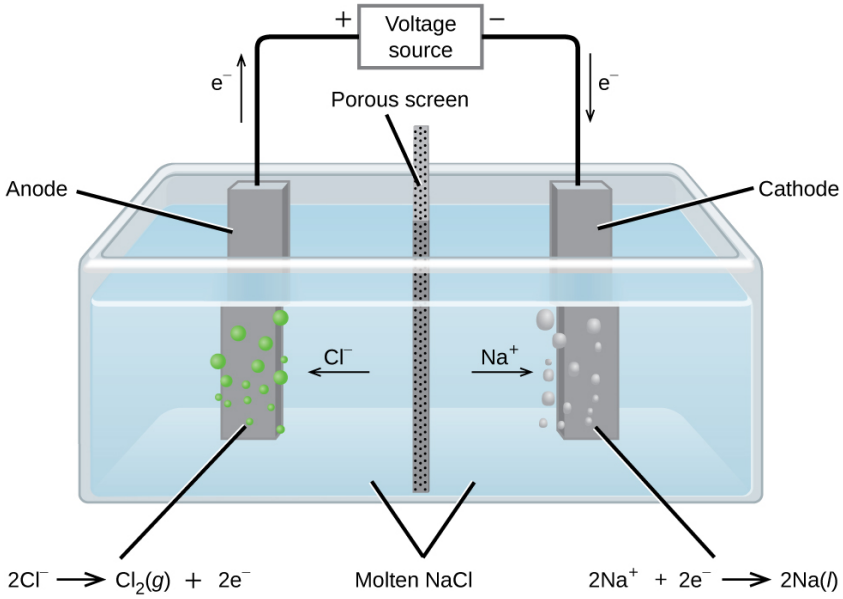


Figure 18.7.1: Cells of this sort (a cell for the electrolysis of molten sodium chloride) are used in the *Downs process* for production of sodium and chlorine, and they typically use iron cathodes and carbon anodes.

The Electrolysis of Water

Water may be electrolytically decomposed in a cell similar to the one illustrated in Figure 18.7.2 To improve electrical conductivity without introducing a different redox species, the hydrogen ion concentration of the water is typically increased by addition of a strong acid. The redox processes associated with this cell are

	oxidation:	$2\,H_2(l)$	$\longrightarrow O_2(g) + 4\,H^+ + 4\,e^-$	$E^o_{anode} = +1.299\ V$
+	reduction:	$2\,H^+(aq) + 2\,e^-$	$\longrightarrow H_2(g)$	$E^o_{anode} = 0\ V$
	cell:	$2\,H_2O(l)$	$\longrightarrow 2\,H_2(g) + O_2(g$	$E^o_{cell} = -1.299\ V$

Again, the cell potential as written is negative, indicating a nonspontaneous cell reaction that must be driven by imposing a cell voltage greater than +1.229 V. Keep in mind that *standard* electrode potentials are used to inform thermodynamic predictions here, though the cell is *not* operating under standard state conditions. Therefore, at best, calculated cell potentials should be considered ballpark estimates.

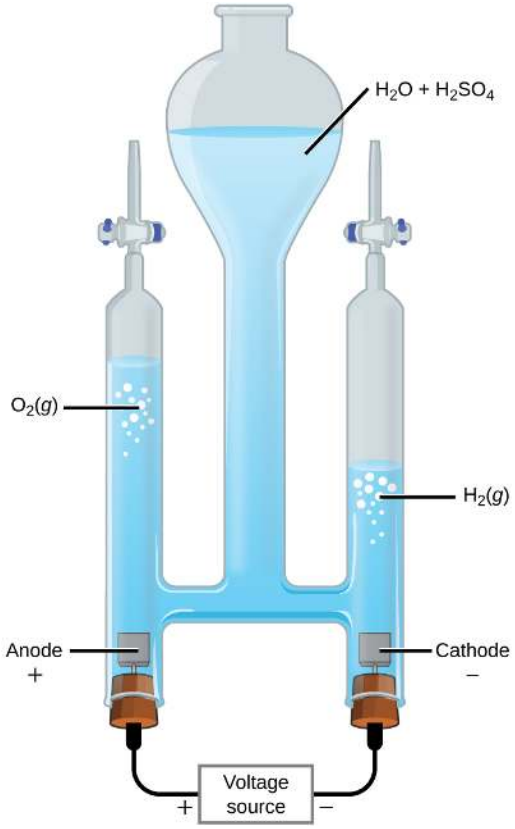


Figure 18.7.2: The electrolysis of water produces stoichiometric amounts of oxygen gas at the anode and hydrogen at the anode.

The Electrolysis of Aqueous Sodium Chloride

When aqueous solutions of ionic compounds are electrolyzed, the anode and cathode half-reactions may involve the electrolysis of either water species (H_2O, H^+, OH^-) or solute species (the cations and anions of the compound). As an example, the electrolysis of aqueous sodium chloride could involve either of these two anode reactions:

$$2\,Cl^-(aq) \longrightarrow Cl_2(g) + 2\,e^- \quad E^\circ_{anode} = +1.35827\ V \tag{i}$$

or

$$2\,H_2O(l) \longrightarrow O_2(g) + 4\,H^+(aq) + 4\,e^- \quad E^\circ_{anode} = +1.229\ V \tag{ii}$$

The standard electrode (*reduction*) potentials of these two half-reactions indicate water may be *oxidized* at a less negative/more positive potential (–1.229 V) than chloride ion (–1.358 V). Thermodynamics thus predicts that water would be more readily

oxidized, though in practice it is observed that both water and chloride ion are oxidized under typical conditions, producing a mixture of oxygen and chlorine gas.

Turning attention to the cathode, the possibilities for reduction are:

$$2\,H^+(aq) + 2\,e^- \longrightarrow H_2(g) \quad E^\circ_{anode} = 0\,V \tag{iii}$$

or

$$2\,H_2O(l) + 2\,e^- \longrightarrow H_2(g) + 4\,OH^-(aq) \quad E^\circ_{anode} = -0.8277\,V \tag{iv}$$

or

$$Na^+(aq) + e^- \longrightarrow Na(s) \quad E^\circ_{anode} = -2.71\,V \tag{v}$$

Comparison of these *standard* half-reaction potentials suggests the reduction of hydrogen ion is thermodynamically favored. However, in a neutral aqueous sodium chloride solution, the concentration of hydrogen ion is far below the standard state value of 1 *M* (approximately 10^{-7} *M*), and so the observed cathode reaction is actually reduction of water. The net cell reaction in this case is then

$$2\,H_2O(l) + 2\,Cl^-(aq) \longrightarrow H_2(g) + Cl_2(g) + 2\,OH^-(aq) \quad E^\circ_{cell} = -2.186\,V$$

This electrolysis reaction is part of the *chlor-alkali process* used by industry to produce chlorine and sodium hydroxide (lye).

Chemistry in Everyday Life: Electroplating

An important use for electrolytic cells is in electroplating. Electroplating results in a thin coating of one metal on top of a conducting surface. Reasons for electroplating include making the object more corrosion resistant, strengthening the surface, producing a more attractive finish, or for purifying metal. The metals commonly used in electroplating include cadmium, chromium, copper, gold, nickel, silver, and tin. Common consumer products include silver-plated or gold-plated tableware, chrome-plated automobile parts, and jewelry. The silver plating of eating utensils is used here to illustrate the process. (Figure 18.7.3).

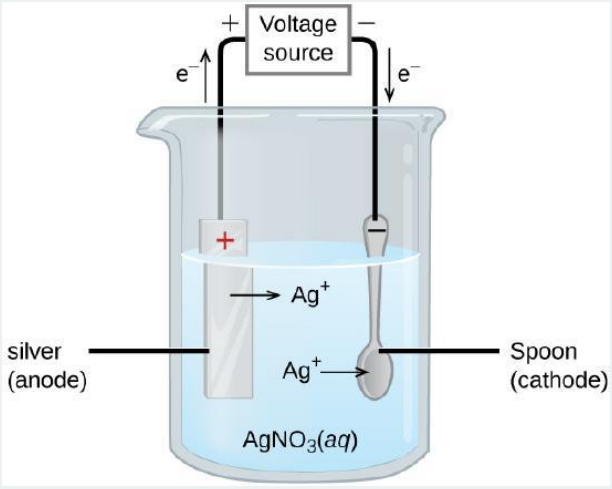


Figure 18.7.3: This schematic shows an electrolytic cell for silver plating eating utensils.

In the figure, the anode consists of a silver electrode, shown on the left. The cathode is located on the right and is the spoon, which is made from inexpensive metal. Both electrodes are immersed in a solution of silver nitrate. Applying a sufficient potential results in the oxidation of the silver anode

$$Ag(s) \longrightarrow Ag^+(aq) + e^- \qquad (anode)$$

and reduction of silver ion at the (spoon) cathode:

$$Ag^+(aq) + e^- \longrightarrow Ag(s) \qquad (cathode)$$

The net result is the transfer of silver metal from the anode to the cathode. Several experimental factors must be carefully controlled to obtain high-quality silver coatings, including the exact composition of the electrolyte solution, the cell voltage applied, and the rate of the electrolysis reaction (electrical current).

Quantitative Aspects of Electrolysis

Electrical current is defined as the rate of flow for any charged species. Most relevant to this discussion is the flow of electrons. Current is measured in a composite unit called an ampere, defined as one coulomb per second (A = 1 C/s). The charge transferred, *Q*, by passage of a constant current, *I*, over a specified time interval, *t*, is then given by the simple mathematical product

$$Q = It$$

When electrons are transferred during a redox process, the stoichiometry of the reaction may be used to derive the total amount of (electronic) charge involved. For example, the generic reduction process

$$\mathrm{M^{n^+}(aq) + n\,e^- \longrightarrow M(s)}$$

involves the transfer of n mole of electrons. The total charge transferred is, therefore,

$$Q = nF$$

where F is Faraday's constant, the charge in coulombs for one mole of electrons. If the reaction takes place in an electrochemical cell, the current flow is conveniently measured, and it may be used to assist in stoichiometric calculations related to the cell reaction.

Example 18.7.1: Converting Current to Moles of Electrons

In one process used for electroplating silver, a current of 10.23 A was passed through an electrolytic cell for exactly 1 hour. How many moles of electrons passed through the cell? What mass of silver was deposited at the cathode from the silver nitrate solution?

Solution

Faraday's constant can be used to convert the charge (*Q*) into moles of electrons (*n*). The charge is the current (*I*) multiplied by the time

$$n = \frac{Q}{F} = \frac{\frac{10.23C}{s} \times 1hr \times \frac{60min}{hr} \times \frac{60s}{\text{min}}}{96,485C/mole^-} = \frac{36,830C}{96,485C/mole^-} = 0.3817mole^-$$

From the problem, the solution contains $AgNO_3$, so the reaction at the cathode involves 1 mole of electrons for each mole of silver

$$\mathrm{Ag^+(aq) + e^- \longrightarrow Ag(s)} \qquad \text{(cathode)}$$

The atomic mass of silver is 107.9 g/mol, so

$$\text{mass } Ag = 0.3817 \text{ mol } \mathrm{e^-} \times \frac{1 \text{ mol Ag}}{1 \text{ mol } \mathrm{e^-}} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 41.19 \text{ g Ag}$$

Exercise 18.7.1

Aluminum metal can be made from aluminum(III) ions by electrolysis. What is the half-reaction at the cathode? What mass of aluminum metal would be recovered if a current of 25.0 A passed through the solution for 15.0 minutes?

Answer

$$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$$

0.0777 mol Al = 2.10 g Al.

Example 18.7.2: Time Required for Deposition

In one application, a 0.010-mm layer of chromium must be deposited on a part with a total surface area of 3.3 m^2 from a solution of containing chromium(III) ions. How long would it take to deposit the layer of chromium if the current was 33.46 A? The density of chromium (metal) is 7.19 g/cm^3.

Solution

First, compute the volume of chromium that must be produced (equal to the product of surface area and thickness):

$$\text{volume } = \left(0.010mm \times \frac{1cm}{10mm}\right) \times \left(3.3m^2 \times \left(\frac{10,000cm^2}{1m^2}\right)\right) = 33cm^3$$

Use the computed volume and the provided density to calculate the molar amount of chromium required:

$$\text{mass } = \text{ volume } \times \text{ density } = 33cm^3 \times \frac{7.19g}{cm^3} = 237gCr$$

$$molCr = 237gCr \times \frac{1molCr}{52.00gCr} = 4.56molCr$$

The stoichiometry of the chromium(III) reduction process requires three moles of electrons for each mole of chromium(0) produced, and so the total charge required is:

$$Q = 4.56molCr \times \frac{3mole^-}{1molCr} \times \frac{96485C}{mole^-} = 1.32 \times 10^6 C$$

Finally, if this charge is passed at a rate of 33.46 C/s, the required time is:

$$t = \frac{Q}{I} = \frac{1.32 \times 10^6 C}{33.46 C/s} = 3.95 \times 10^4 s = 11.0hr$$

Exercise 18.7.2

What mass of zinc is required to galvanize the top of a 3.00 m 5.50 m sheet of iron to a thickness of 0.100 mm of zinc? If the zinc comes from a solution of $Zn(NO_3)_2$ and the current is 25.5 A, how long will it take to galvanize the top of the iron? The density of zinc is 7.140 g/cm^3.

Answer

11.8 kg Zn requires 382 hours.

This page titled 18.7: Electrolysis is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar via source content that was edited to the style and standards of the LibreTexts platform.

- **17.7: Electrolysis** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

18.8: Key Terms

active electrode | electrode that participates as a reactant or product in the oxidation-reduction reaction of an electrochemical cell; the mass of an active electrode changes during the oxidation-reduction reaction

alkaline battery | primary battery similar to a *dry cell* that uses an alkaline (often potassium hydroxide) electrolyte; designed to be an improved replacement for the dry cell, but with more energy storage and less electrolyte leakage than typical dry cell

anode | electrode in an electrochemical cell at which oxidation occurs

battery | single or series of galvanic cells designed for use as a source of electrical power

cathode | electrode in an electrochemical cell at which reduction occurs

cathodic protection | approach to preventing corrosion of a metal object by connecting it to a *sacrificial anode* composed of a more readily oxidized metal

cell notation (schematic) | symbolic representation of the components and reactions in an electrochemical cell

cell potential (E_{cell}) | difference in potential of the cathode and anode half-cells

concentration cell | galvanic cell comprising half-cells of identical composition but for the concentration of one redox reactant or product

corrosion | degradation of metal via a natural electrochemical process

dry cell | primary battery, also called a zinc-carbon battery, based on the spontaneous oxidation of zinc by manganese(IV)

electrode potential (E_X) | the potential of a cell in which the half-cell of interest acts as a cathode when connected to the standard hydrogen electrode

electrolysis | process using electrical energy to cause a nonspontaneous process to occur

electrolytic cell | electrochemical cell in which an external source of electrical power is used to drive an otherwise nonspontaneous process

Faraday's constant (F) | charge on 1 mol of electrons; F = 96,485 C/mol e^-

fuel cell | devices similar to galvanic cells that require a continuous feed of redox reactants; also called a *flow battery*

galvanic (voltaic) cell | electrochemical cell in which a spontaneous redox reaction takes place; also called a *voltaic cell*

galvanization | method of protecting iron or similar metals from corrosion by coating with a thin layer of more easily oxidized zinc.

half cell | component of a cell that contains the redox conjugate pair ("couple") of a single reactant

inert electrode | electrode that conducts electrons to and from the reactants in a half-cell but that is not itself oxidized or reduced

lead acid battery | rechargeable battery commonly used in automobiles; it typically comprises six galvanic cells based on Pb half-reactions in acidic solution

lithium ion battery | widely used rechargeable battery commonly used in portable electronic devices, based on lithium ion transfer between the anode and cathode

Nernst equation | relating the potential of a redox system to its composition

nickel-cadmium battery | rechargeable battery based on Ni/Cd half-cells with applications similar to those of lithium ion batteries

primary cell | nonrechargeable battery, suitable for single use only

sacrificial anode | electrode constructed from an easily oxidized metal, often magnesium or zinc, used to prevent corrosion of metal objects via cathodic protection

salt bridge | tube filled with inert electrolyte solution

secondary cell | battery designed to allow recharging

standard cell potential | the cell potential when all reactants and products are in their standard states (1 bar or 1 atm or gases; 1 *M* for solutes), usually at 298.15 K

standard electrode potential () | electrode potential measured under standard conditions (1 bar or 1 atm for gases; 1 *M* for solutes) usually at 298.15 K

standard hydrogen electrode (SHE) | half-cell based on hydrogen ion production, assigned a potential of exactly 0 V under standard state conditions, used as the universal reference for measuring electrode potential

This page titled 18.8: Key Terms is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.8: Key Terms** is licensed CC BY 4.0.

18.9: Key Equations

$$\Delta G = -nFE_{\text{cell}}$$

$$Q = I\,t = n\,F$$

This page titled 18.9: Key Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.9: Key Equations** is licensed CC BY 4.0.

18.10: Summary

Redox reactions are defined by changes in reactant oxidation numbers, and those most relevant to electrochemistry involve actual transfer of electrons. Aqueous phase redox processes often involve water or its characteristic ions, H^+ and OH^-, as reactants in addition to the oxidant and reductant, and equations representing these reactions can be challenging to balance. The half-reaction method is a systematic approach to balancing such equations that involves separate treatment of the oxidation and reduction half-reactions.

Galvanic cells are devices in which a spontaneous redox reaction occurs indirectly, with the oxidant and reductant redox couples contained in separate half-cells. Electrons are transferred from the reductant (in the anode half-cell) to the oxidant (in the cathode half-cell) through an external circuit, and inert solution phase ions are transferred between half-cells, through a salt bridge, to maintain charge neutrality. The construction and composition of a galvanic cell may be succinctly represented using chemical formulas and others symbols in the form of a cell schematic (cell notation).

The property of potential, E, is the energy associated with the separation/transfer of charge. In electrochemistry, the potentials of cells and half-cells are thermodynamic quantities that reflect the driving force or the spontaneity of their redox processes. The cell potential of an electrochemical cell is the difference in between its cathode and anode. To permit easy sharing of half-cell potential data, the standard hydrogen electrode (SHE) is assigned a potential of exactly 0 V and used to define a single electrode potential for any given half-cell. The electrode potential of a half-cell, E_X, is the cell potential of said half-cell acting as a cathode when connected to a SHE acting as an anode. When the half-cell is operating under standard state conditions, its potential is the standard electrode potential, $E°_X$. Standard electrode potentials reflect the relative oxidizing strength of the half-reaction's reactant, with stronger oxidants exhibiting larger (more positive) $E°_X$ values. Tabulations of standard electrode potentials may be used to compute standard cell potentials, $E°_{cell}$, for many redox reactions. The arithmetic sign of a cell potential indicates the spontaneity of the cell reaction, with positive values for spontaneous reactions and negative values for nonspontaneous reactions (spontaneous in the reverse direction).

Potential is a thermodynamic quantity reflecting the intrinsic driving force of a redox process, and it is directly related to the free energy change and equilibrium constant for the process. For redox processes taking place in electrochemical cells, the maximum (electrical) work done by the system is easily computed from the cell potential and the reaction stoichiometry and is equal to the free energy change for the process. The equilibrium constant for a redox reaction is logarithmically related to the reaction's cell potential, with larger (more positive) potentials indicating reactions with greater driving force that equilibrate when the reaction has proceeded far towards completion (large value of K). Finally, the potential of a redox process varies with the composition of the reaction mixture, being related to the reactions standard potential and the value of its reaction quotient, Q, as described by the Nernst equation.

Galvanic cells designed specifically to function as electrical power supplies are called batteries. A variety of both single-use batteries (primary cells) and rechargeable batteries (secondary cells) are commercially available to serve a variety of applications, with important specifications including voltage, size, and lifetime. Fuel cells, sometimes called flow batteries, are devices that harness the energy of spontaneous redox reactions normally associated with combustion processes. Like batteries, fuel cells enable the reaction's electron transfer via an external circuit, but they require continuous input of the redox reactants (fuel and oxidant) from an external reservoir. Fuel cells are typically much more efficient in converting the energy released by the reaction to useful work in comparison to internal combustion engines.

Spontaneous oxidation of metals by natural electrochemical processes is called corrosion, familiar examples including the rusting of iron and the tarnishing of silver. Corrosion process involve the creation of a galvanic cell in which different sites on the metal object function as anode and cathode, with the corrosion taking place at the anodic site. Approaches to preventing corrosion of metals include use of a protective coating of zinc (galvanization) and the use of sacrificial anodes connected to the metal object (cathodic protection).

Nonspontaneous redox processes may be forced to occur in electrochemical cells by the application of an appropriate potential using an external power source—a process known as electrolysis. Electrolysis is the basis for certain ore refining processes, the industrial production of many chemical commodities, and the electroplating of metal coatings on various products. Measurement of the current flow during electrolysis permits stoichiometric calculations.

This page titled 18.10: Summary is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **17.10: Summary** is licensed CC BY 4.0.

18.11: Exercises

These are homework exercises to accompany the Textmap created for "General Chemistry: Principles and Modern Applications " by Petrucci et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Chemical reactions can be used to produce electricity. This electricity can be used in several different ways, and is applied in a variety of methods in our everyday lives. In studying theoretical cells and reactions, we can get a better understanding of the flow of electrons. The following practice problems are to assist in your mastery of the topic of Electrochemistry.

Q19.1A

From the following observations, estimate the E° from the half reaction $M^+(aq) + e^- \rightarrow M(s)$:

a. The metal M reacts with $H_2SO_4(aq)$; but not with $HI(aq)$; M displaces $Au^+(aq)$, but not $Fe^{3+}(aq)$.
b. The metal M reacts with $HI(aq)$, producing $H_2(g)$, but displaces neither $Al^{3+}(aq)$ nor $Na^+(aq)$.

S19.1A

a. If the metal dissolves in H_2SO_4, it has a reduction potential that is smaller than $E^\circ_{SO_4^{2-}(aq)/SO_2(g)} = 0.17\ V$. If it does not dissolve in HI, it has a reduction potential that is larger than $E^\circ_{H^+(aq)/H_2(g)} = 0\ V$. If it displaces $Au^+(aq)$ from solution, then it has a reduction potential smaller than $E^\circ_{Au^+(aq)/Au(s)} = 1.68\ V$. But if it does not displace $Fe^{3+}(aq)$ from solution, then its reduction potential is larger than

$$E^\circ_{Fe^{3+}(aq)/Fe^{2+}(s)} = 0.769\ V.$$

Therefore, $0\ V < E^\circ < 0.17\ V$.

b. If the metal dissolves in $HI(aq)$, it has a reduction potential that is smaller than $E^\circ_{H^+(aq)/H_2(g)} = 0\ V$. If it does not displace $Al^{3+}(aq)$ from solution, its reduction potential is larger than $E^\circ_{Al^{3+}(aq)/Al(s)} = -1.676\ V$. If it also does not displace $Na^+(aq)$ from solution, its reduction potential is larger than $E^\circ_{Na^+(aq)/Na(s)} = -2.7144\ V$. Therefore, $-1.7676\ V < E^\circ < 0\ V$.

Q19.1B

Estimate the E° for half reaction $M^{2+}(aq) + 2e^- \rightarrow M(s)$. Metal M reacts with $HF(aq)$, but not with $HCl(aq)$; M displaces $Fe^{3+}(aq)$, but not $Sn^{4+}(aq)$.

S19.1B

$+0.13\ V < E^\circ < +0.77\ V$

When M react (including displace) with any substance, its E° is less than that substance's E° and when it doesn't react, its E° is more than that substance's. Thus the answer is: $+0.13\ V < E^\circ < +0.77\ V$.

Q19.1C

For the half reaction, $M^{2+}(aq) + 2e^- \rightarrow M(s)$, what range of E_0 is possible if:

M displaces $MnO_4^-(aq)$, but not $AgCl(s)$. M is unable to react with HCl, but reacts with HNO_3.

S19.1C

Since the reaction can displace $MnO_4^-(aq)$, it would be able to have a standard reaction up to +0.56. However, since it cannot displace $AgCl(s)$, its standard reaction potential would be above +0.2223.

Therefore:
$+0.2223\ V < E_0 < +0.560\ V$

Q19.1D

From the observation listed, estimate the value of E° for the half reaction.
$M^{2+}(aq) + 2e^{2-} \rightarrow M(s)$

a. The metal M reacts with HNO_3, and M can replace Ag^+ but not Cu^{2+}.
b. The metal reacts with HNO_3 and it can not produce H_2; however, it can interact with Fe.

S19.1D

a. $0.223 < x < 0.340$
b. $-0.440 < x < 0$

Q19.1E

Using the table of Standard Reduction Potential and the following clues, determine the metal M.

a. M will be oxidized by $F_2(g)$, but will be reduced by $Fe^{2+}(aq)$
b. M will reduce $Br_2(l)$, but will oxidize $I_2(s)$

S19.1E

a. Because M will be oxidized by $F_2(g)$ and reduced by $Fe^{3+}(aq)$ the reduction potential will be between 2.866 and 0.771.
b. Because M will reduce $Br_2(l)$, the reduction potential of M must be less then 1.065. M will have a reduction potential of 1.065 to 0.077.

Q19.1F

From the observations listed, estimate the value of E° for the imaginary half reaction $M^{2+} + 2e^- \rightarrow M(s)$, where M is an unknown metal to be determined under the following conditions:

a. The metal M reacts with $HCl(aq)$.
b. The metal displaces Fe^{3+} but does not displace Sn^{4+}.
c. The metal reacts with $HNO_3(aq)$.
d. The metal can displace $K^+(aq)$.

S19.1F

a. The E° value will be estimated to be less than 0. In Table P2, metals listed below hydrogen (Pb through Li) should react with HCl. Metals with an E° above hydrogen will not. These metals must be able to displace $H_2(g)$ from acidic solutions.
b. The E° value will be between +0.771 and +0.154. Being displaced means that the ion is forced out of solution and into metal state. The metal must be above Sn^{4+} indicating it is a stronger reducing agent (more willing to give up electrons) and below Fe^{3+}.
c. The E° value must be lower than 0.956, the E° for NO_3. This follows the same reasoning in 1a, but this time the metal must be able to displace $NO(g)$ from acidic solution. Thus, the predicted E° value will be lower than 0.956.
d. The E° must be lower than -2.924, meaning that the metal M will get oxidized and will push the K^+ ion into solution. Thus, the E° will be lower than -2.924.

For review on this topic, visit the page ***"Electrochemistry 1: Introduction"***.

Q19.2A

A Lithium-ion battery has a theoretical E_0 cell of 1.5 V. Find the E_0 for the reduction half-reaction:

$$Li^+(aq) + e^- \rightarrow Li(s)$$

S19.2A

The E° for the reaction $Li^+(aq) + e^- \rightarrow Li(s)$ equals -3.040 according to the Standard Electrode Potential Table.

Q19.2B

Use the table of Standard Reduction Potential to predict whether these reactions will happen spontaneously.

a. $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$
b. $Fe^{3+}(aq) + Na(s) \rightarrow Fe^{2+}(aq) + Na^+(aq)$
c. $Zn^{2+}(aq) + 2I^-(aq) \rightarrow Zn(s) + I_2(l)$

S19.2B

a. The reduction potential of $Ag^+(aq)$ is 0.8
The oxidation potential of $Cu(s)$ is -0.34
$E^\circ_{cell} = E^\circ_{Cat} + E^\circ_{An}$: $0.8 + (-0.34) = 0.46$
E°_{cell} is positive, Yes it will happen spontaneously
b. The reduction potential of $Fe^{3+}(aq)$ is 0.77
The oxidation potential of $Na(s)$ is 2.71

$E^\circ_{cell} = E^\circ_{Cat} + E^\circ_{An} : 0.77 + 2.71 = 3.48$
E°_{cell} is positive, Yes it will happen spontaneously

c. The reduction potential of $Zn^{2+}(aq)$ is -0.763
The oxidation potential of is -0.54
$E^\circ_{cell} = E^\circ_{Cat} + E^\circ_{An} : (-0.763) + (-0.54) = -1.303$
E°_{cell} is Negative, No it will not happen spontaneously

Q19.3A

Write the voltaic cell for the reaction between $I_2(s)$ and $Sn^{4+}(aq)$. Determine which half reaction is the cathode and the anode and calculate the reaction standard potential.

Assuming all the reactants and products are in their standard state, use Table P2 to predict which reactions are spontaneous (in the forward direction).

a. $Zn^{2+} + Cu(s) \rightarrow Cu^{2+} + Zn(s)$
b. $Fe^{3+}(aq) + Cu(s) \rightarrow Fe^{2+} + Cu^{2+}(aq)$
c. $2Fe^{2+}(aq) + Cl_2(g) \rightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$
d. $Cu^{2+} + e^- \rightarrow Cu^{2+}$

S19.3A

a. Half reaction:

Reduction: $Zn^{2+} + Cu(s) \rightarrow Cu^{2+} + Zn(s) \quad E^\circ = -0.763\ V$

Oxidation: $Cu(s) \rightarrow Cu^{2+} + 2e^- \quad E^\circ = -0.340\ V$

$$E^\circ_{cell} = E_{cathode} - E_{anode} \tag{18.11.1}$$
$$= -0.763 - 0.340 \tag{18.11.2}$$
$$= -1.103\ V\ (\text{Spontaneous}) \tag{18.11.3}$$

b. Half reaction:

Reduction: $Fe^{3+} + e^- \rightarrow Fe^{2+} \quad E = 0.771\ V$

Oxidation: $\mCathrm{Cu}(s) \rightarrow Cu^{2+} + 2e^- \quad E = 0.340\ V$

$$E^\circ_{cell} = E_{cathode} - E_{anode} \tag{18.11.4}$$
$$= 0.771 - 0.340 \tag{18.11.5}$$
$$= 0.431\ V\ (N.S) \tag{18.11.6}$$

c. Half reaction:

Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^- \quad E = +1.358\ V$

Oxidation: $2Fe^{2+} \rightarrow 2Fe^{3+} + 2e^-$

$$E^\circ_{cell} = E_{cathode} - E_{anode} \tag{18.11.7}$$
$$= 1.358 - 0.771\ V \tag{18.11.8}$$
$$= 0.587\ V\ (N.S) \tag{18.11.9}$$

d. It's already half reaction

Reduction: $Cu^{2+} + e^- \rightarrow Cu^+ \quad E = +0.159\ V\ (N.S)$

S19.3B

$I_2(s) + 2e^- \rightarrow 2I^-(aq)$
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$

$E^\circ : +0.535\ V$ Reduction $\Rightarrow$ Cathode
$E^\circ : +0.154\ V$ Oxidation $\Rightarrow$ Anode

$\Rightarrow Pt(s)|Sn^{2+}(aq),\ Sn^{4+}(aq)||I_2(s)|2I^-(aq)$

Note: $Pt(s)$ serves as an electrode conductor and does not affect the overall reaction. Since both the Sn reactant and product are both aqueous solutions, Pt or other metal conductors can be used in place.

$$E^\circ = E^\circ_{(\text{cathode})} - E^\circ_{(\text{anode})}$$
$$E^\circ = (+0.535\ \text{V}) - (+0.154\ \text{V})$$
$$E^\circ = +0.0381\ \text{V}$$

Q19.3B

Write cell reactions for the electrochemical cells diagrammed here, and use data from Table P2 to calculate E°_{cell} for each reaction.

a. $\text{Al(s)}|\text{Al}^{3+}||\text{Zn}^{2+}\text{(aq)}|\text{Zn(s)}$
b. $\text{Pt(s)}|\text{Fe}^{2+}\text{(aq)}, \text{Fe}^{3+}\text{(aq)}||\text{Cu}^{2+}\text{(aq)}|\text{Cu(s)}$
c. $\text{Pt(s)}|\text{Cr}_2\text{O}_7, \text{Cr}^{3+}\text{(aq)}||\text{Ag}^{+}\text{(aq)}|\text{Ag(s)}$
d. $\text{O}_2^{-}\text{(aq)}|\text{O}_2\text{(g)}||\text{H}^{+}\text{(aq)}|\text{H}_2\text{(g)}|\text{C(s)}$

S19.3B

a. Half reaction:

Reduction: $\text{Al(s)} + 3\text{e}^- \rightarrow \text{Al}^{3+} \quad E = -2.310\ \text{V}$

Oxidation: $\text{Zn}^{2+} \rightarrow \text{Zn(s)} + 2\text{e}^- \quad E = -0.763\ \text{V}$

$$\begin{aligned} E &= E_{\text{cathode}} - E_{\text{anode}} && (18.11.10)\\ &= -2.310 - -0.763\ \text{V} && (18.11.11)\\ &= 1.447\ \text{V} && (18.11.12) \end{aligned}$$

b. Oxidation: $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+} \quad E = 0.771\ \text{V}$

Reduction: $\text{Cu(s)} \rightarrow \text{Cu}^{2+} + 2\text{e}^- \quad E = 0.520\ \text{V}$

$$\begin{aligned} E &= E_{\text{cathode}} - E_{\text{anode}} && (18.11.13)\\ &= 0.520 - 0.771\ \text{V} && (18.11.14)\\ &= 0.149\ \text{V} && (18.11.15) \end{aligned}$$

c. $\text{Cr}^{6+} + 3\text{e}^- \rightarrow \text{Cr}^{3+} \quad E = 1.33\ \text{V}$

$\text{Ag(s)} \rightarrow \text{Ag}^{+} + \text{e}^- \quad E = 0.800\ \text{V}$

$$\begin{aligned} E &= E_{\text{cathode}} - E_{\text{anode}} && (18.11.16)\\ &= 0.800\ \text{V} - 1.33\ \text{V} && (18.11.17)\\ &= -0.53\ \text{V} && (18.11.18) \end{aligned}$$

d. $\text{O}_2 + 4\text{H}^{+} + 4\text{e}^- \rightarrow 2\text{H}_2\text{O(l)} \quad E = 1.229\ \text{V}$

$\text{H}_2 \rightarrow 2\text{H}^{+} + \text{e}^- \quad E = 0\ \text{V}$

$$\begin{aligned} E &= E_{\text{cathode}} - E_{\text{anode}} && (18.11.19)\\ &= 0 - 1.229\ \text{V} && (18.11.20)\\ &= -1.229\ \text{V} && (18.11.21) \end{aligned}$$

Q19.3C

For the electrochemical cells shown below, write the cell reactions and find E°_{cell} using the Standard Reduction Potential table

a. $\text{Cu(s)}|\text{Cu}^{2+}\text{(aq)}||\text{Ag}^{+}\text{(aq)}|\text{Ag(s)}$
b. $\text{Cu(s)}|\text{Cu}^{2+}\text{(aq)}||\text{Fe}^{2+}\text{(aq)}, \text{Fe}^{3+}\text{(aq)}|\text{Pt(s)}$
c. $\text{PbO}_2\text{(s)}|\text{Pb}^{2+}\text{(aq)}||\text{Mn}^{2+}\text{(aq)}|\text{MnO}_2\text{(s)}$
d. $\text{Cu(s)}|\text{Cu}^{2+}\text{(aq)}||\text{O}_2\text{(g)}|\text{H}_2\text{O}_2\text{(aq)}|\text{Pt(s)}$

S19.3C

a. $2\,\text{Ag}^{+}\text{(aq)} + \text{Cu(s)} \rightarrow 2\,\text{Ag(s)} + \text{Cu}^{2+}\text{(aq)}$
The reduction potential of $\text{Ag}^{+}\text{(aq)}$ is 0.8
The oxidation potential of Cu(s) is -0.34
$E^\circ_{\text{cell}} = E^\circ_{\text{Cat}} + E^\circ_{\text{An}} : 0.8 + (-0.34) = 0.46$

b. $2\,\text{Fe}^{3+}\text{(aq)} + \text{Cu(s)} \rightarrow 2\,\text{Fe}^{2+}\text{(aq)} + \text{Cu}^{2+}\text{(aq)}$
The reduction potential of $2\,\text{Fe}^{3+}\text{(aq)}$ is 0.77
The oxidation potential of Cu(s) is -0.34
$E^\circ_{\text{cell}} = E^\circ_{\text{Cat}} + E^\circ_{\text{An}} : 0.77 + (-0.34) = 0.43$

c. $PbO_2(s) + Mn^{2+}(aq) \rightarrow Pb^{2+}(aq) + MnO_2(s)$
The reduction potental of $PbO_2(s)$ is 1.455
The oxidation potental of $Mn^{2+}(aq)$ is -1.229
$E^\circ_{cell} = E^\circ_{Cat} + E^\circ_{An} : 1.455 + (-1.229) = 0.43$
d. $Cu(s) + O_2(g) + 2\,H_3O(aq) \rightarrow H_2O_2(aq) + Cu^{2+}(aq)$
The reduction potential of $O_2(g)$ is 0.695
The oxidation potential of $Cu(s)$ is -0.34
$E^\circ_{cell} = E^\circ_{Cat} + E^\circ_{An} : 0.695 + (-0.34) = 0.355$

Q19.4A

Given the reaction:

$$2MnO_4^-(aq) + PbO(s) + 6H^+(aq) \rightarrow 2MnO_2(s) + 5PbO_2(s) + 3H_2O(l)$$

$E^\circ_{cell} : +1.23\ V$

What does $E^\circ\ PbO_2(s)$ equal?

S19.4A

$MnO_4^- \rightarrow MnO_2(s)$

$PbO(s) + 6H^+(aq) \rightarrow 2MnO_2(s) + 3H_2O(l)$

Mn oxidation number:
Reactant: +7 Product: +4
$\Rightarrow$ Mn is the reduction reaction: Mn is the cathode
$\Rightarrow E^\circ = 1.51$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
$1.23 = 1.51 - E^\circ_{anode}$
$E^\circ_{anode} = +0.28\ V$

Q19.4B

Use thermal dynamic data from Appendix D to calculate a theoretical voltage of Electrolysis cell
$Cu(s) + Zn^{2+}(aq) \rightarrow Cu^{2+}(aq) + Zn(s) \quad E^\circ_{cell} = -1.103\ V$

S19.4B

The electrolysis is not thermodynamically favored.

$E_{cell} = E_{Zn^{2+}/Zn} - E_{Cu^{2+}/Cu^+} = -0.763 - 0.340\ V = -1.103\ V$

Q19.4C

Sketch a voltaic cell for each of the following conditions, labeling the anode, cathode, and electron flow. Balance the redox equation if necessary. Calculate E°_{cell} for each cell.

a. $Cu^{2+}(aq) + Sn^{2+}(aq) \rightarrow Cu(s) + Sn^{4+}(aq)$
b. $Mg(s)$ displaces $Sn^{2+}(aq)$ from solution
c. $Zn(s)$ donates e^- to acidic solution to form $H_2(g)$
d. $Pb^{2+}(aq) + Sn(s) \rightarrow Pb(s) + Sn^{2+}(aq)$

Q19.5A

Determine whether a forward reaction is spontaneous or non-spontaneous:

a. $Ag^{2+}(aq) + Cr^{2+}(aq) \rightarrow Ag^+(aq)$
b. $Sn^{4+}(aq) + 2I^-(aq) \rightarrow Sn^{2+}(aq) + I_2(aq)$

S19.5A

a. $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$
$E^\circ_{cell} = (E^\circ_{Ag^{2+}/Ag^-}) - (E^\circ_{Cr^{3+}/Cr^{2+}})$
$E^\circ_{cell} = 1.98\ V - (-0.424\ V)$

$E^\circ_{cell} = +2.404\ V$
$\Rightarrow$ Spontaneous

b. $Sn^{4+}(aq)$ reduced, $2I^-(aq)$ oxidized

$E^\circ_{cell} = (+0.154\ V) - (+0.535\ V)$
$E^\circ_{cell} = -0.381\ V$
$\Rightarrow$ Non-spontaneous

Q19.5B

A voltaic cell represented by the following cell diagram has $E_{cell} = -0.119\ V$

$$Ag(s)|Ag^+(0.075\ M)||Hg(?\ M)|Hg(l) \tag{18.11.22}$$

S19.5B

Oxidation: $2Ag(s) \rightarrow 2Ag^+(aq) + 2e^-$

Reduction: $Hg(aq) + 2e^- \rightarrow Hg(s)$

Overall: $2Ag(s) + Hg^{2+}(aq) \rightarrow 2Ag^+(aq) + Hg(l)$

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} \tag{18.11.23}$$
$$= 0.845\ V - 0.800\ V \tag{18.11.24}$$
$$= 0.054\ V \tag{18.11.25}$$
$$E^\circ_{cell} = E^\circ_{cathode} - E_{anode} \tag{18.11.26}$$
$$= 0.854\ V - 0.800\ V \tag{18.11.27}$$
$$= 0.054\ V \tag{18.11.28}$$

$$E = E^\circ_{cell} - \frac{0.0592\ V}{2} \log \frac{[Ag^+]^2}{?}$$

$$-0.173\ V = -0.029 \log \frac{[0.005625]}{x}$$

$$5.844 = \log \frac{[0.005625]}{x}$$

$$x = 0.706$$

Q19.5C

Determine ΔG° for the following voltaic cell reactions:

a. $Pb^{2+}(aq) + Sn(s) \rightarrow Pb(s) + Sn^{2+}(aq)$
b. $O_2(g) + 2H^+(aq) + 2F^-(aq) \rightarrow H_2O_2(aq) + F_2(g)$
c. $Br_2(l) + 2Fe^{2+}(aq) \rightarrow 2Br^-(aq) + 2Fe^{3+}(aq)$

Q19.6A

Given the following voltaic cell diagram and using the standard gold electrode as the reduction half reaction, and metal M as an oxidizing half reaction, find the E° reduction half reaction:

$$Au^{3+}(1\ M) + 3e^- \rightarrow Au(s) \qquad E^\circ = +1.52 \tag{18.11.29}$$
$$M^{3+}(1\ M) + 3e^- \qquad E^\circ = ? \tag{18.11.30}$$

a. In $E^\circ = 1.858\ V$
b. Al $E^\circ = 3.20\ V$
c. La $E^\circ = 3.90\ V$
d. U $E^\circ = 3.18\ V$

S19.6A

Au is the reduction half reaction
$E^\circ = +1.52 = E^\circ_{cathode}$
$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

a. In

$$1.858 = +1.52 - E^\circ_{anode}$$
$$1.858 + E^\circ_{anode} = 1.52$$
$$E^\circ_{anode} = -0.338\ V$$

b. Al

$$3.20 = 1.52 - E^\circ_{anode}$$
$$3.20 + E^\circ_{anode} = 1.52$$
$$E^\circ_{anode} = -1.68\ V$$

c. La

$$3.90 = 1.52 - E^\circ_{anode}$$
$$3.90 + E^\circ_{anode} = 1.52$$
$$E^\circ_{anode} = -2.38\ V$$

d. U

$$3.18 = 1.52 - E^\circ_{anode}$$
$$3.18 + E^\circ_{anode} = 1.52$$
$$E^\circ_{anode} = -1.66\ V$$

Q19.6B

In each of the following examples, write the cell diagram and cathode. Indicate the direction of the electron flow; write a balanced equation.

a. $Al(s) + Cu^{2+} \rightarrow Cu(s) + Al^{3+}$
b. $Fe^{3+} + Ag(s) \rightarrow Ag^{+} + Fe^{2+}$
c. $Cu(s) + 2Fe^{3+} \rightarrow Cu^{2+} + 2Fe^{2+}$

S19.6B

a. Half reaction:
Reduction: $Al(s) + 3e^- \rightarrow Al^{3+} \quad E = -2.310\ V$
Oxidation: $Zn^{2+} \rightarrow Zn(s) + 2e^- \quad E = -0.763\ V$

$$E = E_{cathode} - E_{anode} \tag{18.11.31}$$
$$= -2.310 - -0.763\ V \tag{18.11.32}$$
$$= 1.447\ V \tag{18.11.33}$$

b. Oxidation: $Fe^{3+} + e^- \rightarrow Fe^{2+} \quad E = 0.771\ V$
Reduction: $Cu(s) \rightarrow Cu^{2+} + 2e^- \quad E = 0.520\ V$

$$E = E_{cathode} - E_{anode} \tag{18.11.34}$$
$$= 0.520 - 0.771\ V \tag{18.11.35}$$
$$= 0.149\ V \tag{18.11.36}$$

c. Oxidation: $Cu(s) \rightarrow Cu^{2+} + 2e^- \quad E = 0.340\ V$
$Fe^{3+} + e^- \rightarrow Fe^{2+} \quad E = 0.771\ V$

$$E = E_{cathode} - E_{anode} \tag{18.11.37}$$
$$= 0.771\ V - 0.340\ V \tag{18.11.38}$$
$$= 0.431\ V \tag{18.11.39}$$

Q19.7A

Calculate ΔG° for the following reactions and determine if spontaneous:

a. $NO_3^-(aq) + Al(s) + 4H^+(aq) \rightarrow NO(g) + Al^{3+}(aq) + 2H_2O(l)$
b. $F_2(g) + 2Li(s) \rightarrow 2F^-(aq) + 2Li^+(aq)$

S19.7A

a. $\Delta G^\circ = -nFE^\circ_{cell}$

$$E^\circ_{cell} = (E^\circ_{NO_3^-/NO^-}) - (E^\circ_{Al^{3+}/Al})$$
$$E^\circ_{cell} = (0.956) - (-1.676)$$
$$E^\circ_{cell} = 2.63\ V$$

$N = 3$ since:

N oxidation: reactant = +5 product= +2
Al oxidation: reactant = 0 product= +3

$F = 96,485\ C/mol\ e^-$

$\Delta G^\circ = -[(3)(96,485\ C/mol\ e^-)(2.63\ V)]$
$\Delta G^\circ = -761\ kJ$ Spontaneous

b. $E^\circ_{cell} = (E^\circ_{F_2/F^-}) - (E^\circ_{Li^+/Li})$

$E^\circ_{cell} = 2.866 + 3.040 = 5.906\ V$

$N = 2$ since:
$F_2 + 2e^- \rightarrow 2F^-$
$2Li^+ + 2e^- \rightarrow 2Li(s)$

$\Delta G^\circ = -[(2\ mol\ e^-)(96,485\ C/mol\ e^-)(5.906\ V)]$
$\Delta G^\circ = -1,140\ kJ$
Spontaneous

Q19.7B

Determine the value of the ΔG° for the following reactions carried out in voltaic cells. [Use the equations of 22]

a. $Al(s) + Cu^{2+} \rightarrow Cu(s) + Al^{3+}$
b. $Fe^{3+} + Ag(s) \rightarrow Ag^+ + Fe^{2+}$
c. $Cu(s) + 2Fe^{3+} \rightarrow Cu^{2+} + 2Fe^{2+}(aq)$

S19.7B

To solve the problem, you have first write the balanced equation.

a. $\Delta G^\circ = -nFE^\circ_{cell}$
$= -6\ mol * 96485\ C\ mol^{-1} * 2.106\ V$
$= -1.2 * 10^6\ J$

b. $\Delta G^\circ = -nFE^\circ_{cell}$
$= -1\ mol * 96485\ C\ mol^{-1} * 0.029\ V$
$= -2605.095\ V$

c. $\Delta G^\circ = -nFE^\circ_{cell}$
$= -2\ mol * 96485\ C\ mol^{-1} * 0.029\ V$
$= -83170.07\ J$

Q19.7

Write and balance the net reaction using the two following half reactions. Find the Voltage for this reaction

$BrO^-(aq) + H_2O(l) + 2e^- \rightarrow Br^-(aq) + 2OH^-(aq)$

$Mg(s) + 2OH^-(aq) \rightarrow Mg(OH)_2(s) + 2e^-$

Q19.8A

Find ΔG° by combining the following half reactions to calculate the E° cell at:

$$Ag^{2+} + 2e^- \rightarrow Ag(s) \tag{18.11.40}$$
$$Ag^{2+}(aq) + e^- \rightarrow Ag^+(aq) \qquad E^\circ = +1.98\ V \tag{18.11.41}$$
$$Ag^+(aq) + e^- \rightarrow Ag(s) \qquad E^\circ = +0.800\ V \tag{18.11.42}$$

S19.8A

$$Ag^{2+}(aq) + e^- \rightarrow Ag^+(aq) \qquad \Delta G^\circ = -(1 \times F \times (1.98\ V)) \tag{18.11.43}$$
$$\underline{Ag^+(aq) + e^- \rightarrow Ag(s) \qquad \Delta G^\circ = -(1 \times F \times (0.800\ V))} \tag{18.11.44}$$
$$Ag^{2+}(aq) + 2e^- \rightarrow Ag(s) \qquad \Delta G^\circ = -(1.98\ V)F - (0.800\ V)F \tag{18.11.45}$$
$$\Delta G^\circ = (-2.78\ V)F \tag{18.11.46}$$

$\Delta G^\circ = -nFE^\circ_{cell}$
$(-2.78\ V)F = -2FE^\circ_{cell}$
$E^\circ_{cell} = \dfrac{(-2.78\ V)F}{-2F}$
$E^\circ_{cell} = +1.39\ V$

Q19.8B

For the reaction

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Ag(s) \rightarrow 2Cr^{3+} + 6Ag^+(aq) + 7H_2O(l) \tag{18.11.47}$$

Use the data table to determine

a. E°_{cell}
b. K
c. ΔG°
d. Whether the reaction goes substantially to completion when the reactants and products are initially in their standard state.

S19.8B

a. Half reactions:

Cathode: $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + H_2O \quad E = +1.33\ V$

Anode: $Ag(s) \rightarrow Ag^+ + e^- \quad E = 0.800\ V$

$$E_{cell} = E_{cathode} - E_{anode} \tag{18.11.48}$$
$$= 1.33 - 0.800\ V \tag{18.11.49}$$
$$= 0.53\ V \tag{18.11.50}$$

b. Standard conditions:

$$\Delta G^\circ = -nFE^\circ_{cell} \tag{18.11.51}$$
$$= -6\ mol * 96485\ C\ mol^{-1} * 0.53\ V \tag{18.11.52}$$
$$= -30682.23\ V \tag{18.11.53}$$

c. $E_{cell} = -\dfrac{RT}{nF}\ln K$

$K = e^{(-nFE_{cell})/(RT)}$

$K = 6.43 * 10^{54}$

d. The reaction can be spontaneous. ΔG° **is very negative.**

Q19.9A

A voltaic cell with $E_{cell} = 1.500\ V$ has what $[Ag^+]$ in the cell?

$Zn(s)|Zn^{2+}(1.50\ M)||Ag^+(?\ M)|Ag(s)$

S19.9A

Solve for E_{cell}:

Oxidation reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \quad E^\circ = -0.763$

Reduction reaction: $(Ag^+(aq) + e^- \rightarrow Ag(s)) \times 2 \quad E^\circ = 0.800\ V$

Net Ionic Equation: $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s) \quad E^\circ_{cell} = 1.563\ V$

Use Nernst equation to solve for $[Ag^+]$:

$E = E^\circ_{cell} - \dfrac{0.0592}{2}\log\dfrac{[Zn^{2+}]}{[Ag^+]^2}$

$1.500\ V = 1.563 - \dfrac{0.0592}{2}\log\dfrac{1.5\ M}{x^2}$

$\log\dfrac{1.5\ M}{x^2} = \dfrac{-2(1.500 - 1.563)}{0.0592}$

$$\log \frac{1.5\,\mathrm{M}}{x^2} = 2.13$$

$x = 0.4222$

$[Ag^+] = 0.4222\,\mathrm{M}$

Q19.9B

Calculate E_{cell} using the Nernst equation for the following cells.

a. $Zn(s)|Zn^{2+}(aq)(0.1\,M)||Sn^{2+}(aq)(.8\,M)|Sn(s)$
b. $Cu(s)|Cu^{+}(aq)(0.4\,M)||F^{-}(aq)(O.9\,M)|F_2(g)(0.5\,atm)|Pt(s)$

S19.9B

Step 1: Make half reactions and determine the oxidation and reduction equation. (Make sure equation is balanced!)

Step 2: Find the $E°$ for each half reactions using the standard reduction potential table.

Step 3: Add the $E°$ of both half reactions together to find the $E°_{cell}$.

Step 4: Determine the number of moles of electrons transferred (n).

Step 5: Calculate the reaction quotient (Q). $Q = \frac{[Products]}{[Reactants]}$

Step 6: Calculate the E_{cell} (Plug and chug). $E_{cell} = E°_{cell} - \frac{0.0592V}{n}\log Q$ at 298K.

a) $Zn(s)|Zn^{2+}(aq)(0.1\,M)||Sn^{2+}(aq)(.8\,M)|Sn(s)$

ox: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ $E°$: -0.763

Red: $Sn^{2+}_{(aq)} + 2e^- \rightarrow Sn_{(s)}$ $E°$: -0.137

$E°_{cell}$: -0.9

n=2

$Q = \frac{[Zn^{2+}]}{[Sn^{2+}]} = \frac{0.1}{0.8} = 0.125$

$E_{cell} = E°_{cell} - \frac{0.0592V}{n}\log Q = -0.9 - \frac{0.0592V}{n}\log(0.125)$

E_{cell} = -0.8733V

b) $Cu(s)|Cu^{+}(aq)(0.4\,M)||F^{-}(aq)(O.9\,M)|F_2(g)(0.5\,atm)|Pt(s)$

$Cu_{(s)} \rightarrow Cu^{+}_{(aq)} + e^-$ $E°$ = -0.340

$2F^-_{(aq)} \rightarrow F_{2(g)} + 2e^-$ $E°$ = -2.866

$E°_{cell}$ = 2.186

n=2

$Q = \frac{[Cu^{2+}]}{[F^-]} = \frac{0.4}{0.9} = 0.4444$

$E_{cell} = E°_{cell} - \frac{0.0592V}{n}\log Q$

$= 2.186 - \frac{0.0592V}{n}\log(0.4444)$

E_{cell} = 2.196V

Q19.10A

$[Zn^{2+}] = 1.2\,\mathrm{M}$

What $[Cu^{2+}]$ will allow a forward spontaneous reaction?

Net Reaction: $Zn(s) + Cu^{2+} \rightarrow Cu(s) + Zn^{2+} \quad E^{\circ}_{cell} = 1.100\ V$

S19.10A

Nernst equation setting $E = 0.000\ V$ & $[Zn^{2+}] = 1.0\ M$

$$E = E^{\circ}_{cell} - \frac{0.0592}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$0.00\ V = 1.100 - \frac{0.0592}{2} \log \frac{1.2\ M}{[Cu^{2+}]}$$

$$\frac{0.00 - 1.100}{-0.0296} = \log \frac{1.2\ M}{[Cu^{2+}]}$$

$$37.2 = \log \frac{1.2\ M}{[Cu^{2+}]}$$

$$[Cu^{2+}] = 8.4 \times 10^{-17}\ M$$

Q19.10B

$Pb(s)|Pb^{2+}(aq)(0.07\ M)||Fe^{3+}(aq)(0.64\ M), Fe^{2+}(aq)(0.85\ M)|Pt(s)$

a. Will the cell reaction be spontaneous?

b. What ratio of $\frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$ will reaction will not be spontaneous in either direction?

S19.10B

a. We are asked about the spontaneity of the reaction, which is determined by cell potentials.

We can get the standard cell potential of Pb and Fe from the chart.

Oxidation: $Pb(s) \rightarrow Pb^{2+}_{(aq)} + 2e^-$ $E^{\circ}_{cell} = -0.125V$

Reduction: $Fe^{3+}_{(aq)} + e^- \rightarrow Fe^{2+}_{(aq)}$ $E^{\circ}_{cell} = +0.771V$

Then we can get the differences between the standard cell potentials by using the equation:

$$E^{\circ}_{cell} = E^{\circ}_{reduction} - E^{\circ}_{oxidation}$$

$$E_{cell} = +0.771V - (-0.125V) = 0.896V$$

Then we use the Nerst Equation to get the cell potential in this reaction:

$$E_{cell} = E^{\circ}_{cell} - 0.0592/2 \cdot \log [Pb^{2+}][Fe^{2+}]^2/[Fe^{3+}]^2$$

$$= 0.896V - 0.0596/2 \cdot \log [0.07M][0.85M]^2/[0.64M]^2$$

$$= 0.896V$$

Thus the cell potential is positive, which indicates that the reaction is positive.

b. If the reaction is not spontaneous, then E_{cell} must less than 0.

Therefore, we can change the Nerst Equation into this form:

$$0 < E^{\circ}_{cell} - 0.0592/n \cdot \log [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

$$-E^{\circ}_{cell} < -0.0592/n \cdot \log [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

$$E^{\circ}_{cell} < 0.0592/2 \cdot \log [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

$$1250/37(E^{o}_{cell}) < \log [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

$$10^{(1250/37)(E_{cell})} < [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

$$10^{(1250/37)(0.869V)} < [Pb^{2+}][Fe^{2+}]^2 / [Fe^{3+}]^2$$

Thus the ratio that the reaction is not spontanoues is $[Pb^{2+}][Fe^{3+}]^2 / [Fe^{3+}]^2 > 10^{(1250/37)(0.869V)}$ or precisely $[Pb^{2+}][Fe^{3+}]^2 / [Fe^{3+}]^2 > 2.281 * 10^{29}$

a. Oxidation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^- \qquad E^{\circ}_{cell} = -0.125\ V$

Reduction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq) \quad E^{\circ}_{cell} = +0.771\ V$

$E^{\circ}_{cell} = +0.771\ V - (-0.125\ V) = 0.896\ V$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \log \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2} \tag{18.11.54}$$

$$= 0.896\ V - \frac{0.0592}{2} \log \frac{(0.07\ M)(0.85\ M)^2}{(0.64\ M)^2} \tag{18.11.55}$$

$$= 0.869\ V \tag{18.11.56}$$

Thus since E_{cell} is positive, the reaction is spontaneous.

b. Ratio of $\frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$ for reaction to not be spontaneous is when:

E_{cell} is less than 0. Thus creating:

$$0 < E^{\circ}_{cell} - \frac{0.0592}{n} \log \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$-E^{\circ}_{cell} < -\frac{0.0592}{n} \log \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$E^{\circ}_{cell} < \frac{0.0592}{2} \log \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$\frac{1250}{37}(E^{\circ}_{cell}) < \log \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$10^{(1250/37)(E^{\circ}_{cell})} < \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$10^{(1250/37)(0.869\ V)} < \frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

Or precisely for this cell reaction, ratio $\frac{[Pb^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2} > 2.281 \times 10^{29}$.

Q19.10A

a. $Sn(s)|Sn^{2+}(0.080\ M)||Pb^{2+}(0.700\ M)|Pb(s)$

Given the following cell reaction, will the cell proceed spontaneously as written?

b. For what ratio of $\frac{[Sn^{2+}]}{[Pb^{2+}]}$ for the reaction above in a) will the cell not be spontaneous in either direction?

S19.10A

a. Find the E°_{cell} first by $E^{\circ}_{(reduction)} - E^{\circ}_{(oxidation)}$.

Oxidation: $Sn(s) \rightarrow Sn^{2+}(0.080\ M) + 2e^-$ with an E°_{cell} of 0.137 V

Reduction: $Pb^{2+}(0.700\ M) + 2e^- \rightarrow Pb(s)$ with an E°_{cell} of -0.125

Thus the overall E°_{cell} is 0.012 V. But this cell has specific concentrations given, so we know to use the Nernst equation to find the exact E°_{cell}.

Thus, the Nernst equation sets up as $E_{cell} = 0.012\ V - \frac{0.0592}{2} \times \log \frac{[Sn^{2+}]}{[Pb^{2+}]}$

$$X = 0.012\ V - 0.0296 \log \frac{[0.080\ M]}{[0.700\ M]}$$

$$X = 0.0399\ V$$

Thus the E_{cell} is positive, greater than 0, and will move spontaneously in the forward direction.

For review on this topic, visit the page ***"Electrochemistry 4: The Nernst Equation"***.

b. The E_{cell} must be 0. In order for the cell not to be spontaneous in either direction, the cell must be at equilibrium.

Thus, we set up our Nernst equation again, this time designating the entire Q as a variable (this is the ratio we are seeking) and set our E_{cell} to 0.

This results in the following equation:

$$0 = 0.012\ V - \frac{0.0592}{2} \log[x]$$

$$-0.012\ \text{V} = -0.0296 \log[\text{x}]$$

$$2.54335 = \text{x}$$

Thus the ratio must be 2.54 to ensure the cell is in equilibrium.

For any additional information, consult "Electrochemistry".

Q19.11A

A voltaic cell:

$\text{Ag(s)}|\text{Ag}^+(\text{saturated } \text{Ag}_2\text{CrO}_4)||\text{Ag}^+(0.125\ \text{M})|\text{Ag(s)}$

What is E_{cell} if $\text{K}_{\text{sp}} = 1.2 \times 10^{-11}$ for Ag_2CrO_4?

S19.11B

First must solve for $[\text{Ag}^+]$ in saturated solution

$$\text{K}_{\text{sp}} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] = (2\text{s})^2(\text{s}) = 4\text{s}^3$$

$$\text{S} = \sqrt[3]{\frac{1.2 \times 10^{-11}}{4}} = 1.44 \times 10^{-4}\ \text{M}$$

$$[\text{Ag}^+]_{\text{anode}} = 2\text{s} = 2.88 \times 10^{-4}\ \text{M}$$

$$\text{Ag(s)} + \text{Ag}^+(0.125\ \text{M}) \rightarrow \text{Ag(s)} + \text{Ag}^+(2.88 \times 10^{-4}\ \text{M})$$

$$\text{E}_{\text{cell}} = \text{E}^\circ_{\text{cell}} - \frac{0.0592}{2} \log \frac{2.88 \times 10^{-4}\ \text{M}}{0.125\ \text{M}}$$

$$\text{E}_{\text{cell}} = 0.000 + 0.078\ \text{V} = 0.078\ \text{V}$$

Q19.11B

Predict whether spontaneous reaction in forward direction will occur at 298K for reaction:

a. $\text{Cu(s)} + \text{Br}_2\text{(l)} \rightarrow \text{Cu}^{2+}\text{(aq)} + 2\text{Br}^-\text{(aq)}$
b. $\text{Sn}^{2+}\text{(aq)} + \text{Pb}^{2+}\text{(aq)} \rightarrow \text{Pb(s)} + \text{Sn}^{4+}\text{(aq)}$
c. $2\text{Ag(s)} + 2\text{H}^+\text{(aq)} \rightarrow 2\text{Ag}^+\text{(aq)} + \text{H}_2\text{(g)}$
d. $2\text{Cl}_2\text{(g)} + 2\text{H}_2\text{O(l)} \rightarrow 4\text{HCl(aq)} + \text{O}_2\text{(g)}$

S19.11B

Formula: $\text{E}^\circ_{\text{cell}} = \text{E}^\circ_{\text{cathode}} - \text{E}^\circ_{\text{anode}}$

Oxidation occurs at anode. Reduction occurs at cathode.

Positive $\text{E}^\circ_{\text{cell}}$ = spontaneous reaction at temperature given.

a. Anode: $\text{Cu(s)} \rightarrow \text{Cu}^{2+}\text{(aq)} + 2\text{e}^-$ $\quad \text{E}^\circ_{\text{cell}} = +0.340\ \text{V}$

Cathode: $\text{Br}_2\text{(l)} + 2\text{e}^- \rightarrow 2\text{Br}^-\text{(aq)}$ $\quad \text{E}^\circ_{\text{cell}} = +1.065\ \text{V}$

$\text{E}^\circ_{\text{cell}} = +1.065\ \text{V} - 0.340\ \text{V} = +0.725\ \text{V}$. Spontaneous reaction at 298K.

b. Anode: $\text{Sn}^{2+}\text{(aq)} \rightarrow \text{Sn}^{4+}\text{(aq)} + 2\text{e}^-$ $\quad \text{E}^\circ_{\text{cell}} = +0.154\ \text{V}$

Cathode: $\text{Pb}^{2+}\text{(aq)} + 2\text{e}^- \rightarrow \text{Pb(s)}$ $\quad \text{E}^\circ_{\text{cell}} = -0.125\ \text{V}$

$\text{E}^\circ_{\text{cell}} = -0.125\ \text{V} - 0.154\ \text{V} = -0.279\ \text{V}$. Non-spontaneous at 298K.

c. Anode: $2\text{Ag(s)} \rightarrow 2\text{Ag}^+\text{(aq)} + 2\text{e}^-$ $\quad \text{E}^\circ_{\text{cell}} = +0.800\ \text{V}$

Cathode: $2\text{H}^+\text{(aq)} + 2\text{e}^- \rightarrow \text{H}_2\text{(g)}$ $\quad \text{E}^\circ_{\text{cell}} = 0.000\ \text{V}$

$\text{E}^\circ_{\text{cell}} = 0.000\ \text{V} - 0.800\ \text{V} = -0.800\ \text{V}$. Non-spontaneous at 298K.

d. Anode: $2\text{H}_2\text{O(l)} \rightarrow 4\text{H}^+\text{(aq)} + \text{O}_2\text{(g)} + 4\text{e}^-$ $\quad \text{E}^\circ_{\text{cell}} = +1.229\ \text{V}$

Cathode: $2\text{Cl}_2\text{(l)} + 4\text{e}^- \rightarrow 4\text{Cl}^-\text{(aq)}$ $\quad \text{E}^\circ_{\text{cell}} = +1.358\ \text{V}$

$\text{E}^\circ_{\text{cell}} = +1.358\ \text{V} - (+1.229\ \text{V}) = +0.129\ \text{V}$. Spontaneous reaction at 298K.

Q19.11C

All reactants and products are in their standard states, and use data from the standard electrode reduction potentials table to predict whether the reaction is spontaneous in the forward reaction:

a. $\mathrm{Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)}$
b. $\mathrm{2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)}$
c. $\mathrm{Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)}$
d. $\mathrm{2Fe^{2+}(aq) + I_2(s) \rightarrow 2Fe^{3+}(aq) + 2I^{-}(aq)}$

S19.11C

a. Cell reaction: $\mathrm{Cu(s) + 2Ag^{+}(aq) \rightarrow 2Cu^{2+}(aq) + 2Ag(s)}$

We must split up the overall reaction into balanced half reactions, which is why we multiply Ag by 2. We have to have equal amounts of electrons on both sides so that they cancel. The placement of electrons on the charged sides of the half reactions balances the charges to 0. We must reach 0 because the elemental versions of Cu and Ag have a charge of 0. After we form the half reactions, we must determine the cathode and the anode because E^ocell=E^o(cathode)-E^o(anode). In other words, the anode is the oxidized portion and the cathode is the reduced portion. Oxidized means we lost electrons and reduced means we gained electrons.

Oxidation: $\mathrm{Cu(s) \rightarrow Cu^{2+} + 2e^{-}}$ Lost electrons
Reduction: $\mathrm{(Ag^{+}(aq) + e^{-} \rightarrow Ag(s)) \times 2}$ Gained Electrons

$$\mathrm{E^{\circ}_{cell}} = \mathrm{E^{\circ}_{(reduction)}} - \mathrm{E^{\circ}_{(oxidation)}} \quad (18.11.57)$$
$$= \mathrm{E^{\circ}_{Ag^{+}/Ag}} - \mathrm{E^{\circ}_{Cu/Cu^{2+}}} \quad (18.11.58)$$
$$= 0.7991 - 0.3394 = +0.4597\ \mathrm{V} \quad (18.11.59)$$

These values are taken off the Reduction Half-Reactions chart. To use this chart, we must find the half reaction used and enter the value in V shown depending on the orientation listed. In this case we take the 0.7991 of Ag+ and subtract 0.3394 of Cu. Our result is a positive 0.4597 V.
The cell potential is positive, therefore the spontaneous reaction will occur in the forward reaction.

b. Cell reaction: $\mathrm{2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)}$

We must split up the overall reaction into balanced half reactions, which is why we multiply Al by 2 and Zn by 3. We have to have equal amounts of electrons on both sides so that they cancel. The placement of electrons on the charged sides of the half reactions balances the charges to 0. We must reach 0 because the elemental versions of Al and Zn have a charge of 0. After we form the half reactions, we must determine the cathode and the anode because E^ocell=E^o(cathode)-E^o(anode). In other words, the anode is the oxidized portion and the cathode is the reduced portion. Oxidized means we lost electrons and reduced means we gained electrons.

Oxidation: $\mathrm{(Al(s) \rightarrow Al^{3+} + 3e^{-}) \times 2}$ Lost Electrons
Reduction: $\mathrm{(Zn^{2+}(aq) + 2e^{-} \rightarrow Zn(s)) \times 3}$

$$Gained Electrons \mathrm{E^{\circ}_{cell}} = \mathrm{E^{\circ}_{(reduction)}} - \mathrm{E^{\circ}_{(oxidation)}} \quad (18.11.60)$$
$$= \mathrm{E^{\circ}_{Zn^{2+}/Zn}} - \mathrm{E^{\circ}_{Al/Al^{3+}}} \quad (18.11.61)$$
$$= -0.7621 - (-1.676) = +0.9139\ \mathrm{V} \quad (18.11.62)$$

These values are taken off the Reduction Half-Reactions chart. To use this chart, we must find the half reaction used and enter the value in V shown depending on the orientation listed. In this case we take the -0.7621 of Zn2+ and subtract -1.676 of Al. Our result is a positive 0.09139V.
The cell potential is positive, therefore the spontaneous reaction will occur in the forward reaction.

c. Cell reaction: $\mathrm{Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)}$

We must split up the overall reaction into balanced half reactions. We have to have equal amounts of electrons on both sides so that they cancel. The placement of electrons on the charged sides of the half reactions balances the charges to 0. We must reach 0 because the elemental versions of Ag and Fe have a charge of 0. After we form the half reactions, we must determine the cathode and the anode because E^ocell=E^o(cathode)-E^o(anode). In other words, the anode is the oxidized portion and the cathode is the reduced portion. Oxidized means we lost electrons and reduced means we gained electrons.

Oxidation: $\mathrm{Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^{-}}$ Lost electrons
Reduction: $\mathrm{Ag(s) \rightarrow Ag^{+}(aq) + e^{-}}$

$$Gain e^- \text{ electrons } E^\circ_{\text{cell}} = E^\circ_{\text{(reduction)}} - E^\circ_{\text{(oxidation)}} \quad (18.11.63)$$

$$= E^\circ_{Ag^+/Ag} - E^\circ_{Fe^{2+}/Fe^{3+}} \quad (18.11.64)$$

$$= 0.7991 - 0.769 = +0.0301 \text{ V} \quad (18.11.65)$$

These values are taken off the Reduction Half-Reactions chart. To use this chart, we must find the half reaction used and enter the value in V shown depending on the orientation listed. In this case we take the 0.7991 of Ag+ and subtract 0.769 of Fe2+. Our result is a positive 0.0301V.
The cell potential is positive, therefore the spontaneous reaction will occur in the forward reaction.

d. Cell reaction: $2Fe^{2+}(aq) + I_2(s) \rightarrow 2Fe^{3+}(aq) + 2I^-(aq)$

We must split up the overall reaction into balanced half reactions, which is why we must multiply Fe by 2. We have to have equal amounts of electrons on both sides so that they cancel. The placement of electrons on the charged sides of the half reactions balances the charges to 0. We must reach 0 because the elemental versions of I and Fe have a charge of 0. After we form the half reactions, we must determine the cathode and the anode because E^ocell=E^o(cathode)-E^o(anode). In other words, the anode is the oxidized portion and the cathode is the reduced portion. Oxidized means we lost electrons and reduced means we gained electrons.

Oxidation: $(Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)) \times 2$ Lost electrons
Reduction: $I_2(s) + 2e^- \rightarrow 2I^-(aq)$

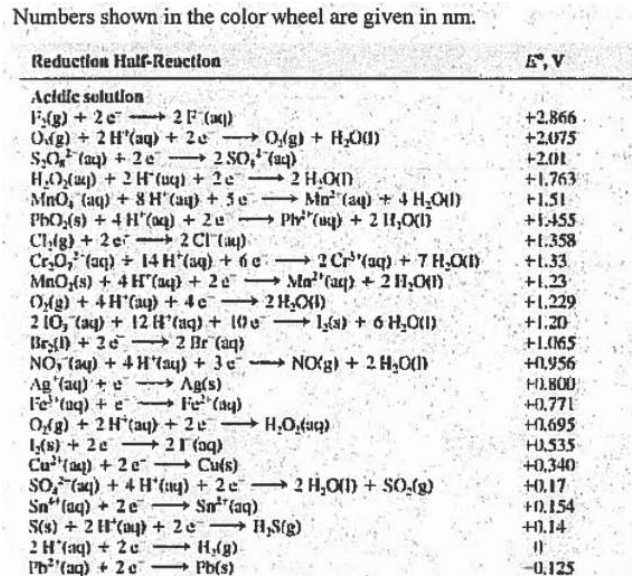
$$Gain e^- \text{ electrons } E^\circ_{\text{cell}} = E^\circ_{\text{(reduction)}} - E^\circ_{\text{(oxidation)}} \quad (18.11.66)$$

$$= E^\circ_{I_2/2I^-} - E^\circ_{Fe^{2+}/Fe^{3+}} \quad (18.11.67)$$

$$= 0.535 - 0.769 = -0.234 \text{ V} \quad (18.11.68)$$

These values are taken off the Reduction Half-Reactions chart. To use this chart, we must find the half reaction used and enter the value in V shown depending on the orientation listed. In this case we take the 0.535 of I2 and subtract 0.769 of Fe2+. Our result is a negative -0.234V. Since the half reaction for Fe2+ is flipped on the chart and yields a negative V, we must change it to a positive. This gives us the correct value for the correct direction of the reaction and results in an overall negative -0.234V.
The cell potential is negative, therefore the spontaneous reaction will not occur in the forward reaction.

Below is a partial image of the chart we must use to come to these conclusions:


Numbers shown in the color wheel are given in nm.

Reduction Half-Reaction	E°, V
Acidic solution	
$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	+2.866
$O_3(g) + 2H^+(aq) + 2e^- \longrightarrow O_2(g) + H_2O(l)$	+2.075
$S_2O_8^{2-}(aq) + 2e^- \longrightarrow 2SO_4^{2-}(aq)$	+2.01
$H_2O_2(aq) + 2H^+(aq) + 2e^- \longrightarrow 2H_2O(l)$	+1.763
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$PbO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Pb^{2+}(aq) + 2H_2O(l)$	+1.455
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.358
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$MnO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Mn^{2+}(aq) + 2H_2O(l)$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	+1.229
$2IO_3^-(aq) + 12H^+(aq) + 10e^- \longrightarrow I_2(s) + 6H_2O(l)$	+1.20
$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$	+1.065
$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$	+0.956
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.800
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.771
$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$	+0.695
$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.535
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.340
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow 2H_2O(l) + SO_2(g)$	+0.17
$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$	+0.154
$S(s) + 2H^+(aq) + 2e^- \longrightarrow H_2S(g)$	+0.14
$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	−0.125

Here is a link to a video that solves examples of E^o cell under nonstandard conditions for additional help:

www.youtube.com/watch?v=62rEVSBBJMw

Q19.11D

You may assume that the reactants and products in the equations are in their standard states. Use the information from Table P2 to predict if a spontaneous reaction will occur in the forward direction *as drawn* for the following cases.

a. $Zn(s) + Cu^{2+} \rightarrow Zn^{2+} + Cu(s)$
b. $2Hg^{2+}(aq) + 2Br^{-}(aq) \rightarrow Hg_2^{2+}(aq) + Br_2(l)$
c. $2Fe^{2+}(aq) + Cl_2(g) \rightarrow 2Fe^{3+}(aq) + 2Cl^{-}(aq)$

S19.11D

a. The $E^{\circ}_{cell} = E^{\circ}_{(reduction\ half\text{-}cell)} - E^{\circ}_{(oxidation\ half\text{-}cell)} = 0.340V - (-0.763V) = 1.103V$. Because the E°_{cell} is positive, the direction of spontaneous change is in the forward reaction. Recognize that Cu^{2+} is the stronger oxidizing agent (see Table P2) and will be more likely to be reduced. Thus, the Cu^{2+} half reaction is where copper is reduced and will be the reduction half-cell. The Zn^{2+}half reaction will be the oxidation half-cell because Zinc gets oxidized at the anode.
b. The $E^{\circ}_{cell} = E^{\circ}_{(reduction\ half\text{-}cell)} - E^{\circ}_{(oxidation\ half\text{-}cell)} = 0.911V - 1.087V = -0.126V$. Because the E°_{cell} is negative, the direction of spontaneous change is in the *reverse* direction.
c. The $E^{\circ}_{cell} = E^{\circ}_{(reduction\ half\text{-}cell)} - E^{\circ}_{(oxidation\ half\text{-}cell)} = 1.358V - 0.771V = 0.587V$. Because the E° is positive, the direction of spontaneous change is in the forward direction. Recognize that the reduction of chlorine takes place at the cathode, making it the reduction half-cell. The oxidation of iron occurs at the anode, making it the oxidation half-cell.

Q19.12A

Describe what would happen to a nail that is in a mixture which demonstrates a pink color if there is a base present and a blue color in the presence of Turbull's blue $KFe[Fe(CN)_6]$ in each situation:

a. Copper wire is wrapped around the tip of the nail
b. The center of the nail has the outer layer filed off
c. The nail is coated with Zn (galvination)

S19.12A

a. Iron is a stronger reducing agent than copper and so blue precipitate will form in the areas not covered by copper (oxidation will occur). Hydroxide would be produced at the cathode (copper wire) and pink will form.
b. Newly exposed metal will oxidize producing a blue precipitate.
c. The zinc will oxidize rather than the iron producing hydroxide ions (pink color).

Q19.13A

An aqueous solution containing the following ions has 2.15 A of current passed through for 60 min, how much metal (in g) is formed at the cathode in each solution?

a. Zn^{2+}
b. Al^{3+}
c. Ag^{+}
d. Ni^{2+}]

S19.13A

First calculate charge passed through and number of moles of electrons transferred:

$$\text{Mol } e^- = 60 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{2.15 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mol } e^-}{96485 \text{ C}} = 0.08 \text{ mol } e^-$$

a. Zn^{2+}

$$0.08 \text{ mol } e^- \times \frac{1 \text{ mol } Zn^{2+}}{2 \text{ mol } e^-} \times \frac{1 \text{ mol Zn}}{1 \text{ mol } Zn^{2+}} \times \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}} = 2.62 \text{ g Zn}$$

b. Al^{3+}

$$0.08 \text{ mol } e^- \times \frac{1 \text{ mol } Al^{3+}}{3 \text{ mol } e^-} \times \frac{1 \text{ mol Al}}{1 \text{ mol } Al^{3+}} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 0.72 \text{ g Al}$$

c. Ag^{+}

$$0.08 \text{ mol } e^- \times \frac{1 \text{ mol } Ag^{+}}{1 \text{ mol } e^-} \times \frac{1 \text{ mol Ag}}{1 \text{ mol } Ag^{+}} \times \frac{107.9 \text{ g Al}}{1 \text{ mol Ag}} = 8.63 \text{ g Ag}$$

d. Ni^{2+}

$$0.08 \text{ mol } e^- \times \frac{1 \text{ mol } Ni^{2+}}{2 \text{ mol } e^-} \times \frac{1 \text{ mol Ni}}{1 \text{ mol } Ni^{2+}} \times \frac{58.69 \text{ g Ni}}{1 \text{ mol Ni}} = 2.35 \text{ g Ni}$$

Q19.14A

What voltage is required for the electrolysis of the following reactions? All reactants are in standards states.

a. $Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$
b. $2Fe^{2+}(aq) + Hg^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Hg(l)$
c. $Cu(s) + Sr^{2+}(aq) \rightarrow Cu^{2+}(aq) + Sr(s)$

S19.14A

a. Oxidation reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad E^\circ_{cell} = -0.763\ V$

Reduction reaction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ $\quad E^\circ_{cell} = -0.137\ V$

Net reaction: $Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$ $\quad E^\circ_{cell} = 0.626\ V$

This reaction is spontaneous

b. Oxidation reaction: $2Fe^{2+} \rightarrow 2Fe^{3+} + 2e^-$ $\quad E^\circ_{cell} = 0.771\ V$

Reduction reaction: $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$ $\quad E^\circ_{cell} = 0.854\ V$

Net reaction: $2Fe^{2+}(aq) + Hg^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Hg(l)$ $\quad E^\circ_{cell} = 0.083\ V$

This reaction is spontaneous

c. Oxidation reaction: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $\quad E^\circ_{cell} = 0.337\ V$

Reduction reaction: $Sr^{2+}(aq) + 2e^-(aq) \rightarrow Sr(s)$ $\quad E^\circ_{cell} = -2.89\ V$

Net reaction: $Cu(s) + Sn^{4+}(aq) \rightarrow Cu^{2+}(aq) + Sn^{2+}(aq)$ $\quad E^\circ_{cell} = -3.227\ V$

Reaction requires voltages $> 3.227\ V$

Q19.15A

$N_2H_4(aq) + O_2(g) \rightarrow N_2(g) + 2H_2O(l)$ $\quad E^\circ_{cell} = 1.559\ V$

Calculate ΔG°_f for $[N_2H_4(aq)]$

S19.15A

Calculate $\Delta G^\circ_f = -nFE^\circ_{cell} = -4\ mol\ e^- \times \dfrac{96485\ C}{1\ mol\ e^-} \times 1.559\ V = -6.017 \times 10^5\ J = -601.7\ kJ$

Find ΔG°_f for (N_2H_4):

$-601.7\ kJ = \Delta G^\circ_f[N_2(g)] + 2\Delta G^\circ_f[H_2O(l)] - \Delta G^\circ_f[N_2H_4(aq)] - \Delta G^\circ_f[O_2(g)]$

$-601.7\ kJ = 0.00\ kJ + 2(-237.2) + 601.7 = 127.3\ kJ$

For more information on calculating E'_{cell} check out ***The Cell Potential****.*

For more info on calculating E_{cell} check out ***Nernst Equation*** *and* ***Electrochemical Cells under Nonstandard Conditions****.*

Q19.16A

Find the weight of MnO_2 in grams from the following overall reaction knowing that the voltmeter reading at non standard state is 0.25V, the pH of the solution is 6 and the solution is diluted to 0.6L.

$Pt(s)|Cr^{3+}(0.2\ M), Cr_2O_7^{2-}(0.15\ M)||MnO_4(x\ M)|MnO_2(s)$

S19.16A

$$\text{Reduction: } 2\,MnO_4^-(aq) + 8\,H^+(aq) + 6\,e^- \rightarrow 2\,MnO_2(s) + 2\,H_2O(l) \qquad E^\circ = +1.51\ V \qquad (18.11.69)$$

$$\text{Oxidation: } 2\,Cr^{3+}(aq) + 7\,H_2O(l) \rightarrow Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,e^- \qquad E^\circ = +1.33\ V \qquad (18.11.70)$$

$$E^\circ_{cell} = E^\circ_{reduction} - E^\circ_{oxidation} \qquad (18.11.71)$$

$$= 1.51 - 1.33 \qquad (18.11.72)$$

$$= 0.18\ V \qquad (18.11.73)$$

Apply Nernst Equation:

$$E_{cell} = E^\circ_{cell} - \frac{0.0592\ V}{n} \log Q$$

$$0.25 = 0.18 - \frac{0.0592\ \text{V}}{6}\log\frac{[\text{H}^+]^6[\text{Cr}_2\text{O}_7^{2-}]}{[\text{Cr}^{3+}][\text{MnO}_4^-]^2}$$

$$-10.135 = \log\frac{[\text{H}^+]^6[\text{Cr}_2\text{O}_7^{2-}]}{[\text{Cr}^{3+}][\text{MnO}_4^-]^2}$$

$$7.3\times10^{-11} = \frac{[\text{H}^+]^6[\text{Cr}_2\text{O}_7^{2-}]}{[\text{Cr}^{3+}][\text{MnO}_4^-]^2}$$

$$7.3\times10^{-11} = \frac{[10^{-6}]^6[0.15]}{[0.2][\text{MnO}_4^-]^2}$$

$$[\text{MnO}_4^-]^2 = 2.2\times10^{-48}$$

$$[\text{MnO}_4^-] = 1.48\times10^{-24}$$

Mass of MnO_2 produced: $0.6\ \text{L}\times\frac{1.48\times10^{-24}}{1\ \text{L of solution}}\times\frac{2\ \text{mol MnO}_2}{2\ \text{mol MnO}_4^-}\times\frac{86.94\ \text{g MnO}_2}{1\ \text{mol MnO}_2}$

$= 7.73\times10^{-23}$ g of MnO_2 produced

Q19.17A

a. Identify the value of the $\text{E}^\circ_{\text{cell}}$ of the following reaction, and indicate if the reaction takes place spontaneous or nonspontaneously.

$$\text{I}_2(\text{s})|\text{I}^-(0.6\ \text{M})||\text{Br}_2(\text{l})|\text{Br}^-(0.80\ \text{M})|\text{Pt}(\text{s})$$

b. What is the minimum ratio of $\frac{[\text{Br}^-]^2}{[\text{I}^-]^2}$ for this reaction to be spontaneous?

S19.17A

a. Reduction: $\text{Br}_2(\text{l}) + 2\,\text{e}^- \rightarrow \text{Br}^-(\text{aq})$ $\quad \text{E}^\circ = +1.065$ V
Oxidation: $2\,\text{I}^-(\text{aq}) \rightarrow \text{I}_2(\text{s}) + 2\,\text{e}^-$ $\quad \text{E}^\circ = +0.535$ V
Overall: $\text{Br}_2(\text{l}) + 2\,\text{I}^-(\text{aq}) \rightarrow 2\,\text{Br}^-(\text{aq}) + \text{I}_2(\text{s})$

$$\begin{aligned}\text{E}^\circ_{\text{cell}} &= \text{E}^\circ_{\text{reduction}} - \text{E}^\circ_{\text{oxidation}} && (18.11.74)\\ &= 1.065 - 0.535 && (18.11.75)\\ &= 0.530\ \text{V} && (18.11.76)\end{aligned}$$

Apply Nernst Equation:

$$\begin{aligned}\text{E}_{\text{cell}} &= \text{E}^\circ_{\text{cell}} - \frac{0.0592\ \text{V}}{\text{n}}\log\text{Q} && (18.11.77)\\ &= 0.53 - \frac{0.0592\ \text{V}}{2}\log\frac{[\text{Br}^-]^2}{[\text{I}^-]^2} && (18.11.78)\\ &= 0.53 - \frac{0.0592\ \text{V}}{2}\log\frac{[0.8]^2}{[0.6]^2} && (18.11.79)\\ &= 0.523\ \text{V} && (18.11.80)\end{aligned}$$

Since $\text{E}_{\text{cell}} > 0$, $\Delta\text{G}^\circ < 0$. Therefore, the reaction is spontaneous.

b. Set $\text{E}_{\text{cell}} = 0$

Apply Nernst Equation:

$$\text{E}_{\text{cell}} = \text{E}^\circ_{\text{cell}} - \frac{0.0592\ \text{V}}{\text{n}}\log\text{Q}$$

$$0 = 0.53\ \text{V} - \frac{0.0592\ \text{V}}{2}\log\frac{[\text{Br}^-]^2}{[\text{I}^-]^2}$$

$$17.9 = \log\frac{[\text{Br}^-]^2}{[\text{I}^-]^2}$$

$$\frac{[\text{Br}^-]^2}{[\text{I}^-]^2} = 8.04\times10^{17}\ \text{M}$$

Q19.19A

Write cell reactions for the electrochemical cells diagrammed, and use data from the table of standard electrode potentials to calculate E°_{cell} for each reaction.

a. $Fe(s)|Fe^{2+}(aq)||Zn^{2+}|Zn(s)(aq)$
b. $Na(s)|Na^{+}(aq)||Cl_2(g)|Cl^{-}(aq)$
c. $Co(s)|Co^{2+}(aq)||Mn^{2+}(aq)|Mn(s)$
d. $Pt(s)|Fe^{2+}(aq), Fe^{3+}(aq)||Pb^{2+}(aq)|Pb(s)$

S19.19A

a. Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ $\quad E^\circ = -(-0.440\ V)$

Reduction: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E^\circ = -0.7621\ V$
Net Reaction: $Fe(s) + Zn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Zn(s)$ $\quad E^\circ_{cell} = -0.3321\ V$

b. Oxidation: $(Na(s) \rightarrow Na^{+}(aq) + e^-) \times 2$ $\quad E^\circ = -(-2.7144\ V)$

Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ $\quad E^\circ = 1.3601\ V$
Net Reaction: $Na(s) + Cl_2(g) \rightarrow Na^{+}(aq) + 2Cl^-(aq)$ $\quad E^\circ_{cell} = +4.0745\ V$

c. Oxidation: $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$ $\quad E^\circ = -(-0.277\ V)$

Reduction: $Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$ $\quad E^\circ = -1.182\ V$
Net Reaction: $Co(s) + Mn^{2+}(aq) \rightarrow Co^{2+}(aq) + Mn(s)$ $\quad E^\circ_{cell} = -0.905\ V$

d. Oxidation: $(Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-) \times 2$ $\quad E^\circ = -(+0.771\ V)$

Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $\quad E^\circ = -0.125\ V$
Net Reaction: $2Fe^{2+} + Pb^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Pb(s)$ $\quad E^\circ_{cell} = -0.896\ V$

Q19.19B

Write cell reaction and calculate its E°_{cell}.

a. $Al(s)|Al^{3+}(aq)||Pb^{2+}(aq)|Pb(s)$
b. $Na(s)|Na^{+}(aq)||Mn^{2+}(aq)|Mn(s)$
c. $Li(s)|Li^{+}(aq)||F_2(g)|F^{-}(aq)|Pt(s)$
d. $Pt(s)|Fe^{2+}(aq),\ Fe^{3+}(aq)||Cu^{2+}(aq)|Cu(s)$

S19.19B

a. Anode/Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$ $\quad E^\circ_{cell} = -1.676\ V$

Cathode/Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$ $\quad E^\circ_{cell} = -0.125\ V$

Overall: $3Pb^{2+}(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$

$E^\circ_{cell} = -0.125\ V - (-1.676\ V) = +1.551\ V$

b. Anode: $Na(s) \rightarrow Na^{+}(aq) + e^-$ $\quad E^\circ_{cell} = -2.7144\ V$

Cathode: $Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$ $\quad E^\circ_{cell} = -1.182\ V$

Overall: $Mn^{2+}(aq) + 2Na(s) \rightarrow Mn(s) + 2Na^{+}(aq)$

$E^\circ_{cell} = -1.182\ V - (-2.7144\ V) = +1.5324\ V$

c. Anode: $Li(s) \rightarrow Li^{+}(aq) + e^-$ $\quad E^\circ_{cell} = -3.040\ V$

Cathode: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$ $\quad E^\circ_{cell} = +2.866\ V$

Overall: $2Li(s) + F_2(g) + \rightarrow 2F^-(aq) + 2Li^{+}(aq)$

$E^\circ_{cell} = +2.866\ V - (-3.040\ V) = +5.906\ V$

d. Anode: $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$ $\quad E^\circ_{cell} = 0.771\ V$

Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E^\circ_{cell} = +0.340\ V$

Overall: $Cu^{2+}(aq) + 2Fe^{2+}(aq) \rightarrow Cu(s) + 2Fe^{3+}(aq)$

$E^\circ_{cell} = +0.340\ V - (0.771\ V) = -0.431\ V$

Q19.19C

Write the cell reactions for the electrochemical cells provided below. Then use data from Table Table P2 calculate E°_{cell} for each of the reactions.

a. $Ag(s)|Ag^+(aq)||Hg^{2+}(aq)|Hg(l)$
b. $Al(s)|Al^{3+}(aq)||Zn^{2+}(aq)|Zn(s)$
c. $Pt|Ce^{4+}(aq), Ce^{3+}(aq)||I^-(aq), I_2(s)|C(s)$
d. $Mg(s)|MCg^{2+}||2H^+(aq)|H_2(g)$

S19.19

a. Oxidation: $2Ag(s) \rightarrow 2Ag^+(aq) + 2e^-$

Reduction: $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$

Overall: $2Ag(s) + Hg^{2+}(aq) \rightarrow 2Ag^+(aq) + Hg(l)$

$$E^\circ_{cell} = E^\circ_{(\text{reduction half-cell})} - E^\circ_{(\text{oxidation half-cell})} \quad (18.11.81)$$

$$= 0.854\ V - (0.800\ V) = 0.054\ V \quad (18.11.82)$$

b. Oxidation: $[Al(s) \rightarrow Al^{3+}(aq) + 3e^-] \times 2$

Reduction: $[Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)] \times 3$

Overall: $2Al(s) + 3Zn^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Zn(s)$

$$E^\circ_{cell} = E^\circ_{(\text{reduction half-cell})} - E^\circ_{(\text{oxidation half-cell})} \quad (18.11.83)$$

$$= -0.763 - (-1.676) = 0.913\ V \quad (18.11.84)$$

c. Oxidation: $I_2(s) \rightarrow 2I^-(aq) + 2e^-$

Reduction: $[Ce^{4+} + e^- \rightarrow Ce^{3+}] \times 2$

Overall: $2Ce^{4+} + I_2(s) \rightarrow 2Ce^{3+} + 2I^-(aq)$

$E^\circ = 1.44 - 0.54 = 0.90\ V$

d. Oxidation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Reduction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overall: $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

$E^\circ = 0 - (-2.372) = 2.372\ V$

For review on this topic, visit the page ***"Electrochemistry 2: Galvanic cells and electrodes"***.

Q19.22A

For each of the following reactions below, draw a voltaic cell. In your drawing include the anode, cathode, and show the flow of electrons. Balance the equation and calculate the E°_{cell}.

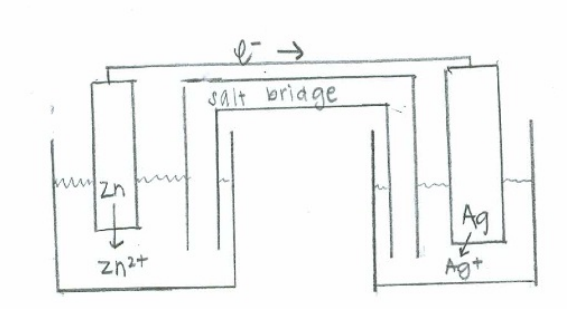
a. $Zn(s) + Ag^+(aq) \rightarrow Ag(s) + Zn^{2+}(aq)$
b. $Rb(s) + Hg^{2+}(aq) \rightarrow Rb^+(aq) + Hg(s)$
c. $F_2(g) + H_2O(l) \rightarrow 4F^-(aq) + O_2(g) + H^+(aq)$
d. $Zn(s) + H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

S19.22A

a. Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad E^\circ = 0.763\ V$

Reduction: $(e^- + Ag^+(aq) \rightarrow Ag(s)) \times 2$ $\quad E^\circ = 0.337\ V$

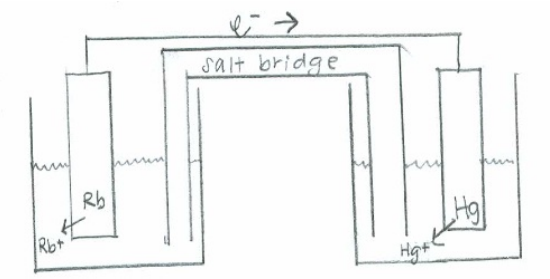
Net: $Zn(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Zn^{2+}(aq)$ $\quad E^\circ_{cell} = 1.100\ V$



b. Oxidation: $(\mathrm{Rb(s) \rightarrow Rb^+(aq) + e^-}) \times 2$ $\mathrm{E^\circ = +2.93\ V}$

Reduction: $\mathrm{Hg^{2+}(aq) + 2e^- \rightarrow Hg(s)}$ $\mathrm{E^\circ = 0.86\ V}$

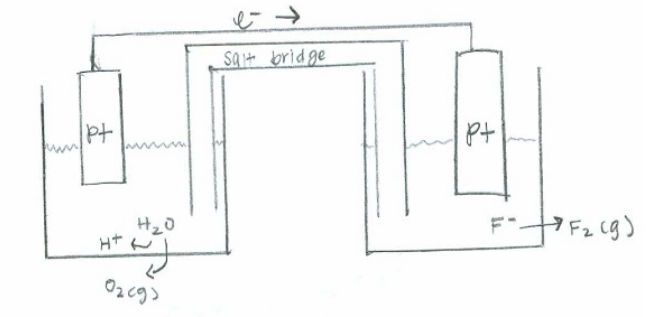
Net: $\mathrm{2Rb(s) + Hg^{2+}(aq) \rightarrow 2Rb^+(aq) + Hg(s)}$ $\mathrm{E^\circ_{cell} = 3.79\ V}$



c. Oxidation: $\mathrm{2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-}$ $\mathrm{E^\circ = -1.229\ V}$

Reduction: $\mathrm{(2e^- + F_2(g) \rightarrow 2F^-(aq)) \times 2}$ $\mathrm{E^\circ = 2.866\ V}$

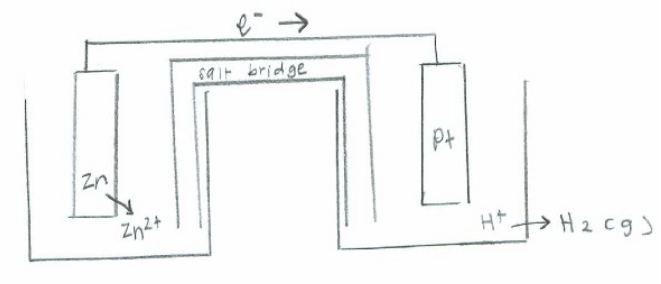
Net: $\mathrm{2F_2(g) + 2H_2O(l) \rightarrow 4F^-(aq) + O_2(g) + 4H^+(aq)}$ $\mathrm{E^\circ_{cell} = 1.637\ V}$



d. Oxidation: $\mathrm{Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-}$ $\mathrm{E^\circ = +0.76\ V}$

Reduction: $\mathrm{2H^+(aq) + 2e^- \rightarrow H_2(g)}$ $\mathrm{E^\circ = 0}$

Net: $\mathrm{Zn(s) + H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)}$ $\mathrm{E^\circ_{cell} = 0.76\ V}$



Q19.22B

Sketch a voltaic cell. Label anode and cathode. Indicate direction of electron flow, write balanced equation for reaction and calculate E°_{cell}.

a. $Fe^{2+}(aq)$ displaced from solution by $Li(s)$
b. $Sn(s) + H^+(aq) + NO_3^-(aq) \rightarrow Sn^{2+}(aq) + H_2O(l) + NO(g)$
c. $F_2(g) + H_2O(l) \rightarrow F^-(aq) + O_2(g) + H^+(aq)$
d. $Cu(s) + F_2(g) \rightarrow Cu^{2+}(aq) + 2F^-(aq)$

S19.22B

a. Anode = Li electrode. Cathode = Fe electrode

For voltaic cell, e- flow from anode to cathode. Salt bridge connects the two solutions.

Anode: $Li(s) \rightarrow Li^+(aq) + e^-$ $\quad E^\circ_{cell} = -3.040\ V$

Cathode: $Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ $\quad E^\circ_{cell} = -0.440\ V$

$E^\circ_{cell} = -0.440\ V - (-3.040\ V) = +2.60\ V$

Balanced equation: $2Li(s) + Fe^{2+}(aq) \rightarrow Fe(s) + 2Li^+(aq)$

b. Anode = Sn electrode. Cathode = Pt electrode

For voltaic cell, e- flow from anode to cathode. Salt bridge connects the two solutions.

Anode: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$ $\quad E^\circ_{cell} = +0.137\ V$

Cathode: $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$ $\quad E^\circ_{cell} = +0.956\ V$

$E^\circ_{cell} = +0.956\ V - (+0.137\ V) = +0.819\ V$

Balanced equation: $3Sn(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Sn^{2+}(aq) + 2NO(g) + 4H_2O(l)$

Cell diagram: $Sn(s)|Sn^{2+}(aq)||H^+(aq), NO_3^-(aq)|NO(g)|Pt(s)$

c. Anode = Pt electrode. Cathode = Pt electrode

For voltaic cell, e- flow from anode to cathode. Salt bridge connects the two solutions.

Anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $\quad E^\circ_{cell} = +1.229\ V$

Cathode: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$ $\quad E^\circ_{cell} = +2.866\ V$

$E^\circ_{cell} = +2.866\ V - (+1.229\ V) = +1.637\ V$

Balanced equation: $2F_2(g) + 2H_2O(l) \rightarrow 4F^-(aq) + O_2(g) + 4H^+(aq)$

d. Anode = Cu electrode. Cathode = Pt electrode

For voltaic cell, e- flow from anode to cathode. Salt bridge connects the two solutions.

Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $\quad E^\circ_{cell} = +0.340\ V$

Cathode: $F_2(g) + 2e^- \rightarrow 2F^-(aq)$ $\quad E^\circ_{cell} = +2.866\ V$

$E^\circ_{cell} = +2.866\ V - (+0.340\ V) = +2.526\ V$

Balanced equation: $Cu(s) + F_2(g) \rightarrow Cu^{2+}(aq) + 2F^-(aq)$

Q19.22C

Identify where each reaction will take place (on the cathode or anode), balance the equation if necessary, and calculate the E°_{cell}.

a. $Fe^{3+}(aq) + Ag(s) \rightarrow Fe^{2+}(aq) + Ag^+(aq)$
b. $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$
c. $Cd(s) + Cu^{2+}(aq) \rightarrow Cd^{2+} + Cu(s)$
d. $Fe^{2+}(aq) + Cl_2(g) \rightarrow Fe^{3+}(aq) + 2Cl^-(aq)$

S19.22C

a. Reduction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ at the cathode (reduction occurs at the cathode)

Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$ at the anode

The equation is already balanced.

$E^\circ_{cell} = (0.771 - 0.800) = -0.029\ V$

b. Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ at the cathode, where reduction occurs

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ at the anode

Reaction is already balanced

$E^\circ_{cell} = 0.340 - (-0.763) = 1.103\ V$

c. Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ at the cathode, where reduction occurs

Oxidation: $Cd(s) \rightarrow Cd^{2+} + 2e^-$ at the anode

Reaction is already balanced

$E^\circ_{cell} = E^\circ_{(right)} - E^\circ_{(left)} = 0.340 - 0.743 = -0.403$

d. Balance the equation, multiplying the iron half equation by 2 to equal the electrons

Reduction: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ at the cathode

Oxidation: $2(Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + e^-)$ at the anode

$E^\circ_{cell} = 1.358 - (-0.440) = 1.798\ V$

For review on this topic, visit the page ***"Electrochemistry 2: Galvanic cells and electrodes"***.

Q19.25A

What are the ΔG° for the reactions below?

a. $Ti(s) + Cr^{2+}(aq) \rightarrow Ti^{2+}(aq) + Cr(s)$
b. $2Cu^+(aq) + Sn^{4+}(aq) \rightarrow 2Cu^{2+}(aq) + Sn^{2+}(aq)$
c. $2Cl_2(g) + 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4Cl^-(aq)$

S19.25A

a. Oxidation: $Ti(s) \rightarrow Ti^{2+}(aq) + 2e^-$ $\quad -E^\circ = +1.63\ V$

Reduction: $Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$ $\quad E^\circ = -0.90\ V$

Net: $Ti(s) + Cr^{2+}(aq) \rightarrow Ti^{2+}(aq) + Cr(s)$ $\quad E^\circ_{cell} = 0.73\ V$

$\Delta G^\circ = -nFE^\circ_{cell} = -(2\ mole\ e^-)(96,485\ C/mol\ e^-)(0.73\ V) = -140,868.1\ J = -1.41 \times 10^2\ KJ$

b. Oxidation: $Sn^{2+}(aq) \rightarrow 2e^- + Sn^{4+}(aq)$ $\quad -E^\circ = -0.154\ V$

Reduction: $(Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)) \times 2$ $\quad E^\circ = 0.159\ V$

Net: $2Cu^+(aq) + Sn^{4+}(aq) \rightarrow 2Cu^{2+}(aq) + Sn^{2+}(aq)$ $\quad E^\circ_{cell} = 0.005\ V$

$\Delta G^\circ = -nFE^\circ_{cell} = -(2\ mole\ e^-)(96,485\ C/mol\ e^-)(0.005\ V) = -946.85\ J = -0.965\ KJ$

c. Oxidation: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $\quad -E^\circ = -1.229\ V$

Reduction: $(Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)) \times 2$ $\quad E^\circ = 1.358$

Net: $2Cl_2(g) + 2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4Cl^-(aq)$ $\quad E^\circ_{cell} = 0.129\ V$

$\Delta G^\circ = -nFE^\circ_{cell} = -(4\ mole\ e^-)(96,485\ C/mol\ e^-)(0.129\ V) = -49,786.26\ J = -4.98 \times 10^4\ KJ$

Q19.25B

Determine ΔG° for reactions:

a. $O_2(g) + 4H^+(aq) + 4Br^-(aq) \rightarrow 2H_2O(l) + 2Br_2(g)$
b. $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6K(s) \rightarrow 2Cr^{3+}(aq) + 6K^+(aq) + 7H_2O(l)$
c. $Cu(s) + Br_2(l) \rightarrow Cu^{2+}(aq) + 2Br^-(aq)$

S19.25B

$\Delta G^\circ = -nFE^\circ_{cell}$

a. Oxidation: $4Br^-(aq) \rightarrow 2Br_2(g) + 4e^-$ $\quad E^\circ_{cell} = +1.065\ V$

Reduction: $O_2(g) + 4e^- + 4H^+(aq) \rightarrow 2H_2O(l)$ $\quad E^\circ_{cell} = +1.229\ V$

$E^\circ_{cell} = +1.229\ V - +1.065\ V = 0.164\ V$

$n = 4\ e^-$
$F = 96,485\ C/mol\ e^-$
$\Delta G^\circ = -(4e^-)(96,485\ C/mol\ e^-)(0.164\ V) = -63294.16\ J = -63\ kJ$

b. Oxidation: $6K(s) \rightarrow 6K^+(aq) + 6e^-$ $E^\circ_{cell} = -2.924\ V$

Reduction: $Cr_2O_7^{2-}(aq) + 6e^- + 14H^+(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ $E^\circ_{cell} = +1.33\ V$
$E^\circ_{cell} = +1.33\ V - (-2.924\ V) = 4.254\ V$
$n = 6\ e^-$
$F = 96,485\ C/mol\ e^-$
$\Delta G^\circ = \dfrac{-(6\ e^-)(96,485\ C/mol\ e^-)(4.254\ V)}{1000} = -2462.68\ kJ$

c. Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $E^\circ_{cell} = +0.340\ V$

Reduction: $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ $E^\circ_{cell} = +1.065\ V$

$E^\circ_{cell} = +1.065\ V - +0.340\ V = +0.725\ V$

$n = 2\ e^-$

$F = 96,485\ C/mol\ e^-$

$\Delta G^\circ = -(2e^-)(96,485\ C/mol\ e^-)(0.725\ V) = -139903\ J = -139.9\ kJ$

Q19.25C

Find the values of ΔG for the following reactions in the voltaic cells

a. $2V^{3+}(aq) + Ni(s) \rightarrow 2V^{2+}(aq) + Ni^{2+}(aq)$
b. $2Al(s) + 3Br_2(l) \rightarrow 2Al^{3+}(aq) + 6Br^-$
c. $2Cu^{2+}(aq) + Sn^{2+}(aq) \rightarrow 2Cu + Sn^{4+}$

S19.25C

a. The z (moles of electrons) $= 2$

$$E^\circ_{cell} = E^\circ_{(reduction)} - E^\circ_{(oxidation)} \tag{18.11.85}$$
$$= -0.255 - (-0.257) = 0.002\ V \tag{18.11.86}$$

$\Delta G = -zFE^\circ_{cell} = -2\ mol\ e^- \times 96485\ C \times 0.002\ V = -385.94 J$

b. The z (moles of electrons) $= 6$

$E^\circ_{cell} = E^\circ_{(reduction)} - E^\circ_{(oxidation)} = 1.065 - (-1.676) = 2.741\ V$

$\Delta G = -zFE^\circ_{cell} = -6\ mol\ e^- \times 96485\ C \times 2.741\ V = -1.59 \times 10^6 J$

c. The z (moles of electrons transferred) $= 2$

$E^\circ_{cell} = 0.337 - 0.150 = 0.187\ V$

$\Delta G = -zFE^\circ_{cell} = -2\ mol\ e^- \times 96485\ C \times 0.187\ V = -3.61 \times 10^4 J$

Q19.27A

Find E°, ΔG°, K, and given that the reactants and products are in their standard state, if they go to completion or not for the following reaction:

$$2BrO_3^-(aq) + 12H^+(aq) + 10SO_4^{2-}(aq) \rightarrow 5S_2O_8^{2-}(aq) + Br_2(l) + 6H_2O(l)$$

S19.27A

Oxidation: $(2SO_4^{2-}(aq) \rightarrow S_2O_8^{2-}(aq) + 2e^-) \times 5$ $-E^\circ = -2.01\ V$

Reduction: $2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(l) + 6H_2O(l)$ $E^\circ = 1.478\ V$

Net: $2BrO_3^-(aq) + 12H^+(aq) + 10SO_4^{2-}(aq) \rightarrow 5S_2O_8^{2-}(aq) + Br_2(l) + 6H_2O(l)$ $E^\circ_{cell} = -0.532\ V$

$\Delta G^\circ = -nFE^\circ_{cell} = -(10\ mole\ e^-)(96,485\ C/mol\ e^-)(-0.532\ V) = 5.13 \times 10^2\ KJ/mol$

$\Delta G^\circ = -RT \ln K$

$5.13 \times 10^2\ KJ/mol = -(-8.314\ KJ/mol)(298.15\ K) \ln K$

$K = 8.13 \times 10^{-1}$

Since K is small (< 1) the reaction will not go to completion.

Q19.27B

Find E°_{cell}, ΔG°, K, and whether reaction goes to completion at 298 K and all substances at standard states for:
$3HClO_2(aq) + 2Cr^{3+}(aq) + 12H_2O(l) \rightarrow 3HClO(aq) + Cr_2O_7^{2-}(aq) + 8H_3O^+(aq)$

S19.27B

Note:

$$\Delta G^\circ = -nFE^\circ_{cell}$$

$$\ln K_{eq} = -\frac{\Delta G^\circ}{RT}$$

Large K_{eq} = reaction goes to completion

Oxidation: $2Cr^{3+}(aq) + 21H_2O(l) \rightarrow Cr_2O_7^{2-}(aq) + 14H_3O^+(aq) + 6e^- \quad E^\circ_{cell} = +1.33\ V$
Reduction: $3HClO_2(aq) + 6H_3O^+(aq) + 6e^- \rightarrow 3HClO(aq) + 9H_2O(l) \quad E^\circ_{cell} = +1.64\ V \quad E^\circ_{cell} = +1.64\ V - +1.33\ V = +0.31\ V$
$\Delta G^\circ = \frac{-(6e^-)(96,485\ C/mol\ e^-)(0.31\ V)}{1000} = -179.46\ kJ\ \ln K_{eq} = -\frac{-179.46\ kJ \times 1000}{(8.3145\ J\ K^{-1}mol^{-1})(298\ K)} = 72.4\ K = 2.86 \times 10^{31}$ Thus since K is very large, reaction will go to completion.

Q19.27C

In the reaction given, calculate the a) E°_{cell} b) ΔG c) K d) and if the reaction goes towards completion when the reactants and products are in their standard states:

$O_2(g) + 4I^-(aq) + 4H^+(aq) \rightarrow 2H_2O(l) + 2I_2(s)$

S19.27C

a. $E^\circ_{cell} = \text{reduction} - \text{oxidation}$
$= (1.229) - (0.535)$
$= 0.694\ V$

b.
$$\Delta G = -zFE^\circ_{cell} = -\frac{4\ mol\ e^- \times (96485\ C)}{1\ mol\ e^- \times 0.694\ V} \tag{18.11.87}$$
$$= -267842\ J \tag{18.11.88}$$

c. $E^\circ_{cell} = \frac{0.025693\ V}{z}\ln(K)$
$0.694\ V = \frac{0.025693\ V}{4}\ln(K)$
$K = 8.38 \times 10^{46}$

d. Since k is very large, the reaction will go substantially towards completion at standard states.

For review on this topic, visit the page "Electrochemistry 3: Cell potentials and thermodynamics".

Q19.31A

Calculate the theoretical cell voltage for the reaction between copper and zinc given that the overall reaction is:
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

S19.31A

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \quad E^\circ = -(-0.763\ V)$
Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \quad E^\circ = 0.340\ V$
$E^\circ_{cell} = 0.763\ V + 0.340\ V = +1.103\ V$

Q19.31B

Calculate a theoretical voltage of silver-lead button cell at 298K. Cell reaction: $Pb(s) + Ag_2O(s) \rightarrow PbO(s) + 2Ag(s) \quad \Delta G^\circ_f$ for $PbO(s) = -187.9\ kJ/mol\ \Delta G^\circ_f$ for $Ag_2O(s) = -11.20\ kJ/mol\ \Delta G^\circ_f$ for $2Ag(s) = 0\ kJ/mol\ \Delta G^\circ_f$ for $Pb(s) = 0\ kJ/mol$

S19.31B

$$\Delta G^\circ = \Delta G_f^\circ[PbO(s)] + 2\Delta G_f^\circ[Ag(s)] - \Delta G_f^\circ[Pb(s)] - \Delta G_f^\circ[Ag_2O(s)]$$

$$\Delta G \text{ of } PbO(s) = -187.9\ kJ/mol + 2(0) - 0 - (-11.20\ kJ/mol) \quad (18.11.89)$$

$$= -176.7 kJ/mol \quad (18.11.90)$$

Use $\Delta G^\circ = -nFE_{cell}^\circ$ to find E_{cell}°. $E_{cell}^\circ = -\dfrac{\Delta G^\circ}{nF} = -\dfrac{(-176.7\ kJ/mol \times 1000)}{(2e^-)(96,485\ C/mol\ e^-)} = 0.9157\ V$

Q19.31C

From the information in appendix D, calculate the theoretical voltage for the hydrogen fuel cell reaction depicted below, and calculate the E_{cell}°:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

S19.31C

$2 \times \Delta G(H_2O) - [\Delta G(O_2) + 2\Delta G(H_2)]$

$= 2(-237.1\ KJ) - 0 + 2(0)\ KJ$ (numbers found from appendix D in the textbook)

$\Delta G = -474.2 - 0$

$\Delta G = -474.2$

Then, use ΔG to calculate E_{cell}°

$$-474.2 = -zFE_{cell}^\circ = -(2\ mol\ e^- \times \frac{96485\ C}{1\ mol\ e^-} \times E_{cell}^\circ)$$

$E_{cell}^\circ = 0.0025\ V$

Q19.35A

A voltaic cell has an E_{cell} value of 1.536 V. What is the concentration of Ag^+ in the cell?

$$Zn(s)|Zn^{2+}(2.00\ M)||Ag^+(x\ M)|Ag(s)$$

S19.35A

First, determine E_{cell}°:

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad -E^\circ = 0.763\ V$

Reduction: $2Ag^+(aq) + 2e^- \rightarrow Ag(s)$ $\quad E^\circ = 0.800\ V$

Net Reaction: $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$ $\quad E_{cell}^\circ = 1.563\ V$

Second, use Nernst equation to calculate $[Ag^+]$

$$E = E_{cell}^\circ - \frac{0.0592}{n}\log\frac{[Zn^{2+}]}{[Ag^+]^2}$$

$$1.536 = 1.563 - \frac{0.0592}{2}\log\frac{2}{x^2}$$

$$-0.027\left(\frac{-2}{0.0592}\right) = \log\frac{2}{x^2}$$

$$8.168 = \frac{2}{x^2}$$

$$x^2 = 0.2448$$

$$x = 0.495 = [Ag^+]$$

Q19.35B

Find $[H_2O_2(aq)]$ in cell if the voltaic cell's $E_{cell} = 1.500$ V. $Zn(s)|Zn^{2+}(aq)(1.00\ M)||H_2O_2(aq)(xM)|H_2O(l)$ Hint: use Nernst's Equation.

S19.35B

Anode: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad E_{cell}^\circ = -0.763\ V$

Cathode: $H_2O_2(aq) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$ $\quad E_{cell}^\circ = +1.763\ V$

$$E_{cell}^\circ = +1.763\ V - (-0.763\ V) = 2.526\ V \quad E = E_{cell}^\circ - \frac{0.0592}{n}\log\frac{[Zn^{2+}]}{[H_2O_2]} \quad 1.5\ V = 2.526\ V - \frac{0.0592}{2}\log\frac{1}{x}$$

$x = 2.18 \times 10^{-35}\ M$

$[H_2O_2(aq)] = 2.18 \times 10^{-35}\ M$

Q19.35C

The voltaic cell in the following diagram has an $E^\circ_{cell} = 0.5464$ V. Solve for the $[Cl^-]$ in the cell

$Ag(s)|Ag^+(0.40\ M)||Cl^-(x\ M),\ Cl_2(0.60\ atm)|Pt(s)$

S19.35C

First calculate E°_{cell}, the theoretical potential. This value is 0.558 V. Then plug in given values to the Nernst formula

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \times \log Q$$

Plugging in values gives us

$$0.5465 = 0.558 - \frac{0.0592}{6} \times \log \frac{(0.40^2)(x^2)}{0.60}$$

Using algebra to isolate the variable on one side, we find that $[Cl^-] = 0.50$.

For review on this topic, visit the page ***"Electrochemistry 4: The Nernst Equation"***.

Q19.37A

Using the Nernst equation, find E_{cell} for the following cells:

a. $Al(s)|Al^{3+}(0.30\ M)||Fe^{2+}(1.1\ M)|Fe(s)$
b. $Co(s)|Co^{2+}(0.6\ M)||Fe^{3+}(1.0\ M), Fe^{2+}(0.9\ M)|Pt(s)$

S19.37A

a. First, determine E°_{cell}

Oxidation: $2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$	$-E^\circ_{cell} = 1.676$ V
Reduction: $3Fe^{2+}(aq) + 6e^- \rightarrow 3Fe(s)$	$E^\circ_{cell} = -0.440$ V
Net Reaction: $2Al(s) + 3Fe^{2+}(1.1\ M) \rightarrow 2Al^{3+}(0.30\ M) + 3Fe(s)$	$E^\circ_{cell} = 1.236$ V

Second, use Nernst equation to calculate E_{cell}

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Al^{3+}]^2}{[Fe^{2+}]^3}$$

$$E_{cell} = 1.236 - \frac{0.0592}{6} \log \frac{(0.3)^2}{(1.1)^3}$$

$$E_{cell} = 1.248$$

b. First, determine E°_{cell}

Oxidation: $Co(s) \rightarrow Co^{2+}(aq) + 2e^-$	$-E^\circ_{cell} = 0.277$ V
Reduction: $2Fe^{3+}(aq) + 2e^- \rightarrow 2Fe^{2+}(aq)$	$E^\circ_{cell} = 0.771$
Net Reaction: $Co(s) + 2Fe^{3+}(1.0\ M) \rightarrow Co^{2+}(0.6\ M) + 2Fe^{2+}(0.9\ M)$	$E^\circ_{cell} = 1.048$ V

Second, use Nernst equation to calculate E_{cell}

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Fe^{2+}]^2[Co^{2+}]}{[Fe^{3+}]^2}$$

$$E_{cell} = 1.048 - \frac{0.0592}{2} \log \frac{(0.9)^2(0.6)}{(1.0)^2}$$

$$E_{cell} = 1.057\ V$$

Q19.37B

Use Nernst's Equation to calculate E_{cell} for:

a. $Al(s)|Al^{3+}(aq)(0.20\ M)||Sn^{2+}(aq)(0.90\ M)|Sn(s)$
b. $Zn(s)|Zn^{2+}(aq)(0.42\ M)||Cl^-(0.02\ M), Cl_2(g,\ 0.5\ atm)|Pt(s)$

S19.37B

Calculate E°_{cell}, then use Nernst equation to calculate E_{cell}

a. Oxidation: $Al(s) \rightarrow Al^{3+}(aq)(0.20\ M) + 3e^-$ $\quad E^\circ_{cell} = -1.676\ V$

Reduction: $Sn^{2+}(aq)(0.90\ M) + 2e^- \rightarrow Sn(s)$ $\quad E^\circ_{cell} = -0.137\ V$

$E^\circ_{cell} = -0.137\ V - (-1.676\ V) = 1.539\ V$

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Al^{3+}]^2}{[Sn^{2+}]^3}$$

$$E_{cell} = 1.539\ V - \frac{0.0592}{6} \log \frac{(0.20\ M)^2}{(0.90\ M)^3}$$

$E_{cell} = 1.551\ V$

b. Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq)(0.42\ M) + 2e^-$ $\quad E^\circ_{cell} = -0.763\ V$

Reduction: $Cl_2(g,\ 0.5\ atm) + 2e^- \rightarrow 2Cl^-(0.02\ M)$ $\quad E^\circ_{cell} = +1.358\ V$

$E^\circ_{cell} = +1.358\ V - (-0.763\ V) = 2.121\ V$

$$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Zn^{2+}][Cl^-]^2}{P_{Cl_2(g)}}$$

$$E_{cell} = 2.121\ V - \frac{0.0592}{2} \log \frac{[0.42\ M][0.02\ M]^2}{0.5\ atm}$$

$E_{cell} = 2.223\ V$

Q19.37C

Using the Nernst equation and a list of E°_{cell} values, calculate the E_{cell} for the following cells:

a. $Al(s)|Al^{3+}(0.36\ M)||Sn^{4+}(0.086\ M),\ Sn^{2+}(0.54\ M)|Pt$
b. $Sn(s)|Sn^{2+}(0.01\ M)||Pb^{2+}(0.700\ M)|Pb(s)$

S19.37C

a. $E^\circ_{cell} = 0.154 - (-1.676) = 1.83$

Using the Nernst equation $E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \times \log Q$, we simply need to set up the Q and establish that our n (moles of electrons transferred) is 2.

Thus, the $Q = \frac{(0.54)(0.36)^2}{(0.086)}$ and plugging in the found E°_{cell} gives us an E_{cell} value of 1.827 V

b. $E^\circ_{cell} = -0.125 - (-0.137) = 0.012\ V$

Using the Nernst equation $E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \times \log Q$, we now need to establish our Q. The overall equation would be $Pb^{2+}(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + Pb(s)$

Thus, the Q would be $\frac{0.01}{0.700}$ and the E_{cell} calculates out to 0.067V

For review on this topic, visit the page ***"Electrochemistry 4: The Nernst Equation"***.

Q19.41A

If $[Cu^{2+}]$ is maintained at 1.0 M, what is the minimum $[Ag^+]$ for which the reaction from 19.2, given below, is spontaneous in the forward direction?

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s) \qquad (18.11.91)$$

S19.41A

Remember - in order for the reaction to be spontaneous, $\Delta G < 0$.

Important equations:

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$

$$0 = E^\circ - \frac{RT}{nF} \ln Q \ \Rightarrow\ E^\circ = \frac{RT}{nF} \ln Q$$

$$Q = \frac{[Cu^{2+}]}{[Ag^+]^2}$$

Write the equation as a cell diagram:

$$Cu(s)|Cu^{2+}(aq)||Ag^{+}(aq)|Ag(s)$$

Oxidation: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-} \qquad E^{\circ} = -0.340$

Reduction: $(Ag^{+}(aq) + e^{-} \rightarrow Ag(s)) \times 2 \quad E^{\circ} = 0.800$

Net reaction: $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

$E^{\circ}_{cell} = 0.800 - 0.340 = 0.46\ V$

$$0.46 = \frac{RT}{2F} \ln \frac{[1]}{[Ag^{+}]^2}$$

$$0.46 = \frac{RT}{2F}(\ln 1 - \ln[Ag^{+}]^2)$$

$$0.46 = -\frac{RT}{2F} \ln[Ag^{+}]^2$$

$$0.46 = -\frac{(8.3145)(298)}{2(96485)} \ln[Ag^{+}]^2$$

Solve for $[Ag^{+}]$

$[Ag^{+}] = 1.66 \times 10^{-8}\ M$

Q19.41B

When $[Pb^{2+}]$ is maintained at 1.1 M for reaction:

$$Pb(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Pb^{2+}(aq)$$

a. What is the minimum $[Cu^{2+}]$ for reaction to be spontaneous in forward direction?

b. Will the displacement of $Cu^{2+}(aq)$ by $Pb(s)$ go to completion? Explain reasoning.

S19.41B

Oxidation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^{-} \quad E^{\circ}_{cell} = -0.125\ V$ Reduction: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s) \quad E^{\circ}_{cell} = +0.340\ V$

$E^{\circ}_{cell} = +0.340\ V - (-0.125\ V) = 0.465\ V$

a. Minimum $[Cu^{2+}]$ for spontaneous $= E_{cell} = 0\ V$ (positive E_{cell})

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \log \frac{[Pb^{2+}(aq)]}{[Cu^{2+}(aq)]}$$

$$0\ V = 0.465\ V - \frac{0.0592}{2} \log \frac{[1.1M]}{x}$$

$x = 2.15 \times 10^{-16}\ M$

Minimum $[Cu^{2+}{}_{(aq)}]$ for spontaneous reaction $= 2.15 \times 10^{-16}\ M$

b. Work the problem backwards,

Assume: $[Cu^{2+}]_{initial} = 1.1\ M$ and $[Pb^{2+}]_{initial} = 0.0\ M$,

then: $[Cu^{2+}]_{final} = 2.15 \times 10^{-16}\ M$ and $[Pb^{2+}]_{final} = 1.1\ M$.

Thus, reaction will go to completion.

Q19.41C

If $[Cu^{2+}]$ is maintained at 2.0 M

What is the minimum $[Ag^{+}]$ must be at to push the forward direction spontaneously. Use equation $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$.

S19.41C

To push a reaction forward spontaneously, the E_{cell} must be greater than 0. We find the E°_{cell} = 0.460 V.

Next, plug in the values given to the Nernst equation. The number of moles of electrons being transferred is 2. The equation will be set up as:

$$0 > 0.460 - \frac{0.0592}{2} \times \log \frac{2}{x^2}$$

Remember to account for coefficients in setting up the Q. Now solve for x $[Ag^{+}]$ and we find that $x = 2.4 \times 10^{-8}\ M$.

For review on this topic, visit the page ***"Electrochemistry 4: The Nernst Equation"***.

Q19.45A

Solve for E_{cell} of the following voltaic cell

$$Cu(s)|Cu^{2+}(0.01\ M)||Cu^{2+}(0.1\ M)|Cu(s)$$

S19.45A

First solve for E°_{cell}

Oxidation: $Cu(s) \rightarrow Cu^{2+} + 2e^- \quad E^\circ = -0.340$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \quad E^\circ = +0.340$

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$0.340 - 0.340 = 0$

Then use the Nernst Equation to solve for E_{cell}

$$E_{cell} = E^\circ_{cell} - \frac{0.0257}{2} \ln Q \tag{18.11.92}$$

$$= 0 - \frac{0.0257}{2} \ln \frac{0.01\ M}{0.1\ M} \tag{18.11.93}$$

$$E_{cell} = 0.0296\ V \tag{18.11.94}$$

Q19.45B

Voltaic cell: What is the value of E_{cell} when K_{sp} of $Na_2CrO_4 = 1.3 \times 10^{-12}$? $Na(s)|Na^+(aq)(satd\ Na_2CrO_4)||Na^+(0.130\ M)|Na(s)$

S19.45B

Find $[Na^+(aq)]$ in saturated solution of Na_2CrO_4 $K_{sp} = [Na^+]^2[CrO_4^{2-}] = (2s)^2(s) = 4s^3 = 1.3 \times 10^{-12} \quad s = 6.9 \times 10^{-5}\ M$
Concentration cell $= E^\circ_{cell} = 0\ V$ and $n = 1$, $[Na^+]_{anode} = 2s = 1.38 \times 10^{-4}\ M$ Cell reaction:

$Na(s) + Na^+(0.130\ M) \times Na(s) + Na^+(aq)(1.38 \times 10^{-4}\ M) \quad E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Na^+(satd\ Na_2CrO_4)]}{[Na^+]}$

$E_{cell} = 0\ V - \frac{0.0592}{1} \log \frac{1.38 \times 10^{-4}\ M}{0.130\ M} \quad E_{cell} = 0.176\ V$

Q19.45C

The following voltaic cell is constructed:

$Pb(s)||Pb^{2+}(satd\ PbI_2)||Pb^{2+}(0.100\ M)|Pb(s)$ Given that the $E_{cell} = 0.0567$ V, find the K_{sp}.

S19.45C

Set up the Nernst equation with the values appropriate for this equation. The moles of electrons transferred are 2. The E°_{cell} for this would be 0.

Thus, the Nernst equation sets up as $0.0567 = E^\circ_{cell} - \frac{0.0592}{n} \times \log Q$

$0.0567 = E^\circ_{cell} - \frac{0.0592}{n} \times \log \frac{x}{0.100}$

Solve for the value of x algebraically; this will yield an X value of 0.001215.

This value is not the answer; it is how much Pb^{2+} needs to be saturated in the equation. To find the K_{sp}, note the stoichiometric values for $[Pb]$ and $[I^-]^2$. The Pb and I concentrations have been found. Now plug in your x value and $K_{sp} = [Pb][I^-]^2$ and $K_{sp} = 1.79 \times 10^{-9}$.

For review on this topic, visit the page ***"Electrochemistry 4: The Nernst Equation"***.

Q19.59A

By referring back to figure 19-20, explain what would happen at each individual circumstance

a. zinc is wrapped around the head and tip of the iron nail
b. a hole is poked at the center of an iron nail
c. the nail is completely covered with copper

S19.59A

a. Through cathodic protection, zinc would get oxidized first. The zinc would protect the nail from oxidation.
b. Oxidation would occur even more because there would now be another head and tip. With more strained regions, the air would be able to oxidize the nail more.

c. The entire nail will be oxidized since copper isn't a sacrificial anode. It won't protect it from corrosion.

Q19.59B

What would happen to the corrosion of iron nail under phenolphtalein indicator, if:

a. Lead wire are wrapped around the nail
b. Deep scratch occurred at the iron nail
c. Galvanized nail is substituted for the iron nail

S19.59B

a. *Weaker reducing agent (lower tendency to get oxidized) = less reactive metal*. Since lead, Pb, is a weaker reducing agent than iron, oxidation of $\mathrm{Fe(s)}$ would be enhanced into $\mathrm{Fe^{2+}(aq)}$ - blue precipitate occurs. $\mathrm{OH^-(aq)}$ would be produced near the lead wire - pink color occurs (electron lost in oxidation).
b. Scratch exposes the metal, more susceptible to corrosion. Blue precipitate around the scratch should occur.
c. Zinc protect iron from corrosion as it is a stronger reducing agent (better tendency to be oxidized). Thus there should be no blue precipitate. Zinc corrodes instead. Pink color of $\mathrm{OH^-(aq)}$ would still continue to form.

Q19.59C

Imagine an iron nail that is corroding in a solution. Predict the appearance of the nail under the given conditions using your knowledge of corrosion in voltaic cells. Assume standard temperature conditions.

a. The head and tip of the nail are covered in magnesium.
b. The nail is galvanized, but there is a break in plating.
c. The nail is plated in copper, but there is a break in plating.

S19.59C

a. Magnesium is a sacrificial anode, meaning it is extremely willing to be oxidized. The $\mathrm{E^\circ_{cell}}$ for magnesium is smaller than the $\mathrm{E^\circ_{cell}}$ for iron, and it will be the metal that is targeted for oxidation. Thus, the head and tail of the nail would not corrode as long as magnesium is present.
b. If a metal is galvanized, that means it is coated with zinc, which is more active than the iron nail and will be oxidized. Even with a break in plating, the zinc is still more active and the iron won't corrode.
c. If the metal is plated in copper, once the copper is broken, the iron is the more active metal and the nail will be oxidized and begin corrosion.

For review on this topic, visit the page ***"Electrochemistry 7: Electrochemical Corrosion"***.

Q19.63A

Calculate the amount in grams of metal that is deposited at the cathode by running a current of 3.15 A for 78 min in an electrolysis reaction for an aqueous solution containing a) $\mathrm{Zn^{2+}}$ b) $\mathrm{Sn^{2+}}$ c) $\mathrm{Fe^{3+}}$ d) $\mathrm{Ni^{2+}}$

S19.63A

$$\#\text{ of mols of electrons} = \frac{(3.15\text{ A})(4680\text{ sec})}{(96485\text{ C})} = 0.153\text{ mol e}^-, \quad \text{since } 78\text{ min} \times \frac{60\text{ sec}}{1\text{ min}}$$

a. $0.153\text{ mol electrons} \times \frac{1\text{ mol Zn}}{2\text{ mol e}^-} = 0.0765\text{ mol Zn} \times \frac{65.41\text{ g}}{1\text{ mol}} = 5.00\text{ g Zn}$
b. $0.153\text{ mol electrons} \times \frac{1\text{ mol Sn}}{2\text{ mol e}^-} = 0.0765\text{ mol Sn} \times \frac{118.7\text{ g}}{1\text{ mol}} = 9.07\text{ g Sn}$
c. $0.153\text{ mol electrons} \times \frac{1\text{ mol Fe}}{3\text{ mol e}^-} = 0.051\text{ mol Fe} \times \frac{2.85\text{ g}}{1\text{ mol}} = 8.55\text{ g Fe}$
d. $0.153\text{ mol electrons} \times \frac{1\text{ mol Ni}}{2\text{ mol e}^-} = 0.0765\text{ mol Ni} \times \frac{58.69\text{ g}}{1\text{ mol}} = 4.49\text{ g Ni}$

Q19.63B

How many grams of metal are deposited at cathode by passage of 2.30 A of current for 75 minutes in electrolysis of aqueous solution containing:

a. $\mathrm{Mg^{2+}}$
b. $\mathrm{Sn^{4+}}$
c. $\mathrm{Fe^{3+}}$
d. $\mathrm{Ni^{2+}}$

S19.63B

$\text{mol e}^- = 75 \text{ minutes} \times \dfrac{60 \text{ sec}}{1 \text{ min}} \times \dfrac{2.30 \text{ C}}{1 \text{ sec}} \times \dfrac{1 \text{ mol e}^-}{96,485 \text{ C}} = 0.11 \text{ mol e}^-$ Mr for Mg = 24.31 Mr for Sn = 118.71 Mr for Fe = 55.85 Mr for Ni = 58.69

a. $\text{Mass Mg} = 0.11 \text{ mol e}^- \times \dfrac{1 \text{ mol Mg}^{2+}}{2 \text{ mol e}^-} \times (\text{Mr for Mg}) = 1.34 \text{ g Mg}$

b. $\text{Mass Sn} = 0.11 \text{ mol e}^- \times \dfrac{1 \text{ mol Sn}^{4+}}{4 \text{ mol e}^-} \times (\text{Mr for Sn}) = 3.26 \text{ g Sn}$

c. $\text{Mass Fe} = 0.11 \text{ mol e}^- \times \dfrac{1 \text{ mol Fe}^{3+}}{3 \text{ mol e}^-} \times (\text{Mr for Fe}) = 2.05 \text{ g Fe}$

d. $\text{Mass Ni} = 0.11 \text{ mol e}^- \times \dfrac{1 \text{ mol Ni}^{2+}}{2 \text{ mol e}^-} \times (\text{Mr for Ni}) = 3.23 \text{ g Ni}$

Q19.65A

Assuming all reactants and products are currently in their standard states, determine which of the following reactions occur spontaneously and which can occur only through the use of electrolysis. Also, for those that require electrolysis, determine what the minimum voltage required is.

a. $\text{Cu(s)} + \text{Zn}^{2+}\text{(aq)} \rightarrow \text{Cu}^{2+}\text{(aq)} + \text{Zn(s)}$
b. $2\text{Al} + 3\text{Cu}^{2+} \rightarrow 3\text{Cu} + 2\text{Al}^{3+}$
c. $\text{Zn} + \text{Cl}_2\text{(g)} \rightarrow \text{ZnCl}_2\text{(aq)}$
d. $2\text{Fe}^{3+} + 2\text{Cl}^- \rightarrow 2\text{Fe}^{2+} + \text{Cl}_2\text{(g)}$

S19.65A

a. Ox: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$ $\quad E^\circ = -0.340$ V

Red: $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$ $\quad E^\circ = -0.763$ V

Overall: $\text{Cu} + \text{Zn}^{2+} \rightarrow \text{Cu}^{2+} + \text{Zn}$ $\quad E^\circ = -1.103$

Since voltage is negative, it requires electrolysis with an applied voltage of greater than 1.103 V

b. Red: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ $\quad E^\circ = 0.340$ V

Ox: $\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$ $\quad E^\circ = 1.676$ V

Overall: $\text{Cu}^{2+} + \text{Al} \rightarrow \text{Cu} + \text{Al}^{3+}$ $\quad E^\circ = 2.016$ V

This is a spontaneous reaction under standard conditions because it has positive V.

c. Red: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$ $\quad E^\circ = 1.358$ V

Ox: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ $\quad E^\circ = 0.763$ V

Overall: $\text{Cl}_2 + \text{Zn} \rightarrow 2\text{Cl}^- + \text{Zn}^{2+}$ $\quad E^\circ = 2.121$ V

This is a spontaneous reaction under standard conditions because it has positive V.

d. Red: $\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$ $\quad E^\circ = 0.771$ V

Ox: $2\text{Cl}^- \rightarrow \text{Cl}_2\text{(g)} + 2\text{e}^-$ $\quad E^\circ = -1.358$ V

Overall: $\text{Fe}^{3+} + 2\text{Cl}^- \rightarrow \text{Fe}^{2+} + \text{Cl}_2\text{(g)}$ $\quad E^\circ = -0.587$ V

Since the V is negative, it requires electrolysis with an applied V of greater than 0.587 V.

Q19.65B

Which of the following reactions are spontaneous under standard conditions? If the reaction is not spontaneous, determine its minimum voltage applied to undergo electrolysis.

a. $\text{Sn}^{2+}\text{(aq)} + \text{Zn}^{2+}\text{(aq)} \rightarrow \text{Zn(s)} + \text{Sn}^{4+}\text{(aq)}$
b. $\text{Li}^+\text{(aq)} + \text{Fe}^{2+}\text{(aq)} \rightarrow \text{Li(s)} + \text{Fe}^{3+}\text{(aq)}$
c. $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)}\ [\text{in 1 M H}^+\text{(aq)}] \rightarrow 2\text{H}_2\text{O(l)}$
d. $2\text{Al}^{3+}\text{(aq)} + 3\text{Pb(s)} \rightarrow 3\text{Pb}^{2+}\text{(aq)} + 2\text{Al(s)}$

S19.65B

a. Anode: $Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^- \quad E^\circ_{cell} = +0.154\ V$

Cathode: $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \quad E^\circ_{cell} = -0.763\ V$
$E^\circ_{cell} = -0.763\ V - (+0.154\ V) = -0.917\ V$
Minimum voltage needed to be applied to undergo electrolysis $= +0.917\ V$

b. Anode: $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^- \quad E^\circ_{cell} = +0.771\ V$

Cathode: $Li^+(aq) + e^- \rightarrow Li(s) \quad E^\circ_{cell} = -3.040\ V$
$E^\circ_{cell} = -3.040\ V - (+0.771\ V) = -3.811\ V$
Minimum voltage needed to be applied to undergo electrolysis $= +3.811\ V$

c. Anode: $2H_2(g) \rightarrow 4H^+(aq) + 4e^- \quad E^\circ_{cell} = +0.000\ V$

Cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l) \quad E^\circ_{cell} = +1.229\ V$
$E^\circ_{cell} = +1.229\ V - (+0.000\ V) = +1.229\ V$
The reaction is spontaneous. No need undergo electrolysis.

d. Anode: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^- \quad E^\circ_{cell} = -0.125\ V$

Cathode: $Al^{3+}(aq) + 3e^- \rightarrow Al(s) \quad E^\circ_{cell} = -1.676\ V$
$E^\circ_{cell} = -1.676\ V - (-0.125\ V) = -1.551\ V$
Minimum voltage needed to be applied to undergo electrolysis $= +1.551\ V$

Q19.65C

Out of the reactions given below, which of the following can occur spontaneously only through electrolysis. For those requiring electrolysis, what is the minimum voltage required?

a. $I_2(s) + Ni(s) \rightarrow 2I^- + Ni^{2+}$
b. $Br_2(aq) + 2Fe^{2+} \rightarrow 2Br^- + 2Fe^{3+}$
c. $Ni^{2+} + 2Cl^-(aq) \rightarrow Ni(s) + Cl_2(g)$
d. $Cr^{2+}(aq) + Fe^{3+}(aq) \rightarrow Cr^{3+}(aq) + Fe^{2+}(aq)$

S19.65C

a. First, set up reduction and oxidation half reactions and calculate the E°_{cell} for the equation.

Reduction: $I_2 + 2e^- \rightarrow 2I^-$ and the value is 0.535

Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$ with a value of -0.257

Thus, the E°_{cell} is $0.535 - (-0.257) = 0.792$and it is spontaneous under current conditions.

b. Set up reduction and oxidation half reactions and calculate the E°_{cell} for the equation.

Reduction: $Br_2(aq) + 2e^- \rightarrow 2Br^-$

Oxidation: $2Fe^{2+} \rightarrow 2e^- + 2Fe^{3+}$

Thus, the E°_{cell} is 0.32 V and the E°_{cell} is positive, so no voltage needs to be applied and the cell is spontaneous under current conditions

c. Set up reduction and oxidation half reactions and calculate the E°_{cell} for the equation.

Reduction: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

Oxidation: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Thus, the E°_{cell} is -1.59. This E°_{cell} is negative. This means that we would need to apply a minimum of 1.59 V to force the reaction to start, and get the E°_{cell} to be positive.

d. Set up reduction and oxidation half reactions and calculate the E°_{cell} for the equation.

Reduction: $Fe^{3+} + e^- \rightarrow Fe^{2+}(aq)$

Oxidation: $Cr^{2+}(aq) \rightarrow Cr^{3+}(aq) + e^-$

Thus, the E°_{cell} is 1.195 V so it is spontaneous at present conditions and does not need any additional voltage.

For review on this topic, visit the page ***"Electrochemistry 8: Electrolytic cells and electrolysis"***.

Q19.72A

Consider the two following reduction half reactions:

$$\mathrm{Sn^{4+} + 2e^- \rightarrow Sn^{2+}} \tag{18.11.95}$$

$$\mathrm{Sn^{2+} + 2e^- \rightarrow Sn(s)} \tag{18.11.96}$$

Calculate the $\mathrm{E^\circ_{reduction}}$ for the reaction

$$\mathrm{Sn^{4+} + 4e^- \rightarrow Sn(s)} \tag{18.11.97}$$

S19.72A

We will need to combine these equations, but we can't simply add the $\mathrm{E^\circ_{cell}}$ values together. We will need to convert and find the $\Delta \mathrm{G}$ values for each equation. First, find the $\mathrm{E^\circ_{cell}}$ of the two given half reactions: they are 0.154 V, and -0.137 V respectively from Table P2.

The $\Delta \mathrm{G}$ of the first equation

$$\mathrm{\Delta G^\circ = n\,F\,E^\circ_{cell}} \tag{18.11.98}$$

Plugging in the values given will simplify to

$$\mathrm{= -2\,F\,(0.154)} \tag{18.11.99}$$

$$\mathrm{\Delta G^\circ = 2\,F\,(0.154)} \tag{18.11.100}$$

where F is Faraday's constant.

2nd equation $= -2^{*}\mathrm{F}^{*}(-0.137)$

$\Delta \mathrm{G}$ for the desired equation $= (-0.308\mathrm{F})\mathrm{V} + (0.274\mathrm{F})\mathrm{V}$

$\Delta \mathrm{G} = (-0.034\mathrm{F})\mathrm{V}$

Now to find the $\mathrm{E^\circ_{cell}} = \left(\dfrac{(-0.034\mathrm{F})\mathrm{V}}{4\mathrm{F}}\right)^{*}\mathrm{V} = -0.0085\ \mathrm{V} = \mathrm{E^\circ_{cell}}$

For review on this topic, visit the page ***"Electrochemistry 5: Applications of the Nernst Equation"***.

Q19.75A

In the two galvanic cells created the following variations occur:

$$\mathrm{L^{2+}(aq) + LO^{2+}(aq) + 2H^+(aq) \rightarrow 2L^{3+}(aq) + H_2O(l)} \qquad \mathrm{E^\circ = 0.616\ V} \tag{18.11.101}$$

$$\mathrm{L^{3+}(aq) + Ag^+(aq) + H_2O \rightarrow LO^{2+}(aq) + 2H^+ + Ag(s)} \qquad \mathrm{E^\circ = 0.439\ V} \tag{18.11.102}$$

Use other values as well as the data presented here to determine $\mathrm{E^\circ}$ for the half reaction:

$$\mathrm{L^{3+} + e^- \rightarrow L^{2+}} \tag{18.11.103}$$

S19.75A

Ox: $\mathrm{L^{3+} + H_2O \rightarrow LO^{2+} + 2H^+ + e^-}$ $\quad -\mathrm{E^\circ_a}$
Red: $\mathrm{Ag^+ + e^- \rightarrow Ag(s)}$ $\quad \mathrm{E^\circ = 0.800\ V}$

Overall: $\quad \mathrm{E^\circ = 0.439\ V}$

$0.439\ \mathrm{V} = -\mathrm{E_a} + 0.800\ \mathrm{V}$

$\mathrm{E_a} = 0.361\ \mathrm{V}$

Ox: $\mathrm{L^{2+} \rightarrow L^{3+} + e^-}$ $\quad -\mathrm{E^\circ_b}$
Red: $\mathrm{LO^{2+} + 2H^+ + e^- \rightarrow L^{3+} + H_2O}$ $\quad \mathrm{E^\circ = 0.361\ V}$

Overall: $\quad \mathrm{E^\circ = 0.616\ V}$

$0.616\ \mathrm{V} = -\mathrm{E_b} + 0.361\ \mathrm{V}$

$\mathrm{E_b} = -0.255\ \mathrm{V}$

Q19.75B

Using reactions below, calculate E_{cell} for $X^{3+}(aq) + e^- \rightarrow X^{2+}(aq)$.

$X^{2+}(aq) + XO^{2+}(aq) + 2H^+(aq) \rightarrow 2X^{3+}(aq) + H_2O(l)$ $\quad E^\circ_{cell} = 0.789\ V$

$X^{3+}(aq) + Fe^{3+}(aq) + H_2O(l) \rightarrow XO^{2+}(aq) + 2H^+(aq) + Fe^{2+}(aq)$ $\quad E^\circ_{cell} = 0.243\ V$

S19.75B

First we must know that what are the reactions on the both side

Oxidation: $X^{3+}(aq) + H_2O(l) \rightarrow XO^{2+}(aq) + 2H^+(aq) + e^-$ $\quad -E^\circ_a$

Reduction: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $\quad E^\circ_{cell} = +0.771\ V$

Then we can get the net equation by combing these two reactions:

Net: $X^{3+}(aq) + H_2O(l) + Fe^{3+}(aq) \rightarrow XO^{2+}(aq) + 2H^+(aq) + Fe^{2+}(aq)$ $\quad E^\circ_{cell} = +0.243\ V$

As we know that $E^\circ_{cell} = E^\circ_c - E^\circ_a$, and in this case

$$E^\circ_{cell} = +0.771\ V - E^\circ_a$$
$$0.243\ V = +0.771\ V - E^\circ_a$$
$$E^\circ_a = 0.528\ V$$

Thus, now we get: Oxidation: $X^{2+}(aq) \rightarrow X^{3+}(aq) + e^-$ $\quad -E^\circ_b$

Reduction: $XO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow X^{3+}(aq) + H_2O(l)$ $\quad E^\circ_a = +0.528\ V$

Net: $X^{2+}(aq) + XO^{2+}(aq) + 2H^+(aq) \rightarrow 2X^{3+}(aq) + H_2O(l)$ $\quad E^\circ_{cell} = +0.789\ V$

Then we use the equation again:

$E^\circ_{cell} = 0.528\ V - E^\circ_b$ $0.789\ V = 0.528\ V - E^\circ_b E^\circ_b = -0.261\ V$ Thus $X^{3+}(aq) + e^- \rightarrow X^{2+}(aq)$ $\quad E^\circ_{cell} = -0.261\ V$

Q19.82A

A fuel cells is based on the reaction

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \tag{18.11.104}$$

The theoretical E°_{cell} of this fuel cells is 0.92 V. Calculate a value of ΔG°_f for $[CH_4(g)]$.

S19.82A

Calculate ΔG°: $n = 8$ (8 moles of e^- transferred)

$\Delta G^\circ = -nFE^\circ_{cell}$

$$\Delta G^\circ = -8\ mol\ e^-(96845\ C/mol\ e^-)(0.92\ V) \tag{18.11.105}$$
$$= -7.101 \times 10^5\ J = -710.1\ KJ \tag{18.11.106}$$

Now to determine ΔG°_f for $[CH_4(g)]$

$-710.0\ KJ = \Delta G^\circ_f[CO_2(g)] + 2\Delta G^\circ_f[H_2O(l)] - \Delta G^\circ_f[CH_4(g)] - 2\Delta G^\circ_f[O_2(g)]$

$-710.0\ KJ = -394.4\ KJ + 2(-237.1\ KJ) - \Delta G^\circ_f[CH_4(g)] - 0$

$-158.5\ KJ = \Delta G^\circ_f[CH_4(g)]$

Q19.82B

Calculate ΔG°_f for $[CH_4]$ using $E^\circ_{cell} = 1.191\ V$ for fuel cell reaction: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ Given that:

$\Delta G^\circ_f\ [CO_2(g)] = -394.4\ kJ\ mol^{-1}$

$\Delta G^\circ_f\ [H_2O(l)] = -237.1\ kJ\ mol^{-1}$ $\Delta G^\circ_f\ [O_2(g)] = 0.00\ kJ\ mol^{-1}$

S19.82B

$\Delta G^\circ = -nFE^\circ_{cell} = \sum \Delta G^\circ_f$ $\quad \Delta G^\circ = -(8e^-)(96,485\ C/mol\ e^-)(1.191\ V)(1/1000\ kJ) = -919.31\ kJ/mol$

$\Delta G^\circ = \Delta G^\circ_f\ [CH_4(g)] + 2\Delta G^\circ_f\ [O_2(g)] - \Delta G^\circ_f\ [CO_2(g)] - 2\Delta G^\circ_f\ [H_2O(l)]$ $\quad \Delta G^\circ = \Delta G^\circ_f\ [CH_4(g)] + 2(0) - 394.4 + 2(-237.1)$

$-919.31 = \Delta G^\circ_f\ [CH_4(g)] + -868.6 \Delta G^\circ_f\ [CH_4(g)] = -50.71\ kJ\ mol^{-1}$

Q19.82C

The E°_{cell} is 0.323 V for the following reaction under non-standard conditions

$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s) \tag{18.11.107}$$

Calculate a ΔG value for $[Fe^{2+}]$.

S19.82C

Using the Nernst equation, we find the ΔG for the overall reaction.

The Nernst equation sets up as $\Delta G = -zFE^{\circ}_{cell}$ and we can plug in our values. The moles of electrons transferred in the reaction is 2. After plugging in 2 for z and the given E°_{cell} along with Faraday's constant, we find that $\Delta G = -62.3283$ KJ.

Next, we find the ΔG stepwise by using our reaction along with a table of ΔG values found in the back of your textbook.

We must realize that the half reactions are Zn(s)->Zn 2+ +2e and 2e+ Fe2+->Fe(s). According to the table we can find the the values for the reactions in regard to their direction which are 0.763V of the cathode (Zn being oxidized) and -0.440V of the anode (Fe2+ being reduced).

This sets up as $\Delta G = \Delta G[Zn^{2+}] - \Delta G[Fe^{2+}]$ because we don't account for the solids.

We must then plug in -62.3283 kj for Delta G and multiply by these concentrations. Solving for Fe2+, we can conclude that it's Delta G is -84.77 kj. We cannot come to this conclusion without using the Half reactions table in the back. We are also unable to find these answers without finding the overall Delta G for the reaction as a whole.

The resulting $[Fe^{2+}]$ value is -84.77 KJ so in this specific reaction, this is the desired ΔG.

The link below provides additional examples of solving Nearst Equations in various forms:

www.youtube.com/watch?v=iKNFEjYDnts

For review on this topic, visit the page ***"Electrochemistry 8: Electrolytic cells and electrolysis"***.

Q19.90A

$Pb^{2+}(aq)$ and $Ag^{+}(aq)$ were produced when 0.978 g sample of nitric acid was dissolved. After diluting the solution to 250.0 mL of deionized water, a silver electrode was placed in solution. The potential difference read 0.273 V. Calculate the mass of Ag in the sample as a percent.

S19.90A

Oxidation: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ $\quad -E^{\circ} = 0.00$ V
Reduction: $(Ag^+(aq) + e^- \rightarrow Ag(s)) \times 2$ $\quad E^{\circ} = 0.800$ V
Net: $H_2(g) + 2Ag^+(aq) \rightarrow 2H^+(aq) + 2Ag(s)$ $\quad E^{\circ}_{cell} = 0.800$ V

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{2}\log\frac{[H^+]^2}{[Ag^+]^2} \tag{18.11.108}$$

$$= 0.800\text{ V} - \frac{0.0592}{2}\log\frac{[1.00]^2}{[Ag^+]^2} = 0.273\text{ V} \tag{18.11.109}$$

$$\log\frac{1}{[Ag^+]^2} = 17.804$$

$$[Ag^+] = 1.25 \times 10^{-9}$$

$$\text{mass Ag} = 250\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{1.25 \times 10^{-9}\text{ mol Ag}^+}{1\text{ L}} \times \frac{1\text{ mol Ag}}{1\text{ mol Ag}^+} \times \frac{107.87\text{ Ag}}{1\text{ mol Ag}} = 3.38 \times 10^{-8}\text{ g Ag}$$

$$\%\text{ Ag} = \frac{3.38 \times 10^{-8}\text{ g Ag}}{0.978\text{ g}} \times 100 = 3.46 \times 10^{-6}\%$$

Q19.90B

% of Ag in Pb was: 1.065 g sample was dissolved in nitric acid to produce Pb^{2+} and Ag^+. The solution was then diluted to 400 mL with water. An Ag electrode was immersed in solution and the potential difference between electrode and SHE was 0.604 V. What was % Ag by mass in the Pb material?

S19.90B

Oxidation: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ $\quad E^{\circ}_{cell} = +0.00$ V
Reduction: $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$ $\quad E^{\circ}_{cell} = +0.800$ V

$$E^{\circ}_{cell} = +0.800\ V - (+0.00\ V) = 0.800\ V$$

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \log \frac{[H^+]^2}{[Ag^+]^2}$$

$$0.604\ V = 0.800\ V - \frac{0.0592}{2} \log \frac{[1.00]^2}{[Ag^+]^2} [Ag^+] = 4.89 \times 10^{-4}\ M$$

$$\text{Mass Ag } = 0.400\ L \times \frac{[Ag^+]}{1\ L} \times \frac{1\ mol\ Ag}{1\ mol\ Ag^+} \times \text{Mr Ag} \tag{18.11.110}$$

$$= 0.400\ L \times (4.89 \times 10^{-4}\ M) \times \frac{Ag^+}{1\ L} \times \frac{1\ mol\ Ag}{1\ mol\ Ag^+} \times 107.87\ g \tag{18.11.111}$$

$$= 0.0211\ \text{grams} \tag{18.11.112}$$

$$\%Ag = \frac{\text{mass Ag}}{\text{mass of sample}} \times 100\% = \frac{0.0211}{1.065} \times 100\% = 1.98\%$$

Q19.96A

A source of Cu is recovered from lead. The percentage of Cu in lead is determined from the following information: A 2 gram sample is dissolved in nitric acid to make Cu^{2+} and Pb^{2+}. A copper electrode was immersed in 750 mL of water. The potential was found to be 0.250 V. What was the % Copper recovered from the lead metal? Assume that Pb^{2+} is ignored by the copper electrode.

S19.96A

Identify the oxidation/reduction reactions first.

Oxidation: $H_2(g) \rightarrow 2H^+(aq) + 2e^-$ with an E°_{cell} value of 0

Reduction: $Cu^{2+} + 2e^- \rightarrow Cu(s)$ with an E°_{cell} value of 0.340

Thus the E°_{cell} is 0.340. We know that the E_{cell} is 0.250 V so we can use the Nernst equation.

This will set up as $0.250\ V = E^{\circ}_{cell} - \frac{0.0592}{n} \times \log Q$

Further values inputted will result in $0.250 = 0.340 - \frac{0.0592}{2} \times \log \frac{[I]^2}{[Cu^{2+}]}$

Solve for the value of Cu^{2+} using algebraic methods and we find that the concentration is 0.000911. We use this concentration and account for the mass in grams.

$$0.750\ L\ (\text{volume of solution}) \times 0.000911 \times \frac{1\ mol\ Cu}{1\ mol\ Cu^{2+}} \times \frac{63.546\ g}{1\ mol\ Cu} \times \frac{1}{2}\ \text{gram sample present} \times 100\% \tag{18.11.113}$$

$$= 2.17\%\ \text{copper results by mass} \tag{18.11.114}$$

For review on this topic, visit the page ***"Electrochemistry 5: Applications of the Nernst Equation"***.

This page titled 18.11: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang via source content that was edited to the style and standards of the LibreTexts platform.

CHAPTER OVERVIEW

19: Transition Metals and Coordination Chemistry

Transition metals are defined as those elements that have (or readily form) partially filled *d* orbitals. These include the *d*-block (groups 3–11) and f-block element elements. The variety of properties exhibited by transition metals is due to their complex valence shells. Unlike most main group metals where one oxidation state is normally observed, the valence shell structure of transition metals means that they usually occur in several different stable oxidation states. In addition, electron transitions in these elements can correspond with absorption of photons in the visible electromagnetic spectrum, leading to colored compounds. Because of these behaviors, transition metals exhibit a rich and fascinating chemistry.

Faculty Feature

We highlight Dr. Bernie Kraatz's innovative research, which spans peptide hydrogels, prebiotic self-assembly, enzymatic sensors, and CO_2 electrochemical reduction. His work demonstrates the critical role of transition metals in coordination chemistry, catalysis, and sensor development. Dr. Kraatz's research exemplifies the transformative applications of transition metal chemistry, showcasing its relevance to cutting-edge scientific advancements.

Read more about Dr. Bernie Kraatz and his research here: https://www.utsc.utoronto.ca/people/bkraatz/the-group/



19.1: Introduction
19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds
19.3: Coordination Chemistry of Transition Metals
19.4: Spectroscopic and Magnetic Properties of Coordination Compounds
19.5: Key Terms
19.6: Summary
19.7: Exercises

This page titled 19: Transition Metals and Coordination Chemistry is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis, Anna Galang, and Amir Kanbar.

19.1: Introduction

In this chapter, we delve into the captivating chemistry of transition metals. Their intricate valence shell structures result in a variety of unique properties, such as multiple stable oxidation states and the ability to form colored compounds through electron transitions that absorb visible light. These distinctive characteristics make the study of transition metals a fundamental and intriguing aspect of inorganic chemistry.

As you delve into the intricate world of transition metals, consider how the principles you learn are applied in real-world scientific research. Dr. Kraatz's work exemplifies the practical and transformative impact of inorganic chemistry, providing a window into the innovative applications of transition metal chemistry in various scientific fields.

This page titled 19.1: Introduction is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds

Transition metals are defined as those elements that have (or readily form) partially filled *d* orbitals. As shown in Figure 19.2.1, the *d*-block elements in groups 3–11 are transition elements. The *f*-block elements, also called *inner transition metals* (the lanthanides and actinides), also meet this criterion because the *d* orbital is partially occupied before the *f* orbitals. The *d* orbitals fill with the copper family (group 11); for this reason, the next family (group 12) are technically not transition elements. However, the group 12 elements do display some of the same chemical properties and are commonly included in discussions of transition metals. Some chemists do treat the group 12 elements as transition metals.

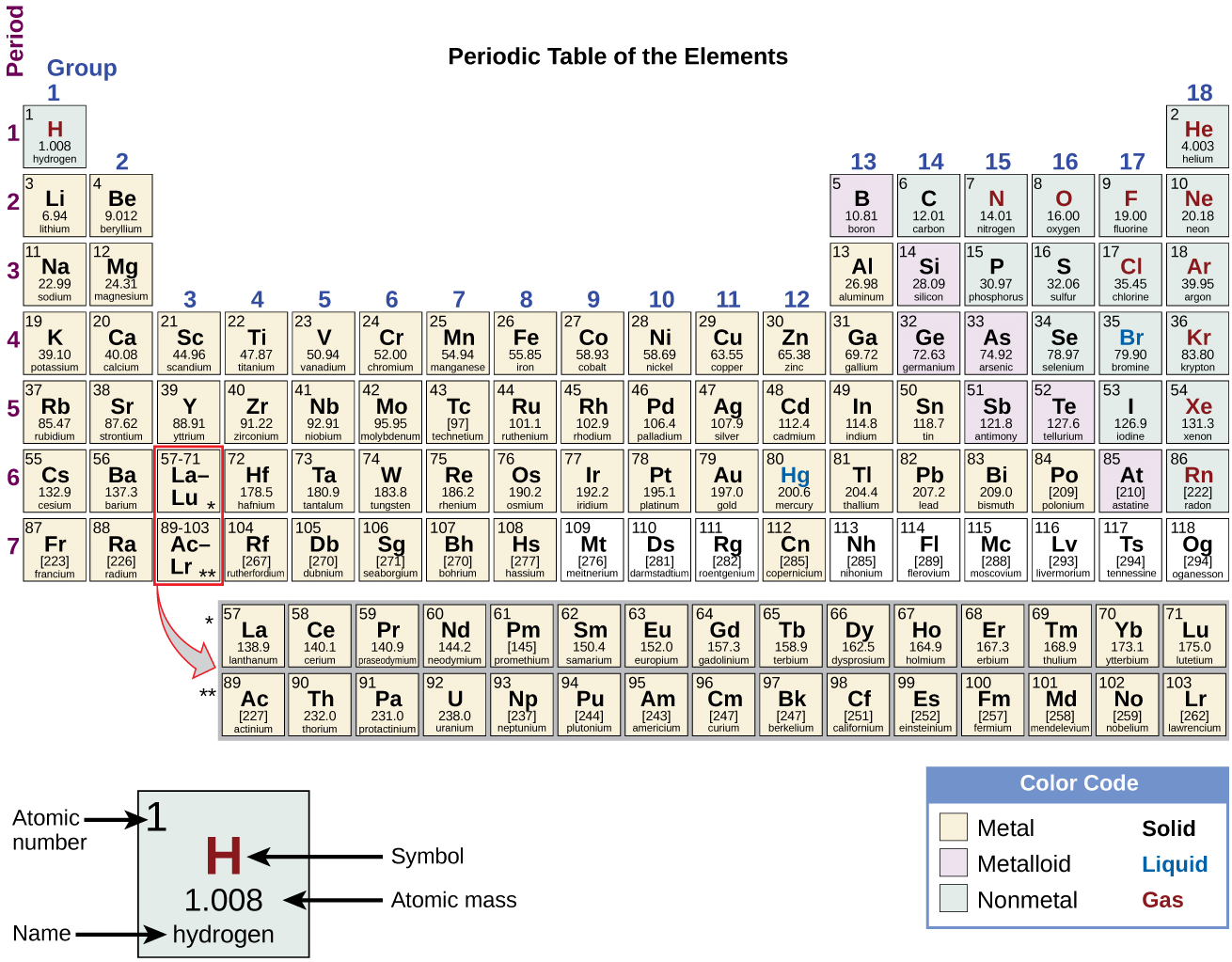


Figure 19.2.1: The transition metals are located in groups 3–11 of the periodic table. The inner transition metals are in the two rows below the body of the table.

The *d*-block elements are divided into the first transition series (the elements Sc through Cu), the second transition series (the elements Y through Ag), and the third transition series (the element La and the elements Hf through Au). Actinium, Ac, is the first member of the fourth transition series, which also includes Rf through Rg.

The *f*-block elements are the elements Ce through Lu, which constitute the lanthanide series (or lanthanoid series), and the elements Th through Lr, which constitute the actinide series (or actinoid series). Because lanthanum behaves very much like the lanthanide elements, it is considered a lanthanide element, even though its electron configuration makes it the first member of the third transition series. Similarly, the behavior of actinium means it is part of the actinide series, although its electron configuration makes it the first member of the fourth transition series.

Example 19.2.1: Valence Electrons in Transition Metals

Review how to write electron configurations, covered in the chapter on electronic structure and periodic properties of elements. Recall that for the transition and inner transition metals, it is necessary to remove the *s* electrons before the *d* or *f* electrons. Then, for each ion, give the electron configuration:

a. cerium(III)
b. lead(II)
c. Ti^{2+}
d. Am^{3+}
e. Pd^{2+}

For the examples that are transition metals, determine to which series they belong.

Solution

For ions, the *s*-valence electrons are lost prior to the *d* or *f* electrons.

a. $Ce^{3+}[Xe]4f^1$; Ce^{3+} is an inner transition element in the lanthanide series.
b. $Pb^{2+}[Xe]6s^25d^{10}4f^{14}$; the electrons are lost from the *p* orbital. This is a main group element.
c. titanium(II) $[Ar]3d^2$; first transition series
d. americium(III) $[Rn]5f^6$; actinide
e. palladium(II) $[Kr]4d^8$; second transition series

Exercise 19.2.1

Give an example of an ion from the first transition series with no *d* electrons.

Answer

V^{5+} is one possibility. Other examples include Sc^{3+}, Ti^{4+}, Cr^{6+}, and Mn^{7+}.

Chemistry in Everyday Life: Uses of Lanthanides in Devices

Lanthanides (elements 57–71) are fairly abundant in the earth's crust, despite their historic characterization as rare earth elements. Thulium, the rarest naturally occurring lanthanoid, is more common in the earth's crust than silver (4.5×10^{-5} % versus 0.79×10^{-5} % by mass). There are 17 rare earth elements, consisting of the 15 lanthanoids plus scandium and yttrium. They are called rare because they were once difficult to extract economically, so it was rare to have a pure sample; due to similar chemical properties, it is difficult to separate any one lanthanide from the others. However, newer separation methods, such as ion exchange resins similar to those found in home water softeners, make the separation of these elements easier and more economical. Most ores that contain these elements have low concentrations of all the rare earth elements mixed together.

The commercial applications of lanthanides are growing rapidly. For example, europium is important in flat screen displays found in computer monitors, cell phones, and televisions. Neodymium is useful in laptop hard drives and in the processes that convert crude oil into gasoline (Figure 19.2.2). Holmium is found in dental and medical equipment. In addition, many alternative energy technologies rely heavily on lanthanoids. Neodymium and dysprosium are key components of hybrid vehicle engines and the magnets used in wind turbines.

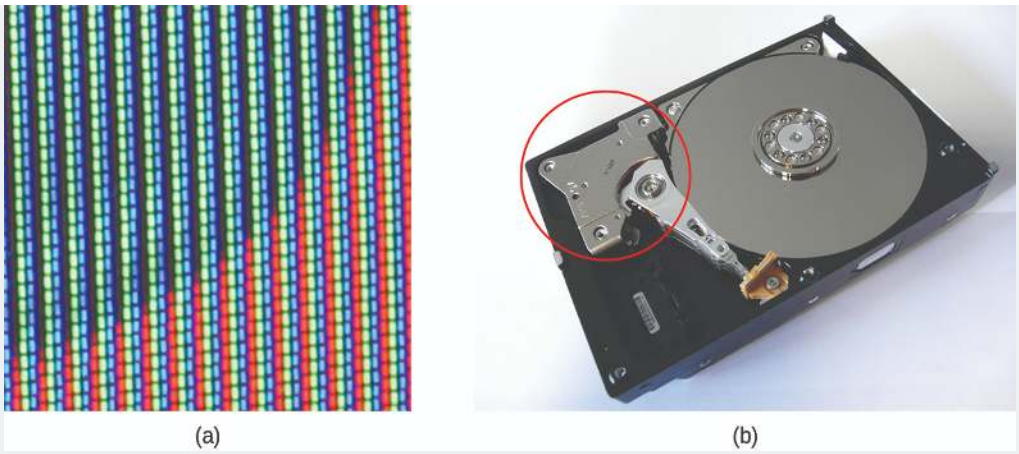


Figure 19.2.2: (a) Europium is used in display screens for televisions, computer monitors, and cell phones. (b) Neodymium magnets are commonly found in computer hard drives. (credit b: modification of work by "KUERT Datenrettung"/Flickr)

As the demand for lanthanide materials has increased faster than supply, prices have also increased. In 2008, dysprosium cost $110/kg; by 2014, the price had increased to $470/kg. Increasing the supply of lanthanoid elements is one of the most significant challenges facing the industries that rely on the optical and magnetic properties of these materials.

The transition elements have many properties in common with other metals. They are almost all hard, high-melting solids that conduct heat and electricity well. They readily form alloys and lose electrons to form stable cations. In addition, transition metals form a wide variety of stable **coordination compounds**, in which the central metal atom or ion acts as a Lewis acid and accepts one or more pairs of electrons. Many different molecules and ions can donate lone pairs to the metal center, serving as Lewis bases. In this chapter, we shall focus primarily on the chemical behavior of the elements of the first transition series.

Properties of the Transition Elements

Transition metals demonstrate a wide range of chemical behaviors. As can be seen from their reduction potentials (see Appendix H), some transition metals are strong reducing agents, whereas others have very low reactivity. For example, the lanthanides all form stable 3+ aqueous cations. The driving force for such oxidations is similar to that of alkaline earth metals such as Be or Mg, forming Be^{2+} and Mg^{2+}. On the other hand, materials like platinum and gold have much higher reduction potentials. Their ability to resist oxidation makes them useful materials for constructing circuits and jewelry.

Ions of the lighter *d*-block elements, such as Cr^{3+}, Fe^{3+}, and Co^{2+}, form colorful hydrated ions that are stable in water. However, ions in the period just below these (Mo^{3+}, Ru^{3+}, and Ir^{2+}) are unstable and react readily with oxygen from the air. The majority of simple, water-stable ions formed by the heavier *d*-block elements are oxyanions such as and

Ruthenium, osmium, rhodium, iridium, palladium, and platinum are the **platinum metals**. With difficulty, they form simple cations that are stable in water, and, unlike the earlier elements in the second and third transition series, they do not form stable oxyanions.

Both the *d*- and *f*-block elements react with nonmetals to form binary compounds; heating is often required. These elements react with halogens to form a variety of halides ranging in oxidation state from 1+ to 6+. On heating, oxygen reacts with all of the transition elements except palladium, platinum, silver, and gold. The oxides of these latter metals can be formed using other reactants, but they decompose upon heating. The *f*-block elements, the elements of group 3, and the elements of the first transition series except copper react with aqueous solutions of acids, forming hydrogen gas and solutions of the corresponding salts.

Transition metals can form compounds with a wide range of oxidation states. Some of the observed oxidation states of the elements of the first transition series are shown in Figure 19.2.3. As we move from left to right across the first transition series, we see that the number of common oxidation states increases at first to a maximum towards the middle of the table, then decreases. The values in the table are typical values; there are other known values, and it is possible to synthesize new additions. For example, in 2014, researchers were successful in synthesizing a new oxidation state of iridium (9+).

21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn
								1+	
		2+	2+	2+	2+	2+	2+	2+	2+
3+	3+	3+	3+	3+	3+	3+	3+	3+	
	4+	4+	4+	4+					
		5+							
			6+	6+	6+				
				7+					

Figure 19.2.3: Transition metals of the first transition series can form compounds with varying oxidation states.

For the elements scandium through manganese (the first half of the first transition series), the highest oxidation state corresponds to the loss of all of the electrons in both the *s* and *d* orbitals of their valence shells. The titanium(IV) ion, for example, is formed when the titanium atom loses its two 3*d* and two 4*s* electrons. These highest oxidation states are the most stable forms of scandium, titanium, and vanadium. However, it is not possible to continue to remove all of the valence electrons from metals as we continue through the series. Iron is known to form oxidation states from 2+ to 6+, with iron(II) and iron(III) being the most common. Most of the elements of the first transition series form ions with a charge of 2+ or 3+ that are stable in water, although those of the early members of the series can be readily oxidized by air.

The elements of the second and third transition series generally are more stable in higher oxidation states than are the elements of the first series. In general, the atomic radius increases down a group, which leads to the ions of the second and third series being larger than are those in the first series. Removing electrons from orbitals that are located farther from the nucleus is easier than removing electrons close to the nucleus. For example, molybdenum and tungsten, members of group 6, are limited mostly to an oxidation state of 6+ in aqueous solution. Chromium, the lightest member of the group, forms stable Cr^{3+} ions in water and, in the absence of air, less stable Cr^{2+} ions. The sulfide with the highest oxidation state for chromium is Cr_2S_3, which contains the Cr^{3+} ion. Molybdenum and tungsten form sulfides in which the metals exhibit oxidation states of 4+ and 6+.

Example 19.2.2: Activity of the Transition Metals

Which is the strongest oxidizing agent in acidic solution: dichromate ion, which contains chromium(VI), permanganate ion, which contains manganese(VII), or titanium dioxide, which contains titanium(IV)?

Solution

First, we need to look up the reduction half reactions (in Appendix L) for each oxide in the specified oxidation state:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O \quad +1.33V$$
$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + H_2O \quad +1.51V$$
$$TiO_2 + 4H^+ + 2e^- \longrightarrow Ti^{2+} + 2H_2O \quad -0.50V$$

A larger reduction potential means that it is easier to reduce the reactant. Permanganate, with the largest reduction potential, is the strongest oxidizer under these conditions. Dichromate is next, followed by titanium dioxide as the weakest oxidizing agent (the hardest to reduce) of this set.

Exercise 19.2.2

Predict what reaction (if any) will occur between HCl and Co(*s*), and between HBr and Pt(*s*). You will need to use the standard reduction potentials from Appendix L.

Answer

$$\mathrm{Co(s) + 2\,HCl \longrightarrow H_2 + CoCl_2(aq)}$$

no reaction because Pt(*s*) will not be oxidized by H^+

Preparation of the Transition Elements

Ancient civilizations knew about iron, copper, silver, and gold. The time periods in human history known as the Bronze Age and Iron Age mark the advancements in which societies learned to isolate certain metals and use them to make tools and goods.

Naturally occurring ores of copper, silver, and gold can contain high concentrations of these metals in elemental form (Figure 19.2.4). Iron, on the other hand, occurs on earth almost exclusively in oxidized forms, such as rust (Fe_2O_3). The earliest known iron implements were made from iron meteorites. Surviving iron artifacts dating from approximately 4000 to 2500 BC are rare, but all known examples contain specific alloys of iron and nickel that occur only in extraterrestrial objects, not on earth. It took thousands of years of technological advances before civilizations developed iron **smelting**, the ability to extract a pure element from its naturally occurring ores and for iron tools to become common.



Figure 19.2.4: Transition metals occur in nature in various forms. Examples include (a) a nugget of copper, (b) a deposit of gold, and (c) an ore containing oxidized iron. (credit a: modification of work by http://images-of-elements.com/copper-2.jpg; credit c: modification of work by http://images-of-elements.com/iron-ore.jpg)

Generally, the transition elements are extracted from minerals found in a variety of ores. However, the ease of their recovery varies widely, depending on the concentration of the element in the ore, the identity of the other elements present, and the difficulty of reducing the element to the free metal.

In general, it is not difficult to reduce ions of the *d*-block elements to the free element. Carbon is a sufficiently strong reducing agent in most cases. However, like the ions of the more active main group metals, ions of the *f*-block elements must be isolated by electrolysis or by reduction with an active metal such as calcium.

We shall discuss the processes used for the isolation of iron, copper, and silver because these three processes illustrate the principal means of isolating most of the *d*-block metals. In general, each of these processes involves three principal steps: preliminary treatment, smelting, and refining.

1. Preliminary treatment. In general, there is an initial treatment of the ores to make them suitable for the extraction of the metals. This usually involves crushing or grinding the ore, concentrating the metal-bearing components, and sometimes treating these substances chemically to convert them into compounds that are easier to reduce to the metal.
2. Smelting. The next step is the extraction of the metal in the molten state, a process called smelting, which includes reduction of the metallic compound to the metal. Impurities may be removed by the addition of a compound that forms a slag—a substance with a low melting point that can be readily separated from the molten metal.
3. Refining. The final step in the recovery of a metal is refining the metal. Low boiling metals such as zinc and mercury can be refined by distillation. When fused on an inclined table, low melting metals like tin flow away from higher-melting impurities. Electrolysis is another common method for refining metals.

Isolation of Iron

The early application of iron to the manufacture of tools and weapons was possible because of the wide distribution of iron ores and the ease with which iron compounds in the ores could be reduced by carbon. For a long time, charcoal was the form of carbon used in the reduction process. The production and use of iron became much more widespread about 1620, when coke was introduced as the reducing agent. Coke is a form of carbon formed by heating coal in the absence of air to remove impurities.

The first step in the metallurgy of iron is usually roasting the ore (heating the ore in air) to remove water, decomposing carbonates into oxides, and converting sulfides into oxides. The oxides are then reduced in a blast furnace that is 80–100 feet high and about 25 feet in diameter (Figure 19.2.5) in which the roasted ore, coke, and limestone (impure $CaCO_3$) are introduced continuously into the top. Molten iron and slag are withdrawn at the bottom. The entire stock in a furnace may weigh several hundred tons.

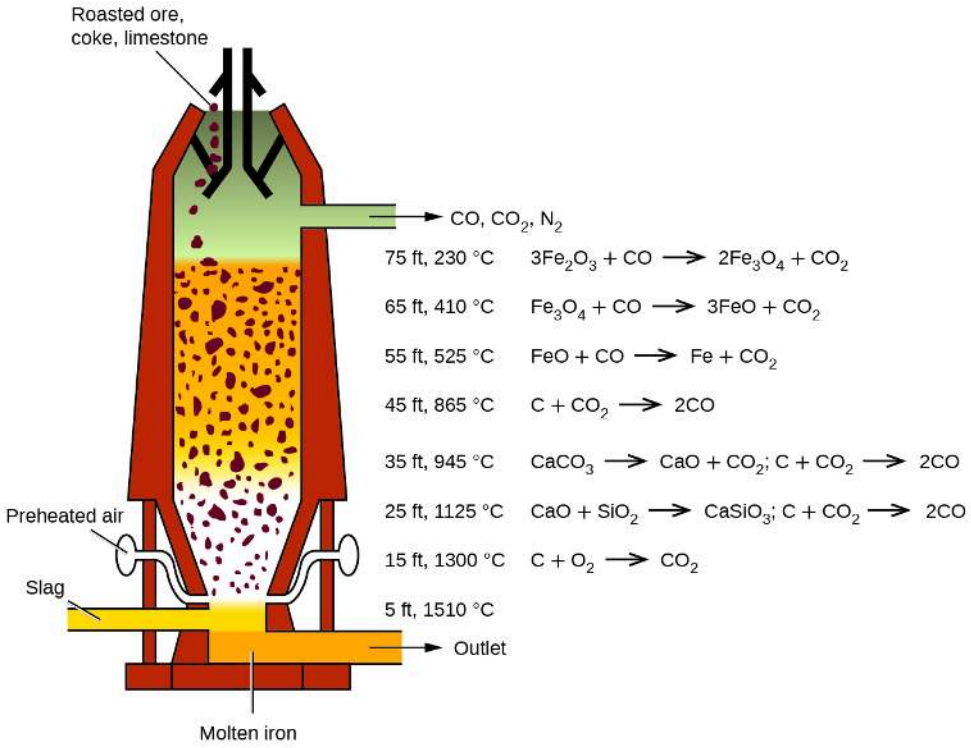


Figure 19.2.5: Within a blast furnace, different reactions occur in different temperature zones. Carbon monoxide is generated in the hotter bottom regions and rises upward to reduce the iron oxides to pure iron through a series of reactions that take place in the upper regions.

Near the bottom of a furnace are nozzles through which preheated air is blown into the furnace. As soon as the air enters, the coke in the region of the nozzles is oxidized to carbon dioxide with the liberation of a great deal of heat. The hot carbon dioxide passes upward through the overlying layer of white-hot coke, where it is reduced to carbon monoxide:

$$\mathrm{CO_2(g) + C(s) \longrightarrow 2\,CO(g)}$$

The carbon monoxide serves as the reducing agent in the upper regions of the furnace. The individual reactions are indicated in Figure 19.2.5

The iron oxides are reduced in the upper region of the furnace. In the middle region, limestone (calcium carbonate) decomposes, and the resulting calcium oxide combines with silica and silicates in the ore to form slag. The slag is mostly calcium silicate and contains most of the commercially unimportant components of the ore:

$$\mathrm{CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)}$$

Just below the middle of the furnace, the temperature is high enough to melt both the iron and the slag. They collect in layers at the bottom of the furnace; the less dense slag floats on the iron and protects it from oxidation. Several times a day, the slag and molten iron are withdrawn from the furnace. The iron is transferred to casting machines or to a steelmaking plant (Figure 19.2.6).



Figure 19.2.6: Molten iron is shown being cast as steel. (credit: Clint Budd)

Much of the iron produced is refined and converted into steel. **Steel** is made from iron by removing impurities and adding substances such as manganese, chromium, nickel, tungsten, molybdenum, and vanadium to produce alloys with properties that make the material suitable for specific uses. Most steels also contain small but definite percentages of carbon (0.04%–2.5%). However, a large part of the carbon contained in iron must be removed in the manufacture of steel; otherwise, the excess carbon would make the iron brittle.

Link to Learning

You can watch an animation of steelmaking that walks you through the process.

Isolation of Copper

The most important ores of copper contain copper sulfides (such as covellite, CuS), although copper oxides (such as tenorite, CuO) and copper hydroxycarbonates [such as malachite, $Cu_2(OH)_2CO_3$] are sometimes found. In the production of copper metal, the concentrated sulfide ore is roasted to remove part of the sulfur as sulfur dioxide. The remaining mixture, which consists of Cu_2S, FeS, FeO, and SiO_2, is mixed with limestone, which serves as a flux (a material that aids in the removal of impurities), and heated. Molten slag forms as the iron and silica are removed by Lewis acid-base reactions:

$$\begin{aligned} \mathrm{CaCO_3(s) + SiO_2(s)} &\longrightarrow \mathrm{CaSiO_3(l) + CO_2(g)} \\ \mathrm{FeO(s) + SiO_2(s)} &\longrightarrow \mathrm{FeSiO_3(l)} \end{aligned}$$

In these reactions, the silicon dioxide behaves as a Lewis acid, which accepts a pair of electrons from the Lewis base (the oxide ion).

Reduction of the Cu_2S that remains after smelting is accomplished by blowing air through the molten material. The air converts part of the Cu_2S into Cu_2O. As soon as copper(I) oxide is formed, it is reduced by the remaining copper(I) sulfide to metallic copper:

$$\begin{aligned} \mathrm{2\,Cu_2S(l) + 3\,O_2(g)} &\longrightarrow \mathrm{2\,Cu_2O(l) + 2\,SO_2(g)} \\ \mathrm{2\,Cu_2O(l) + Cu_2S(l)} &\longrightarrow \mathrm{6\,Cu(l) + SO_2(g)} \end{aligned}$$

The copper obtained in this way is called blister copper because of its characteristic appearance, which is due to the air blisters it contains (Figure 19.2.7). This impure copper is cast into large plates, which are used as anodes in the electrolytic refining of the metal (which is described in the chapter on electrochemistry).



Figure 19.2.7: Blister copper is obtained during the conversion of copper-containing ore into pure copper. (credit: “Tortie tude”/Wikimedia Commons)

Isolation of Silver

Silver sometimes occurs in large nuggets (Figure 19.2.8) but more frequently in veins and related deposits. At one time, panning was an effective method of isolating both silver and gold nuggets. Due to their low reactivity, these metals, and a few others, occur in deposits as nuggets. The discovery of platinum was due to Spanish explorers in Central America mistaking platinum nuggets for silver. When the metal is not in the form of nuggets, it often useful to employ a process called **hydrometallurgy** to separate silver from its ores. Hydrology involves the separation of a metal from a mixture by first converting it into soluble ions and then extracting and reducing them to precipitate the pure metal. In the presence of air, alkali metal cyanides readily form the soluble dicyanoargentate(I) ion, from silver metal or silver-containing compounds such as Ag_2S and AgCl. Representative equations are:

$$\begin{align*}
\ce{4 Ag(s) + 8 CN^{-}(aq) + O2(g) + 2 H2O (l) \longrightarrow 4\left[Ag(CN)2 \right]^{-}(aq) + 4 OH^{-}(aq)} \\
\ce{2 Ag2S(s) + 8 CN^{-}(aq) + O2(g) + 2 H2O(l) \longrightarrow 4\left[Ag(CN)2 \right]^{-}(aq) + 2 S(s) + 4 OH^{-}(aq)} \\
\ce{AgCl(s) + 2 CN^{-}(aq) \longrightarrow \left[Ag(CN)2 \right]^{-}(aq) + Cl^{-}(aq)}
\end{align*}$$

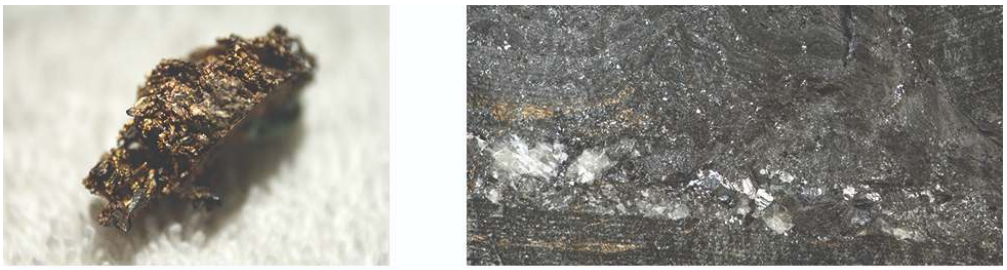


Figure 19.2.8: Naturally occurring free silver may be found as nuggets (a) or in veins (b). (credit a: modification of work by “Teravolt”/Wikimedia Commons; credit b: modification of work by James St. John)

The silver is precipitated from the cyanide solution by the addition of either zinc or iron(II) ions, which serves as the reducing agent:

$$\ce{ 2\left[Ag(CN)2\right]^{-}(aq) + Zn(s) \longrightarrow 2 Ag(s) + \left[Zn(CN)4\right]^{2-}(aq)} \nonumber$$

Example 19.2.3: Refining Redox

One of the steps for refining silver involves converting silver into dicyanoargenate(I) ions:

$$\ce{4 Ag(s) + 8 CN^{-}(aq) + O_2(g) + 2 H_2O(l) \longrightarrow 4 \left[Ag(CN)2 \right]^{-}(aq) + 4 OH^{-}(aq)} \nonumber$$

Explain why oxygen must be present to carry out the reaction. Why does the reaction not occur as:

$$\ce{4 Ag(s) + 8 CN^{-}(aq) \longrightarrow 4\left[Ag(CN)2\right]^{-}(aq)} \; ? \nonumber$$

Solution

The charges, as well as the atoms, must balance in reactions. The silver atom is being oxidized from the 0 oxidation state to the 1+ state. Whenever something loses electrons, something must also gain electrons (be reduced) to balance the equation. Oxygen is a good oxidizing agent for these reactions because it can gain electrons to go from the 0 oxidation state to the 2− state.

Exercise 19.2.3

During the refining of iron, carbon must be present in the blast furnace. Why is carbon necessary to convert iron oxide into iron?

Answer

The carbon is converted into CO, which is the reducing agent that accepts electrons so that iron(III) can be reduced to iron(0).

Transition Metal Compounds

The bonding in the simple compounds of the transition elements ranges from ionic to covalent. In their lower oxidation states, the transition elements form ionic compounds; in their higher oxidation states, they form covalent compounds or polyatomic ions. The variation in oxidation states exhibited by the transition elements gives these compounds a metal-based, oxidation-reduction chemistry. The chemistry of several classes of compounds containing elements of the transition series follows.

Halides

Anhydrous halides of each of the transition elements can be prepared by the direct reaction of the metal with halogens. For example:

$$2\,\mathrm{Fe(s)} + 3\,\mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{FeCl_3(s)}$$

Heating a metal halide with additional metal can be used to form a halide of the metal with a lower oxidation state:

$$\mathrm{Fe(s)} + 2\,\mathrm{FeCl_3(s)} \longrightarrow 3\,\mathrm{FeCl_2(s)}$$

The stoichiometry of the metal halide that results from the reaction of the metal with a halogen is determined by the relative amounts of metal and halogen and by the strength of the halogen as an oxidizing agent. Generally, fluorine forms fluoride-containing metals in their highest oxidation states. The other halogens may not form analogous compounds.

In general, the preparation of stable water solutions of the halides of the metals of the first transition series is by the addition of a hydrohalic acid to carbonates, hydroxides, oxides, or other compounds that contain basic anions. Sample reactions are:

$$\mathrm{NiCO_3(s)} + 2\,\mathrm{HF(aq)} \longrightarrow \mathrm{NiF_2(aq)} + \mathrm{H_2O(l)} + \mathrm{CO_2(g)}$$
$$\mathrm{Co(OH)_2(s)} + 2\,\mathrm{HBr(aq)} \longrightarrow \mathrm{CoBr_2(aq)} + 2\,\mathrm{H_2O(l)}$$

Most of the first transition series metals also dissolve in acids, forming a solution of the salt and hydrogen gas. For example:

$$\mathrm{Cr(s)} + 2\,\mathrm{HCl(aq)} \longrightarrow \mathrm{CrCl_2(aq)} + \mathrm{H_2(g)}$$

The polarity of bonds with transition metals varies based not only upon the electronegativities of the atoms involved but also upon the oxidation state of the transition metal. Remember that bond polarity is a continuous spectrum with electrons being shared evenly (covalent bonds) at one extreme and electrons being transferred completely (ionic bonds) at the other. No bond is ever 100%

ionic, and the degree to which the electrons are evenly distributed determines many properties of the compound. Transition metal halides with low oxidation numbers form more ionic bonds. For example, titanium(II) chloride and titanium(III) chloride ($TiCl_2$ and $TiCl_3$) have high melting points that are characteristic of ionic compounds, but titanium(IV) chloride ($TiCl_4$) is a volatile liquid, consistent with having covalent titanium-chlorine bonds. All halides of the heavier *d*-block elements have significant covalent characteristics.

The covalent behavior of the transition metals with higher oxidation states is exemplified by the reaction of the metal tetrahalides with water. Like covalent silicon tetrachloride, both the titanium and vanadium tetrahalides react with water to give solutions containing the corresponding hydrohalic acids and the metal oxides:

$$\mathrm{SiCl_4(l) + 2\,H_2O(l) \longrightarrow SiO_2(s) + 4\,HCl(aq)}$$
$$\mathrm{TiCl_4(l) + 2\,H_2O(l) \longrightarrow TiO_2(s) + 4\,HCl(aq)}$$

Oxides

As with the halides, the nature of bonding in oxides of the transition elements is determined by the oxidation state of the metal. Oxides with low oxidation states tend to be more ionic, whereas those with higher oxidation states are more covalent. These variations in bonding are because the electronegativities of the elements are not fixed values. The electronegativity of an element increases with increasing oxidation state. Transition metals in low oxidation states have lower electronegativity values than oxygen; therefore, these metal oxides are ionic. Transition metals in very high oxidation states have electronegativity values close to that of oxygen, which leads to these oxides being covalent.

The oxides of the first transition series can be prepared by heating the metals in air. These oxides are Sc_2O_3, TiO_2, V_2O_5, Cr_2O_3, Mn_3O_4, Fe_3O_4, Co_3O_4, NiO, and CuO.

Alternatively, these oxides and other oxides (with the metals in different oxidation states) can be produced by heating the corresponding hydroxides, carbonates, or oxalates in an inert atmosphere. Iron(II) oxide can be prepared by heating iron(II) oxalate, and cobalt(II) oxide is produced by heating cobalt(II) hydroxide:

$$\mathrm{FeC_2O_4(s) \longrightarrow FeO(s) + CO(g) + CO_2(g)}$$
$$\mathrm{Co(OH)_2(s) \longrightarrow CoO(s) + H_2O(g)}$$

With the exception of CrO_3 and Mn_2O_7, transition metal oxides are not soluble in water. They can react with acids and, in a few cases, with bases. Overall, oxides of transition metals with the lowest oxidation states are basic (and react with acids), the intermediate ones are amphoteric, and the highest oxidation states are primarily acidic. Basic metal oxides at a low oxidation state react with aqueous acids to form solutions of salts and water. Examples include the reaction of cobalt(II) oxide accepting protons from nitric acid, and scandium(III) oxide accepting protons from hydrochloric acid:

\[\begin{align*}
\ce{CoO(s) + 2 HNO3(aq) &\longrightarrow Co(NO3)2(aq) + H2O(l)} \\
\ce{Sc2O3(s) + 6 HCl(aq) &\longrightarrow 2 ScCl3(aq) + 3 H2O(l)}
\end{align*}]

The oxides of metals with oxidation states of 4+ are amphoteric, and most are not soluble in either acids or bases. Vanadium(V) oxide, chromium(VI) oxide, and manganese(VII) oxide are acidic. They react with solutions of hydroxides to form salts of the oxyanions and \ce{VO4^{3-}}, \ce{CrO4^{2-}}, and $\mathrm{MnO_4-}$. For example, the complete ionic equation for the reaction of chromium(VI) oxide with a strong base is given by:

$$\mathrm{CrO_3(s) + 2\,Na^+(aq) + 2\,OH^-(aq) \longrightarrow 2\,Na^+(aq) + CrO_4^{2-}(aq) + H_2O(l)}$$

Chromium(VI) oxide and manganese(VII) oxide react with water to form the acids H_2CrO_4 and $HMnO_4$, respectively.

Hydroxides

When a soluble hydroxide is added to an aqueous solution of a salt of a transition metal of the first transition series, a gelatinous precipitate forms. For example, adding a solution of sodium hydroxide to a solution of cobalt sulfate produces a gelatinous pink or blue precipitate of cobalt(II) hydroxide. The net ionic equation is:

$$Co^{2+}(aq) + 2OH^-(aq) \longrightarrow Co(OH)_2(s)$$

In this and many other cases, these precipitates are hydroxides containing the transition metal ion, hydroxide ions, and water coordinated to the transition metal. In other cases, the precipitates are hydrated oxides composed of the metal ion, oxide ions, and water of hydration:

$$4Fe^{3+}(aq) + 6OH^{-}(aq) + nH_2O(l) \longrightarrow 2Fe_2O_3 \cdot (n+3)H_2O(s)$$

These substances do not contain hydroxide ions. However, both the hydroxides and the hydrated oxides react with acids to form salts and water. When precipitating a metal from solution, it is necessary to avoid an excess of hydroxide ion, as this may lead to complex ion formation as discussed later in this chapter. The precipitated metal hydroxides can be separated for further processing or for waste disposal.

Carbonates

Many of the elements of the first transition series form insoluble carbonates. It is possible to prepare these carbonates by the addition of a soluble carbonate salt to a solution of a transition metal salt. For example, nickel carbonate can be prepared from solutions of nickel nitrate and sodium carbonate according to the following net ionic equation:

$$Ni^{2+}(aq) + {CO_3}^{2-} \longrightarrow NiCO_3(s)$$

The reactions of the transition metal carbonates are similar to those of the active metal carbonates. They react with acids to form metals salts, carbon dioxide, and water. Upon heating, they decompose, forming the transition metal oxides.

Other Salts

In many respects, the chemical behavior of the elements of the first transition series is very similar to that of the main group metals. In particular, the same types of reactions that are used to prepare salts of the main group metals can be used to prepare simple ionic salts of these elements.

A variety of salts can be prepared from metals that are more active than hydrogen by reaction with the corresponding acids: Scandium metal reacts with hydrobromic acid to form a solution of scandium bromide:

$$2Sc(s) + 6HBr(aq) \longrightarrow 2ScBr_3(aq) + 3H_2(g)$$

The common compounds that we have just discussed can also be used to prepare salts. The reactions involved include the reactions of oxides, hydroxides, or carbonates with acids. For example:

$$Ni(OH)_2(s) + 2H_3O^{+}(aq) + 2ClO_4^{-}(aq) \longrightarrow Ni^{2+}(aq) + 2ClO_4^{-}(aq) + 4H_2O(l)$$

Substitution reactions involving soluble salts may be used to prepare insoluble salts. For example:

$$Ba^{2+}(aq) + 2Cl^{-}(aq) + 2K^{+}(aq) + CrO_4^{2-}(aq) \longrightarrow BaCrO_4(s) + 2K^{+}(aq) + 2Cl^{-}(aq)$$

In our discussion of oxides in this section, we have seen that reactions of the covalent oxides of the transition elements with hydroxides form salts that contain oxyanions of the transition elements.

How Sciences Interconnect: High Temperature Superconductors

A **superconductor** is a substance that conducts electricity with no resistance. This lack of resistance means that there is no energy loss during the transmission of electricity. This would lead to a significant reduction in the cost of electricity.

Most currently used, commercial superconducting materials, such as NbTi and Nb_3Sn, do not become superconducting until they are cooled below 23 K (−250 °C). This requires the use of liquid helium, which has a boiling temperature of 4 K and is expensive and difficult to handle. The cost of liquid helium has deterred the widespread application of superconductors.

One of the most exciting scientific discoveries of the 1980s was the characterization of compounds that exhibit superconductivity at temperatures above 90 K. (Compared to liquid helium, 90 K is a high temperature.) Typical among the high-temperature superconducting materials are oxides containing yttrium (or one of several rare earth elements), barium, and copper in a 1:2:3 ratio. The formula of the ionic yttrium compound is $YBa_2Cu_3O_7$.

The new materials become superconducting at temperatures close to 90 K (Figure 19.2.9), temperatures that can be reached by cooling with liquid nitrogen (boiling temperature of 77 K). Not only are liquid nitrogen-cooled materials easier to handle, but the cooling costs are also about 1000 times lower than for liquid helium.

Further advances during the same period included materials that became superconducting at even higher temperatures and with a wider array of materials. The DuPont team led by Uma Chowdry and Arthur Sleight identified Bismouth-Strontium-Copper-Oxides that became superconducting at temperatures as high as 110 K and, importantly, did not contain rare earth elements. Advances continued through the subsequent decades until, in 2020, a team led by Ranga Dias at University of Rochester announced the development of a room-temperature superconductor, opening doors to widespread applications. More research and development is needed to realize the potential of these materials, but the possibilities are very promising.

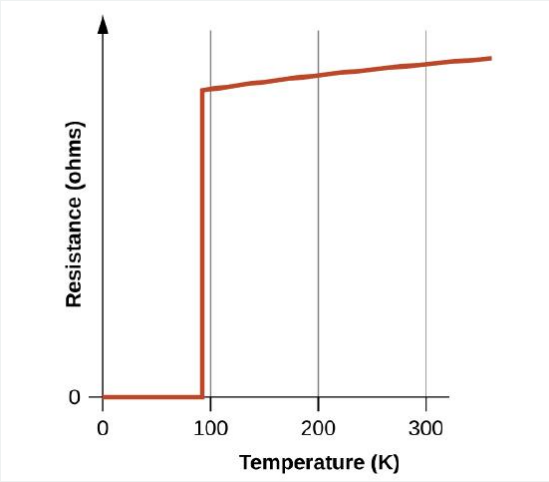


Figure 19.2.9: The resistance of the high-temperature superconductor $YBa_2Cu_3O_7$ varies with temperature. Note how the resistance falls to zero below 92 K, when the substance becomes superconducting.

Although the brittle, fragile nature of these materials presently hampers their commercial applications, they have tremendous potential that researchers are hard at work improving their processes to help realize. Superconducting transmission lines would carry current for hundreds of miles with no loss of power due to resistance in the wires. This could allow generating stations to be located in areas remote from population centers and near the natural resources necessary for power production. The first project demonstrating the viability of high-temperature superconductor power transmission was established in New York in 2008.

Researchers are also working on using this technology to develop other applications, such as smaller and more powerful microchips. In addition, high-temperature superconductors can be used to generate magnetic fields for applications such as medical devices, magnetic levitation trains, and containment fields for nuclear fusion reactors (Figure 19.2.10).

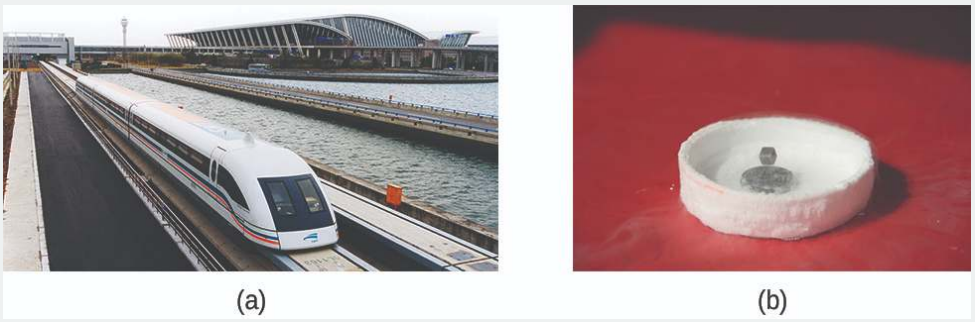


Figure 19.2.10: (a) This magnetic levitation train (or maglev) uses superconductor technology to move along its tracks. (b) A magnet can be levitated using a dish like this as a superconductor. (credit a: modification of work by Alex Needham; credit b: modification of work by Kevin Jarrett)

Link to Learning

Watch how a high-temperature superconductor levitates around a magnetic racetrack in the video.

This page titled 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **19.1: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

19.3: Coordination Chemistry of Transition Metals

Learning Objectives

By the end of this section, you will be able to:

- List the defining traits of coordination compounds
- Describe the structures of complexes containing monodentate and polydentate ligands
- Use standard nomenclature rules to name coordination compounds
- Explain and provide examples of geometric and optical isomerism
- Identify several natural and technological occurrences of coordination compounds

The hemoglobin in your blood, the chlorophyll in green plants, vitamin B-12, and the catalyst used in the manufacture of polyethylene all contain coordination compounds. Ions of the metals, especially the transition metals, are likely to form complexes. Many of these compounds are highly colored (Figure 19.3.1). In the remainder of this chapter, we will consider the structure and bonding of these remarkable compounds.



Figure 19.3.1: Metal ions that contain partially filled *d* subshell usually form colored complex ions; ions with empty *d* subshell (d^0) or with filled *d* subshells (d^{10}) usually form colorless complexes. This figure shows, from left to right, solutions containing $[M(H_2O)_6]^{n+}$ ions with M = $Sc^{3+}(d^0)$, $Cr^{3+}(d^3)$, $Co^{2+}(d^7)$, $Ni^{2+}(d^8)$, $Cu^{2+}(d^9)$, and $Zn^{2+}(d^{10})$. (credit: Sahar Atwa)

Remember that in most main group element compounds, the valence electrons of the isolated atoms combine to form chemical bonds that satisfy the octet rule. For instance, the four valence electrons of carbon overlap with electrons from four hydrogen atoms to form CH_4. The one valence electron leaves sodium and adds to the seven valence electrons of chlorine to form the ionic formula unit NaCl (Figure 19.3.2). Transition metals do not normally bond in this fashion. They primarily form coordinate covalent bonds, a form of the Lewis acid-base interaction in which both of the electrons in the bond are contributed by a donor (Lewis base) to an electron acceptor (Lewis acid). The Lewis acid in coordination complexes, often called a **central metal** ion (or atom), is often a transition metal or inner transition metal, although main group elements can also form **coordination compounds**. The Lewis base donors, called **ligands**, can be a wide variety of chemicals—atoms, molecules, or ions. The only requirement is that they have one or more electron pairs, which can be donated to the central metal. Most often, this involves a **donor atom** with a lone pair of electrons that can form a coordinate bond to the metal.

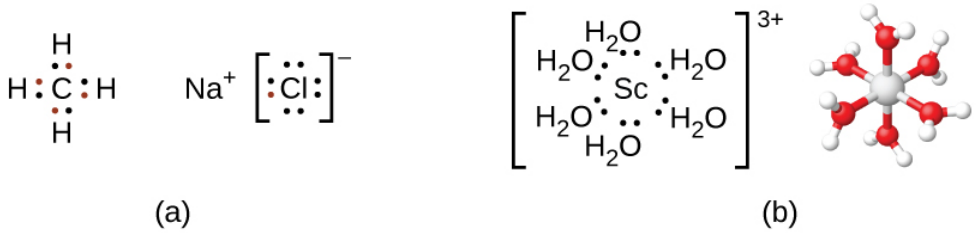


Figure 19.3.2: (a) Covalent bonds involve the sharing of electrons, and ionic bonds involve the transferring of electrons associated with each bonding atom, as indicated by the colored electrons. (b) However, coordinate covalent bonds involve electrons from a Lewis base being donated to a metal center. The lone pairs from six water molecules form bonds to the scandium ion to form an octahedral complex. (Only the donated pairs are shown.)

The coordination sphere consists of the central metal ion or atom plus its attached ligands. Brackets in a formula enclose the coordination sphere; species outside the brackets are not part of the coordination sphere. The coordination number of the central metal ion or atom is the number of donor atoms bonded to it. The coordination number for the silver ion in $[Ag(NH_3)_2]^+$ is two (Figure 19.3.3). For the copper(II) ion in $[CuCl_4]^{2-}$, the coordination number is four, whereas for the cobalt(II) ion in $[Co(H_2O)_6]^{2+}$ the coordination number is six. Each of these ligands is monodentate, from the Greek for "one toothed," meaning that they connect with the central metal through only one atom. In this case, the number of ligands and the coordination number are equal.

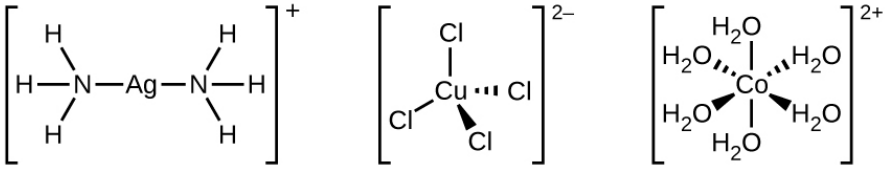


Figure 19.3.3: The complexes (a) $[Ag(NH_3)_2]^+$, (b) $[Cu(Cl)_4]^{2-}$, and (c) $[Co(H_2O)_6]^{2+}$ have coordination numbers of two, four, and six, respectively. The geometries of these complexes are the same as we have seen with VSEPR theory for main group elements: linear, tetrahedral, and octahedral.

Many other ligands coordinate to the metal in more complex fashions. **Bidentate ligands** are those in which two atoms coordinate to the metal center. For example, ethylenediamine (en, $H_2NCH_2CH_2NH_2$) contains two nitrogen atoms, each of which has a lone pair and can serve as a Lewis base (Figure 19.3.4). Both of the atoms can coordinate to a single metal center. In the complex $[Co(en)_3]^{3+}$, there are three bidentate en ligands, and the coordination number of the cobalt(III) ion is six. The most common coordination numbers are two, four, and six, but examples of all coordination numbers from 1 to 15 are known.

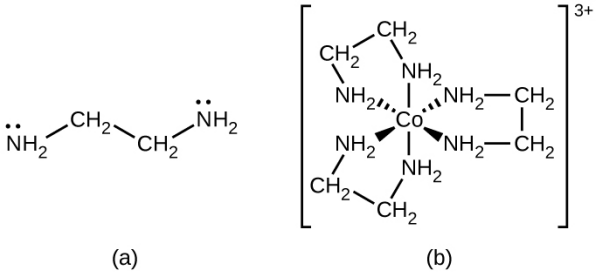


Figure 19.3.4: (a) The ethylenediamine (en) ligand contains two atoms with lone pairs that can coordinate to the metal center. (b) The cobalt(III) complex contains three of these ligands, each forming two bonds to the cobalt ion.

Any ligand that bonds to a central metal ion by more than one donor atom is a **polydentate ligand** (or "many teeth") because it can bite into the metal center with more than one bond. The term **chelate** (pronounced "KEY-late") from the Greek for "claw" is also used to describe this type of interaction. Many polydentate ligands are chelating ligands, and a complex consisting of one or more of these ligands and a central metal is a chelate. A **chelating ligand** is also known as a chelating agent. A chelating ligand holds the metal ion rather like a crab's claw would hold a marble. Figure 19.3.4 showed one example of a chelate. The heme complex in hemoglobin is another important example (Figure 19.3.5). It contains a polydentate ligand with four donor atoms that coordinate to iron.

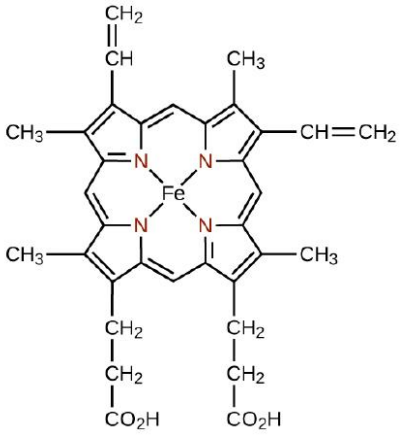


Figure 19.3.5: The single ligand heme contains four nitrogen atoms that coordinate to iron in hemoglobin to form a chelate.

Polydentate ligands are sometimes identified with prefixes that indicate the number of donor atoms in the ligand. As we have seen, ligands with one donor atom, such as NH_3, Cl^-, and H_2O, are monodentate ligands. Ligands with two donor groups are bidentate ligands. Ethylenediamine, \cE$H2NCH2CH2NH2$, and the anion of the acid glycine, $NH_2CH_2CO_2^-$ (Figure 19.3.6) are examples of bidentate ligands. Tridentate ligands, tetradentate ligands, pentadentate ligands, and hexadentate ligands contain three, four, five, and six donor atoms, respectively. The ligand in heme (Figure 19.3.5) is a tetradentate ligand.

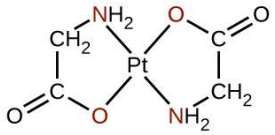


Figure 19.3.6: Each of the anionic ligands shown attaches in a bidentate fashion to platinum(II), with both a nitrogen and oxygen atom coordinating to the metal.

The Naming of Complexes

The nomenclature of the complexes is patterned after a system suggested by Alfred Werner, a Swiss chemist and Nobel laureate, whose outstanding work more than 100 years ago laid the foundation for a clearer understanding of these compounds. The following five rules are used for naming complexes:

1. If a coordination compound is ionic, name the cation first and the anion second, in accordance with the usual nomenclature.
2. Name the ligands first, followed by the central metal. Name the ligands alphabetically. Negative ligands (anions) have names formed by adding *-o* to the stem name of the group. For examples, see Table 19.3.1. For most neutral ligands, the name of the molecule is used. The four common exceptions are *aqua* (H_2O), *ammine* (NH_3), *carbonyl* (CO), and *nitrosyl* (NO). For example, name $[Pt(NH_3)_2Cl_4]$ as diamminetetrachloroplatinum(IV).

Table 19.3.1: Examples of Anionic Ligands

Anionic Ligand	Name
F^-	fluoro
Cl^-	chloro
Br^-	bromo
I^-	iodo
CN^-	cyano
	nitrato
OH^-	hydroxo
O^{2-}	oxo
	oxalato
	carbonato

3. If more than one ligand of a given type is present, the number is indicated by the prefixes *di-* (for two), *tri-* (for three), *tetra-* (for four), *penta-* (for five), and *hexa-* (for six). Sometimes, the prefixes *bis-* (for two), *tris-* (for three), and *tetrakis-* (for four) are used when the name of the ligand already includes *di-*, *tri-*, or *tetra-*, or when the ligand name begins with a vowel. For example, the ion bis(bipyridyl)osmium(II) uses bis- to signify that there are two ligands attached to Os, and each bipyridyl ligand contains two pyridine groups (C_5H_4N).

When the complex is either a cation or a neutral molecule, the name of the central metal atom is spelled exactly like the name of the element and is followed by a Roman numeral in parentheses to indicate its oxidation state (Table 19.3.2 and Table 19.3.3). When the complex is an anion, the suffix -ate is added to the stem of the name of the metal, followed by the Roman numeral designation of its oxidation state (Table 19.3.4). Sometimes, the Latin name of the metal is used when the English name is clumsy. For example, *ferrate* is used instead of *ironate*, *plumbate* instead *leadate*, and *stannate* instead of *tinate*. The oxidation state of the metal is determined based on the charges of each ligand and the overall charge of the coordination compound. For example, in $[Cr(H_2O)_4Cl_2]Br$, the coordination sphere (in brackets) has a charge of 1+ to balance the bromide ion. The water ligands are

neutral, and the chloride ligands are anionic with a charge of 1− each. To determine the oxidation state of the metal, we set the overall charge equal to the sum of the ligands and the metal: +1 = −2 + *x*, so the oxidation state (*x*) is equal to 3+.

Table 19.3.2: Examples in Which the Complex Is a Cation

$[Co(NH_3)_6]Cl_3$	hexaamminecobalt(III) chloride
$[Pt(NH_3)_4Cl_2]^{2+}$	tetraamminedichloroplatinum(IV) ion
$[Ag(NH_3)_2]^+$	diamminesilver(I) ion
$[Cr(H_2O)_4Cl_2]Cl$	tetraaquadichlorochromium(III) chloride
$[Co(H_2NCH_2CH_2NH_2)_3]_2(SO_4)_3$	tris(ethylenediamine)cobalt(III) sulfate

Table 19.3.3: Examples in Which the Complex Is Neutral

$[Pt(NH_3)_2Cl_4]$	diamminetetrachloroplatinum(IV)
$[Ni(H_2NCH_2CH_2NH_2)_2Cl_2]$	dichlorobis(ethylenediamine)nickel(II)

Table 19.3.4: Examples in Which the Complex Is an Anion

$[PtCl_6]^{2-}$	hexachloroplatinate(IV) ion
$Na_2[SnCl_6]$	sodium hexachlorostannate(IV)

Link to Learning

Do you think you understand naming coordination complexes? You can look over more examples and test yourself with online quizzes at the University of Sydney's site.

Example 19.3.1: Coordination Numbers and Oxidation States

Determine the name of the following complexes and give the coordination number of the central metal atom.

a. $Na_2[PtCl_6]$
b. $K_3[Fe(C_2O_4)_3]$
c. $[Co(NH_3)_5Cl]Cl_2$

Solution

(a) There are two Na^+ ions, so the coordination sphere has a negative two charge: $[PtCl_6]^{2-}$. There are six anionic chloride ligands, so −2 = −6 + *x*, and the oxidation state of the platinum is 4+. The name of the complex is sodium hexachloroplatinate(IV), and the coordination number is six.

(b) The coordination sphere has a charge of 3− (based on the potassium) and the oxalate ligands each have a charge of 2−, so the metal oxidation state is given by −3 = −6 + *x*, and this is an iron(III) complex. The name is potassium trisoxalatoferrate(III) (note that tris is used instead of tri because the ligand name starts with a vowel). Because oxalate is a bidentate ligand, this complex has a coordination number of six.

(c) In this example, the coordination sphere has a cationic charge of 2+. The NH_3 ligand is neutral, but the chloro ligand has a charge of 1−. The oxidation state is found by +2 = −1 + *x* and is 3+, so the complex is pentaamminechlorocobalt(III) chloride and the coordination number is six.

Exercise 19.3.1

The complex potassium dicyanoargenate(I) is used to make antiseptic compounds. Give the formula and coordination number.

Answer

$K[Ag(CN)_2]$; coordination number two

The Structures of Complexes

For transition metal complexes, the coordination number determines the geometry around the central metal ion. The most common structures of the complexes in coordination compounds are square planar, tetrahedral, and octahedral, corresponding to coordination numbers of four, four, and six, respectively. Coordination numbers greater than six are less common and yield a variety of structures (see Figure 19.3.7 and Table 19.3.5):

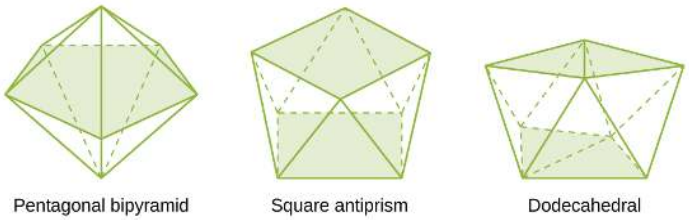


Figure 19.3.7: These are geometries of some complexes with coordination numbers of seven and eight.

Table 19.3.5: Coordination Numbers and Molecular Geometry

Coordination Number	Molecular Geometry	Example
2	linear	$[Ag(NH_3)_2]^+$
3	trigonal planar	$[Cu(CN)_3]^{2-}$
4	tetrahedral(d^0 or d^{10}), low oxidation states for M	$[Ni(CO)_4]$
4	square planar (d^8)	$[Ni(CN)_4]^{2-}$
5	trigonal bipyramidal	$[CoCl_5]^{2-}$
5	square pyramidal	$[VO(CN)_4]^{2-}$
6	octahedral	$[CoCl_6]^{3-}$
7	pentagonal bipyramid	$[ZrF_7]^{3-}$
8	square antiprism	$[ReF_8]^{2-}$
8	dodecahedron	$[Mo(CN)_8]^{4-}$
9 and above	more complicated structures	$[ReH_9]^{2-}$

Unlike main group atoms in which both the bonding and nonbonding electrons determine the molecular shape, the nonbonding *d*-electrons do not change the arrangement of the ligands. Octahedral complexes have a coordination number of six, and the six donor atoms are arranged at the corners of an octahedron around the central metal ion. Examples are shown in Figure 19.3.8 The chloride and nitrate anions in $[Co(H_2O)_6]Cl_2$ and $[Cr(en)_3](NO_3)_3$, and the potassium cations in $K_2[PtCl_6]$, are outside the brackets and are not bonded to the metal ion.

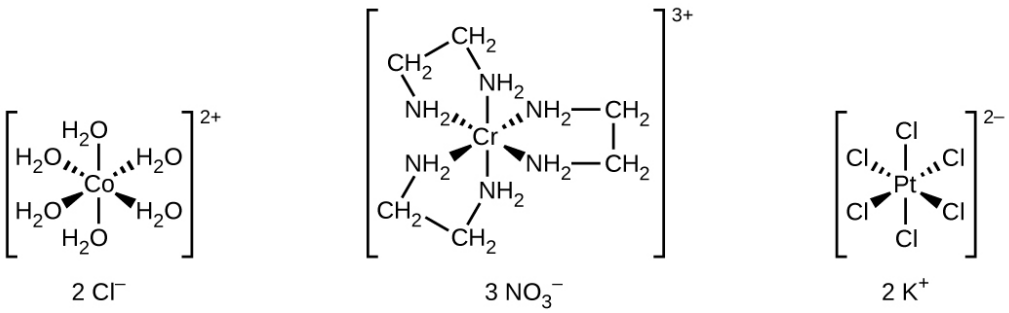


Figure 19.3.8: Many transition metal complexes adopt octahedral geometries, with six donor atoms forming bond angles of 90° about the central atom with adjacent ligands. Note that only ligands within the coordination sphere affect the geometry around the metal center.

For transition metals with a coordination number of four, two different geometries are possible: tetrahedral or square planar. Unlike main group elements, where these geometries can be predicted from VSEPR theory, a more detailed discussion of transition metal orbitals (discussed in the section on Crystal Field Theory) is required to predict which complexes will be tetrahedral and which will be square planar. In tetrahedral complexes such as $[Zn(CN)_4]^{2-}$ (Figure 19.3.9), each of the ligand pairs forms an angle of 109.5°. In square planar complexes, such as $[Pt(NH_3)_2Cl_2]$, each ligand has two other ligands at 90° angles (called the *cis* positions) and one additional ligand at an 180° angle, in the *trans* position.

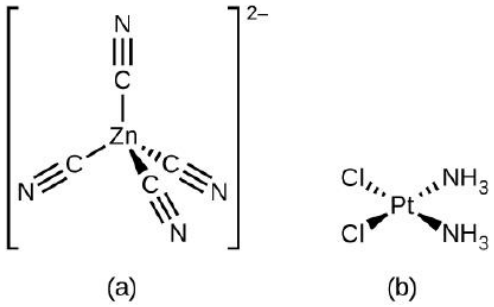


Figure 19.3.9: Transition metals with a coordination number of four can adopt a tetrahedral geometry (a) as in $K_2[Zn(CN)_4]$ or a square planar geometry (b) as shown in $[Pt(NH_3)_2Cl_2]$.

Isomerism in Complexes

Isomers are different chemical species that have the same chemical formula. Transition metal complexes often exist as **geometric isomers**, in which the same atoms are connected through the same types of bonds but with differences in their orientation in space. Coordination complexes with two different ligands in the *cis* and *trans* positions from a ligand of interest form isomers. For example, the octahedral $[Co(NH_3)_4Cl_2]^+$ ion has two isomers. In the ***cis* configuration**, the two chloride ligands are adjacent to each other (Figure 19.3.10). The other isomer, the ***trans* configuration**, has the two chloride ligands directly across from one another.

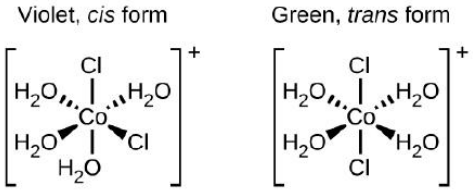


Figure 19.3.10: The *cis* and *trans* isomers of $[Co(H_2O)_4Cl_2]^+$ contain the same ligands attached to the same metal ion, but the spatial arrangement causes these two compounds to have very different properties.

Different geometric isomers of a substance are different chemical compounds. They exhibit different properties, even though they have the same formula. For example, the two isomers of $[Co(NH_3)_4Cl_2]NO_3$ differ in color; the *cis* form is violet, and the *trans* form is green. Furthermore, these isomers have different dipole moments, solubilities, and reactivities. As an example of how the arrangement in space can influence the molecular properties, consider the polarity of the two $[Co(NH_3)_4Cl_2]NO_3$ isomers. Remember that the polarity of a molecule or ion is determined by the bond dipoles (which are due to the difference in electronegativity of the bonding atoms) and their arrangement in space. In one isomer, *cis* chloride ligands cause more electron density on one side of the molecule than on the other, making it polar. For the *trans* isomer, each ligand is directly across from an identical ligand, so the bond dipoles cancel out, and the molecule is nonpolar.

Example 19.3.2: Geometric Isomers

Identify which geometric isomer of $[Pt(NH_3)_2Cl_2]$ is shown in Figure 19.3.9. Draw the other geometric isomer and give its full name.

Solution

In the Figure 19.3.9, the two chlorine ligands occupy *cis* positions. The other form is shown in Figure 19.3.11. When naming specific isomers, the descriptor is listed in front of the name. Therefore, this complex is *trans*-diamminedichloroplatinum(II).

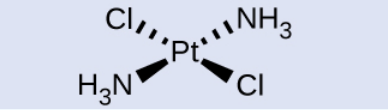
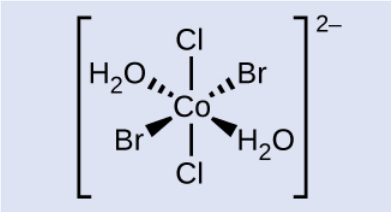


Figure 19.3.11: The *trans* isomer of $[Pt(NH_3)_2Cl_2]$ has each ligand directly across from an adjacent ligand.

? Exercise 19.3.1

Draw the ion *trans*-diaqua-*trans*-dibromo-*trans*-dichlorocobalt(II).

Answer



Another important type of isomers are **optical isomers**, or **enantiomers**, in which two objects are exact mirror images of each other but cannot be lined up so that all parts match. This means that optical isomers are nonsuperimposable mirror images. A classic example of this is a pair of hands, in which the right and left hand are mirror images of one another but cannot be superimposed. Optical isomers are very important in organic and biochemistry because living systems often incorporate one specific optical isomer and not the other. Unlike geometric isomers, pairs of optical isomers have identical properties (boiling point, polarity, solubility, etc.). Optical isomers differ only in the way they affect polarized light and how they react with other optical isomers. For coordination complexes, many coordination compounds such as $[M(en)_3]^{n+}$ [in which M^{n+} is a central metal ion such as iron(III) or cobalt(II)] form enantiomers, as shown in Figure 19.3.12 These two isomers will react differently with other optical isomers. For example, DNA helices are optical isomers, and the form that occurs in nature (right-handed DNA) will bind to only one isomer of $[M(en)_3]^{n+}$ and not the other.

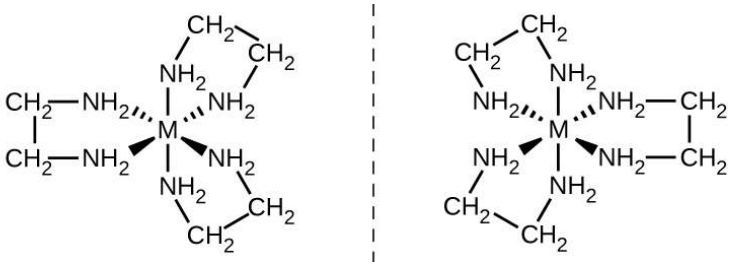


Figure 19.3.12: The complex $[M(en)_3]^{n+}$ (M^{n+} = a metal ion, en = ethylenediamine) has a nonsuperimposable mirror image.

The $[Co(en)_2Cl_2]^+$ ion exhibits geometric isomerism (*cis*/*trans*), and its *cis* isomer exists as a pair of optical isomers (Figure 19.3.13).

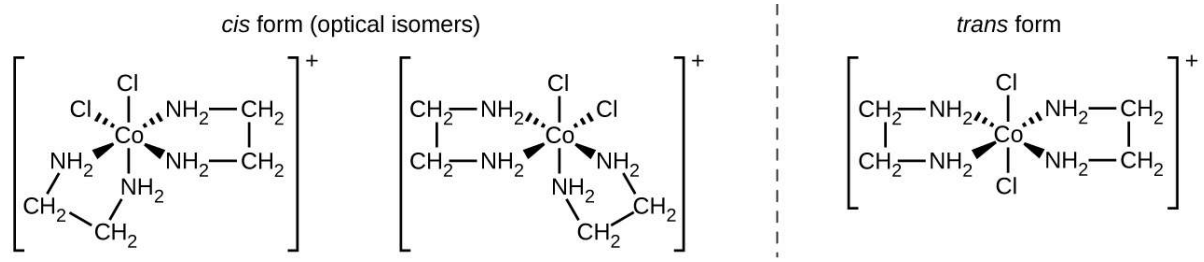


Figure 19.3.13: Three isomeric forms of $[Co(en)_2Cl_2]^+$ exist. The *trans* isomer, formed when the chlorines are positioned at a 180° angle, has very different properties from the *cis* isomers. The mirror images of the *cis* isomer form a pair of optical isomers, which have identical behavior except when reacting with other enantiomers.

Linkage isomers occur when the coordination compound contains a ligand that can bind to the transition metal center through two different atoms. For example, the CN ligand can bind through the carbon atom (cyano) or through the nitrogen atom (isocyano).

Similarly, SCN^- can be bound through the sulfur or nitrogen atom, affording two distinct compounds ($[Co(NH_3)_5SCN]^{2+}$ or $[Co(NH_3)_5NCS]^{2+}$).

Ionization isomers (or **coordination isomers**) occur when one anionic ligand in the inner coordination sphere is replaced with the counter ion from the outer coordination sphere. A simple example of two ionization isomers are $[CoCl_6][Br]$ and $[CoCl_5Br][Cl]$.

Coordination Complexes in Nature and Technology

Chlorophyll, the green pigment in plants, is a complex that contains magnesium (Figure 19.3.14). This is an example of a main group element in a coordination complex. Plants appear green because chlorophyll absorbs red and purple light; the reflected light consequently appears green. The energy resulting from the absorption of light is used in photosynthesis.

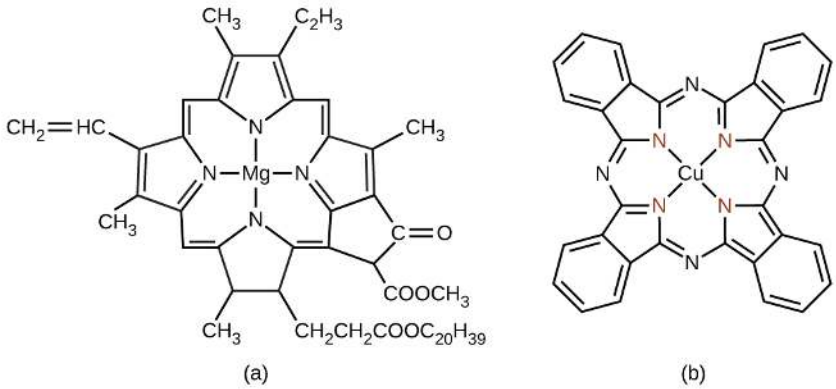


Figure 19.3.14: (a) Chlorophyll comes in several different forms, which all have the same basic structure around the magnesium center. (b) Copper phthalocyanine blue, a square planar copper complex, is present in some blue dyes.

Chemistry in Everyday Life: Transition Metal Catalysts

One of the most important applications of transition metals is as industrial catalysts. As you recall from the chapter on kinetics, a catalyst increases the rate of reaction by lowering the activation energy and is regenerated in the catalytic cycle. Over 90% of all manufactured products are made with the aid of one or more catalysts. The ability to bind ligands and change oxidation states makes transition metal catalysts well suited for catalytic applications. Vanadium oxide is used to produce 230,000,000 tons of sulfuric acid worldwide each year, which in turn is used to make everything from fertilizers to cans for food. Plastics are made with the aid of transition metal catalysts, along with detergents, fertilizers, paints, and more (see Figure 19.3.15). Very complicated pharmaceuticals are manufactured with catalysts that are selective, reacting with one specific bond out of a large number of possibilities. Catalysts allow processes to be more economical and more environmentally friendly. Developing new catalysts and better understanding of existing systems are important areas of current research.



Figure 19.3.15: (a) Detergents, (b) paints, and (c) fertilizers are all made using transition metal catalysts. (credit a: modification of work by “Mr. Brian”/Flickr; credit b: modification of work by Ewen Roberts; credit c: modification of work by “osseous”/Flickr)

Portrait of a Chemist: Deanna D'Alessandro

Dr. Deanna D'Alessandro develops new metal-containing materials that demonstrate unique electronic, optical, and magnetic properties. Her research combines the fields of fundamental inorganic and physical chemistry with materials engineering. She is working on many different projects that rely on transition metals. For example, one type of compound she is developing captures carbon dioxide waste from power plants and catalytically converts it into useful products (see Figure 19.3.16).

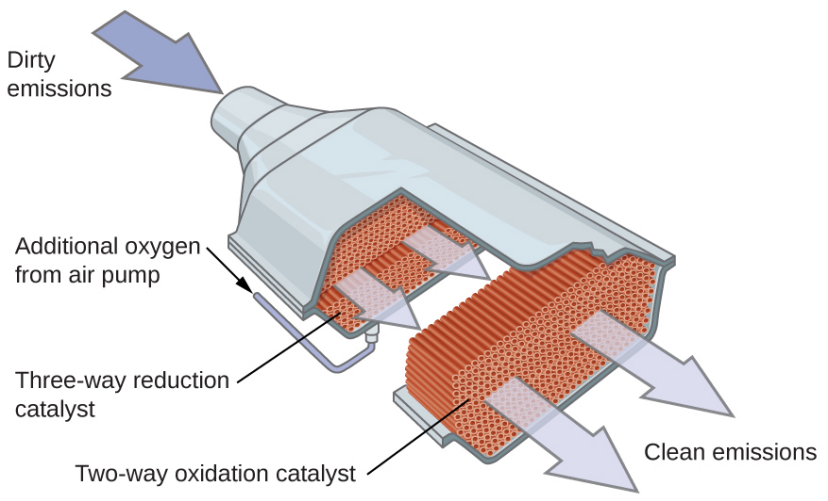


Figure 19.3.16: Catalytic converters change carbon dioxide emissions from power plants into useful products, and, like the one shown here, are also found in cars.

Another project involves the development of porous, sponge-like materials that are "photoactive." The absorption of light causes the pores of the sponge to change size, allowing gas diffusion to be controlled. This has many potential useful applications, from powering cars with hydrogen fuel cells to making better electronics components. Although not a complex, self-darkening sunglasses are an example of a photoactive substance.

Watch this video to learn more about this research and listen to Dr. D'Alessandro (shown in Figure 19.3.17) describe what it is like being a research chemist.



Figure 19.3.17: Dr. Deanna D'Alessandro is a functional materials researcher. Her work combines the inorganic and physical chemistry fields with engineering, working with transition metals to create new systems to power cars and convert energy (credit: image courtesy of Deanna D'Alessandro).

Many other coordination complexes are also brightly colored. The square planar copper(II) complex phthalocyanine blue (from Figure 19.3.14) is one of many complexes used as pigments or dyes. This complex is used in blue ink, blue jeans, and certain blue paints.

The structure of heme (Figure 19.3.18), the iron-containing complex in hemoglobin, is very similar to that in chlorophyll. In hemoglobin, the red heme complex is bonded to a large protein molecule (globin) by the attachment of the protein to the heme ligand. Oxygen molecules are transported by hemoglobin in the blood by being bound to the iron center. When the hemoglobin loses its oxygen, the color changes to a bluish red. Hemoglobin will only transport oxygen if the iron is Fe^{2+}; oxidation of the iron to Fe^{3+} prevents oxygen transport.

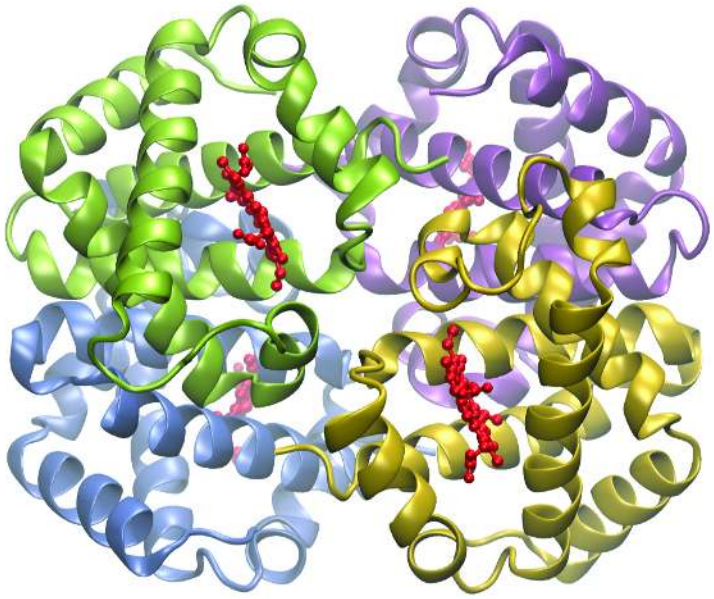


Figure 19.3.18: Hemoglobin contains four protein subunits, each of which has an iron center attached to a heme ligand (shown in red), which is coordinated to a globin protein. Each subunit is shown in a different color.

Complexing agents often are used for water softening because they tie up such ions as Ca^{2+}, Mg^{2+}, and Fe^{2+}, which make water hard. Many metal ions are also undesirable in food products because these ions can catalyze reactions that change the color of food. Coordination complexes are useful as preservatives. For example, the ligand EDTA, $(HO_2CCH_2)_2NCH_2CH_2N(CH_2CO_2H)_2$, coordinates to metal ions through six donor atoms and prevents the metals from reacting (Figure 19.3.19). This ligand also is used to sequester metal ions in paper production, textiles, and detergents, and has pharmaceutical uses.

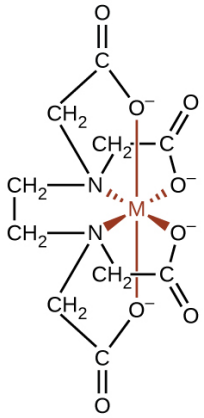


Figure 19.3.19: The ligand EDTA binds tightly to a variety of metal ions by forming hexadentate complexes.

Complexing agents that tie up metal ions are also used as drugs. British Anti-Lewisite (BAL), $HSCH_2CH(SH)CH_2OH$, is a drug developed during World War I as an antidote for the arsenic-based war gas Lewisite. BAL is now used to treat poisoning by heavy metals, such as arsenic, mercury, thallium, and chromium. The drug is a ligand and functions by making a water-soluble chelate of the metal; the kidneys eliminate this metal chelate (Figure 19.3.20). Another polydentate ligand, enterobactin, which is isolated from certain bacteria, is used to form complexes of iron and thereby to control the severe iron buildup found in patients suffering from blood diseases such as Cooley's anemia, who require frequent transfusions. As the transfused blood breaks down, the usual metabolic processes that remove iron are overloaded, and excess iron can build up to fatal levels. Enterobactin forms a water-soluble complex with excess iron, and the body can safely eliminate this complex.

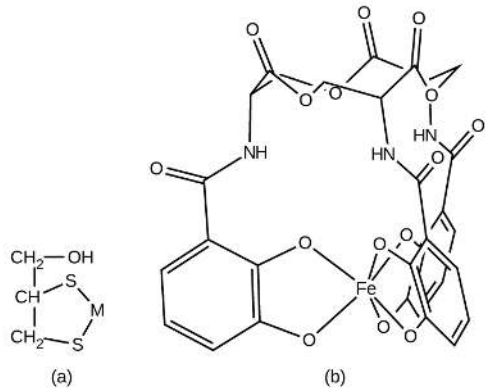


Figure 19.3.20: Coordination complexes are used as drugs. (a) British Anti-Lewisite is used to treat heavy metal poisoning by coordinating metals (M), and enterobactin (b) allows excess iron in the blood to be removed.

Example 19.3.3: Chelation Therapy

Ligands like BAL and enterobactin are important in medical treatments for heavy metal poisoning. However, chelation therapies can disrupt the normal concentration of ions in the body, leading to serious side effects, so researchers are searching for new chelation drugs. One drug that has been developed is dimercaptosuccinic acid (DMSA), shown in Figure 19.3.21. Identify which atoms in this molecule could act as donor atoms.

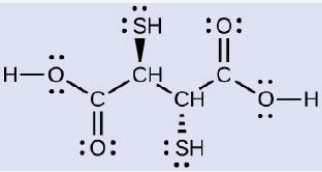


Figure 19.3.21: Dimercaptosuccinic acid is used to treat heavy metal poisoning.

Solution

All of the oxygen and sulfur atoms have lone pairs of electrons that can be used to coordinate to a metal center, so there are six possible donor atoms. Geometrically, only two of these atoms can be coordinated to a metal at once. The most common binding mode involves the coordination of one sulfur atom and one oxygen atom, forming a five-member ring with the metal.

Exercise 19.3.1

Some alternative medicine practitioners recommend chelation treatments for ailments that are not clearly related to heavy metals, such as cancer and autism, although the practice is discouraged by many scientific organizations.[1] Identify at least two biologically important metals that could be disrupted by chelation therapy.

Answer

Ca, Fe, Zn, and Cu

Ligands are also used in the electroplating industry. When metal ions are reduced to produce thin metal coatings, metals can clump together to form clusters and nanoparticles. When metal coordination complexes are used, the ligands keep the metal atoms isolated from each other. It has been found that many metals plate out as a smoother, more uniform, better-looking, and more adherent surface when plated from a bath containing the metal as a complex ion. Thus, complexes such as $[Ag(CN)_2]^-$ and $[Au(CN)_2]^-$ are used extensively in the electroplating industry.

In 1965, scientists at Michigan State University discovered that there was a platinum complex that inhibited cell division in certain microorganisms. Later work showed that the complex was *cis*-diamminedichloroplatinum(II), $[Pt(NH_3)_2(Cl)_2]$, and that the *trans* isomer was not effective. The inhibition of cell division indicated that this square planar compound could be an anticancer agent. In 1978, the US Food and Drug Administration approved this compound, known as cisplatin, for use in the treatment of certain forms of cancer. Since that time, many similar platinum compounds have been developed for the treatment of cancer. In all cases, these

are the *cis* isomers and never the *trans* isomers. The diammine $(NH_3)_2$ portion is retained with other groups, replacing the dichloro $[(Cl)_2]$ portion. The newer drugs include carboplatin, oxaliplatin, and satraplatin.

Footnotes

- 1National Council against Health Fraud, *NCAHF Policy Statement on Chelation Therapy*, (Peabody, MA, 2002).

This page titled 19.3: Coordination Chemistry of Transition Metals is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **19.2: Coordination Chemistry of Transition Metals** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

19.4: Spectroscopic and Magnetic Properties of Coordination Compounds

Learning Objectives

By the end of this section, you will be able to:

- Outline the basic premise of crystal field theory (CFT)
- Identify molecular geometries associated with various d-orbital splitting patterns
- Predict electron configurations of split d orbitals for selected transition metal atoms or ions
- Explain spectral and magnetic properties in terms of CFT concepts

The behavior of coordination compounds cannot be adequately explained by the same theories used for main group element chemistry. The observed geometries of coordination complexes are not consistent with hybridized orbitals on the central metal overlapping with ligand orbitals, as would be predicted by valence bond theory. The observed colors indicate that the *d* orbitals often occur at different energy levels rather than all being degenerate, that is, of equal energy, as are the three *p* orbitals. To explain the stabilities, structures, colors, and magnetic properties of transition metal complexes, a different bonding model has been developed. Just as valence bond theory explains many aspects of bonding in main group chemistry, crystal field theory is useful in understanding and predicting the behavior of transition metal complexes.

Crystal Field Theory

To explain the observed behavior of transition metal complexes (such as how colors arise), a model involving electrostatic interactions between the electrons from the ligands and the electrons in the unhybridized *d* orbitals of the central metal atom has been developed. This electrostatic model is **crystal field theory** (CFT). It allows us to understand, interpret, and predict the colors, magnetic behavior, and some structures of coordination compounds of transition metals.

CFT focuses on the nonbonding electrons on the central metal ion in coordination complexes not on the metal-ligand bonds. Like valence bond theory, CFT tells only part of the story of the behavior of complexes. However, it tells the part that valence bond theory does not. In its pure form, CFT ignores any covalent bonding between ligands and metal ions. Both the ligand and the metal are treated as infinitesimally small point charges.

All electrons are negative, so the electrons donated from the ligands will repel the electrons of the central metal. Let us consider the behavior of the electrons in the unhybridized *d* orbitals in an octahedral complex. The five *d* orbitals consist of lobe-shaped regions and are arranged in space, as shown in Figure 19.4.1. In an octahedral complex, the six ligands coordinate along the axes.

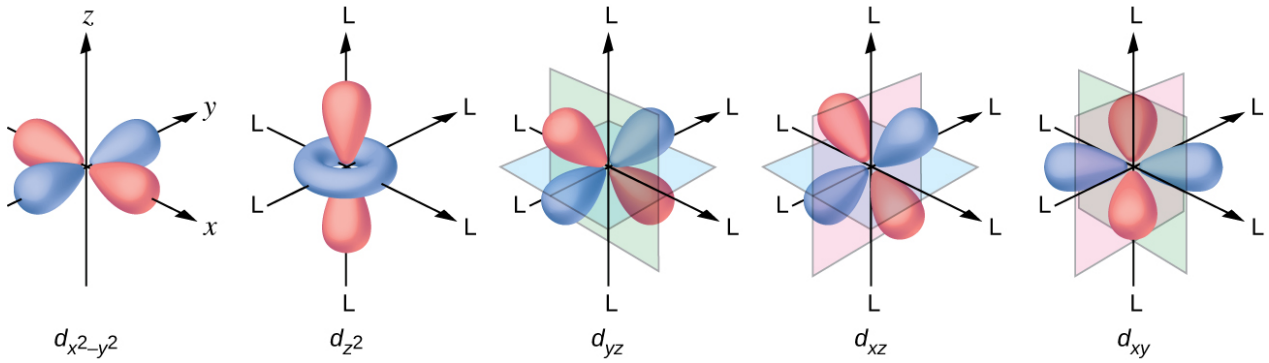


Figure 19.4.1: The directional characteristics of the five *d* orbitals are shown here. The shaded portions indicate the phase of the orbitals. The ligands (L) coordinate along the axes. For clarity, the ligands have been omitted from the orbital so that the axis labels could be shown.

In an uncomplexed metal ion in the gas phase, the electrons are distributed among the five *d* orbitals in accord with Hund's rule because the orbitals all have the same energy. However, when ligands coordinate to a metal ion, the energies of the *d* orbitals are no longer the same.

In octahedral complexes, the lobes in two of the five d orbitals, the d_{z^2} and $d_{x^2-y^2}$ orbitals, point toward the ligands (Figure 19.4.1). These two orbitals are called the e_g **orbitals** (the symbol actually refers to the symmetry of the orbitals, but we will use it as a convenient name for these two orbitals in an octahedral complex). The other three orbitals, the d_{xy}, d_{xz}, and d_{yz} orbitals, have

lobes that point between the ligands and are called the t_{2g} **orbitals** (again, the symbol really refers to the symmetry of the orbitals). As six ligands approach the metal ion along the axes of the octahedron, their point charges repel the electrons in the *d* orbitals of the metal ion. However, the repulsions between the electrons in the e_g orbitals (the d_{z^2} and $d_{x^2-y^2}$ orbitals) and the ligands are greater than the repulsions between the electrons in the t_{2g} orbitals (the d_{zy}, d_{xz}, and d_{yz} orbitals) and the ligands. This is because the lobes of the eg orbitals point directly at the ligands, whereas the lobes of the t_{2g} orbitals point between them. Thus, electrons in the eg orbitals of the metal ion in an octahedral complex have higher potential energies than those of electrons in the t_{2g} orbitals. The difference in energy may be represented as shown in Figure 19.4.2

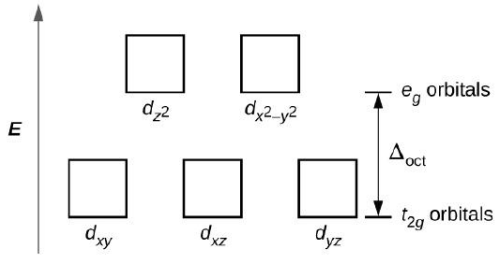


Figure 19.4.2: In octahedral complexes, the e_g orbitals are destabilized (higher in energy) compared to the t_{2g} orbitals because the ligands interact more strongly with the *d* orbitals at which they are pointed directly.

The difference in energy between the e_g and the t_{2g} orbitals is called the **crystal field splitting** and is symbolized by Δ_{oct}, where oct stands for octahedral.

The magnitude of Δ_{oct} depends on many factors, including the nature of the six ligands located around the central metal ion, the charge on the metal, and whether the metal is using 3*d*, 4*d*, or 5*d* orbitals. Different ligands produce different crystal field splittings. The increasing crystal field splitting produced by ligands is expressed in the **spectrochemical series**, a short version of which is given here:

$$\underset{\text{a few ligands of the spectrochemical series, in order of increasing field strength of the ligand}}{\underset{\longrightarrow}{\mathrm{I^- < Br^- < Cl^- < F^- < H_2O < NH_3 < en < NO_2^- < CN^-}}}$$

In this series, ligands on the left cause small crystal field splittings and are **weak-field ligands**, whereas those on the right cause larger splittings and are **strong-field ligands**. Thus, the Δ_{oct} value for an octahedral complex with iodide ligands (I^-) is much smaller than the Δ_{oct} value for the same metal with cyanide ligands (CN^-).

Electrons in the *d* orbitals follow the aufbau ("filling up") principle, which says that the orbitals will be filled to give the lowest total energy, just as in main group chemistry. When two electrons occupy the same orbital, the like charges repel each other. The energy needed to pair up two electrons in a single orbital is called the **pairing energy** (P). Electrons will always singly occupy each orbital in a degenerate set before pairing. P is similar in magnitude to Δ_{oct}. When electrons fill the *d* orbitals, the relative magnitudes of Δ_{oct} and P determine which orbitals will be occupied.

In $[Fe(CN)_6]^{4-}$, the strong field of six cyanide ligands produces a large Δ_{oct}. Under these conditions, the electrons require less energy to pair than they require to be excited to the e_g orbitals (Δ_{oct} > P). The six 3*d* electrons of the Fe^{2+} ion pair in the three t_{2g} orbitals (Figure 19.4.3). Complexes in which the electrons are paired because of the large crystal field splitting are called **low-spin complexes** because the number of unpaired electrons (spins) is minimized.

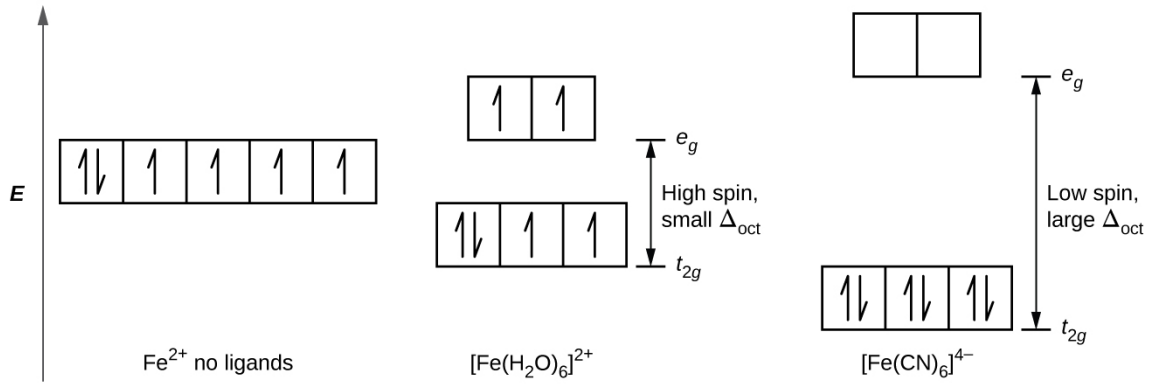


Figure 19.4.3: Iron(II) complexes have six electrons in the 5*d* orbitals. In the absence of a crystal field, the orbitals are degenerate. For coordination complexes with strong-field ligands such as $[Fe(CN)_6]^{4-}$, Δ_{oct} is greater than P, and the electrons pair in the lower energy t_{2g} orbitals before occupying the *eg* orbitals. With weak-field ligands such as H_2O, the ligand field splitting is less than the pairing energy, Δ_{oct} less than P, so the electrons occupy all *d* orbitals singly before any pairing occurs.

In $[Fe(H_2O)_6]^{2+}$, on the other hand, the weak field of the water molecules produces only a small crystal field splitting (Δ_{oct} < P). Because it requires less energy for the electrons to occupy the e_g orbitals than to pair together, there will be an electron in each of the five 3*d* orbitals before pairing occurs. For the six *d* electrons on the iron(II) center in $[Fe(H_2O)_6]^{2+}$, there will be one pair of electrons and four unpaired electrons (Figure 19.4.3). Complexes such as the $[Fe(H_2O)_6]^{2+}$ ion, in which the electrons are unpaired because the crystal field splitting is not large enough to cause them to pair, are called high-spin complexes because the number of unpaired electrons (spins) is maximized.

A similar line of reasoning shows why the $[Fe(CN)_6]^{3-}$ ion is a low-spin complex with only one unpaired electron, whereas both the $[Fe(H_2O)_6]^{3+}$ and $[FeF_6]^{3-}$ ions are high-spin complexes with five unpaired electrons.

✔ Example 19.4.1: High- and Low-Spin Complexes

Predict the number of unpaired electrons.

a. $K_3[CrI_6]$
b. $[Cu(en)_2(H_2O)_2]Cl_2$
c. $Na_3[Co(NO_2)_6]$

Solution

The complexes are octahedral.

a. Cr^{3+} has a d^3 configuration. These electrons will all be unpaired.
b. Cu^{2+} is d^9, so there will be one unpaired electron.
c. Co^{3+} has d^6 valence electrons, so the crystal field splitting will determine how many are paired. Nitrite is a strong-field ligand, so the complex will be low spin. Six electrons will go in the t_{2g} orbitals, leaving 0 unpaired.

? Exercise 19.4.1

The size of the crystal field splitting only influences the arrangement of electrons when there is a choice between pairing electrons and filling the higher-energy orbitals. For which *d*-electron configurations will there be a difference between high- and low-spin configurations in octahedral complexes?

Answer

d^4, d^5, d^6, and d^7

✔ Example 19.4.2: CFT for Other Geometries

CFT is applicable to molecules in geometries other than octahedral. In octahedral complexes, remember that the lobes of the e_g set point directly at the ligands. For tetrahedral complexes, the *d* orbitals remain in place, but now we have only four ligands located between the axes (Figure 19.4.4). None of the orbitals points directly at the tetrahedral ligands. However, the e_g set

(along the Cartesian axes) overlaps with the ligands less than does the t_{2g} set. By analogy with the octahedral case, predict the energy diagram for the *d* orbitals in a tetrahedral crystal field. To avoid confusion, the octahedral e_g set becomes a tetrahedral *e* set, and the octahedral t_{2g} set becomes a t_2 set.

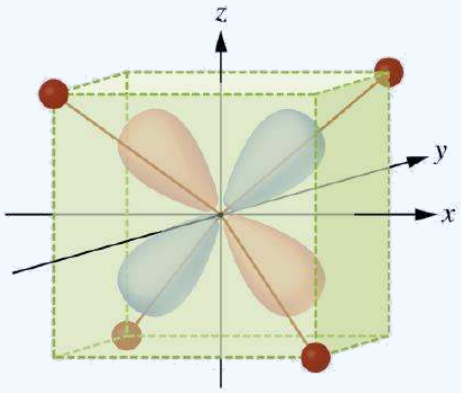
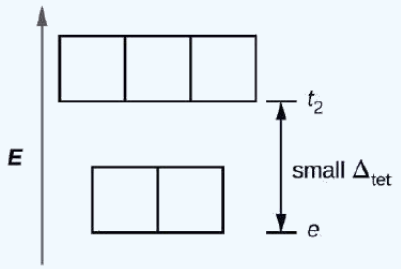


Figure 19.4.4: This diagram shows the orientation of the tetrahedral ligands with respect to the axis system for the orbitals.

Solution

Since CFT is based on electrostatic repulsion, the orbitals closer to the ligands will be destabilized and raised in energy relative to the other set of orbitals. The splitting is less than for octahedral complexes because the overlap is less, so Δ_{tet} is usually small $\left(\Delta_{\text{tet}} = \frac{4}{9}\Delta_{oct}\right)$



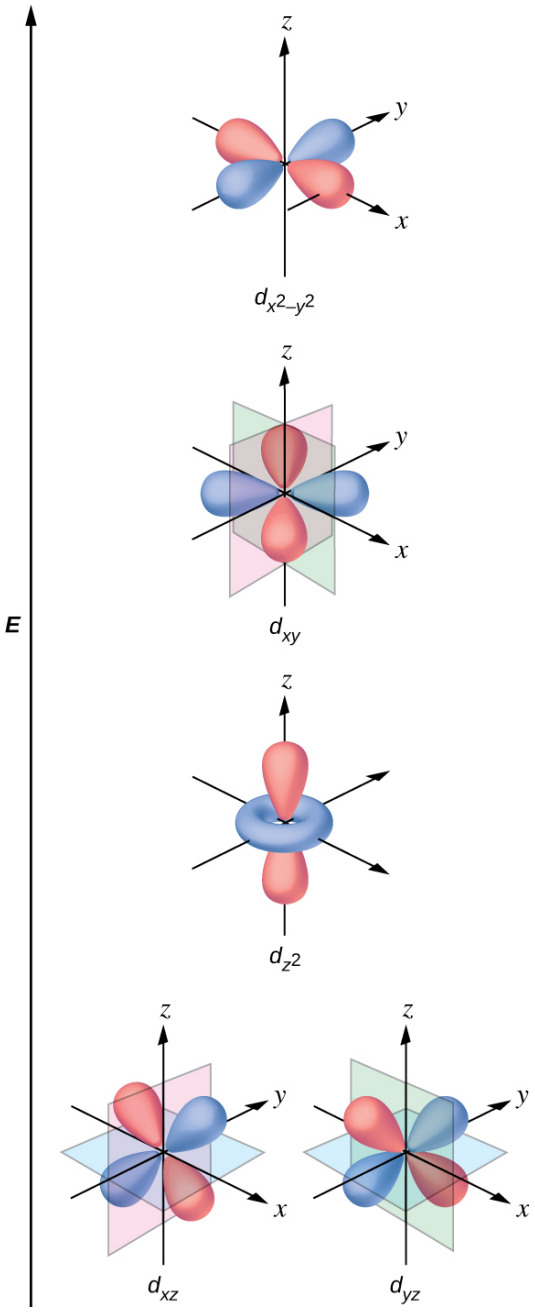
Exercise 19.4.2

Explain how many unpaired electrons a tetrahedral d^4 ion will have.

Answer

4; because Δ_{tet} is small, all tetrahedral complexes are high spin and the electrons go into the t_2 orbitals before pairing

The other common geometry is square planar. It is possible to consider a square planar geometry as an octahedral structure with a pair of *trans* ligands removed. The removed ligands are assumed to be on the *z*-axis. This changes the distribution of the *d* orbitals, as orbitals on or near the *z*-axis become more stable, and those on or near the *x*- or *y*-axes become less stable. This results in the octahedral t_{2g} and the e_g sets splitting and gives a more complicated pattern, as depicted below:



Magnetic Moments of Molecules and Ions

Experimental evidence of magnetic measurements supports the theory of high- and low-spin complexes. Remember that molecules such as O_2 that contain unpaired electrons are paramagnetic. Paramagnetic substances are attracted to magnetic fields. Many transition metal complexes have unpaired electrons and hence are paramagnetic. Molecules such as N_2 and ions such as Na^+ and $[Fe(CN)_6]^{4-}$ that contain no unpaired electrons are diamagnetic. Diamagnetic substances have a slight tendency to be repelled by magnetic fields.

When an electron in an atom or ion is unpaired, the magnetic moment due to its spin makes the entire atom or ion paramagnetic. The size of the magnetic moment of a system containing unpaired electrons is related directly to the number of such electrons: the greater the number of unpaired electrons, the larger the magnetic moment. Therefore, the observed magnetic moment is used to determine the number of unpaired electrons present. The measured magnetic moment of low-spin d^6 $[Fe(CN)_6]^{4-}$ confirms that iron is diamagnetic, whereas high-spin d^6 $[Fe(H_2O)_6]^{2+}$ has four unpaired electrons with a magnetic moment that confirms this arrangement.

Colors of Transition Metal Complexes

When atoms or molecules absorb light at the proper frequency, their electrons are excited to higher-energy orbitals. For many main group atoms and molecules, the absorbed photons are in the ultraviolet range of the electromagnetic spectrum, which cannot be detected by the human eye. For coordination compounds, the energy difference between the *d* orbitals often allows photons in the visible range to be absorbed.

The human eye perceives a mixture of all the colors, in the proportions present in sunlight, as white light. Complementary colors, those located across from each other on a color wheel, are also used in color vision. The eye perceives a mixture of two complementary colors, in the proper proportions, as white light. Likewise, when a color is missing from white light, the eye sees its complement. For example, when red photons are absorbed from white light, the eyes see the color green. When violet photons are removed from white light, the eyes see lemon yellow. The blue color of the $[Cu(NH_3)_4]^{2+}$ ion results because this ion absorbs orange and red light, leaving the complementary colors of blue and green (Figure 19.4.5).

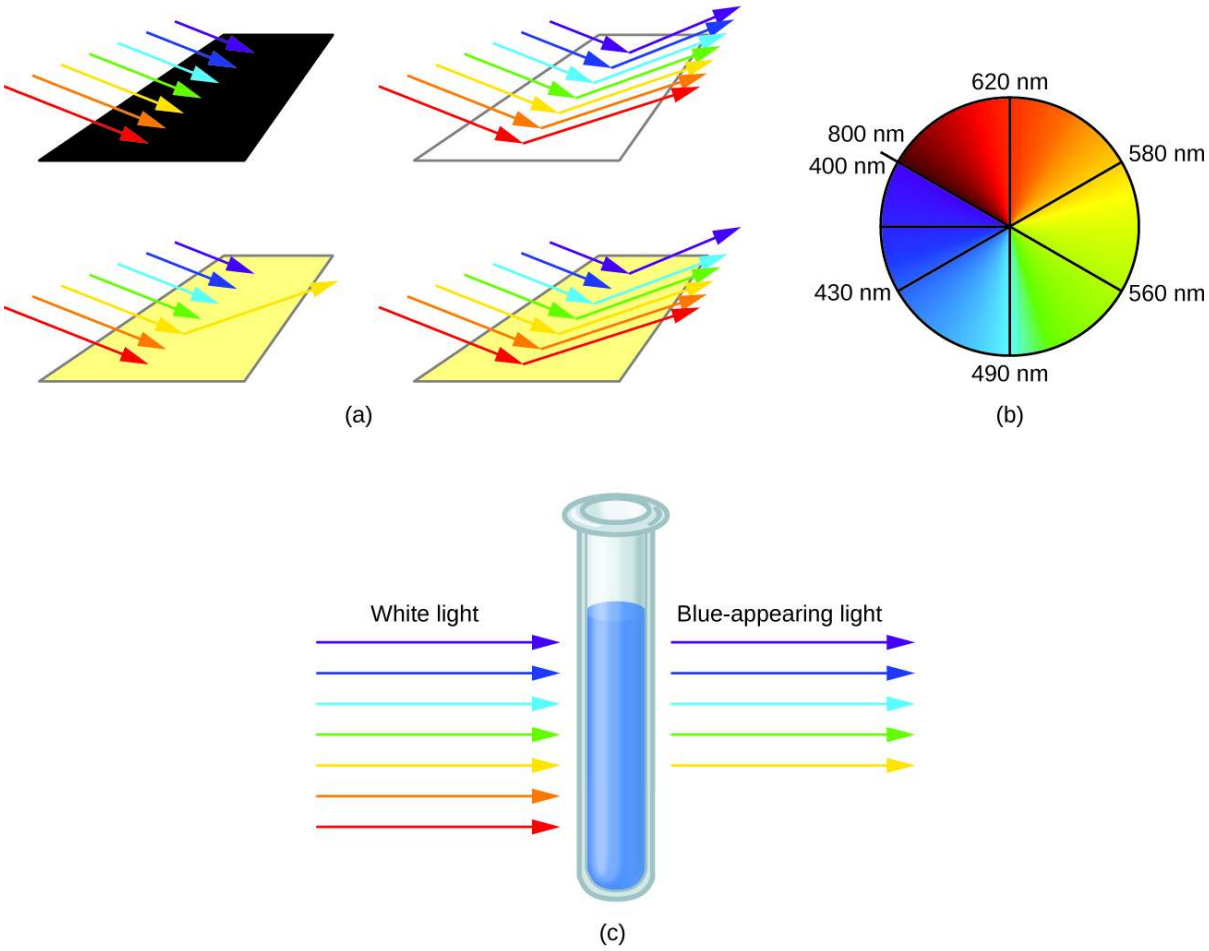


Figure 19.4.5: (a) An object is black if it absorbs all colors of light. If it reflects all colors of light, it is white. An object has a color if it absorbs all colors except one, such as this yellow strip. The strip also appears yellow if it absorbs the complementary color from white light (in this case, indigo). (b) Complementary colors are located directly across from one another on the color wheel. (c) A solution of $[Cu(NH_3)_4]^{2+}$ ions absorbs red and orange light, so the transmitted light appears as the complementary color, blue.

Example 19.4.3: Colors of Complexes

The octahedral complex $[Ti(H_2O)_6]^{3+}$ has a single *d* electron. To excite this electron from the ground state t_{2g} orbital to the e_g orbital, this complex absorbs light from 450 to 600 nm. The maximum absorbance corresponds to Δ_{oct} and occurs at 499 nm. Calculate the value of Δ_{oct} in Joules and predict what color the solution will appear.

Solution

Using Planck's equation (refer to the section on electromagnetic energy), we calculate:

$$\nu = \frac{c}{\lambda} \text{ so } \frac{3.00 \times 10^8 m/s}{\frac{499nm \times 1m}{10^9 nm}} = 6.01 \times 10^{14} Hz$$

$$E = h\nu \text{ so } 6.63 \times 10^{-34} J \cdot s \times 6.01 \times 10^{14} Hz = 3.99 \times 10^{-19} \text{ Joules/ion}$$

Because the complex absorbs 600 nm (orange) through 450 (blue), the indigo, violet, and red wavelengths will be transmitted, and the complex will appear purple.

? Exercise 19.4.3

A complex that appears green, absorbs photons of what wavelengths?

Answer

red, 620–800 nm

Small changes in the relative energies of the orbitals that electrons are transitioning between can lead to drastic shifts in the color of light absorbed. Therefore, the colors of coordination compounds depend on many factors. As shown in Figure 19.4.6, different aqueous metal ions can have different colors. In addition, different oxidation states of one metal can produce different colors, as shown for the vanadium complexes in the link below.



Figure 19.4.6: The partially filled *d* orbitals of the stable ions $Cr^{3+}(aq)$, $Fe^{3+}(aq)$, and $Co^{2+}(aq)$ (left, center and right, respectively) give rise to various colors. (credit: Sahar Atwa)

The specific ligands coordinated to the metal center also influence the color of coordination complexes. For example, the iron(II) complex $[Fe(H_2O)_6]SO_4$ appears blue-green because the high-spin complex absorbs photons in the red wavelengths (Figure 19.4.7). In contrast, the low-spin iron(II) complex $K_4[Fe(CN)_6]$ appears pale yellow because it absorbs higher-energy violet photons.



Figure 19.4.7: Both (a) hexaaquairon(II) sulfate and (b) potassium hexacyanoferrate(II) contain d^6 iron(II) octahedral metal centers, but they absorb photons in different ranges of the visible spectrum.

Link to Learning

Watch this video of the reduction of vanadium complexes to observe the colorful effect of changing oxidation states.

In general, strong-field ligands cause a large split in the energies of *d* orbitals of the central metal atom (large Δ_{oct}). Transition metal coordination compounds with these ligands are yellow, orange, or red because they absorb higher-energy violet or blue light. On the other hand, coordination compounds of transition metals with weak-field ligands are often blue-green, blue, or indigo because they absorb lower-energy yellow, orange, or red light.

A coordination compound of the Cu^+ ion has a d^{10} configuration, and all the e_g orbitals are filled. To excite an electron to a higher level, such as the 4*p* orbital, photons of very high energy are necessary. This energy corresponds to very short wavelengths in the ultraviolet region of the spectrum. No visible light is absorbed, so the eye sees no change, and the compound appears white or colorless. A solution containing $[Cu(CN)_2]^-$, for example, is colorless. On the other hand, octahedral Cu^{2+} complexes have a vacancy in the e_g orbitals, and electrons can be excited to this level. The wavelength (energy) of the light absorbed corresponds to the visible part of the spectrum, and Cu^{2+} complexes are almost always colored—blue, blue-green violet, or yellow (Figure 19.4.8). Although CFT successfully describes many properties of coordination complexes, molecular orbital explanations (beyond the introductory scope provided here) are required to understand fully the behavior of coordination complexes.

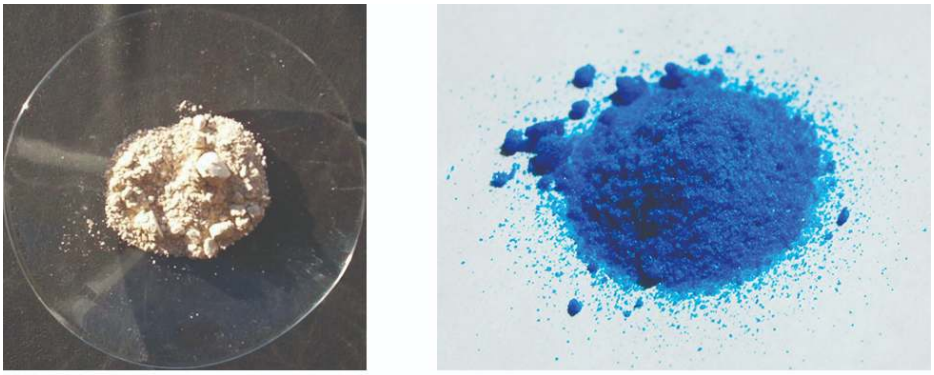


Figure 19.4.8: (a) Copper(I) complexes with d^{10} configurations such as CuI tend to be colorless, whereas (b) d^9 copper(II) complexes such as $Cu(NO_3)_2 \cdot 5H_2O$ are brightly colored.

This page titled 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **19.3: Spectroscopic and Magnetic Properties of Coordination Compounds** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

19.5: Key Terms

actinide series | (also, actinoid series) actinium and the elements in the second row or the *f*-block, atomic numbers 89–103

bidentate ligand | ligand that coordinates to one central metal through coordinate bonds from two different atoms

central metal | ion or atom to which one or more ligands is attached through coordinate covalent bonds

chelate | complex formed from a polydentate ligand attached to a central metal

chelating ligand | ligand that attaches to a central metal ion by bonds from two or more donor atoms

***cis* configuration** | configuration of a geometrical isomer in which two similar groups are on the same side of an imaginary reference line on the molecule

coordination compound | stable compound in which the central metal atom or ion acts as a Lewis acid and accepts one or more pairs of electrons

coordination compound | substance consisting of atoms, molecules, or ions attached to a central atom through Lewis acid-base interactions

coordination number | number of coordinate covalent bonds to the central metal atom in a complex or the number of closest contacts to an atom in a crystalline form

coordination sphere | central metal atom or ion plus the attached ligands of a complex

crystal field splitting (Δ_{oct}) | difference in energy between the t_{2g} and e_g sets or *t* and *e* sets of orbitals

crystal field theory | model that explains the energies of the orbitals in transition metals in terms of electrostatic interactions with the ligands but does not include metal ligand bonding

***d*-block element** | one of the elements in groups 3–11 with valence electrons in *d* orbitals

donor atom | atom in a ligand with a lone pair of electrons that forms a coordinate covalent bond to a central metal

e_g orbitals | set of two *d* orbitals that are oriented on the Cartesian axes for coordination complexes; in octahedral complexes, they are higher in energy than the t_{2g} orbitals

***f*-block element** | (also, inner transition element) one of the elements with atomic numbers 58–71 or 90–103 that have valence electrons in *f* orbitals; they are frequently shown offset below the periodic table

first transition series | transition elements in the fourth period of the periodic table (first row of the *d*-block), atomic numbers 21–29

fourth transition series | transition elements in the seventh period of the periodic table (fourth row of the *d*-block), atomic numbers 89 and 104–111

geometric isomers | isomers that differ in the way in which atoms are oriented in space relative to each other, leading to different physical and chemical properties

high-spin complex | complex in which the electrons maximize the total electron spin by singly populating all of the orbitals before pairing two electrons into the lower-energy orbitals

hydrometallurgy | process in which a metal is separated from a mixture by first converting it into soluble ions, extracting the ions, and then reducing the ions to precipitate the pure metal

ionization isomer | (or coordination isomer) isomer in which an anionic ligand is replaced by the counter ion in the inner coordination sphere

lanthanide series | (also, lanthanoid series) lanthanum and the elements in the first row or the *f*-block, atomic numbers 57–71

ligand | ion or neutral molecule attached to the central metal ion in a coordination compound

linkage isomer | coordination compound that possesses a ligand that can bind to the transition metal in two different ways (CN^- vs. NC^-)

low-spin complex | complex in which the electrons minimize the total electron spin by pairing in the lower-energy orbitals before populating the higher-energy orbitals

monodentate | ligand that attaches to a central metal through just one coordinate covalent bond

optical isomer | (also, enantiomer) molecule that is a nonsuperimposable mirror image with identical chemical and physical properties, except when it reacts with other optical isomers

pairing energy (P) | energy required to place two electrons with opposite spins into a single orbital

platinum metals | group of six transition metals consisting of ruthenium, osmium, rhodium, iridium, palladium, and platinum that tend to occur in the same minerals and demonstrate similar chemical properties

polydentate ligand | ligand that is attached to a central metal ion by bonds from two or more donor atoms, named with prefixes specifying how many donors are present (e.g., hexadentate = six coordinate bonds formed)

rare earth element | collection of 17 elements including the lanthanides, scandium, and yttrium that often occur together and have similar chemical properties, making separation difficult

second transition series | transition elements in the fifth period of the periodic table (second row of the *d*-block), atomic numbers 39–47

smelting | process of extracting a pure metal from a molten ore

spectrochemical series | ranking of ligands according to the magnitude of the crystal field splitting they induce

steel | material made from iron by removing impurities in the iron and adding substances that produce alloys with properties suitable for specific uses

strong-field ligand | ligand that causes larger crystal field splittings

superconductor | material that conducts electricity with no resistance

t_{2g} orbitals | set of three *d* orbitals aligned between the Cartesian axes for coordination complexes; in octahedral complexes, they are lowered in energy compared to the e_g orbitals according to CFT

third transition series | transition elements in the sixth period of the periodic table (third row of the *d*-block), atomic numbers 57 and 72–79

***trans* configuration** | configuration of a geometrical isomer in which two similar groups are on opposite sides of an imaginary reference line on the molecule

weak-field ligand | ligand that causes small crystal field splittings

This page titled 19.5: Key Terms is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **19.4: Key Terms** is licensed CC BY 4.0.

19.6: Summary

The transition metals are elements with partially filled *d* orbitals, located in the *d*-block of the periodic table. The reactivity of the transition elements varies widely from very active metals such as scandium and iron to almost inert elements, such as the platinum metals. The type of chemistry used in the isolation of the elements from their ores depends upon the concentration of the element in its ore and the difficulty of reducing ions of the elements to the metals. Metals that are more active are more difficult to reduce.

Transition metals exhibit chemical behavior typical of metals. For example, they oxidize in air upon heating and react with elemental halogens to form halides. Those elements that lie above hydrogen in the activity series react with acids, producing salts and hydrogen gas. Oxides, hydroxides, and carbonates of transition metal compounds in low oxidation states are basic. Halides and other salts are generally stable in water, although oxygen must be excluded in some cases. Most transition metals form a variety of stable oxidation states, allowing them to demonstrate a wide range of chemical reactivity.

The transition elements and main group elements can form coordination compounds, or complexes, in which a central metal atom or ion is bonded to one or more ligands by coordinate covalent bonds. Ligands with more than one donor atom are called polydentate ligands and form chelates. The common geometries found in complexes are tetrahedral and square planar (both with a coordination number of four) and octahedral (with a coordination number of six). *Cis* and *trans* configurations are possible in some octahedral and square planar complexes. In addition to these geometrical isomers, optical isomers (molecules or ions that are mirror images but not superimposable) are possible in certain octahedral complexes. Coordination complexes have a wide variety of uses including oxygen transport in blood, water purification, and pharmaceutical use.

Crystal field theory treats interactions between the electrons on the metal and the ligands as a simple electrostatic effect. The presence of the ligands near the metal ion changes the energies of the metal *d* orbitals relative to their energies in the free ion. Both the color and the magnetic properties of a complex can be attributed to this crystal field splitting. The magnitude of the splitting (Δ_{oct}) depends on the nature of the ligands bonded to the metal. Strong-field ligands produce large splitting and favor low-spin complexes, in which the t_{2g} orbitals are completely filled before any electrons occupy the e_g orbitals. Weak-field ligands favor formation of high-spin complexes. The t_{2g} and the e_g orbitals are singly occupied before any are doubly occupied.

This page titled 19.6: Summary is shared under a not declared license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

- **19.5: Summary** is licensed CC BY 4.0.

19.7: Exercises

1.

Write the electron configurations for each of the following elements:

(a) Sc

(b) Ti

(c) Cr

(d) Fe

(e) Ru

2.

Write the electron configurations for each of the following elements and its ions:

(a) Ti

(b) Ti^{2+}

(c) Ti^{3+}

(d) Ti^{4+}

3.

Write the electron configurations for each of the following elements and its 3+ ions:

(a) La

(b) Sm

(c) Lu

4.

Why are the lanthanoid elements not found in nature in their elemental forms?

5.

Which of the following elements is most likely to be used to prepare La by the reduction of La_2O_3: Al, C, or Fe? Why?

6.

Which of the following is the strongest oxidizing agent: VO_4^{3}, CrO_4^{2-}, or MnO_4^{-}?

7.

Which of the following elements is most likely to form an oxide with the formula MO_3: Zr, Nb, or Mo?

8.

The following reactions all occur in a blast furnace. Which of these are redox reactions?

(a) $3Fe_2O_3(s) + CO(g) \longrightarrow 2Fe_3O_4(s) + CO_2(g)$

(b) $Fe_3O_4(s) + CO(g) \longrightarrow 3FeO(s) + CO_2(g)$

(c) $FeO(s) + CO(g) \longrightarrow Fe(l) + CO_2(g)$

(d) $C(s) + O_2(g) \longrightarrow CO_2(g)$

(e) $C(s) + CO_2(g) \longrightarrow 2CO(g)$

(f) $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

(g) $CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)$

9.

Why is the formation of slag useful during the smelting of iron?

10.

Would you expect an aqueous manganese(VII) oxide solution to have a pH greater or less than 7.0? Justify your answer.

11.

Iron(II) can be oxidized to iron(III) by dichromate ion, which is reduced to chromium(III) in acid solution. A 2.5000-g sample of iron ore is dissolved and the iron converted into iron(II). Exactly 19.17 mL of 0.0100 *M* $Na_2Cr_2O_7$ is required in the titration. What percentage of the ore sample was iron?

12.

How many cubic feet of air at a pressure of 760 torr and 0 °C is required per ton of Fe_2O_3 to convert that Fe_2O_3 into iron in a blast furnace? For this exercise, assume air is 19% oxygen by volume.

13.

Find the potentials of the following electrochemical cell:

Cd | Cd^{2+}, *M* = 0.10 ‖ Ni^{2+}, *M* = 0.50 | Ni

14.

A 2.5624-g sample of a pure solid alkali metal chloride is dissolved in water and treated with excess silver nitrate. The resulting precipitate, filtered and dried, weighs 3.03707 g. What was the percent by mass of chloride ion in the original compound? What is the identity of the salt?

15.

The standard reduction potential for the reaction $[Co(H_2O)_6]^{3+}(aq) + e^- \longrightarrow [Co(H_2O)_6]^{2+}(aq)$ is about 1.8 V. The reduction potential for the reaction $[Co(NH_3)_6]^{3+}(aq) + e^- \longrightarrow [Co(NH_3)_6]^{2+}(aq)$ is +0.1 V. Calculate the cell potentials to show whether the complex ions, $[Co(H_2O)_6]^{2+}$ and/or $[Co(NH_3)_6]^{2+}$, can be oxidized to the corresponding cobalt(III) complex by oxygen.

16.

Predict the products of each of the following reactions. (Note: In addition to using the information in this chapter, also use the knowledge you have accumulated at this stage of your study, including information on the prediction of reaction products.)

(a) $MnCO_3(s) + HI(aq)$

(b) $CoO(s) + O_2(g)$

(c) $La(s) + O_2(g) \longrightarrow$

(d) $V(s) + VCl_4(s) \longrightarrow$

(e) $Co(s) + xsF_2(g) \longrightarrow$

(f) $CrO_3(s) + CsOH(aq) \longrightarrow$

17.

Predict the products of each of the following reactions. (Note: In addition to using the information in this chapter, also use the knowledge you have accumulated at this stage of your study, including information on the prediction of reaction products.)

(a) $\mathrm{Fe}(s) + \mathrm{H_2SO_4}(aq) \longrightarrow$

(b) $\mathrm{FeCl_3}(aq) + \mathrm{NaOH}(aq) \longrightarrow$

(c) $\mathrm{Mn(OH)_2}(s) + \mathrm{HBr}(aq) \longrightarrow$

(d) $\mathrm{Cr}(s) + \mathrm{O_2}(g) \longrightarrow$

(e) $\mathrm{Mn_2O_3}(s) + \mathrm{HCl}(aq) \longrightarrow$

(f) $\mathrm{Ti}(s) + x s \mathrm{F_2}(g) \longrightarrow$

18.

Describe the electrolytic process for refining copper.

19.

Predict the products of the following reactions and balance the equations.

(a) Zn is added to a solution of $Cr_2(SO_4)_3$ in acid.

(b) $FeCl_2$ is added to a solution containing an excess of $\mathrm{Cr_2O_7}^{2-}$ in hydrochloric acid.

(c) Cr^{2+} is added to $\mathrm{Cr_2O_7}^{2-}$ in acid solution.

(d) Mn is heated with CrO_3.

(e) CrO is added to $2HNO_3$ in water.

(f) $FeCl_3$ is added to an aqueous solution of NaOH.

20.

What is the gas produced when iron(II) sulfide is treated with a nonoxidizing acid?

21.

Predict the products of each of the following reactions and then balance the chemical equations.

(a) Fe is heated in an atmosphere of steam.

(b) NaOH is added to a solution of $Fe(NO_3)_3$.

(c) $FeSO_4$ is added to an acidic solution of $KMnO_4$.

(d) Fe is added to a dilute solution of H_2SO_4.

(e) A solution of $Fe(NO_3)_2$ and HNO_3 is allowed to stand in air.

(f) $FeCO_3$ is added to a solution of $HClO_4$.

(g) Fe is heated in air.

22.

Balance the following equations by oxidation-reduction methods; note that *three* elements change oxidation state.
$\mathrm{Co(NO_3)_2}(s) \longrightarrow \mathrm{Co_2O_3}(s) + \mathrm{NO_2}(g) + \mathrm{O_2}(g)$

23.

Dilute sodium cyanide solution is slowly dripped into a slowly stirred silver nitrate solution. A white precipitate forms temporarily but dissolves as the addition of sodium cyanide continues. Use chemical equations to explain this observation. Silver cyanide is similar to silver chloride in its solubility.

24.

Predict which will be more stable, $[CrO_4]^{2-}$ or $[WO_4]^{2-}$, and explain.

25.

Give the oxidation state of the metal for each of the following oxides of the first transition series. (Hint: Oxides of formula M_3O_4 are examples of *mixed valence compounds* in which the metal ion is present in more than one oxidation state. It is possible to write these compound formulas in the equivalent format $MO{\cdot}M_2O_3$, to permit estimation of the metal's two oxidation states.)

(a) Sc_2O_3

(b) TiO_2

(c) V_2O_5

(d) CrO_3

(e) MnO_2

(f) Fe_3O_4

(g) Co_3O_4

(h) NiO

(i) Cu_2O

26.

Indicate the coordination number for the central metal atom in each of the following coordination compounds:

(a) $[Pt(H_2O)_2Br_2]$

(b) $[Pt(NH_3)(py)(Cl)(Br)]$ (py = pyridine, C_5H_5N)

(c) $[Zn(NH_3)_2Cl_2]$

(d) $[Zn(NH_3)(py)(Cl)(Br)]$

(e) $[Ni(H_2O)_4Cl_2]$

(f) $[Fe(en)_2(CN)_2]^+$ (en = ethylenediamine, $C_2H_8N_2$)

27.

Give the coordination numbers and write the formulas for each of the following, including all isomers where appropriate:

(a) tetrahydroxozincate(II) ion (tetrahedral)

(b) hexacyanopalladate(IV) ion

(c) dichloroaurate(I) ion (note that *aurum* is Latin for "gold")

(d) diamminedichloroplatinum(II)

(e) potassium diamminetetrachlorochromate(III)

(f) hexaamminecobalt(III) hexacyanochromate(III)

(g) dibromobis(ethylenediamine) cobalt(III) nitrate

28.

Give the coordination number for each metal ion in the following compounds:

(a) $[Co(CO_3)_3]^{3-}$ (note that CO_3^{2-} is bidentate in this complex)

(b) $[Cu(NH_3)_4]^{2+}$

(c) $[Co(NH_3)_4Br_2]_2(SO_4)$

(d) $[Pt(NH_3)_4][PtCl_4]$

(e) $[Cr(en)_3](NO_3)_3$

(f) $[Pd(NH_3)_2Br_2]$ (square planar)

(g) $K_3[Cu(Cl)_5]$

(h) $[Zn(NH_3)_2Cl_2]$

29.

Sketch the structures of the following complexes. Indicate any *cis*, *trans*, and optical isomers.

(a) $[Pt(H_2O)_2Br_2]$ (square planar)

(b) $[Pt(NH_3)(py)(Cl)(Br)]$ (square planar, py = pyridine, C_5H_5N)

(c) $[Zn(NH_3)_3Cl]^+$ (tetrahedral)

(d) $[Pt(NH_3)_3Cl]^+$ (square planar)

(e) $[Ni(H_2O)_4Cl_2]$

(f) $[Co(C_2O_4)_2Cl_2]^{3-}$ (note that $C_2O_4{}^{2-}$ is the bidentate oxalate ion, $-O_2CCO_2{}^-$)

30.

Draw diagrams for any *cis*, *trans*, and optical isomers that could exist for the following (en is ethylenediamine):

(a) $[Co(en)_2(NO_2)Cl]^+$

(b) $[Co(en)_2Cl_2]^+$

(c) $[Pt(NH_3)_2Cl_4]$

(d) $[Cr(en)_3]^{3+}$

(e) $[Pt(NH_3)_2Cl_2]$

31.

Name each of the compounds or ions given in Exercise 19.28, including the oxidation state of the metal.

32.

Name each of the compounds or ions given in Exercise 19.30.

33.

Specify whether the following complexes have isomers.

(a) tetrahedral $[Ni(CO)_2(Cl)_2]$

(b) trigonal bipyramidal $[Mn(CO)_4NO]$

(c) $[Pt(en)_2Cl_2]Cl_2$

34.

Predict whether the carbonate ligand $CO_3{}^{2-}$ will coordinate to a metal center as a monodentate, bidentate, or tridentate ligand.

35.

Draw the geometric, linkage, and ionization isomers for $[CoCl_5CN][CN]$.

36.

Determine the number of unpaired electrons expected for $[Fe(NO_2)_6]^{3-}$and for $[FeF_6]^{3-}$ in terms of crystal field theory.

37.

Draw the crystal field diagrams for $[Fe(NO_2)_6]^{4-}$ and $[FeF_6]^{3-}$. State whether each complex is high spin or low spin, paramagnetic or diamagnetic, and compare Δ_{oct} to P for each complex.

38.

Give the oxidation state of the metal, number of *d* electrons, and the number of unpaired electrons predicted for $[Co(NH_3)_6]Cl_3$.

39.

The solid anhydrous solid $CoCl_2$ is blue in color. Because it readily absorbs water from the air, it is used as a humidity indicator to monitor if equipment (such as a cell phone) has been exposed to excessive levels of moisture. Predict what product is formed by this reaction, and how many unpaired electrons this complex will have.

40.

Is it possible for a complex of a metal in the transition series to have six unpaired electrons? Explain.

41.

How many unpaired electrons are present in each of the following?

(a) $[CoF_6]^{3-}$ (high spin)

(b) $[Mn(CN)_6]^{3-}$ (low spin)

(c) $[Mn(CN)_6]^{4-}$ (low spin)

(d) $[MnCl_6]^{4-}$ (high spin)

(e) $[RhCl_6]^{3-}$ (low spin)

42.

Explain how the diphosphate ion, $[O_3P{-}O{-}PO_3]^{4-}$, can function as a water softener that prevents the precipitation of Fe^{2+} as an insoluble iron salt.

43.

For complexes of the same metal ion with no change in oxidation number, the stability increases as the number of electrons in the t_{2g} orbitals increases. Which complex in each of the following pairs of complexes is more stable?

(a) $[Fe(H_2O)_6]^{2+}$ or $[Fe(CN)_6]^{4-}$

(b) $[Co(NH_3)_6]^{3+}$ or $[CoF_6]^{3-}$

(c) $[Mn(CN)_6]^{4-}$ or $[MnCl_6]^{4-}$

44.

Trimethylphosphine, $P(CH_3)_3$, can act as a ligand by donating the lone pair of electrons on the phosphorus atom. If trimethylphosphine is added to a solution of nickel(II) chloride in acetone, a blue compound that has a molecular mass of approximately 270 g and contains 21.5% Ni, 26.0% Cl, and 52.5% $P(CH_3)_3$ can be isolated. This blue compound does not have any isomeric forms. What are the geometry and molecular formula of the blue compound?

45.

Would you expect the complex $[Co(en)_3]Cl_3$ to have any unpaired electrons? Any isomers?

46.

Would you expect the $Mg_3[Cr(CN)_6]_2$ to be diamagnetic or paramagnetic? Explain your reasoning.

47.

Would you expect salts of the gold(I) ion, Au^+, to be colored? Explain.

48.

$[CuCl_4]^{2-}$ is green. $[Cu(H_2O)_6]^{2+}$is blue. Which absorbs higher-energy photons? Which is predicted to have a larger crystal field splitting?

This page titled 19.7: Exercises is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Marco Zimmer-De Iuliis and Anna Galang.

Index

A

active electrode
18.3: Galvanic Cells
alkaline battery
18.6: Batteries and Fuel Cells
amphiprotic
15.4: Self-Ionization of Water and the pH Scale
anode
18.3: Galvanic Cells
atom
1.2: Phases and Classification of Matter
atomic spectra
3.3: Atomic Spectra

B

basic oxides
16.7: Solubility and pH
battery
18.6: Batteries and Fuel Cells
boiling point elevation
13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions

C

cathode
18.3: Galvanic Cells
Cell Notation
18.3: Galvanic Cells
cell potential
18.3: Galvanic Cells
chemical change
1.3: Physical and Chemical Properties
chemical property
1.3: Physical and Chemical Properties
chemical tests
16.9: Qualitative Cation Analysis
circuit
18.2: Review of Redox Chemistry
common ion effect
16.3: Common-Ion Effect in Solubility Equilibria
16.4: Limitations of the K_{sp} Concept
16.7: Solubility and pH
complex ion
16.8: Equilibria Involving Complex Ions
compound
1.2: Phases and Classification of Matter
concentration cell
18.5: Potential, Free Energy, and Equilibrium
concentration units
13.2: Solution Concentration
consumer surplus
13.11: Exercises
current
18.2: Review of Redox Chemistry

D

deadweight loss
13.11: Exercises
Desalinization
13.7: Osmotic Pressure
dry cell
18.6: Batteries and Fuel Cells

E

economic surplus
13.11: Exercises
electrical potential
18.2: Review of Redox Chemistry
electrical work
18.5: Potential, Free Energy, and Equilibrium
electrolysis
18.7: Electrolysis
electrolytic cell
18.7: Electrolysis
electroplating
18.7: Electrolysis
element
1.2: Phases and Classification of Matter
exercises
16.10: Exercises
extensive property
1.3: Physical and Chemical Properties

F

Faraday's constant
18.5: Potential, Free Energy, and Equilibrium
formation constant
16.8: Equilibria Involving Complex Ions
freezing point depression
13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions
fuel cell
18.6: Batteries and Fuel Cells

G

galvanic cell
18.3: Galvanic Cells
gas
1.2: Phases and Classification of Matter

H

heterogeneous mixture
1.2: Phases and Classification of Matter
homogeneous mixture
1.2: Phases and Classification of Matter
hydrohelium
15.9: Molecular Structure and Acid-Base Behavior

I

inductive effect
15.9: Molecular Structure and Acid-Base Behavior
inert electrode
18.3: Galvanic Cells
intensive property
1.3: Physical and Chemical Properties
ion pair
16.4: Limitations of the K_{sp} Concept
ion product
16.1: Solubility Product Constant, Ksp
ionic solvent
18.7: Electrolysis

L

law of conservation of matter
1.2: Phases and Classification of Matter
lead acid battery
18.6: Batteries and Fuel Cells
ligands
16.8: Equilibria Involving Complex Ions
liquid
1.2: Phases and Classification of Matter
lithium ion battery
18.6: Batteries and Fuel Cells

M

mass
1.2: Phases and Classification of Matter
miscibility
13.1: Types of Solutions - Some Terminology
mixture
1.2: Phases and Classification of Matter
molecule
1.2: Phases and Classification of Matter

N

Nernst equation
18.5: Potential, Free Energy, and Equilibrium

O

osmotic pressure
13.7: Osmotic Pressure
overpotential
18.7: Electrolysis

P

pH scale
15.4: Self-Ionization of Water and the pH Scale
physical change
1.3: Physical and Chemical Properties
physical property
1.3: Physical and Chemical Properties
plasma
1.2: Phases and Classification of Matter
primary battery
18.6: Batteries and Fuel Cells
producer surplus
13.11: Exercises
pure substance
1.2: Phases and Classification of Matter

Q

qualitative analysis
16.9: Qualitative Cation Analysis
quantum theory
3.4: Quantum Theory

R

Raoult’s law
13.6: Vapor Pressures of Solutions
reverse osmosis
13.7: Osmotic Pressure

S

secondary battery
18.6: Batteries and Fuel Cells
selective precipitation
16.7: Solubility and pH

social surplus
13.11: Exercises
solid
1.2: Phases and Classification of Matter
solid solution
13.1: Types of Solutions - Some Terminology
solubility
16.2: Relationship Between Solubility and Ksp
Solubility of gases
13.5: Solubilities of Gases
solubility product constant
16.1: Solubility Product Constant, Ksp
solvation
13.3: Intermolecular Forces and the Solution Process
standard cell potential
18.4: Electrode and Cell Potentials
standard hydrogen electrode
18.4: Electrode and Cell Potentials
standard reduction potential
18.4: Electrode and Cell Potentials
strong acids
15.5: Strong Acids and Strong Bases
strong bases
15.5: Strong Acids and Strong Bases
V
voltaic cell
18.3: Galvanic Cells
W
weak acid
15.6: Weak Acids and Weak Bases
weak base
15.6: Weak Acids and Weak Bases
weight
1.2: Phases and Classification of Matter

Glossary

Sample Word 1 | Sample Definition 1

Detailed Licensing

Overview

Title: UTSC: First-Year Chemistry Textbook (Fall 2025)

Webpages: 242

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY 4.0: 49.2% (119 pages)
- Undeclared: 25.6% (62 pages)
- CC BY-NC-SA 4.0: 23.6% (57 pages)
- CC BY-NC-SA 3.0: 1.7% (4 pages)

By Page

- UTSC: First-Year Chemistry Textbook (Fall 2025) - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction to Chemistry - Fundamental Concepts and Skills - *Undeclared*
    - 1.1: Introduction - *Undeclared*
    - 1.2: Phases and Classification of Matter - *CC BY 4.0*
    - 1.3: Physical and Chemical Properties - *CC BY 4.0*
    - 1.4: Measurements - *CC BY 4.0*
    - 1.5: Measurement Uncertainty, Accuracy, and Precision - *CC BY 4.0*
    - 1.6: Mathematical Treatment of Measurement Results - *CC BY 4.0*
    - 1.7: Key Terms - *Undeclared*
    - 1.8: Key Equations - *Undeclared*
    - 1.9: Summary - *Undeclared*
    - 1.10: Exercises - *CC BY 4.0*
  - 2: Atoms, Molecules, and Ions - *CC BY 4.0*
    - 2.1: Introduction - *CC BY 4.0*
    - 2.2: Early Ideas in Atomic Theory - *CC BY 4.0*
    - 2.3: Evolution of Atomic Theory - *CC BY 4.0*
    - 2.4: Atomic Structure and Symbolism - *CC BY 4.0*
    - 2.5: Chemical Formulas - *CC BY 4.0*
    - 2.6: Key Terms - *CC BY 4.0*
    - 2.7: Key Equations - *CC BY 4.0*
    - 2.8: Summary - *CC BY 4.0*
    - 2.9: Exercises - *CC BY 4.0*
  - 3: The Quantum Model of the Atom - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Radiation - *CC BY-NC-SA 4.0*
    - 3.3: Atomic Spectra - *CC BY-NC-SA 4.0*
    - 3.4: Quantum Theory - *CC BY-NC-SA 4.0*
    - 3.5: The Bohr Atom - *CC BY-NC-SA 4.0*
    - 3.6: Two Ideas Leading to a New Quantum Mechanics - *CC BY-NC-SA 4.0*
    - 3.7: Wave Mechanics - *CC BY-NC-SA 4.0*
    - 3.8: Electron Spin- A Fourth Quantum Number - *CC BY-NC-SA 4.0*
    - 3.9: The Shape of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 3.10: Multielectron Atoms - *CC BY-NC-SA 4.0*
    - 3.11: Key Terms - *Undeclared*
    - 3.12: Key Equations - *Undeclared*
    - 3.13: Summary - *Undeclared*
    - 3.14: Exercises - *CC BY 4.0*
  - 4: The Periodic Table and Some Atomic Properties - *CC BY-NC-SA 4.0*
    - 4.1: Classifying the Elements- The Periodic Law and the Periodic Table - *CC BY-NC-SA 4.0*
    - 4.2: Metals and Nonmetals and their Ions - *CC BY-NC-SA 4.0*
    - 4.3: Sizes of Atoms and Ions - *CC BY-NC-SA 4.0*
    - 4.4: Ionization Energy - *CC BY-NC-SA 4.0*
    - 4.5: Electron Affinity - *CC BY-NC-SA 4.0*
    - 4.6: Magnetic Properties - *CC BY-NC-SA 4.0*
    - 4.7: Periodic Properties of the Elements - *CC BY-NC-SA 4.0*
    - 4.8: Key Terms - *CC BY 4.0*
    - 4.9: Key Equations - *CC BY 4.0*
    - 4.10: Summary - *CC BY 4.0*
    - 4.11: Exercises - *CC BY 4.0*
  - 5: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Ionic Bonding - *CC BY-NC-SA 4.0*
    - 5.3: Covalent Bonding - *CC BY 4.0*
    - 5.4: Chemical Nomenclature - *CC BY 4.0*

- 5.5: Lewis Symbols and Structures - *CC BY 4.0*
- 5.6: Formal Charges and Resonance - *CC BY 4.0*
- 5.7: Molecular Structure and Polarity - *CC BY 4.0*
- 5.8: Key Terms - *Undeclared*
- 5.9: Key Equations - *Undeclared*
- 5.10: Summary - *Undeclared*
- 5.11: Exercises - *CC BY 4.0*

6: Advanced Theories of Covalent Bonding - *CC BY 4.0*

- 6.1: Introduction - *CC BY 4.0*
- 6.2: Valence Bond Theory - *CC BY 4.0*
- 6.3: Hybrid Atomic Orbitals - *CC BY 4.0*
- 6.4: Multiple Bonds - *CC BY 4.0*
- 6.5: Molecular Orbital Theory - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*
- 6.9: Exercises - *CC BY 4.0*

7: Thermochemistry - *CC BY 4.0*

- 7.1: Introduction - *CC BY 4.0*
- 7.2: Energy Basics - *CC BY 4.0*
- 7.3: Calorimetry - *CC BY 4.0*
- 7.4: Enthalpy - *CC BY 4.0*
- 7.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 7.6: Key Terms - *Undeclared*
- 7.7: Key Equations - *Undeclared*
- 7.8: Summary - *Undeclared*
- 7.9: Exercises - *CC BY 4.0*

8: Liquids and Solids and Phase Changes - *CC BY 4.0*

- 8.1: Introduction - *CC BY 4.0*
- 8.2: Intermolecular Forces - *CC BY 4.0*
- 8.3: Properties of Liquids - *CC BY 4.0*
- 8.4: Phase Transitions - *CC BY 4.0*
- 8.5: Phase Diagrams - *CC BY 4.0*
- 8.6: The Solid State of Matter - *CC BY 4.0*
- 8.7: Key Terms - *Undeclared*
- 8.8: Key Equations - *Undeclared*
- 8.9: Summary - *Undeclared*
- 8.10: Exercises - *CC BY 4.0*

9: Gases - *CC BY 4.0*

- 9.1: Introduction - *Undeclared*
- 9.2: Gas Pressure - *CC BY 4.0*
- 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 9.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 9.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 9.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 9.8: Key Terms - *Undeclared*
- 9.9: Key Equations - *Undeclared*
- 9.10: Summary - *Undeclared*
- 9.11: Exercises - *CC BY 4.0*

10: Three Major Classes of Chemical Reactions - *Undeclared*

- 10.1: Formula Mass - *CC BY 4.0*
- 10.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 10.3: Reaction Stoichiometry - *CC BY 4.0*
- 10.4: Reaction Yields - *CC BY 4.0*
- 10.5: Molarity - *CC BY 4.0*
- 10.6: Precipitation Reactions - *CC BY-NC-SA 4.0*
- 10.7: Acid-Base Reactions - *CC BY-NC-SA 4.0*
- 10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations - *CC BY-NC-SA 4.0*
- 10.9: Oxidation-Reduction- Some General Principles - *CC BY-NC-SA 4.0*
- 10.10: Balancing Oxidation-Reduction Equations - *CC BY-NC-SA 4.0*
- 10.11: Oxidizing and Reducing Agents - *CC BY-NC-SA 4.0*
- 10.12: Key Terms - *Undeclared*
- 10.13: Key Equations - *Undeclared*
- 10.E: Reactions in Aqueous Solution (Exercises) - *CC BY-NC-SA 3.0*
- 10.S: Summary - *CC BY-NC-SA 3.0*

11: Nuclear Chemistry - *CC BY 4.0*

- 11.1: Introduction - *Undeclared*
- 11.2: Nuclear Structure and Stability - *CC BY 4.0*
- 11.3: Nuclear Equations - *CC BY 4.0*
- 11.4: Radioactive Decay - *CC BY 4.0*
- 11.5: Transmutation and Nuclear Energy - *CC BY 4.0*
- 11.6: Uses of Radioisotopes - *CC BY 4.0*
- 11.7: Biological Effects of Radiation - *CC BY 4.0*
- 11.8: Key Terms - *Undeclared*
- 11.9: Key Equations - *Undeclared*
- 11.10: Summary - *Undeclared*
- 11.11: Exercises - *CC BY 4.0*

12: Kinetics - *CC BY 4.0*

- 12.1: Introduction - *CC BY 4.0*
- 12.2: Chemical Reaction Rates - *CC BY 4.0*
- 12.3: Factors Affecting Reaction Rates - *CC BY 4.0*
- 12.4: Rate Laws - *CC BY 4.0*
- 12.5: Integrated Rate Laws - *CC BY 4.0*
- 12.6: Collision Theory - *CC BY 4.0*
- 12.7: Reaction Mechanisms - *CC BY 4.0*
- 12.8: Catalysis - *CC BY 4.0*
- 12.9: Key Terms - *Undeclared*
- 12.10: Key Equations - *Undeclared*
- 12.11: Summary - *Undeclared*
- 12.12: Exercises - *CC BY 4.0*

- 13: Solutions and their Physical Properties - *CC BY-NC-SA 4.0*
  - 13.1: Types of Solutions - Some Terminology - *CC BY-NC-SA 4.0*
  - 13.2: Solution Concentration - *CC BY-NC-SA 4.0*
  - 13.3: Intermolecular Forces and the Solution Process - *CC BY-NC-SA 4.0*
  - 13.4: Solution Formation and Equilibrium - *CC BY-NC-SA 4.0*
  - 13.5: Solubilities of Gases - *CC BY 4.0*
  - 13.6: Vapor Pressures of Solutions - *CC BY-NC-SA 4.0*
  - 13.7: Osmotic Pressure - *CC BY-NC-SA 4.0*
  - 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions - *CC BY-NC-SA 4.0*
  - 13.9: Solutions of Electrolytes - *CC BY-NC-SA 4.0*
  - 13.10: Colloidal Mixtures - *CC BY-NC-SA 4.0*
  - 13.11: Exercises - *CC BY 4.0*
- 14: Fundamental Equilibrium Concepts - *CC BY 4.0*
  - 14.1: Introduction - *CC BY 4.0*
  - 14.2: Chemical Equilibria - *CC BY 4.0*
  - 14.3: Equilibrium Constants - *CC BY 4.0*
  - 14.4: Shifting Equilibria - Le Chatelier's Principle - *CC BY 4.0*
  - 14.5: Equilibrium Calculations - *CC BY 4.0*
  - 14.6: Key Terms - *Undeclared*
  - 14.7: Key Equations - *Undeclared*
  - 14.8: Summary - *Undeclared*
  - 14.9: Exercises - *CC BY 4.0*
- 15: Acid-Base Equilibria - A More Detailed Look - *CC BY 4.0*
  - 15.1: Introduction - *Undeclared*
  - 15.2: Arrhenius Theory- A Brief Review - *CC BY-NC-SA 4.0*
  - 15.3: Brønsted-Lowry Theory of Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.4: Self-Ionization of Water and the pH Scale - *CC BY-NC-SA 4.0*
  - 15.5: Strong Acids and Strong Bases - *CC BY-NC-SA 4.0*
  - 15.6: Weak Acids and Weak Bases - *CC BY-NC-SA 3.0*
  - 15.7: Polyprotic Acids - *CC BY 4.0*
  - 15.8: Ions as Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.9: Molecular Structure and Acid-Base Behavior - *CC BY-NC-SA 4.0*
  - 15.10: Lewis Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.11: Common-Ion Effect in Acid-Base Equilibria - *CC BY-NC-SA 4.0*
  - 15.12: Buffer Solutions - *CC BY-NC-SA 4.0*
  - 15.13: Acid-Base Indicators - *CC BY-NC-SA 4.0*
  - 15.14: Neutralization Reactions and Titration Curves - *CC BY-NC-SA 4.0*
  - 15.15: Solutions of Salts of Polyprotic Acids - *CC BY-NC-SA 4.0*
  - 15.16: Acid-Base Equilibrium Calculations- A Summary - *CC BY-NC-SA 4.0*
  - 15.17: Key Terms - *Undeclared*
  - 15.18: Key Equations - *Undeclared*
  - 15.19: Summary - *Undeclared*
  - 15.20: Exercises - *CC BY-NC-SA 4.0*
- 16: Solubility and Complex-Ion Equilibria - *CC BY-NC-SA 4.0*
  - 16.1: Solubility Product Constant, Ksp - *CC BY-NC-SA 3.0*
  - 16.2: Relationship Between Solubility and Ksp - *CC BY 4.0*
  - 16.3: Common-Ion Effect in Solubility Equilibria - *CC BY-NC-SA 4.0*
  - 16.4: Limitations of the K_{sp} Concept - *CC BY-NC-SA 4.0*
  - 16.5: Criteria for Precipitation and its Completeness - *CC BY-NC-SA 4.0*
  - 16.6: Fractional Precipitation - *CC BY-NC-SA 4.0*
  - 16.7: Solubility and pH - *CC BY-NC-SA 4.0*
  - 16.8: Equilibria Involving Complex Ions - *CC BY-NC-SA 4.0*
  - 16.9: Qualitative Cation Analysis - *CC BY-NC-SA 4.0*
  - 16.10: Exercises - *CC BY-NC-SA 4.0*
- 17: Thermodynamics - *CC BY 4.0*
  - 17.1: Introduction - *CC BY 4.0*
  - 17.2: Enthalpy - *CC BY 4.0*
  - 17.3: Spontaneity - *CC BY 4.0*
  - 17.4: Entropy - *CC BY 4.0*
  - 17.5: The Second and Third Laws of Thermodynamics - *CC BY 4.0*
  - 17.6: Free Energy - *CC BY 4.0*
  - 17.7: Key Terms - *Undeclared*
  - 17.8: Key Equations - *Undeclared*
  - 17.9: Summary - *Undeclared*
  - 17.10: Exercises - *CC BY 4.0*
- 18: Electrochemistry - *CC BY 4.0*
  - 18.1: Introduction - *CC BY 4.0*
  - 18.2: Review of Redox Chemistry - *CC BY 4.0*
  - 18.3: Galvanic Cells - *CC BY 4.0*
  - 18.4: Electrode and Cell Potentials - *CC BY 4.0*
  - 18.5: Potential, Free Energy, and Equilibrium - *CC BY 4.0*
  - 18.6: Batteries and Fuel Cells - *CC BY 4.0*
  - 18.7: Electrolysis - *CC BY 4.0*
  - 18.8: Key Terms - *CC BY 4.0*

- 18.9: Key Equations - *CC BY 4.0*
- 18.10: Summary - *CC BY 4.0*
- 18.11: Exercises - *CC BY 4.0*

19: Transition Metals and Coordination Chemistry - *CC BY 4.0*

- 19.1: Introduction - *Undeclared*
- 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds - *CC BY 4.0*
- 19.3: Coordination Chemistry of Transition Metals - *CC BY 4.0*
- 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds - *CC BY 4.0*
- 19.5: Key Terms - *Undeclared*
- 19.6: Summary - *Undeclared*
- 19.7: Exercises - *CC BY 4.0*

Back Matter - *Undeclared*

- Index - *Undeclared*
- Glossary - *Undeclared*
- Detailed Licensing - *Undeclared*
- Detailed Licensing - *Undeclared*
- Detailed Licensing - *Undeclared*
- Detailed Licensing - *Undeclared*

Detailed Licensing

Overview

Title: UTSC: First-Year Chemistry Textbook (Fall 2025)

Webpages: 242

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY 4.0: 49.2% (119 pages)
- Undeclared: 25.6% (62 pages)
- CC BY-NC-SA 4.0: 23.6% (57 pages)
- CC BY-NC-SA 3.0: 1.7% (4 pages)

By Page

- UTSC: First-Year Chemistry Textbook (Fall 2025) - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction to Chemistry - Fundamental Concepts and Skills - *Undeclared*
    - 1.1: Introduction - *Undeclared*
    - 1.2: Phases and Classification of Matter - *CC BY 4.0*
    - 1.3: Physical and Chemical Properties - *CC BY 4.0*
    - 1.4: Measurements - *CC BY 4.0*
    - 1.5: Measurement Uncertainty, Accuracy, and Precision - *CC BY 4.0*
    - 1.6: Mathematical Treatment of Measurement Results - *CC BY 4.0*
    - 1.7: Key Terms - *Undeclared*
    - 1.8: Key Equations - *Undeclared*
    - 1.9: Summary - *Undeclared*
    - 1.10: Exercises - *CC BY 4.0*
  - 2: Atoms, Molecules, and Ions - *CC BY 4.0*
    - 2.1: Introduction - *CC BY 4.0*
    - 2.2: Early Ideas in Atomic Theory - *CC BY 4.0*
    - 2.3: Evolution of Atomic Theory - *CC BY 4.0*
    - 2.4: Atomic Structure and Symbolism - *CC BY 4.0*
    - 2.5: Chemical Formulas - *CC BY 4.0*
    - 2.6: Key Terms - *CC BY 4.0*
    - 2.7: Key Equations - *CC BY 4.0*
    - 2.8: Summary - *CC BY 4.0*
    - 2.9: Exercises - *CC BY 4.0*
  - 3: The Quantum Model of the Atom - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Radiation - *CC BY-NC-SA 4.0*
    - 3.3: Atomic Spectra - *CC BY-NC-SA 4.0*
    - 3.4: Quantum Theory - *CC BY-NC-SA 4.0*
    - 3.5: The Bohr Atom - *CC BY-NC-SA 4.0*
    - 3.6: Two Ideas Leading to a New Quantum Mechanics - *CC BY-NC-SA 4.0*
    - 3.7: Wave Mechanics - *CC BY-NC-SA 4.0*
    - 3.8: Electron Spin- A Fourth Quantum Number - *CC BY-NC-SA 4.0*
    - 3.9: The Shape of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 3.10: Multielectron Atoms - *CC BY-NC-SA 4.0*
    - 3.11: Key Terms - *Undeclared*
    - 3.12: Key Equations - *Undeclared*
    - 3.13: Summary - *Undeclared*
    - 3.14: Exercises - *CC BY 4.0*
  - 4: The Periodic Table and Some Atomic Properties - *CC BY-NC-SA 4.0*
    - 4.1: Classifying the Elements- The Periodic Law and the Periodic Table - *CC BY-NC-SA 4.0*
    - 4.2: Metals and Nonmetals and their Ions - *CC BY-NC-SA 4.0*
    - 4.3: Sizes of Atoms and Ions - *CC BY-NC-SA 4.0*
    - 4.4: Ionization Energy - *CC BY-NC-SA 4.0*
    - 4.5: Electron Affinity - *CC BY-NC-SA 4.0*
    - 4.6: Magnetic Properties - *CC BY-NC-SA 4.0*
    - 4.7: Periodic Properties of the Elements - *CC BY-NC-SA 4.0*
    - 4.8: Key Terms - *CC BY 4.0*
    - 4.9: Key Equations - *CC BY 4.0*
    - 4.10: Summary - *CC BY 4.0*
    - 4.11: Exercises - *CC BY 4.0*
  - 5: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Ionic Bonding - *CC BY-NC-SA 4.0*
    - 5.3: Covalent Bonding - *CC BY 4.0*
    - 5.4: Chemical Nomenclature - *CC BY 4.0*

- 5.5: Lewis Symbols and Structures - *CC BY 4.0*
- 5.6: Formal Charges and Resonance - *CC BY 4.0*
- 5.7: Molecular Structure and Polarity - *CC BY 4.0*
- 5.8: Key Terms - *Undeclared*
- 5.9: Key Equations - *Undeclared*
- 5.10: Summary - *Undeclared*
- 5.11: Exercises - *CC BY 4.0*

6: Advanced Theories of Covalent Bonding - *CC BY 4.0*

- 6.1: Introduction - *CC BY 4.0*
- 6.2: Valence Bond Theory - *CC BY 4.0*
- 6.3: Hybrid Atomic Orbitals - *CC BY 4.0*
- 6.4: Multiple Bonds - *CC BY 4.0*
- 6.5: Molecular Orbital Theory - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*
- 6.9: Exercises - *CC BY 4.0*

7: Thermochemistry - *CC BY 4.0*

- 7.1: Introduction - *CC BY 4.0*
- 7.2: Energy Basics - *CC BY 4.0*
- 7.3: Calorimetry - *CC BY 4.0*
- 7.4: Enthalpy - *CC BY 4.0*
- 7.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 7.6: Key Terms - *Undeclared*
- 7.7: Key Equations - *Undeclared*
- 7.8: Summary - *Undeclared*
- 7.9: Exercises - *CC BY 4.0*

8: Liquids and Solids and Phase Changes - *CC BY 4.0*

- 8.1: Introduction - *CC BY 4.0*
- 8.2: Intermolecular Forces - *CC BY 4.0*
- 8.3: Properties of Liquids - *CC BY 4.0*
- 8.4: Phase Transitions - *CC BY 4.0*
- 8.5: Phase Diagrams - *CC BY 4.0*
- 8.6: The Solid State of Matter - *CC BY 4.0*
- 8.7: Key Terms - *Undeclared*
- 8.8: Key Equations - *Undeclared*
- 8.9: Summary - *Undeclared*
- 8.10: Exercises - *CC BY 4.0*

9: Gases - *CC BY 4.0*

- 9.1: Introduction - *Undeclared*
- 9.2: Gas Pressure - *CC BY 4.0*
- 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 9.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 9.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 9.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 9.8: Key Terms - *Undeclared*
- 9.9: Key Equations - *Undeclared*
- 9.10: Summary - *Undeclared*
- 9.11: Exercises - *CC BY 4.0*

10: Three Major Classes of Chemical Reactions - *Undeclared*

- 10.1: Formula Mass - *CC BY 4.0*
- 10.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 10.3: Reaction Stoichiometry - *CC BY 4.0*
- 10.4: Reaction Yields - *CC BY 4.0*
- 10.5: Molarity - *CC BY 4.0*
- 10.6: Precipitation Reactions - *CC BY-NC-SA 4.0*
- 10.7: Acid-Base Reactions - *CC BY-NC-SA 4.0*
- 10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations - *CC BY-NC-SA 4.0*
- 10.9: Oxidation-Reduction- Some General Principles - *CC BY-NC-SA 4.0*
- 10.10: Balancing Oxidation-Reduction Equations - *CC BY-NC-SA 4.0*
- 10.11: Oxidizing and Reducing Agents - *CC BY-NC-SA 4.0*
- 10.12: Key Terms - *Undeclared*
- 10.13: Key Equations - *Undeclared*
- 10.E: Reactions in Aqueous Solution (Exercises) - *CC BY-NC-SA 3.0*
- 10.S: Summary - *CC BY-NC-SA 3.0*

11: Nuclear Chemistry - *CC BY 4.0*

- 11.1: Introduction - *Undeclared*
- 11.2: Nuclear Structure and Stability - *CC BY 4.0*
- 11.3: Nuclear Equations - *CC BY 4.0*
- 11.4: Radioactive Decay - *CC BY 4.0*
- 11.5: Transmutation and Nuclear Energy - *CC BY 4.0*
- 11.6: Uses of Radioisotopes - *CC BY 4.0*
- 11.7: Biological Effects of Radiation - *CC BY 4.0*
- 11.8: Key Terms - *Undeclared*
- 11.9: Key Equations - *Undeclared*
- 11.10: Summary - *Undeclared*
- 11.11: Exercises - *CC BY 4.0*

12: Kinetics - *CC BY 4.0*

- 12.1: Introduction - *CC BY 4.0*
- 12.2: Chemical Reaction Rates - *CC BY 4.0*
- 12.3: Factors Affecting Reaction Rates - *CC BY 4.0*
- 12.4: Rate Laws - *CC BY 4.0*
- 12.5: Integrated Rate Laws - *CC BY 4.0*
- 12.6: Collision Theory - *CC BY 4.0*
- 12.7: Reaction Mechanisms - *CC BY 4.0*
- 12.8: Catalysis - *CC BY 4.0*
- 12.9: Key Terms - *Undeclared*
- 12.10: Key Equations - *Undeclared*
- 12.11: Summary - *Undeclared*
- 12.12: Exercises - *CC BY 4.0*

- 13: Solutions and their Physical Properties - *CC BY-NC-SA 4.0*
  - 13.1: Types of Solutions - Some Terminology - *CC BY-NC-SA 4.0*
  - 13.2: Solution Concentration - *CC BY-NC-SA 4.0*
  - 13.3: Intermolecular Forces and the Solution Process - *CC BY-NC-SA 4.0*
  - 13.4: Solution Formation and Equilibrium - *CC BY-NC-SA 4.0*
  - 13.5: Solubilities of Gases - *CC BY 4.0*
  - 13.6: Vapor Pressures of Solutions - *CC BY-NC-SA 4.0*
  - 13.7: Osmotic Pressure - *CC BY-NC-SA 4.0*
  - 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions - *CC BY-NC-SA 4.0*
  - 13.9: Solutions of Electrolytes - *CC BY-NC-SA 4.0*
  - 13.10: Colloidal Mixtures - *CC BY-NC-SA 4.0*
  - 13.11: Exercises - *CC BY 4.0*
- 14: Fundamental Equilibrium Concepts - *CC BY 4.0*
  - 14.1: Introduction - *CC BY 4.0*
  - 14.2: Chemical Equilibria - *CC BY 4.0*
  - 14.3: Equilibrium Constants - *CC BY 4.0*
  - 14.4: Shifting Equilibria - Le Chatelier's Principle - *CC BY 4.0*
  - 14.5: Equilibrium Calculations - *CC BY 4.0*
  - 14.6: Key Terms - *Undeclared*
  - 14.7: Key Equations - *Undeclared*
  - 14.8: Summary - *Undeclared*
  - 14.9: Exercises - *CC BY 4.0*
- 15: Acid-Base Equilibria - A More Detailed Look - *CC BY 4.0*
  - 15.1: Introduction - *Undeclared*
  - 15.2: Arrhenius Theory- A Brief Review - *CC BY-NC-SA 4.0*
  - 15.3: Brønsted-Lowry Theory of Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.4: Self-Ionization of Water and the pH Scale - *CC BY-NC-SA 4.0*
  - 15.5: Strong Acids and Strong Bases - *CC BY-NC-SA 4.0*
  - 15.6: Weak Acids and Weak Bases - *CC BY-NC-SA 3.0*
  - 15.7: Polyprotic Acids - *CC BY 4.0*
  - 15.8: Ions as Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.9: Molecular Structure and Acid-Base Behavior - *CC BY-NC-SA 4.0*
  - 15.10: Lewis Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.11: Common-Ion Effect in Acid-Base Equilibria - *CC BY-NC-SA 4.0*
  - 15.12: Buffer Solutions - *CC BY-NC-SA 4.0*
  - 15.13: Acid-Base Indicators - *CC BY-NC-SA 4.0*
  - 15.14: Neutralization Reactions and Titration Curves - *CC BY-NC-SA 4.0*
  - 15.15: Solutions of Salts of Polyprotic Acids - *CC BY-NC-SA 4.0*
  - 15.16: Acid-Base Equilibrium Calculations- A Summary - *CC BY-NC-SA 4.0*
  - 15.17: Key Terms - *Undeclared*
  - 15.18: Key Equations - *Undeclared*
  - 15.19: Summary - *Undeclared*
  - 15.E: Exercises - *CC BY-NC-SA 4.0*
- 16: Solubility and Complex-Ion Equilibria - *CC BY-NC-SA 4.0*
  - 16.1: Solubility Product Constant, Ksp - *CC BY-NC-SA 3.0*
  - 16.2: Relationship Between Solubility and Ksp - *CC BY 4.0*
  - 16.3: Common-Ion Effect in Solubility Equilibria - *CC BY-NC-SA 4.0*
  - 16.4: Limitations of the K_{sp} Concept - *CC BY-NC-SA 4.0*
  - 16.5: Criteria for Precipitation and its Completeness - *CC BY-NC-SA 4.0*
  - 16.6: Fractional Precipitation - *CC BY-NC-SA 4.0*
  - 16.7: Solubility and pH - *CC BY-NC-SA 4.0*
  - 16.8: Equilibria Involving Complex Ions - *CC BY-NC-SA 4.0*
  - 16.9: Qualitative Cation Analysis - *CC BY-NC-SA 4.0*
  - 16.10: Exercises - *CC BY-NC-SA 4.0*
- 17: Thermodynamics - *CC BY 4.0*
  - 17.1: Introduction - *CC BY 4.0*
  - 17.2: Enthalpy - *CC BY 4.0*
  - 17.3: Spontaneity - *CC BY 4.0*
  - 17.4: Entropy - *CC BY 4.0*
  - 17.5: The Second and Third Laws of Thermodynamics - *CC BY 4.0*
  - 17.6: Free Energy - *CC BY 4.0*
  - 17.7: Key Terms - *Undeclared*
  - 17.8: Key Equations - *Undeclared*
  - 17.9: Summary - *Undeclared*
  - 17.10: Exercises - *CC BY 4.0*
- 18: Electrochemistry - *CC BY 4.0*
  - 18.1: Introduction - *CC BY 4.0*
  - 18.2: Review of Redox Chemistry - *CC BY 4.0*
  - 18.3: Galvanic Cells - *CC BY 4.0*
  - 18.4: Electrode and Cell Potentials - *CC BY 4.0*
  - 18.5: Potential, Free Energy, and Equilibrium - *CC BY 4.0*
  - 18.6: Batteries and Fuel Cells - *CC BY 4.0*
  - 18.7: Electrolysis - *CC BY 4.0*
  - 18.8: Key Terms - *CC BY 4.0*

- 18.9: Key Equations - *CC BY 4.0*
- 18.10: Summary - *CC BY 4.0*
- 18.11: Exercises - *CC BY 4.0*

- 19: Transition Metals and Coordination Chemistry - *CC BY 4.0*
  - 19.1: Introduction - *Undeclared*
  - 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds - *CC BY 4.0*
  - 19.3: Coordination Chemistry of Transition Metals - *CC BY 4.0*
  - 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds - *CC BY 4.0*
  - 19.5: Key Terms - *Undeclared*
  - 19.6: Summary - *Undeclared*
  - 19.7: Exercises - *CC BY 4.0*

- Back Matter - *Undeclared*
  - Index - *Undeclared*
  - Glossary - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*

Detailed Licensing

Overview

Title: UTSC: First-Year Chemistry Textbook (Fall 2025)

Webpages: 242

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY 4.0: 49.2% (119 pages)
- Undeclared: 25.6% (62 pages)
- CC BY-NC-SA 4.0: 23.6% (57 pages)
- CC BY-NC-SA 3.0: 1.7% (4 pages)

By Page

- UTSC: First-Year Chemistry Textbook (Fall 2025) - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction to Chemistry - Fundamental Concepts and Skills - *Undeclared*
    - 1.1: Introduction - *Undeclared*
    - 1.2: Phases and Classification of Matter - *CC BY 4.0*
    - 1.3: Physical and Chemical Properties - *CC BY 4.0*
    - 1.4: Measurements - *CC BY 4.0*
    - 1.5: Measurement Uncertainty, Accuracy, and Precision - *CC BY 4.0*
    - 1.6: Mathematical Treatment of Measurement Results - *CC BY 4.0*
    - 1.7: Key Terms - *Undeclared*
    - 1.8: Key Equations - *Undeclared*
    - 1.9: Summary - *Undeclared*
    - 1.10: Exercises - *CC BY 4.0*
  - 2: Atoms, Molecules, and Ions - *CC BY 4.0*
    - 2.1: Introduction - *CC BY 4.0*
    - 2.2: Early Ideas in Atomic Theory - *CC BY 4.0*
    - 2.3: Evolution of Atomic Theory - *CC BY 4.0*
    - 2.4: Atomic Structure and Symbolism - *CC BY 4.0*
    - 2.5: Chemical Formulas - *CC BY 4.0*
    - 2.6: Key Terms - *CC BY 4.0*
    - 2.7: Key Equations - *CC BY 4.0*
    - 2.8: Summary - *CC BY 4.0*
    - 2.9: Exercises - *CC BY 4.0*
  - 3: The Quantum Model of the Atom - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Radiation - *CC BY-NC-SA 4.0*
    - 3.3: Atomic Spectra - *CC BY-NC-SA 4.0*
    - 3.4: Quantum Theory - *CC BY-NC-SA 4.0*
    - 3.5: The Bohr Atom - *CC BY-NC-SA 4.0*
    - 3.6: Two Ideas Leading to a New Quantum Mechanics - *CC BY-NC-SA 4.0*
    - 3.7: Wave Mechanics - *CC BY-NC-SA 4.0*
    - 3.8: Electron Spin- A Fourth Quantum Number - *CC BY-NC-SA 4.0*
    - 3.9: The Shape of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 3.10: Multielectron Atoms - *CC BY-NC-SA 4.0*
    - 3.11: Key Terms - *Undeclared*
    - 3.12: Key Equations - *Undeclared*
    - 3.13: Summary - *Undeclared*
    - 3.14: Exercises - *CC BY 4.0*
  - 4: The Periodic Table and Some Atomic Properties - *CC BY-NC-SA 4.0*
    - 4.1: Classifying the Elements- The Periodic Law and the Periodic Table - *CC BY-NC-SA 4.0*
    - 4.2: Metals and Nonmetals and their Ions - *CC BY-NC-SA 4.0*
    - 4.3: Sizes of Atoms and Ions - *CC BY-NC-SA 4.0*
    - 4.4: Ionization Energy - *CC BY-NC-SA 4.0*
    - 4.5: Electron Affinity - *CC BY-NC-SA 4.0*
    - 4.6: Magnetic Properties - *CC BY-NC-SA 4.0*
    - 4.7: Periodic Properties of the Elements - *CC BY-NC-SA 4.0*
    - 4.8: Key Terms - *CC BY 4.0*
    - 4.9: Key Equations - *CC BY 4.0*
    - 4.10: Summary - *CC BY 4.0*
    - 4.11: Exercises - *CC BY 4.0*
  - 5: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Ionic Bonding - *CC BY-NC-SA 4.0*
    - 5.3: Covalent Bonding - *CC BY 4.0*
    - 5.4: Chemical Nomenclature - *CC BY 4.0*

- 5.5: Lewis Symbols and Structures - *CC BY 4.0*
- 5.6: Formal Charges and Resonance - *CC BY 4.0*
- 5.7: Molecular Structure and Polarity - *CC BY 4.0*
- 5.8: Key Terms - *Undeclared*
- 5.9: Key Equations - *Undeclared*
- 5.10: Summary - *Undeclared*
- 5.11: Exercises - *CC BY 4.0*

6: Advanced Theories of Covalent Bonding - *CC BY 4.0*

- 6.1: Introduction - *CC BY 4.0*
- 6.2: Valence Bond Theory - *CC BY 4.0*
- 6.3: Hybrid Atomic Orbitals - *CC BY 4.0*
- 6.4: Multiple Bonds - *CC BY 4.0*
- 6.5: Molecular Orbital Theory - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*
- 6.9: Exercises - *CC BY 4.0*

7: Thermochemistry - *CC BY 4.0*

- 7.1: Introduction - *CC BY 4.0*
- 7.2: Energy Basics - *CC BY 4.0*
- 7.3: Calorimetry - *CC BY 4.0*
- 7.4: Enthalpy - *CC BY 4.0*
- 7.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 7.6: Key Terms - *Undeclared*
- 7.7: Key Equations - *Undeclared*
- 7.8: Summary - *Undeclared*
- 7.9: Exercises - *CC BY 4.0*

8: Liquids and Solids and Phase Changes - *CC BY 4.0*

- 8.1: Introduction - *CC BY 4.0*
- 8.2: Intermolecular Forces - *CC BY 4.0*
- 8.3: Properties of Liquids - *CC BY 4.0*
- 8.4: Phase Transitions - *CC BY 4.0*
- 8.5: Phase Diagrams - *CC BY 4.0*
- 8.6: The Solid State of Matter - *CC BY 4.0*
- 8.7: Key Terms - *Undeclared*
- 8.8: Key Equations - *Undeclared*
- 8.9: Summary - *Undeclared*
- 8.10: Exercises - *CC BY 4.0*

9: Gases - *CC BY 4.0*

- 9.1: Introduction - *Undeclared*
- 9.2: Gas Pressure - *CC BY 4.0*
- 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 9.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 9.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 9.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 9.8: Key Terms - *Undeclared*
- 9.9: Key Equations - *Undeclared*
- 9.10: Summary - *Undeclared*
- 9.11: Exercises - *CC BY 4.0*

10: Three Major Classes of Chemical Reactions - *Undeclared*

- 10.1: Formula Mass - *CC BY 4.0*
- 10.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 10.3: Reaction Stoichiometry - *CC BY 4.0*
- 10.4: Reaction Yields - *CC BY 4.0*
- 10.5: Molarity - *CC BY 4.0*
- 10.6: Precipitation Reactions - *CC BY-NC-SA 4.0*
- 10.7: Acid-Base Reactions - *CC BY-NC-SA 4.0*
- 10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations - *CC BY-NC-SA 4.0*
- 10.9: Oxidation-Reduction- Some General Principles - *CC BY-NC-SA 4.0*
- 10.10: Balancing Oxidation-Reduction Equations - *CC BY-NC-SA 4.0*
- 10.11: Oxidizing and Reducing Agents - *CC BY-NC-SA 4.0*
- 10.12: Key Terms - *Undeclared*
- 10.13: Key Equations - *Undeclared*
- 10.E: Reactions in Aqueous Solution (Exercises) - *CC BY-NC-SA 3.0*
- 10.S: Summary - *CC BY-NC-SA 3.0*

11: Nuclear Chemistry - *CC BY 4.0*

- 11.1: Introduction - *Undeclared*
- 11.2: Nuclear Structure and Stability - *CC BY 4.0*
- 11.3: Nuclear Equations - *CC BY 4.0*
- 11.4: Radioactive Decay - *CC BY 4.0*
- 11.5: Transmutation and Nuclear Energy - *CC BY 4.0*
- 11.6: Uses of Radioisotopes - *CC BY 4.0*
- 11.7: Biological Effects of Radiation - *CC BY 4.0*
- 11.8: Key Terms - *Undeclared*
- 11.9: Key Equations - *Undeclared*
- 11.10: Summary - *Undeclared*
- 11.11: Exercises - *CC BY 4.0*

12: Kinetics - *CC BY 4.0*

- 12.1: Introduction - *CC BY 4.0*
- 12.2: Chemical Reaction Rates - *CC BY 4.0*
- 12.3: Factors Affecting Reaction Rates - *CC BY 4.0*
- 12.4: Rate Laws - *CC BY 4.0*
- 12.5: Integrated Rate Laws - *CC BY 4.0*
- 12.6: Collision Theory - *CC BY 4.0*
- 12.7: Reaction Mechanisms - *CC BY 4.0*
- 12.8: Catalysis - *CC BY 4.0*
- 12.9: Key Terms - *Undeclared*
- 12.10: Key Equations - *Undeclared*
- 12.11: Summary - *Undeclared*
- 12.12: Exercises - *CC BY 4.0*

- 13: Solutions and their Physical Properties - *CC BY-NC-SA 4.0*
  - 13.1: Types of Solutions - Some Terminology - *CC BY-NC-SA 4.0*
  - 13.2: Solution Concentration - *CC BY-NC-SA 4.0*
  - 13.3: Intermolecular Forces and the Solution Process - *CC BY-NC-SA 4.0*
  - 13.4: Solution Formation and Equilibrium - *CC BY-NC-SA 4.0*
  - 13.5: Solubilities of Gases - *CC BY 4.0*
  - 13.6: Vapor Pressures of Solutions - *CC BY-NC-SA 4.0*
  - 13.7: Osmotic Pressure - *CC BY-NC-SA 4.0*
  - 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions - *CC BY-NC-SA 4.0*
  - 13.9: Solutions of Electrolytes - *CC BY-NC-SA 4.0*
  - 13.10: Colloidal Mixtures - *CC BY-NC-SA 4.0*
  - 13.11: Exercises - *CC BY 4.0*
- 14: Fundamental Equilibrium Concepts - *CC BY 4.0*
  - 14.1: Introduction - *CC BY 4.0*
  - 14.2: Chemical Equilibria - *CC BY 4.0*
  - 14.3: Equilibrium Constants - *CC BY 4.0*
  - 14.4: Shifting Equilibria - Le Chatelier's Principle - *CC BY 4.0*
  - 14.5: Equilibrium Calculations - *CC BY 4.0*
  - 14.6: Key Terms - *Undeclared*
  - 14.7: Key Equations - *Undeclared*
  - 14.8: Summary - *Undeclared*
  - 14.9: Exercises - *CC BY 4.0*
- 15: Acid-Base Equilibria - A More Detailed Look - *CC BY 4.0*
  - 15.1: Introduction - *Undeclared*
  - 15.2: Arrhenius Theory- A Brief Review - *CC BY-NC-SA 4.0*
  - 15.3: Brønsted-Lowry Theory of Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.4: Self-Ionization of Water and the pH Scale - *CC BY-NC-SA 4.0*
  - 15.5: Strong Acids and Strong Bases - *CC BY-NC-SA 4.0*
  - 15.6: Weak Acids and Weak Bases - *CC BY-NC-SA 3.0*
  - 15.7: Polyprotic Acids - *CC BY 4.0*
  - 15.8: Ions as Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.9: Molecular Structure and Acid-Base Behavior - *CC BY-NC-SA 4.0*
  - 15.10: Lewis Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.11: Common-Ion Effect in Acid-Base Equilibria - *CC BY-NC-SA 4.0*
  - 15.12: Buffer Solutions - *CC BY-NC-SA 4.0*
  - 15.13: Acid-Base Indicators - *CC BY-NC-SA 4.0*
  - 15.14: Neutralization Reactions and Titration Curves - *CC BY-NC-SA 4.0*
  - 15.15: Solutions of Salts of Polyprotic Acids - *CC BY-NC-SA 4.0*
  - 15.16: Acid-Base Equilibrium Calculations- A Summary - *CC BY-NC-SA 4.0*
  - 15.17: Key Terms - *Undeclared*
  - 15.18: Key Equations - *Undeclared*
  - 15.19: Summary - *Undeclared*
  - 15.E: Exercises - *CC BY-NC-SA 4.0*
- 16: Solubility and Complex-Ion Equilibria - *CC BY-NC-SA 4.0*
  - 16.1: Solubility Product Constant, Ksp - *CC BY-NC-SA 3.0*
  - 16.2: Relationship Between Solubility and Ksp - *CC BY 4.0*
  - 16.3: Common-Ion Effect in Solubility Equilibria - *CC BY-NC-SA 4.0*
  - 16.4: Limitations of the K_{sp} Concept - *CC BY-NC-SA 4.0*
  - 16.5: Criteria for Precipitation and its Completeness - *CC BY-NC-SA 4.0*
  - 16.6: Fractional Precipitation - *CC BY-NC-SA 4.0*
  - 16.7: Solubility and pH - *CC BY-NC-SA 4.0*
  - 16.8: Equilibria Involving Complex Ions - *CC BY-NC-SA 4.0*
  - 16.9: Qualitative Cation Analysis - *CC BY-NC-SA 4.0*
  - 16.10: Exercises - *CC BY-NC-SA 4.0*
- 17: Thermodynamics - *CC BY 4.0*
  - 17.1: Introduction - *CC BY 4.0*
  - 17.2: Enthalpy - *CC BY 4.0*
  - 17.3: Spontaneity - *CC BY 4.0*
  - 17.4: Entropy - *CC BY 4.0*
  - 17.5: The Second and Third Laws of Thermodynamics - *CC BY 4.0*
  - 17.6: Free Energy - *CC BY 4.0*
  - 17.7: Key Terms - *Undeclared*
  - 17.8: Key Equations - *Undeclared*
  - 17.9: Summary - *Undeclared*
  - 17.10: Exercises - *CC BY 4.0*
- 18: Electrochemistry - *CC BY 4.0*
  - 18.1: Introduction - *CC BY 4.0*
  - 18.2: Review of Redox Chemistry - *CC BY 4.0*
  - 18.3: Galvanic Cells - *CC BY 4.0*
  - 18.4: Electrode and Cell Potentials - *CC BY 4.0*
  - 18.5: Potential, Free Energy, and Equilibrium - *CC BY 4.0*
  - 18.6: Batteries and Fuel Cells - *CC BY 4.0*
  - 18.7: Electrolysis - *CC BY 4.0*
  - 18.8: Key Terms - *CC BY 4.0*

- 18.9: Key Equations - *CC BY 4.0*
- 18.10: Summary - *CC BY 4.0*
- 18.11: Exercises - *CC BY 4.0*

- 19: Transition Metals and Coordination Chemistry - *CC BY 4.0*
  - 19.1: Introduction - *Undeclared*
  - 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds - *CC BY 4.0*
  - 19.3: Coordination Chemistry of Transition Metals - *CC BY 4.0*
  - 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds - *CC BY 4.0*
  - 19.5: Key Terms - *Undeclared*
  - 19.6: Summary - *Undeclared*
  - 19.7: Exercises - *CC BY 4.0*

- Back Matter - *Undeclared*
  - Index - *Undeclared*
  - Glossary - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*

Detailed Licensing

Overview

Title: UTSC: First-Year Chemistry Textbook (Fall 2025)

Webpages: 242

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY 4.0: 49.2% (119 pages)
- Undeclared: 25.6% (62 pages)
- CC BY-NC-SA 4.0: 23.6% (57 pages)
- CC BY-NC-SA 3.0: 1.7% (4 pages)

By Page

- UTSC: First-Year Chemistry Textbook (Fall 2025) - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction to Chemistry - Fundamental Concepts and Skills - *Undeclared*
    - 1.1: Introduction - *Undeclared*
    - 1.2: Phases and Classification of Matter - *CC BY 4.0*
    - 1.3: Physical and Chemical Properties - *CC BY 4.0*
    - 1.4: Measurements - *CC BY 4.0*
    - 1.5: Measurement Uncertainty, Accuracy, and Precision - *CC BY 4.0*
    - 1.6: Mathematical Treatment of Measurement Results - *CC BY 4.0*
    - 1.7: Key Terms - *Undeclared*
    - 1.8: Key Equations - *Undeclared*
    - 1.9: Summary - *Undeclared*
    - 1.10: Exercises - *CC BY 4.0*
  - 2: Atoms, Molecules, and Ions - *CC BY 4.0*
    - 2.1: Introduction - *CC BY 4.0*
    - 2.2: Early Ideas in Atomic Theory - *CC BY 4.0*
    - 2.3: Evolution of Atomic Theory - *CC BY 4.0*
    - 2.4: Atomic Structure and Symbolism - *CC BY 4.0*
    - 2.5: Chemical Formulas - *CC BY 4.0*
    - 2.6: Key Terms - *CC BY 4.0*
    - 2.7: Key Equations - *CC BY 4.0*
    - 2.8: Summary - *CC BY 4.0*
    - 2.9: Exercises - *CC BY 4.0*
  - 3: The Quantum Model of the Atom - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Radiation - *CC BY-NC-SA 4.0*
    - 3.3: Atomic Spectra - *CC BY-NC-SA 4.0*
    - 3.4: Quantum Theory - *CC BY-NC-SA 4.0*
    - 3.5: The Bohr Atom - *CC BY-NC-SA 4.0*
    - 3.6: Two Ideas Leading to a New Quantum Mechanics - *CC BY-NC-SA 4.0*
    - 3.7: Wave Mechanics - *CC BY-NC-SA 4.0*
    - 3.8: Electron Spin- A Fourth Quantum Number - *CC BY-NC-SA 4.0*
    - 3.9: The Shape of Atomic Orbitals - *CC BY-NC-SA 4.0*
    - 3.10: Multielectron Atoms - *CC BY-NC-SA 4.0*
    - 3.11: Key Terms - *Undeclared*
    - 3.12: Key Equations - *Undeclared*
    - 3.13: Summary - *Undeclared*
    - 3.14: Exercises - *CC BY 4.0*
  - 4: The Periodic Table and Some Atomic Properties - *CC BY-NC-SA 4.0*
    - 4.1: Classifying the Elements- The Periodic Law and the Periodic Table - *CC BY-NC-SA 4.0*
    - 4.2: Metals and Nonmetals and their Ions - *CC BY-NC-SA 4.0*
    - 4.3: Sizes of Atoms and Ions - *CC BY-NC-SA 4.0*
    - 4.4: Ionization Energy - *CC BY-NC-SA 4.0*
    - 4.5: Electron Affinity - *CC BY-NC-SA 4.0*
    - 4.6: Magnetic Properties - *CC BY-NC-SA 4.0*
    - 4.7: Periodic Properties of the Elements - *CC BY-NC-SA 4.0*
    - 4.8: Key Terms - *CC BY 4.0*
    - 4.9: Key Equations - *CC BY 4.0*
    - 4.10: Summary - *CC BY 4.0*
    - 4.11: Exercises - *CC BY 4.0*
  - 5: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Ionic Bonding - *CC BY-NC-SA 4.0*
    - 5.3: Covalent Bonding - *CC BY 4.0*
    - 5.4: Chemical Nomenclature - *CC BY 4.0*

- 5.5: Lewis Symbols and Structures - *CC BY 4.0*
- 5.6: Formal Charges and Resonance - *CC BY 4.0*
- 5.7: Molecular Structure and Polarity - *CC BY 4.0*
- 5.8: Key Terms - *Undeclared*
- 5.9: Key Equations - *Undeclared*
- 5.10: Summary - *Undeclared*
- 5.11: Exercises - *CC BY 4.0*

6: Advanced Theories of Covalent Bonding - *CC BY 4.0*

- 6.1: Introduction - *CC BY 4.0*
- 6.2: Valence Bond Theory - *CC BY 4.0*
- 6.3: Hybrid Atomic Orbitals - *CC BY 4.0*
- 6.4: Multiple Bonds - *CC BY 4.0*
- 6.5: Molecular Orbital Theory - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*
- 6.9: Exercises - *CC BY 4.0*

7: Thermochemistry - *CC BY 4.0*

- 7.1: Introduction - *CC BY 4.0*
- 7.2: Energy Basics - *CC BY 4.0*
- 7.3: Calorimetry - *CC BY 4.0*
- 7.4: Enthalpy - *CC BY 4.0*
- 7.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 7.6: Key Terms - *Undeclared*
- 7.7: Key Equations - *Undeclared*
- 7.8: Summary - *Undeclared*
- 7.9: Exercises - *CC BY 4.0*

8: Liquids and Solids and Phase Changes - *CC BY 4.0*

- 8.1: Introduction - *CC BY 4.0*
- 8.2: Intermolecular Forces - *CC BY 4.0*
- 8.3: Properties of Liquids - *CC BY 4.0*
- 8.4: Phase Transitions - *CC BY 4.0*
- 8.5: Phase Diagrams - *CC BY 4.0*
- 8.6: The Solid State of Matter - *CC BY 4.0*
- 8.7: Key Terms - *Undeclared*
- 8.8: Key Equations - *Undeclared*
- 8.9: Summary - *Undeclared*
- 8.10: Exercises - *CC BY 4.0*

9: Gases - *CC BY 4.0*

- 9.1: Introduction - *Undeclared*
- 9.2: Gas Pressure - *CC BY 4.0*
- 9.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 9.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 9.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 9.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 9.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 9.8: Key Terms - *Undeclared*
- 9.9: Key Equations - *Undeclared*
- 9.10: Summary - *Undeclared*
- 9.11: Exercises - *CC BY 4.0*

10: Three Major Classes of Chemical Reactions - *Undeclared*

- 10.1: Formula Mass - *CC BY 4.0*
- 10.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 10.3: Reaction Stoichiometry - *CC BY 4.0*
- 10.4: Reaction Yields - *CC BY 4.0*
- 10.5: Molarity - *CC BY 4.0*
- 10.6: Precipitation Reactions - *CC BY-NC-SA 4.0*
- 10.7: Acid-Base Reactions - *CC BY-NC-SA 4.0*
- 10.8: Stoichiometry of Reactions in Aqueous Solutions- Titrations - *CC BY-NC-SA 4.0*
- 10.9: Oxidation-Reduction- Some General Principles - *CC BY-NC-SA 4.0*
- 10.10: Balancing Oxidation-Reduction Equations - *CC BY-NC-SA 4.0*
- 10.11: Oxidizing and Reducing Agents - *CC BY-NC-SA 4.0*
- 10.12: Key Terms - *Undeclared*
- 10.13: Key Equations - *Undeclared*
- 10.E: Reactions in Aqueous Solution (Exercises) - *CC BY-NC-SA 3.0*
- 10.S: Summary - *CC BY-NC-SA 3.0*

11: Nuclear Chemistry - *CC BY 4.0*

- 11.1: Introduction - *Undeclared*
- 11.2: Nuclear Structure and Stability - *CC BY 4.0*
- 11.3: Nuclear Equations - *CC BY 4.0*
- 11.4: Radioactive Decay - *CC BY 4.0*
- 11.5: Transmutation and Nuclear Energy - *CC BY 4.0*
- 11.6: Uses of Radioisotopes - *CC BY 4.0*
- 11.7: Biological Effects of Radiation - *CC BY 4.0*
- 11.8: Key Terms - *Undeclared*
- 11.9: Key Equations - *Undeclared*
- 11.10: Summary - *Undeclared*
- 11.11: Exercises - *CC BY 4.0*

12: Kinetics - *CC BY 4.0*

- 12.1: Introduction - *CC BY 4.0*
- 12.2: Chemical Reaction Rates - *CC BY 4.0*
- 12.3: Factors Affecting Reaction Rates - *CC BY 4.0*
- 12.4: Rate Laws - *CC BY 4.0*
- 12.5: Integrated Rate Laws - *CC BY 4.0*
- 12.6: Collision Theory - *CC BY 4.0*
- 12.7: Reaction Mechanisms - *CC BY 4.0*
- 12.8: Catalysis - *CC BY 4.0*
- 12.9: Key Terms - *Undeclared*
- 12.10: Key Equations - *Undeclared*
- 12.11: Summary - *Undeclared*
- 12.12: Exercises - *CC BY 4.0*

- 13: Solutions and their Physical Properties - *CC BY-NC-SA 4.0*
  - 13.1: Types of Solutions - Some Terminology - *CC BY-NC-SA 4.0*
  - 13.2: Solution Concentration - *CC BY-NC-SA 4.0*
  - 13.3: Intermolecular Forces and the Solution Process - *CC BY-NC-SA 4.0*
  - 13.4: Solution Formation and Equilibrium - *CC BY-NC-SA 4.0*
  - 13.5: Solubilities of Gases - *CC BY 4.0*
  - 13.6: Vapor Pressures of Solutions - *CC BY-NC-SA 4.0*
  - 13.7: Osmotic Pressure - *CC BY-NC-SA 4.0*
  - 13.8: Freezing-Point Depression and Boiling-Point Elevation of Nonelectrolyte Solutions - *CC BY-NC-SA 4.0*
  - 13.9: Solutions of Electrolytes - *CC BY-NC-SA 4.0*
  - 13.10: Colloidal Mixtures - *CC BY-NC-SA 4.0*
  - 13.11: Exercises - *CC BY 4.0*
- 14: Fundamental Equilibrium Concepts - *CC BY 4.0*
  - 14.1: Introduction - *CC BY 4.0*
  - 14.2: Chemical Equilibria - *CC BY 4.0*
  - 14.3: Equilibrium Constants - *CC BY 4.0*
  - 14.4: Shifting Equilibria - Le Chatelier's Principle - *CC BY 4.0*
  - 14.5: Equilibrium Calculations - *CC BY 4.0*
  - 14.6: Key Terms - *Undeclared*
  - 14.7: Key Equations - *Undeclared*
  - 14.8: Summary - *Undeclared*
  - 14.9: Exercises - *CC BY 4.0*
- 15: Acid-Base Equilibria - A More Detailed Look - *CC BY 4.0*
  - 15.1: Introduction - *Undeclared*
  - 15.2: Arrhenius Theory- A Brief Review - *CC BY-NC-SA 4.0*
  - 15.3: Brønsted-Lowry Theory of Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.4: Self-Ionization of Water and the pH Scale - *CC BY-NC-SA 4.0*
  - 15.5: Strong Acids and Strong Bases - *CC BY-NC-SA 4.0*
  - 15.6: Weak Acids and Weak Bases - *CC BY-NC-SA 3.0*
  - 15.7: Polyprotic Acids - *CC BY 4.0*
  - 15.8: Ions as Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.9: Molecular Structure and Acid-Base Behavior - *CC BY-NC-SA 4.0*
  - 15.10: Lewis Acids and Bases - *CC BY-NC-SA 4.0*
  - 15.11: Common-Ion Effect in Acid-Base Equilibria - *CC BY-NC-SA 4.0*
  - 15.12: Buffer Solutions - *CC BY-NC-SA 4.0*
  - 15.13: Acid-Base Indicators - *CC BY-NC-SA 4.0*
  - 15.14: Neutralization Reactions and Titration Curves - *CC BY-NC-SA 4.0*
  - 15.15: Solutions of Salts of Polyprotic Acids - *CC BY-NC-SA 4.0*
  - 15.16: Acid-Base Equilibrium Calculations- A Summary - *CC BY-NC-SA 4.0*
  - 15.17: Key Terms - *Undeclared*
  - 15.18: Key Equations - *Undeclared*
  - 15.19: Summary - *Undeclared*
  - 15.E: Exercises - *CC BY-NC-SA 4.0*
- 16: Solubility and Complex-Ion Equilibria - *CC BY-NC-SA 4.0*
  - 16.1: Solubility Product Constant, Ksp - *CC BY-NC-SA 3.0*
  - 16.2: Relationship Between Solubility and Ksp - *CC BY 4.0*
  - 16.3: Common-Ion Effect in Solubility Equilibria - *CC BY-NC-SA 4.0*
  - 16.4: Limitations of the K_{sp} Concept - *CC BY-NC-SA 4.0*
  - 16.5: Criteria for Precipitation and its Completeness - *CC BY-NC-SA 4.0*
  - 16.6: Fractional Precipitation - *CC BY-NC-SA 4.0*
  - 16.7: Solubility and pH - *CC BY-NC-SA 4.0*
  - 16.8: Equilibria Involving Complex Ions - *CC BY-NC-SA 4.0*
  - 16.9: Qualitative Cation Analysis - *CC BY-NC-SA 4.0*
  - 16.10: Exercises - *CC BY-NC-SA 4.0*
- 17: Thermodynamics - *CC BY 4.0*
  - 17.1: Introduction - *CC BY 4.0*
  - 17.2: Enthalpy - *CC BY 4.0*
  - 17.3: Spontaneity - *CC BY 4.0*
  - 17.4: Entropy - *CC BY 4.0*
  - 17.5: The Second and Third Laws of Thermodynamics - *CC BY 4.0*
  - 17.6: Free Energy - *CC BY 4.0*
  - 17.7: Key Terms - *Undeclared*
  - 17.8: Key Equations - *Undeclared*
  - 17.9: Summary - *Undeclared*
  - 17.10: Exercises - *CC BY 4.0*
- 18: Electrochemistry - *CC BY 4.0*
  - 18.1: Introduction - *CC BY 4.0*
  - 18.2: Review of Redox Chemistry - *CC BY 4.0*
  - 18.3: Galvanic Cells - *CC BY 4.0*
  - 18.4: Electrode and Cell Potentials - *CC BY 4.0*
  - 18.5: Potential, Free Energy, and Equilibrium - *CC BY 4.0*
  - 18.6: Batteries and Fuel Cells - *CC BY 4.0*
  - 18.7: Electrolysis - *CC BY 4.0*
  - 18.8: Key Terms - *CC BY 4.0*

- 18.9: Key Equations - *CC BY 4.0*
- 18.10: Summary - *CC BY 4.0*
- 18.11: Exercises - *CC BY 4.0*

- 19: Transition Metals and Coordination Chemistry - *CC BY 4.0*
  - 19.1: Introduction - *Undeclared*
  - 19.2: Occurrence, Preparation, and Properties of Transition Metals and Their Compounds - *CC BY 4.0*
  - 19.3: Coordination Chemistry of Transition Metals - *CC BY 4.0*
  - 19.4: Spectroscopic and Magnetic Properties of Coordination Compounds - *CC BY 4.0*
  - 19.5: Key Terms - *Undeclared*
  - 19.6: Summary - *Undeclared*
  - 19.7: Exercises - *CC BY 4.0*

- Back Matter - *Undeclared*
  - Index - *Undeclared*
  - Glossary - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*
  - Detailed Licensing - *Undeclared*