VICTOR VALLEY COLLEGE GENERAL CHEMISTRY 201 TEXTBOOK



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Chemistry 201 Textbook

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TABLE OF CONTENTS

Licensing

Work Page

1: Measurements in Chemistry

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

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

3: Electronic Structure and Periodic Properties

- 3.1: Introduction
- 3.2: Electromagnetic Energy
- 3.3: The Bohr Model
- 3.4: Development of Quantum Theory
- 3.5: Electronic Structure of Atoms (Electron Configurations)
- 3.6: Periodic Variations in Element Properties
- 3.7: The Periodic Table
- 3.8: Ionic and Molecular Compounds
- 3.9: Key Terms
- 3.10: Key Equations
- 3.11: Summary

4: Chemical Bonding and Molecular Geometry

- 4.1: Introduction
- 4.2: Ionic Bonding
- 4.3: Covalent Bonding
- 4.4: Chemical Nomenclature
- 4.5: Lewis Symbols and Structures
- 4.6: Formal Charges and Resonance
- 4.7: Molecular Structure and Polarity

- 4.8: Key Terms
- 4.9: Key Equations
- 4.10: Summary

5: Advanced Theories of Covalent Bonding

- 5.1: Introduction
- 5.2: Valence Bond Theory
- 5.3: Hybrid Atomic Orbitals
- 5.4: Multiple Bonds
- 5.5: Molecular Orbital Theory
- 5.6: Key Terms
- 5.7: Key Equations
- 5.8: Summary

6: Composition of Substances and Solutions

- 6.1: Introduction
- 6.2: Formula Mass
- 6.3: Determining Empirical and Molecular Formulas
- 6.4: Molarity
- 6.5: Other Units for Solution Concentrations
- 6.6: Key Terms
- 6.7: Key Equations
- 6.8: Summary

7: Stoichiometry of Chemical Reactions

- 7.1: Introduction
- 7.2: Writing and Balancing Chemical Equations
- 7.3: Classifying Chemical Reactions
- 7.4: Reaction Stoichiometry
- 7.5: Reaction Yields
- 7.6: Quantitative Chemical Analysis
- 7.7: Key Terms
- 7.8: Key Equations
- 7.9: Summary

8: Gases

- 8.1: Introduction
- 8.2: Gas Pressure
- 8.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law
- 8.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions
- 8.5: Effusion and Diffusion of Gases
- 8.6: The Kinetic-Molecular Theory
- 8.7: Non-Ideal Gas Behavior
- 8.8: Key Terms
- 8.9: Key Equations
- 8.10: Summary

9: Thermochemistry

- 9.1: Introduction
- 9.2: Energy Basics
- 9.3: Calorimetry
- 9.4: Enthalpy
- 9.5: Strengths of Ionic and Covalent Bonds
- 9.6: Key Terms
- 9.7: Key Equations
- 9.8: Summary

10: Liquids and Solids

- 10.1: Introduction
- 10.2: Intermolecular Forces
- 10.3: Properties of Liquids
- 10.4: Phase Transitions
- 10.5: Phase Diagrams
- 10.6: The Solid State of Matter
- 10.7: Lattice Structures in Crystalline Solids
- 10.8: Key Terms
- 10.9: Key Equations
- 10.10: Summary

11: Solutions and Colloids

- 11.1: Introduction
- 11.2: The Dissolution Process
- 11.3: Electrolytes
- 11.4: Solubility
- 11.5: Colligative Properties
- 11.6: Colloids
- 11.7: Key Terms
- 11.8: Key Equations
- 11.9: Summary

12: Lab Manual

- 12.1: Lab 1 - Using Excel for Graphical Analysis of Data
- 12.2: Lab 2 - Measurements in the Laboratory
- 12.3: Lab 3 - The Densities of Solutions and Solids
- 12.4: Lab 4 - Paper Chromatography- Separation and Identification of Five Metal Cations
- 12.5: Lab 5 - Flame Tests of Metal Cations
- 12.6: Lab 6 - Atomic Emission
- 12.7: Lab 7 - Chemical Nomenclature Dry Lab
- 12.8: Lab 8 - Measuring Manganese Concentration Using Spectrophotometry
- 12.9: Lab 9 - VSEPR Theory and Shapes of Molecules Dry Lab
- 12.10: Lab 10 - Titration of Vinegar
- 12.11: Lab 11 - Types of Chemical Reactions
- 12.12: Lab 12 - Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions
- 12.13: Lab 13 - Mole Ratios and Reaction Stoichiometry
- 12.14: Lab 14 - Properties of Hydrates
- 12.15: Lab 15 - The Properties of Oxygen Gas
- 12.16: Lab 16 - The Dumas Method

- 12.17: Lab 17 - The Molecular Weight of Carbon Dioxide
- 12.18: Lab 18 - Calorimetry and Hess's Law

Index

Glossary

Detailed Licensing

Detailed Licensing

Licensing

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

Work Page

Learning Objectives

Link withing textbook now external link

- Objective 1
- Objective 2



Figure 1: Image Sample.

Exercise 1

What is 2+2?

Answer

4

Let's watch a video (tell why + QR code/link)



QR code for this video2



Video Title

Video text/introduction.

Video Title

Video text/introduction.



Additional text/introduction.

Additional text/introduction.

Column 1	Column 2	Column 3

google

CHAPTER OVERVIEW

1: Measurements in Chemistry

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

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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.

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1.2: Chemistry in Context

Learning Objectives

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

- Outline the historical development of chemistry
- Provide examples of the importance of chemistry in everyday life
- Describe the scientific method
- Differentiate among hypotheses, theories, and laws
- Provide examples illustrating macroscopic, microscopic, and symbolic domains

Throughout human history, people have tried to convert matter into more useful forms. Our Stone Age ancestors chipped pieces of flint into useful tools and carved wood into statues and toys. These endeavors involved changing the shape of a substance without changing the substance itself. But as our knowledge increased, humans began to change the composition of the substances as well—clay was converted into pottery, hides were cured to make garments, copper ores were transformed into copper tools and weapons, and grain was made into bread.

Humans began to practice chemistry when they learned to control fire and use it to cook, make pottery, and smelt metals. Subsequently, they began to separate and use specific components of matter. A variety of drugs such as aloe, myrrh, and opium were isolated from plants. Dyes, such as indigo and Tyrian purple, were extracted from plant and animal matter. Metals were combined to form alloys—for example, copper and tin were mixed together to make bronze—and more elaborate smelting techniques produced iron. Alkalis were extracted from ashes, and soaps were prepared by combining these alkalis with fats. Alcohol was produced by fermentation and purified by distillation.

Attempts to understand the behavior of matter extend back for more than 2500 years. As early as the sixth century BC, Greek philosophers discussed a system in which water was the basis of all things. You may have heard of the Greek postulate that matter consists of four elements: earth, air, fire, and water. Subsequently, an amalgamation of chemical technologies and philosophical speculations was spread from Egypt, China, and the eastern Mediterranean by alchemists, who endeavored to transform “base metals” such as lead into “noble metals” like gold, and to create elixirs to cure disease and extend life (Figure 1.2.1).

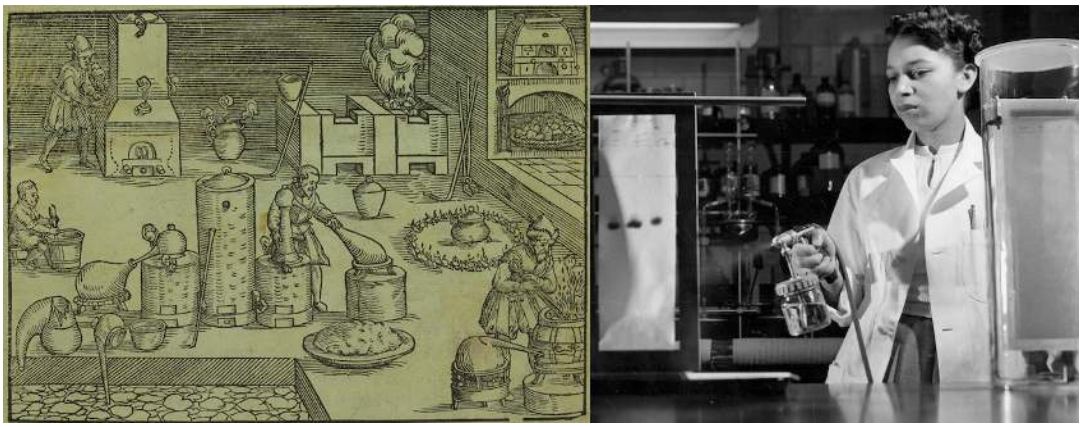


Figure 1.2.1: (a) This portrayal shows an alchemist’s workshop circa 1580. Although alchemy made some useful contributions to how to manipulate matter, it was not scientific by modern standards. (b) While the equipment used by Alma Levant Hayden in this 1952 picture might not seem as sleek as you might find in a lab today, her approach was highly methodical and carefully recorded. A department head at the FDA, Hayden is most famous for exposing an aggressively marketed anti-cancer drug as nothing more than an unhelpful solution of common substances. (credit a: Chemical Heritage Foundation; b: NIH History Office)

From alchemy came the historical progressions that led to modern chemistry: the isolation of drugs from natural sources, such as plants and animals. But while many of the substances extracted or processed from those natural sources were critical in the treatment of diseases, many were scarce. For example, progesterone, which is critical to women's health, became available as a medicine in 1935, but its animal sources produced extremely small quantities, limiting its availability and increasing its expense. Likewise, in the 1940s, cortisone came into use to treat arthritis and other disorders and injuries, but it took a 36-step process to synthesize. Chemist Percy Lavon Julian turned to a more plentiful source: soybeans. Previously, Julian had developed a lab to isolate soy protein, which was used in firefighting among other applications. He focused on using the soy sterols—substances

mostly used in plant membranes—and was able to quickly produce progesterone and later testosterone and other hormones. He later developed a process to do the same for cortisone, and laid the groundwork for modern drug design. Since soybeans and similar plant sources were extremely plentiful, the drugs soon became widely available, saving many lives.

Chemistry: The Central Science

Chemistry is sometimes referred to as “the central science” due to its interconnectedness with a vast array of other STEM disciplines (STEM stands for areas of study in the science, technology, engineering, and math fields). Chemistry and the language of chemists play vital roles in biology, medicine, materials science, forensics, environmental science, and many other fields (Figure 1.2.2). The basic principles of physics are essential for understanding many aspects of chemistry, and there is extensive overlap between many subdisciplines within the two fields, such as chemical physics and nuclear chemistry. Mathematics, computer science, and information theory provide important tools that help us calculate, interpret, describe, and generally make sense of the chemical world. Biology and chemistry converge in biochemistry, which is crucial to understanding the many complex factors and processes that keep living organisms (such as us) alive. Chemical engineering, materials science, and nanotechnology combine chemical principles and empirical findings to produce useful substances, ranging from gasoline to fabrics to electronics. Agriculture, food science, veterinary science, and brewing and wine making help provide sustenance in the form of food and drink to the world’s population. Medicine, pharmacology, biotechnology, and botany identify and produce substances that help keep us healthy. Environmental science, geology, oceanography, and atmospheric science incorporate many chemical ideas to help us better understand and protect our physical world. Chemical ideas are used to help understand the universe in astronomy and cosmology.

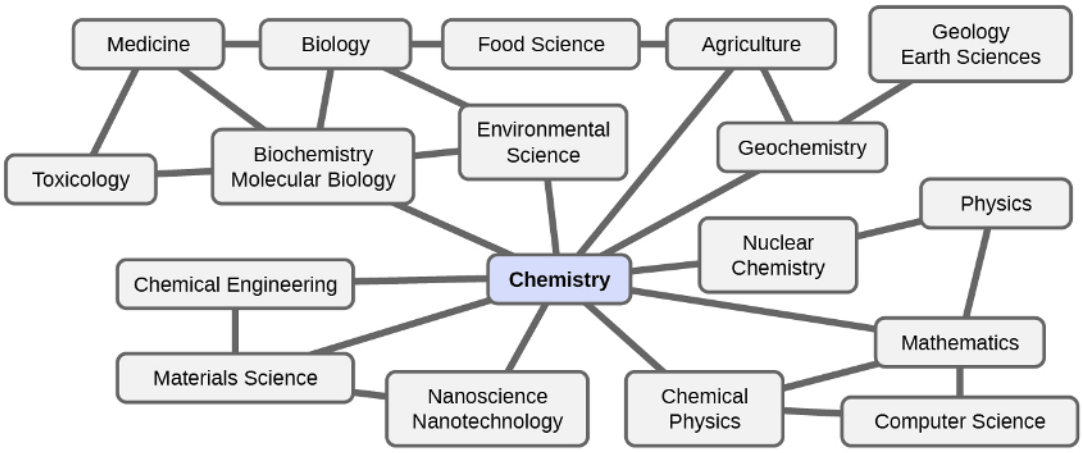


Figure 1.2.2: Knowledge of chemistry is central to understanding a wide range of scientific disciplines. This diagram shows just some of the interrelationships between chemistry and other fields.

What are some changes in matter that are essential to daily life? Digesting and assimilating food, synthesizing polymers that are used to make clothing, containers, cookware, and credit cards, and refining crude oil into gasoline and other products are just a few examples. As you proceed through this course, you will discover many different examples of changes in the composition and structure of matter, how to classify these changes and how they occurred, their causes, the changes in energy that accompany them, and the principles and laws involved. As you learn about these things, you will be learning chemistry, the study of the composition, properties, and interactions of matter. The practice of chemistry is not limited to chemistry books or laboratories: It happens whenever someone is involved in changes in matter or in conditions that may lead to such changes.

The Scientific Method

Chemistry is a science based on observation and experimentation. Doing chemistry involves attempting to answer questions and explain observations in terms of the laws and theories of chemistry, using procedures that are accepted by the scientific community. There is no single route to answering a question or explaining an observation, but there is an aspect common to every approach: Each uses knowledge based on experiments that can be reproduced to verify the results. Some routes involve a hypothesis, a tentative explanation of observations that acts as a guide for gathering and checking information. A hypothesis is tested by experimentation, calculation, and/or comparison with the experiments of others and then refined as needed.

Some hypotheses are attempts to explain the behavior that is summarized in laws. The laws of science summarize a vast number of experimental observations, and describe or predict some facet of the natural world. If such a hypothesis turns out to be capable of

explaining a large body of experimental data, it can reach the status of a theory. Scientific theories are well-substantiated, comprehensive, testable explanations of particular aspects of nature. Theories are accepted because they provide satisfactory explanations, but they can be modified if new data become available. The 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, is called the scientific method (Figure 1.2.3).

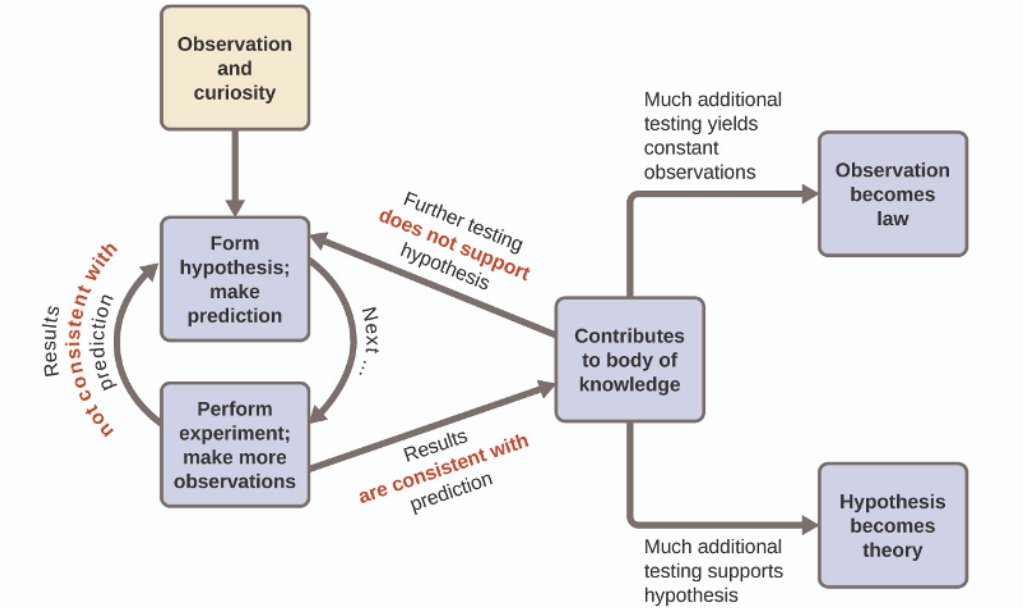


Figure 1.2.3: The scientific method follows a process similar to the one shown in this diagram. All the key components are shown, in roughly the right order. Scientific progress is seldom neat and clean: It requires open inquiry and the reworking of questions and ideas in response to findings.

The Domains of Chemistry

Chemists study and describe the behavior of matter and energy in three different domains: macroscopic, microscopic, and symbolic. These domains provide different ways of considering and describing chemical behavior.

Macro is a Greek word that means “large.” The macroscopic domain is familiar to us: It is the realm of everyday things that are large enough to be sensed directly by human sight or touch. In daily life, this includes the food you eat and the breeze you feel on your face. The macroscopic domain includes everyday and laboratory chemistry, where we observe and measure physical and chemical properties such as density, solubility, and flammability.

Micro comes from Greek and means “small.” The microscopic domain of chemistry is often visited in the imagination. Some aspects of the microscopic domain are visible through standard optical microscopes, for example, many biological cells. More sophisticated instruments are capable of imaging even smaller entities such as molecules and atoms (see Figure 1.2.4: **(b)**).

However, most of the subjects in the microscopic domain of chemistry are too small to be seen even with the most advanced microscopes and may only be pictured in the mind. Other components of the microscopic domain include ions and electrons, protons and neutrons, and chemical bonds, each of which is far too small to see.

The symbolic domain contains the specialized language used to represent components of the macroscopic and microscopic domains. Chemical symbols (such as those used in the periodic table), chemical formulas, and chemical equations are part of the symbolic domain, as are graphs, drawings, and calculations. These symbols play an important role in chemistry because they help interpret the behavior of the macroscopic domain in terms of the components of the microscopic domain. One of the challenges for students learning chemistry is recognizing that the same symbols can represent different things in the macroscopic and microscopic domains, and one of the features that makes chemistry fascinating is the use of a domain that must be imagined to explain behavior in a domain that can be observed.

A helpful way to understand the three domains is via the essential and ubiquitous substance of water. That water is a liquid at moderate temperatures, will freeze to form a solid at lower temperatures, and boil to form a gas at higher temperatures (Figure 1.2.4) are macroscopic observations. But some properties of water fall into the microscopic domain—what cannot be observed with the naked eye. The description of water as comprising two hydrogen atoms and one oxygen atom, and the explanation of

freezing and boiling in terms of attractions between these molecules, is within the microscopic arena. The formula H_2O, which can describe water at either the macroscopic or microscopic levels, is an example of the symbolic domain. The abbreviations (*g*) for gas, (*s*) for solid, and (*l*) for liquid are also symbolic.

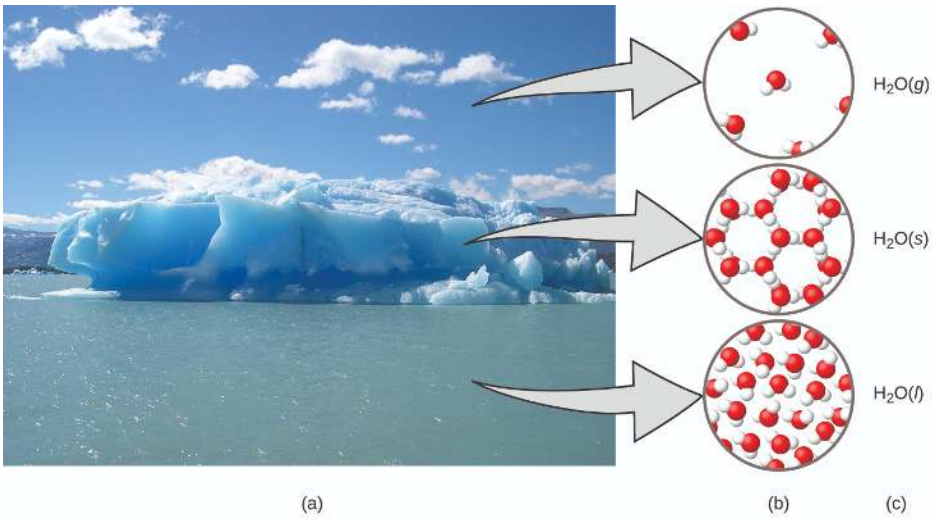


Figure 1.2.4: (a) Moisture in the air, icebergs, and the ocean represent water in the macroscopic domain. (b) At the molecular level (microscopic domain), gas molecules are far apart and disorganized, solid water molecules are close together and organized, and liquid molecules are close together and disorganized. (c) The formula H_2O symbolizes water, and (*g*), (*s*), and (*l*) symbolize its phases. Note that clouds are actually comprised of either very small liquid water droplets or solid water crystals; gaseous water in our atmosphere is not visible to the naked eye, although it may be sensed as humidity. (credit a: modification of work by "Gorkaazk"/Wikimedia Commons)

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1.3: 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 1.3.1). 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.

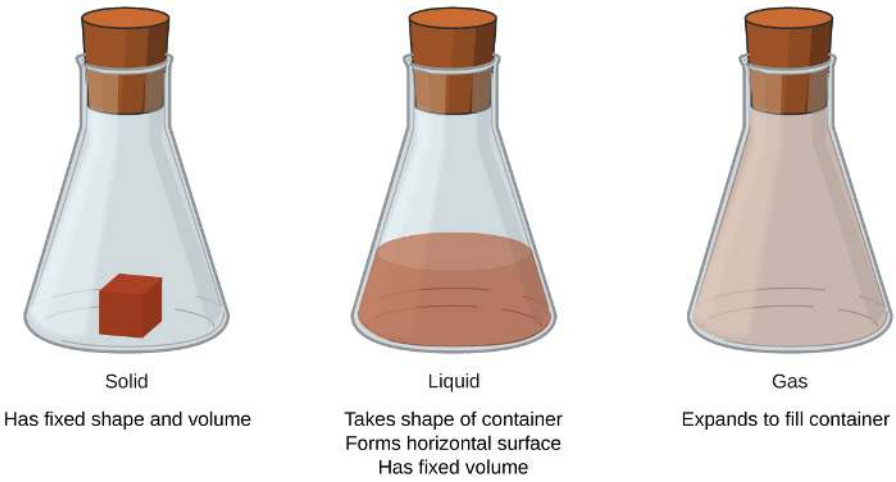


Figure 1.3.1: The three most common states or phases of matter are solid, liquid, and gas.

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 1.3.2). 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 1.3.2: 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 1.3.3). 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.

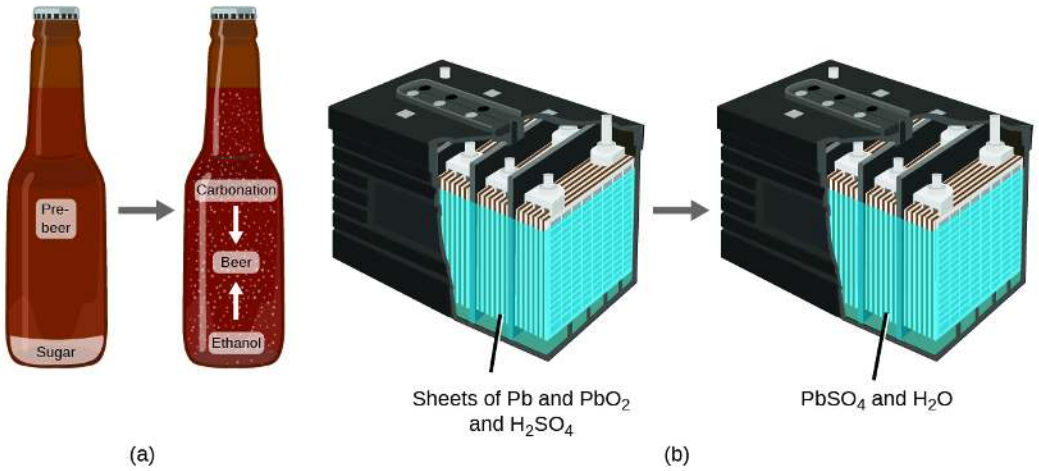


Figure 1.3.3: (a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbon dioxide. (b) The mass of the lead, lead oxide, and sulfuric acid consumed by the production of electricity is exactly equal to the mass of lead sulfate and water that is 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 1.3.4). 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).

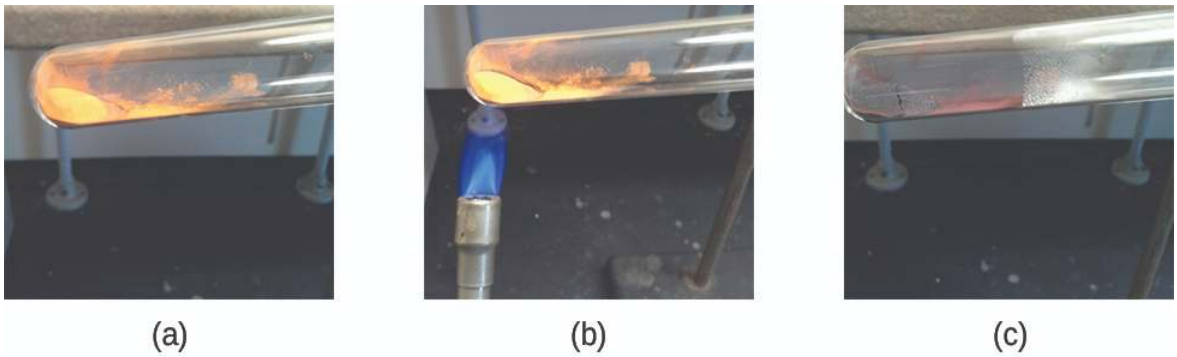


Figure 1.3.4: (a) The compound mercury(II) oxide, (b) when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)

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 1.3.5). 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 1.3.5). 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 1.3.5: (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 1.3.6).

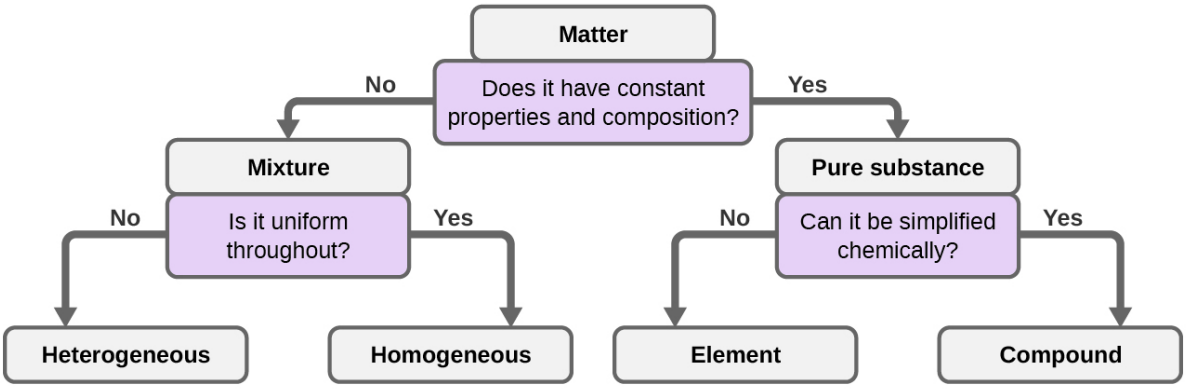


Figure 1.3.6: Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound, or an element.

Eleven elements make up about 99% of the earth's crust and atmosphere (Table 1.3.1). 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 1.3.1: 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 1.3.7). This atom would no longer be gold if it were divided any further.

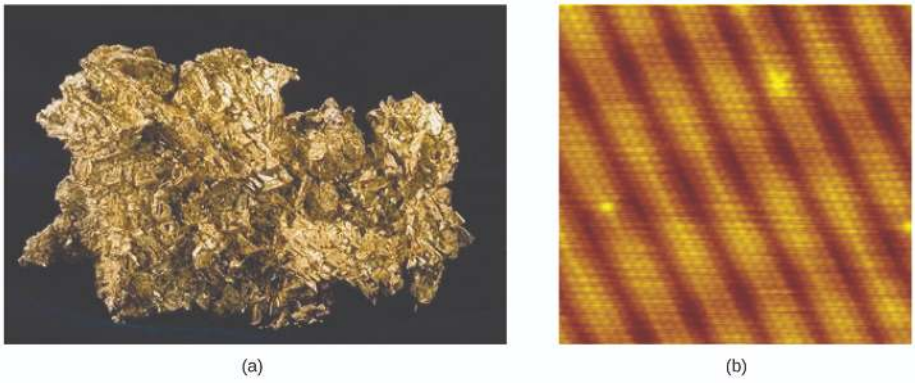


Figure 1.3.7: (a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom. (credit a: modification of work by United States Geological Survey; credit b: modification 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 1.3.8) shows increasingly close microscopic and atomic-level views of ordinary cotton.

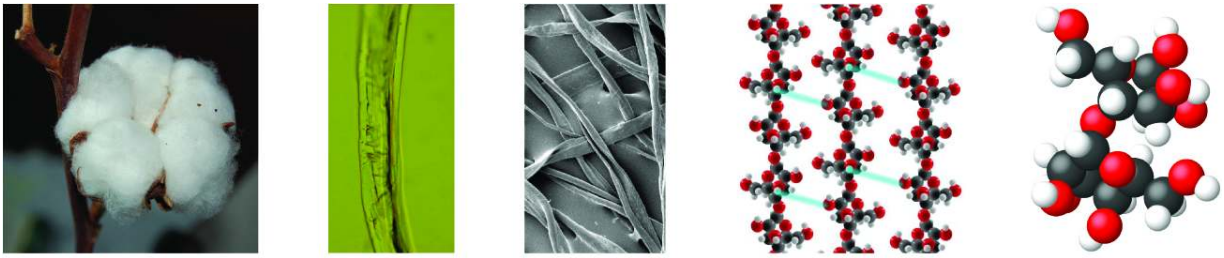


Figure 1.3.8: These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d 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 1.3.9). 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.

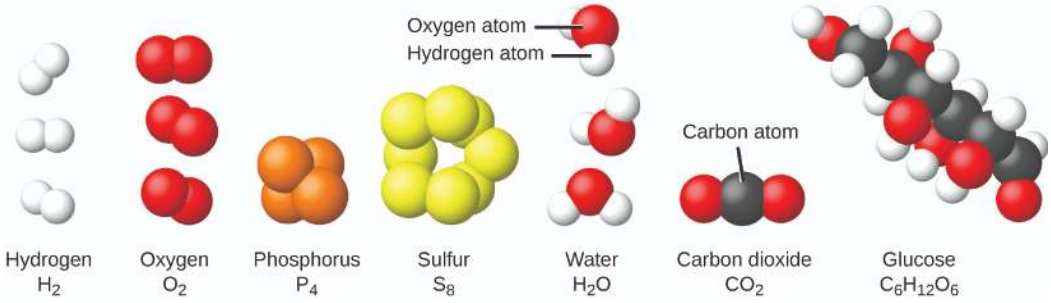


Figure 1.3.9: The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

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 1.3.10).

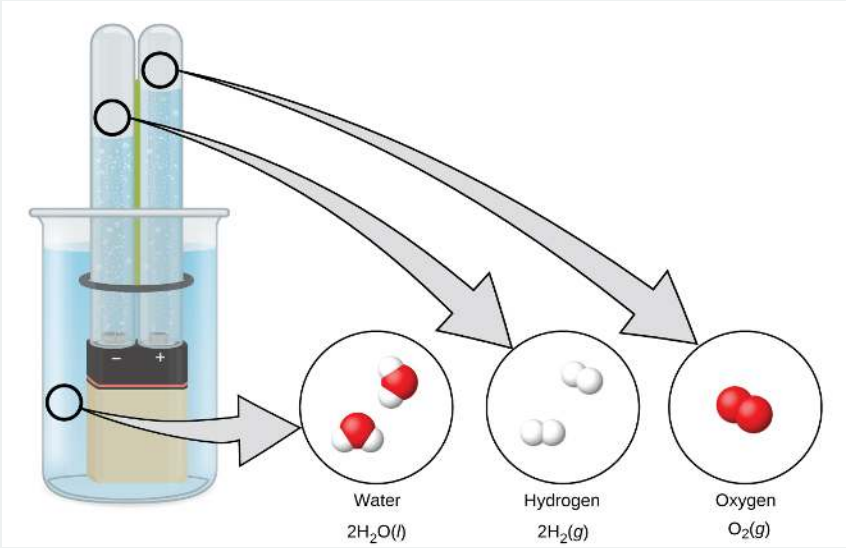


Figure 1.3.10: The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, 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 1.3.11). 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.

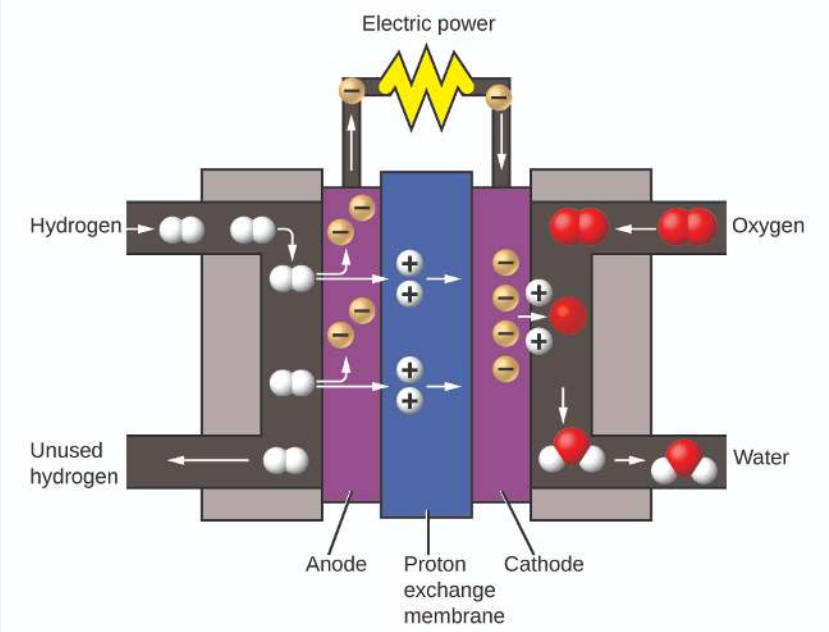


Figure 1.3.11: A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.

Chemistry in Everyday Life: Chemistry of Cell Phones

Imagine how different your life would be without cell phones (Figure 1.3.12) 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 1.3.12: Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)

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1.4: 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 1.4.1). 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 1.4.1: (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 1.4.2). Nitroglycerin is very dangerous because it explodes easily; neon poses almost no hazard because it is very unreactive.



(a) (b)

Figure 1.4.2: (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 1.4.3).

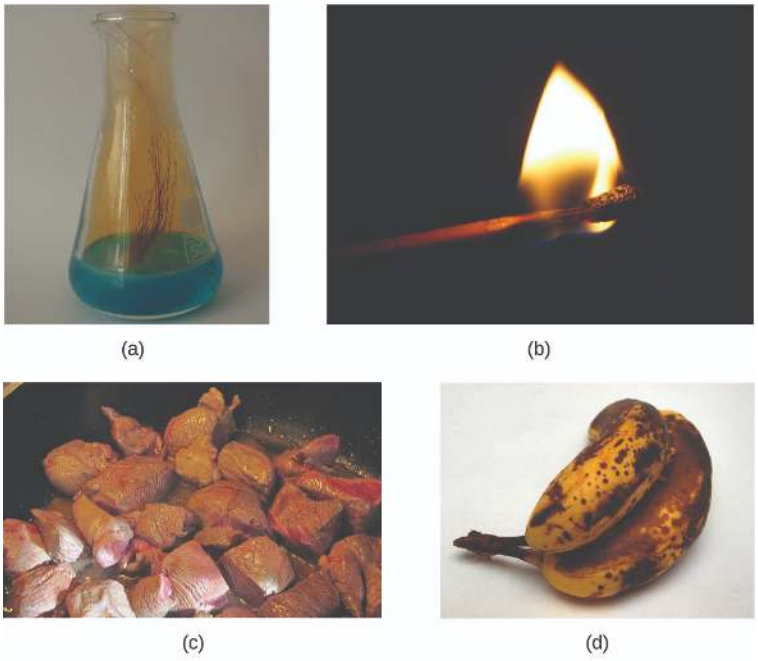


Figure 1.4.3: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes 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 1.4.4: 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.

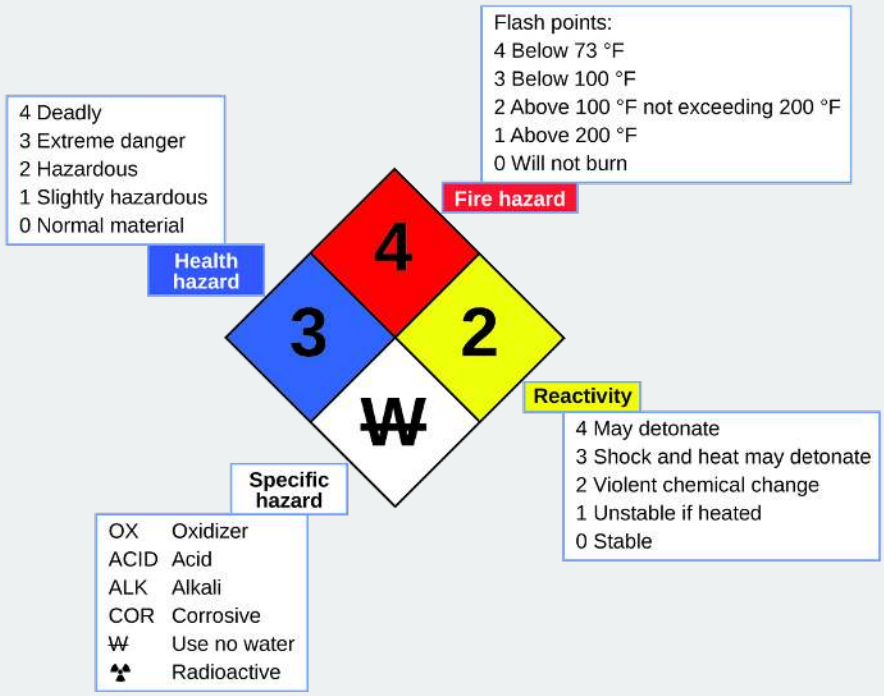


Figure 1.4.4: The National Fire Protection Agency (NFPA) hazard diamond summarizes the major hazards of a chemical substance.

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 1.4.5). You will learn more about the periodic table as you continue your study of chemistry.

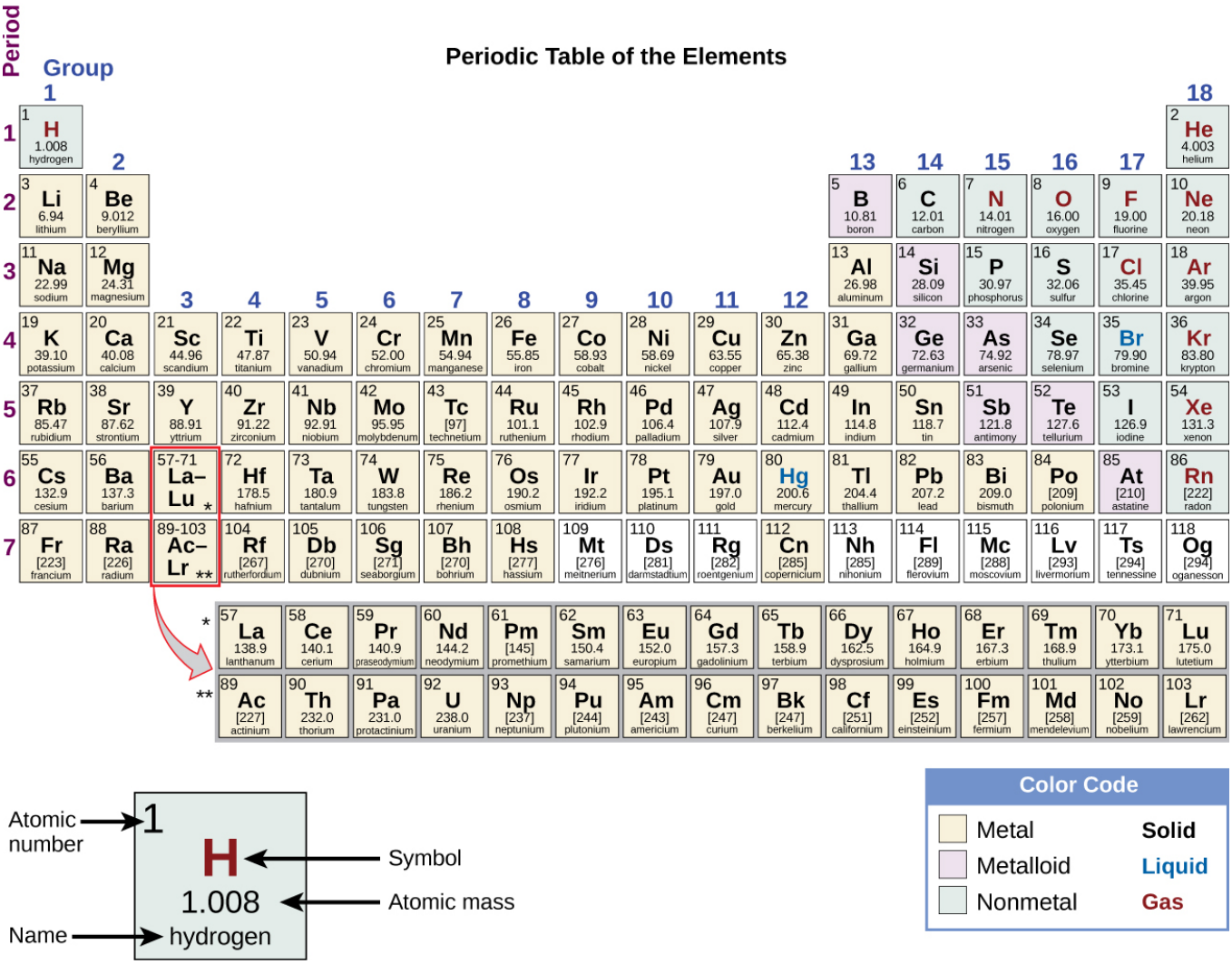


Figure 1.4.5: The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.

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1.5: 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 1.5.1. 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 1.5.1: 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 1.5.2.

Table 1.5.2: 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 1.5.1); 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).

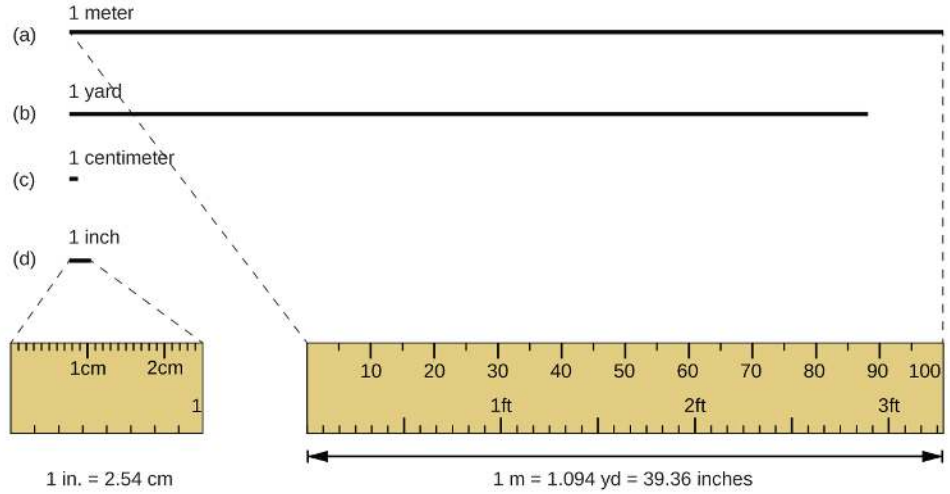


Figure 1.5.1: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

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 1.5.2). 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 1.5.2: 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 1.5.3). 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.

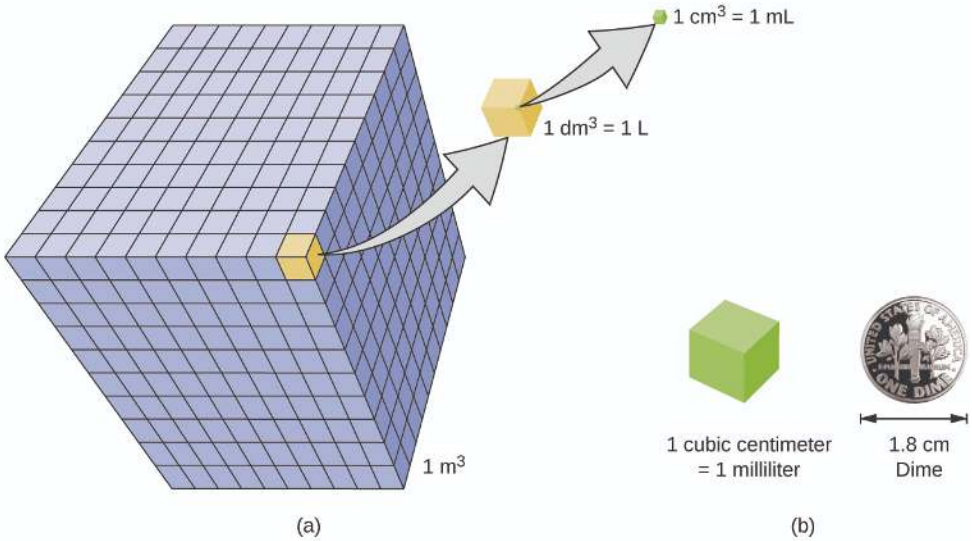


Figure 1.5.3: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

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 1.5.3 shows the densities of some common substances.

Table 1.5.3: 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.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.5.1}$$

Example 1.5.1: 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 1.5.1

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 1.5.2: 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 1.5.2

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

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1.6: 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.

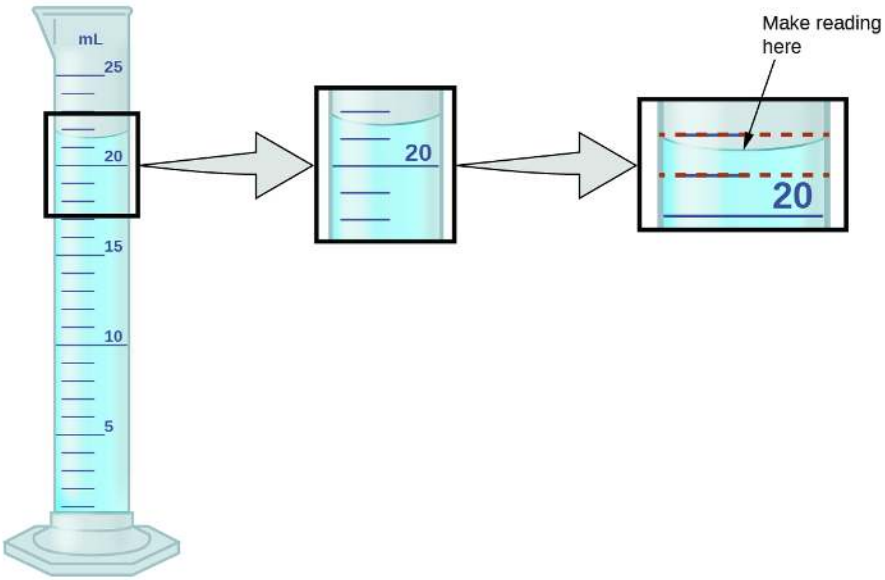
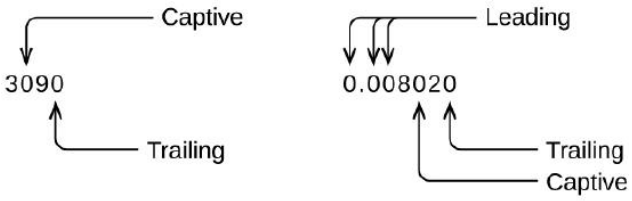


Figure 1.6.1: To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in Figure 1.6.1. 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.

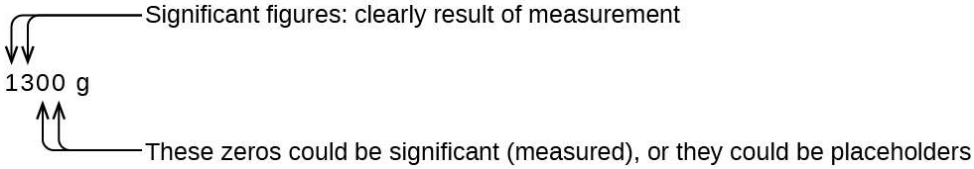


Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.



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 1.6.1: 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 1.6.1

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 1.6.2: 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)

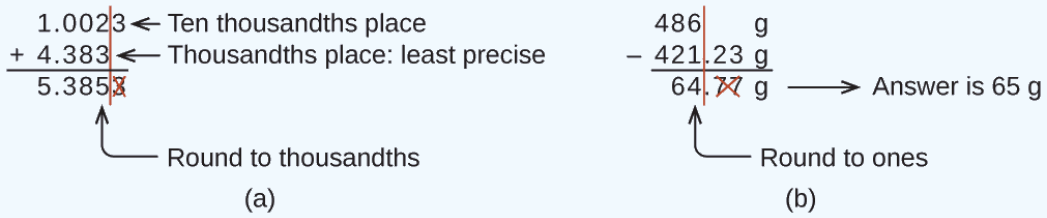
$$\begin{array}{rr} & 1.0023\ \text{g} \\ + & 4.383\ \text{g} \\ \hline & 5.3853\ \text{g} \end{array}$$

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 1.6.2

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 1.6.3: 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 1.6.3

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 1.6.4: 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 1.6.4

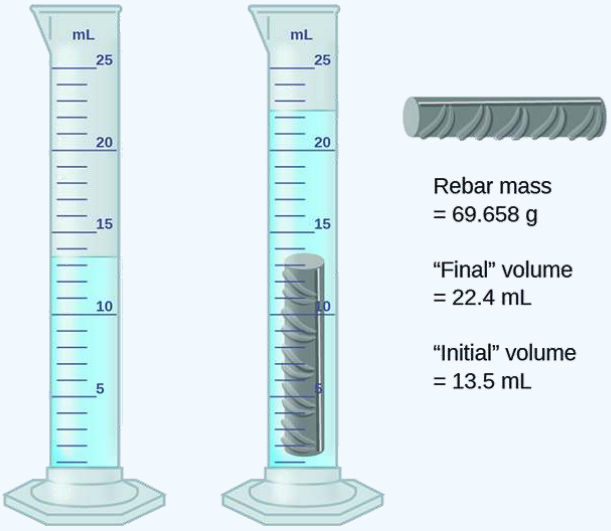
What is the density of a liquid with a mass of 31.1415 g and a volume of 30.13 cm^3?

Answer

1.034 g/mL

Example 1.6.5: 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.



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 1.6.5

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



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 1.6.2).

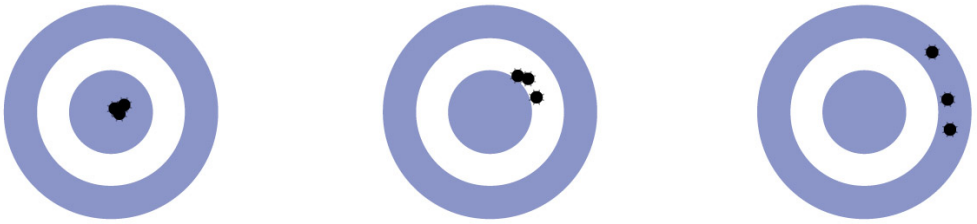


Figure 1.6.2: (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither 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 1.6.1.

Table 1.6.1: 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).

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1.7: 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 1.7.1.

Example 1.7.1: 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 1.7.1: 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 1.7.1).

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 1.7.1

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 1.7.2: 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 1.7.1: 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 1.7.2

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 1.7.3: 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,

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

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 1.7.3

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 1.7.1: shows the relationship among the three temperature scales.

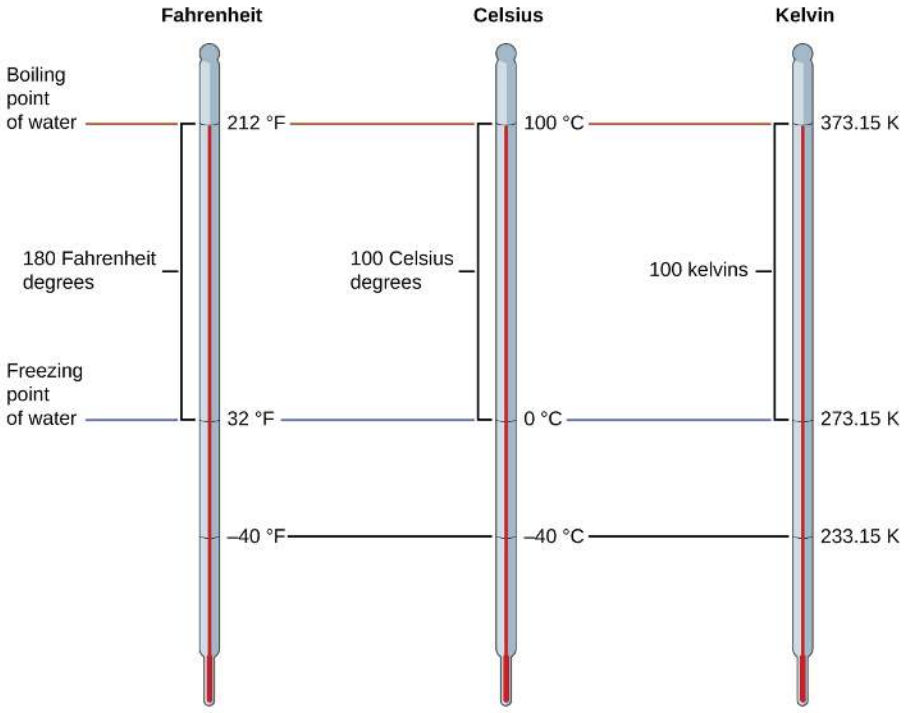


Figure 1.7.1: The Fahrenheit, Celsius, and kelvin temperature scales are compared.

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 1.7.4: 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

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

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

? Exercise 1.7.1

Convert 80.92 °C to K and °F.

Answer

354.07 K, 177.7 °F

✔ Example 1.7.5: 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 1.7.1

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.

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1.8: 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

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1.9: Key Equations

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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

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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?

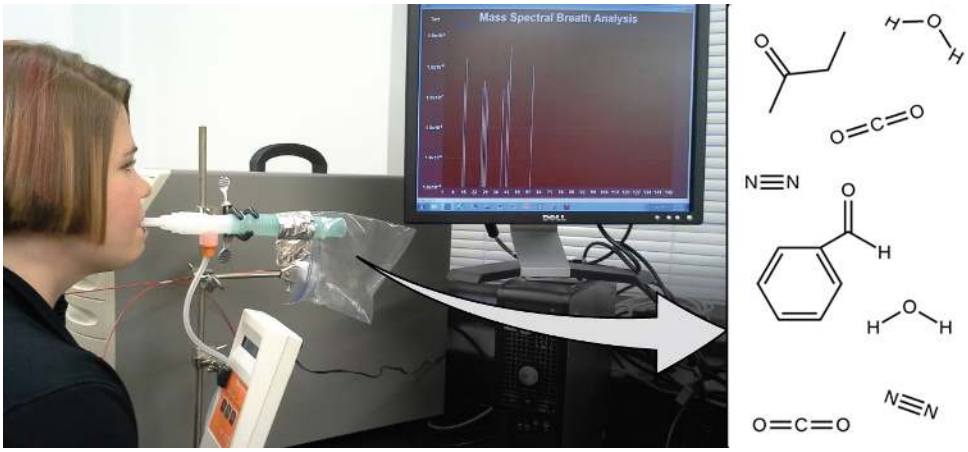


Figure 2.1.1: Analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers)

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.

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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.

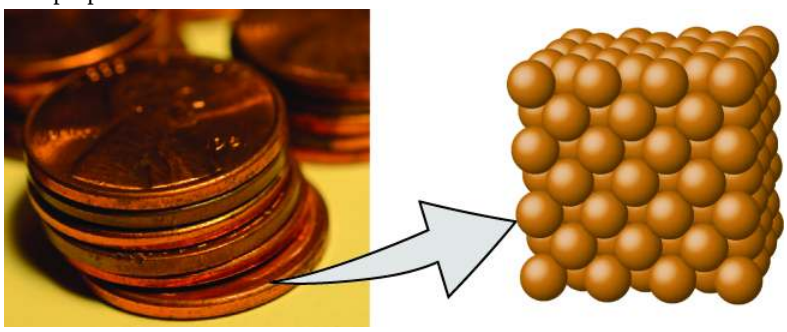


Figure 2.2.1: A pre-1982 copper penny (left) contains approximately 3 10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by "slgckgc"/Flickr)

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).

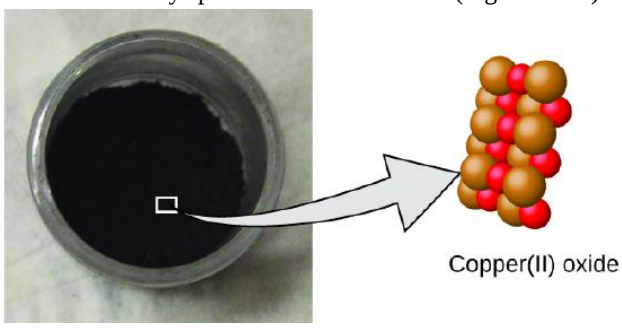


Figure 2.2.2: Copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown spheres) and oxygen (red spheres)—in a 1:1 ratio. (credit: modification of work by "Chemicalinterest"/Wikimedia Commons)

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).

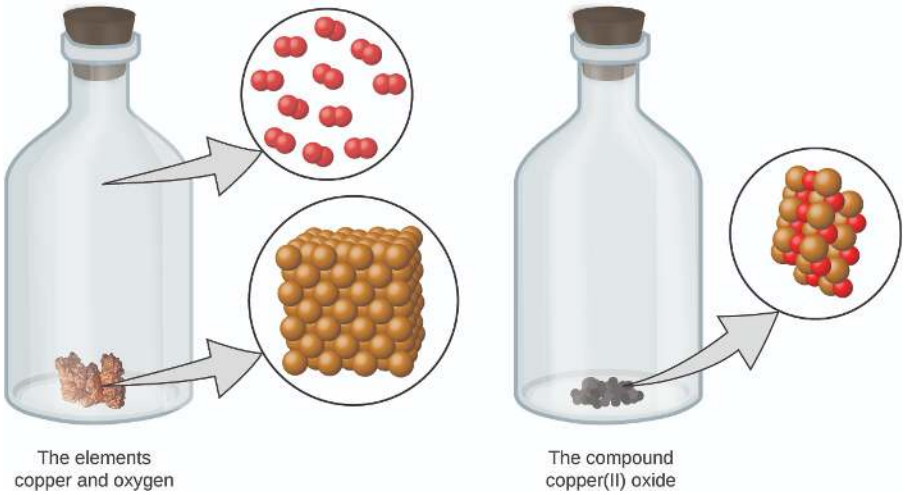


Figure 2.2.3: When the elements copper (a shiny, red-brown solid, shown here as brown spheres) and oxygen (a clear and colorless gas, shown here as red spheres) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work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

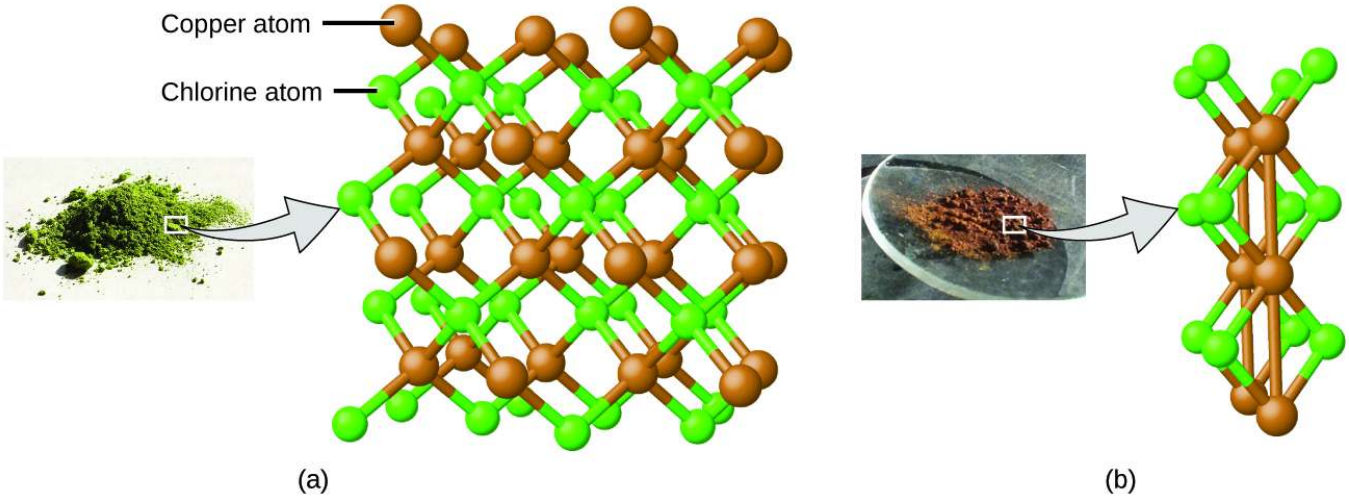


Figure 2.2.4: Compared to the copper chlorine compound in (a), where copper is represented by brown spheres and chlorine by green spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by "Benjah-bmm27"/Wikimedia Commons; credit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

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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).

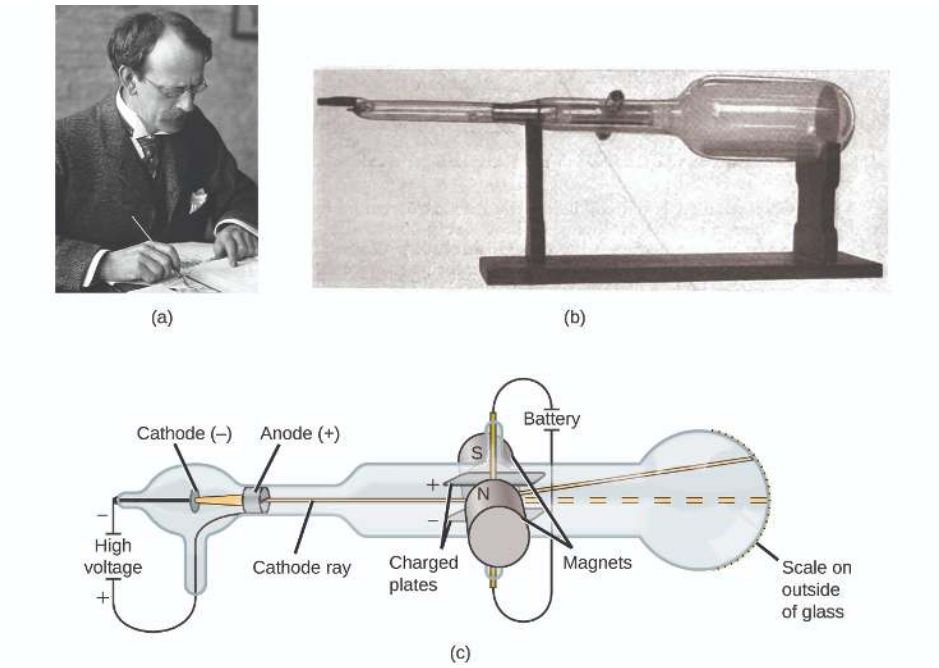


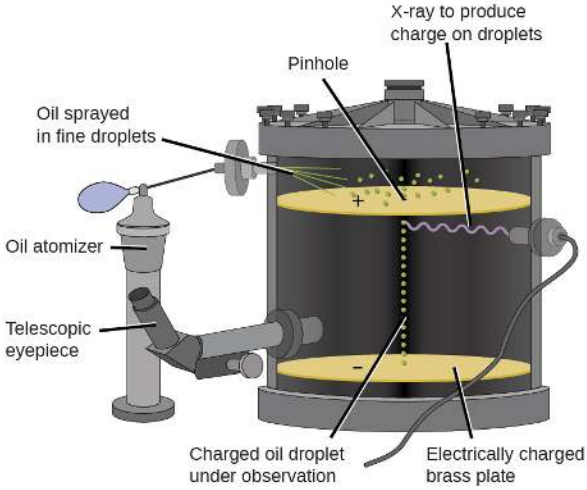
Figure 2.3.1: (a) J. J. Thomson produced a visible beam in a cathode ray tube. (b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun. (c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end 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).

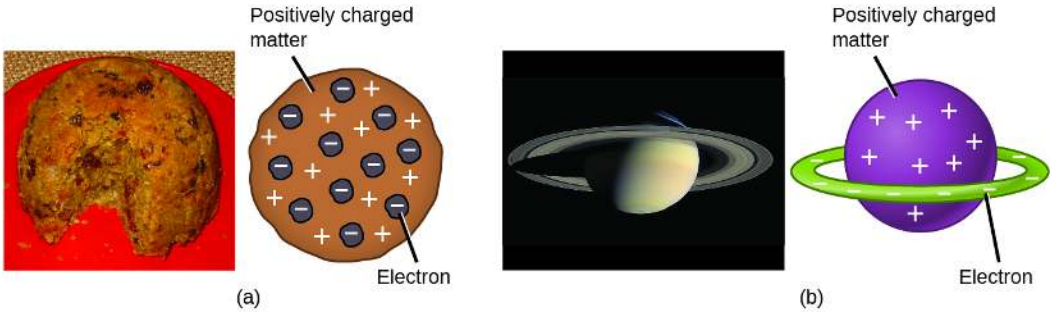


Figure 2.3.3: (a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins ("plums"). (b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive "planet." (credit a: modification of work by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

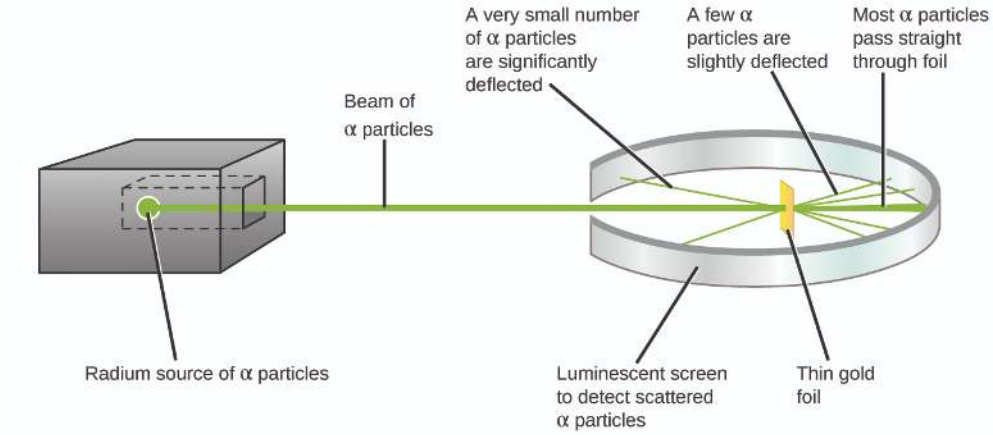


Figure 2.3.4: Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly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

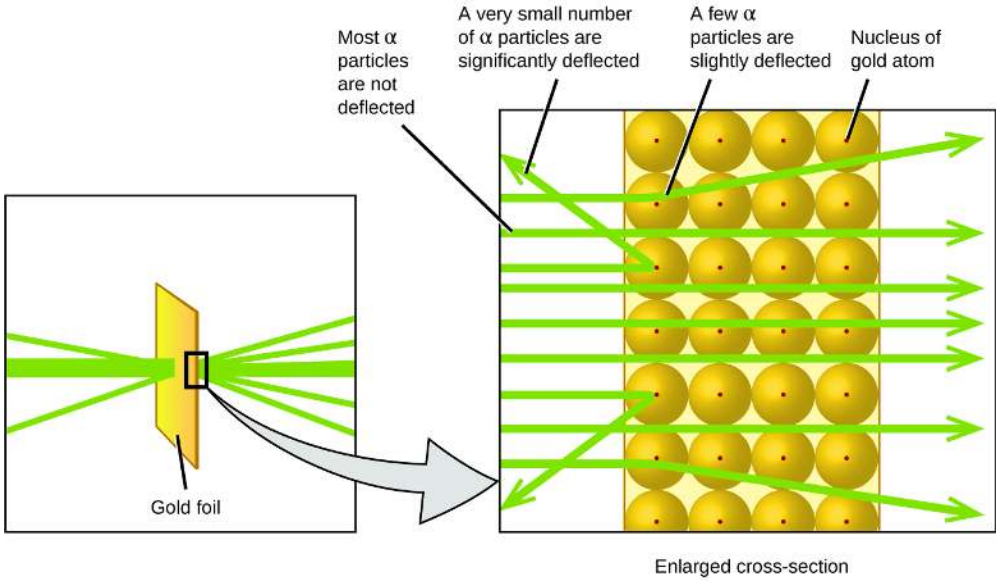


Figure 2.3.5: The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by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

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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.

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

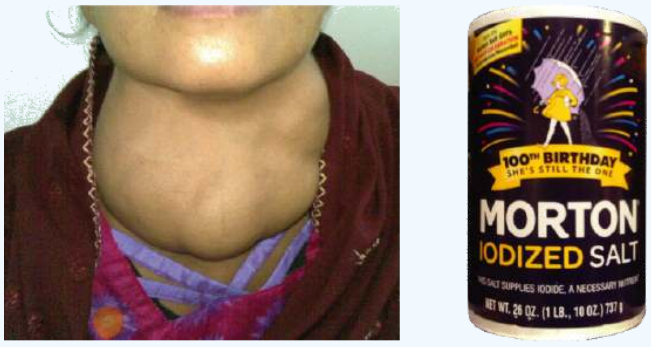
Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an **ion**. The charge of an atom is defined as follows:

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

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.

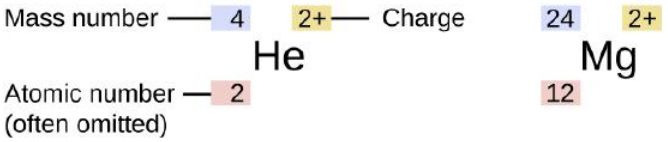


Figure 2.4.4: The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

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.

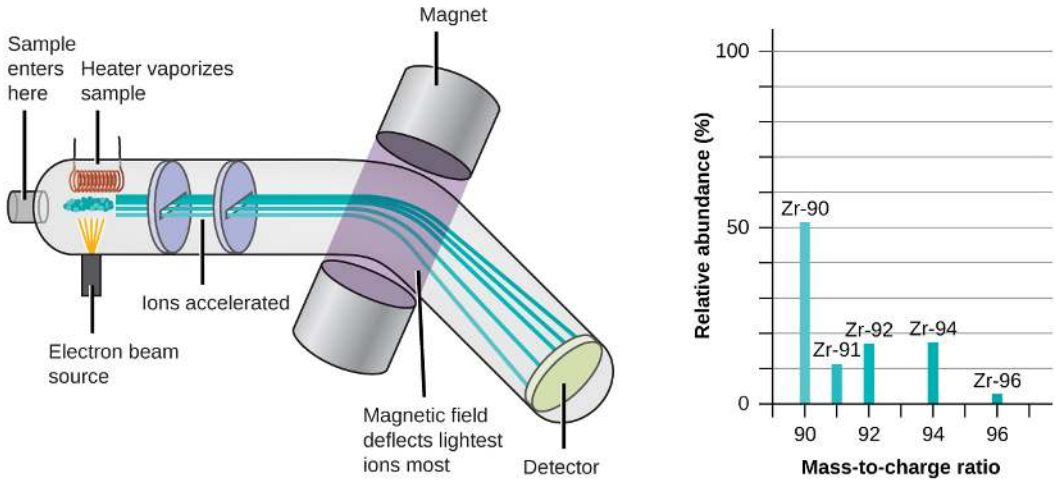


Figure 2.4.5: Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

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.

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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.

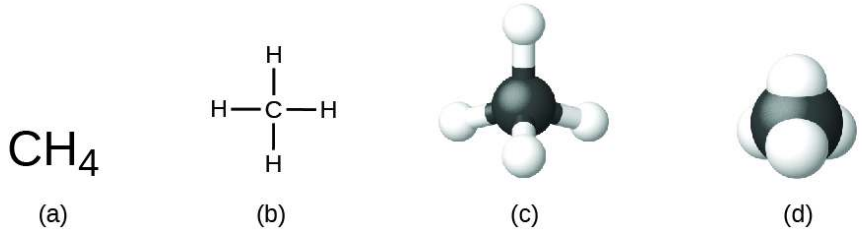


Figure 2.5.1 : A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

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).

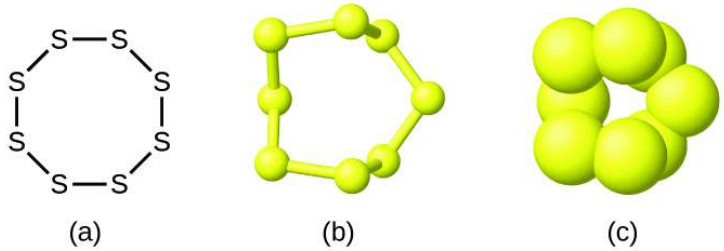


Figure 2.5.2 : A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

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).

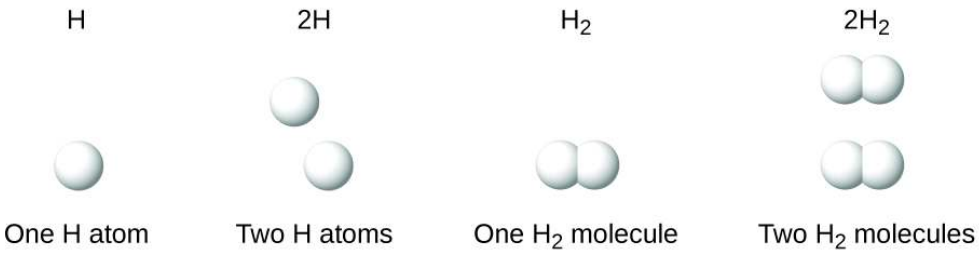


Figure 2.5.3 : The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

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).

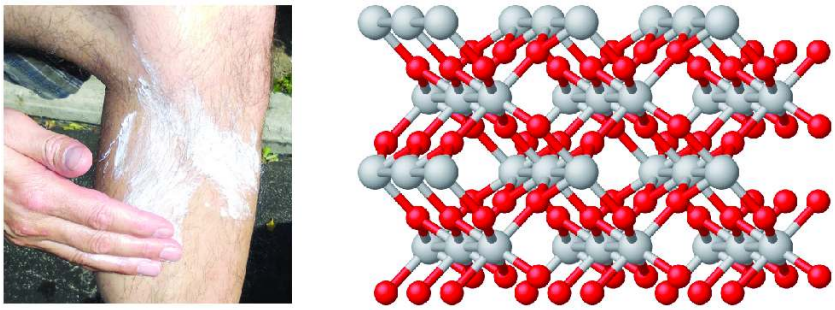


Figure 2.5.4 : (a) The white compound titanium dioxide provides effective protection from the sun. (b) A crystal of titanium dioxide, TiO_2, contains titanium and oxygen in a ratio of 1 to 2. The titanium atoms are gray and the oxygen atoms are red. (credit a: modification of work by “osseous”/Flickr)

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).

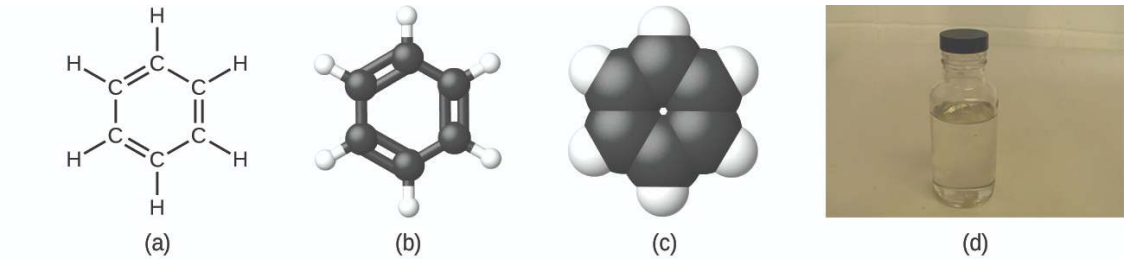
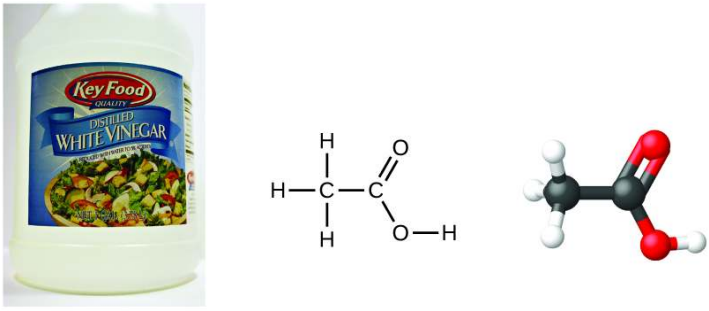


Figure 2.5.5 : Benzene, C_6H_6, is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa)

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.

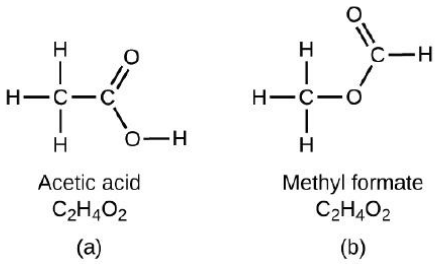


Figure 2.5.8: Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula ($C_2H_4O_2$) but different structures (and therefore different chemical properties).

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.

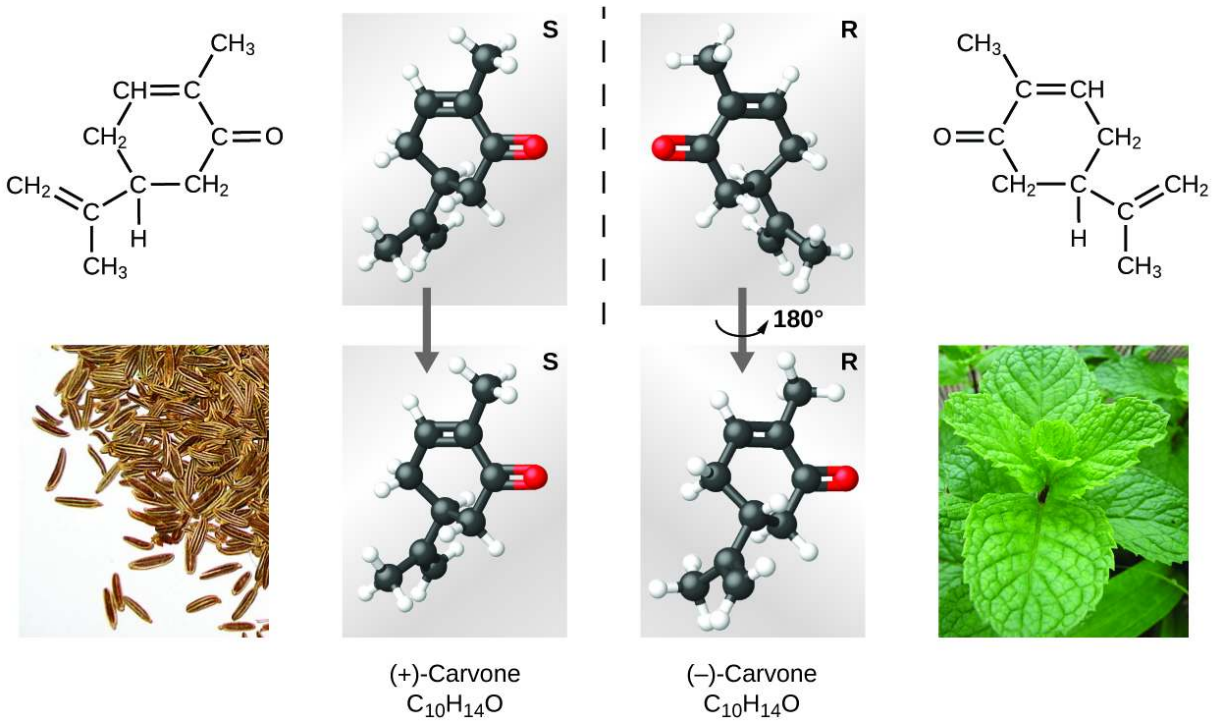


Figure 2.5.0: Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by "Miansari66"/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)

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.

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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

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2.7: Key Equations

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

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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.

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CHAPTER OVERVIEW

3: Electronic Structure and Periodic Properties

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 Energy
3.3: The Bohr Model
3.4: Development of Quantum Theory
3.5: Electronic Structure of Atoms (Electron Configurations)
3.6: Periodic Variations in Element Properties
3.7: The Periodic Table
3.8: Ionic and Molecular Compounds
3.9: Key Terms
3.10: Key Equations
3.11: Summary

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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.

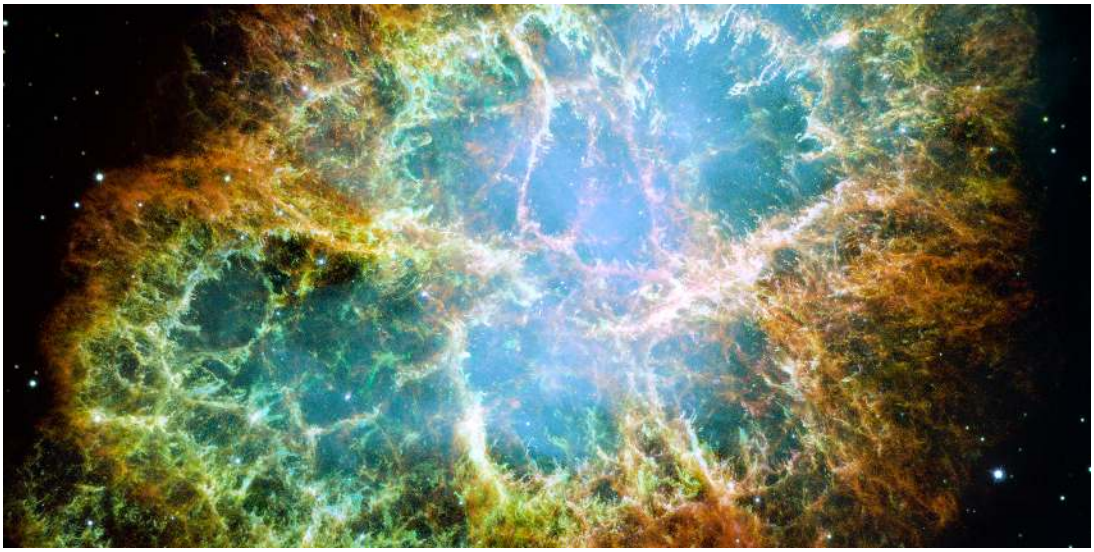


Figure 3.1.1: The Crab Nebula consists of remnants of a supernova (the explosion of a star). NASA’s Hubble Space Telescope produced this composite image. Measurements of the emitted light wavelengths enabled astronomers to identify the elements in the nebula, determining that it contains specific ions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

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3.2: Electromagnetic Energy

Learning Objectives

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

- Explain the basic behavior of waves, including travelling waves and standing waves
- Describe the wave nature of light
- Use appropriate equations to calculate related light-wave properties such as frequency, wavelength, and energy
- Distinguish between line and continuous emission spectra
- Describe the particle nature of light

The nature of light has been a subject of inquiry since antiquity. In the seventeenth century, Isaac Newton performed experiments with lenses and prisms and was able to demonstrate that white light consists of the individual colors of the rainbow combined together. Newton explained his optics findings in terms of a "corpuscular" view of light, in which light was composed of streams of extremely tiny particles travelling at high speeds according to Newton's laws of motion. Others in the seventeenth century, such as Christiaan Huygens, had shown that optical phenomena such as reflection and refraction could be equally well explained in terms of light as waves travelling at high speed through a medium called "luminiferous aether" that was thought to permeate all space. Early in the nineteenth century, Thomas Young demonstrated that light passing through narrow, closely spaced slits produced interference patterns that could not be explained in terms of Newtonian particles but could be easily explained in terms of waves. Later in the nineteenth century, after James Clerk Maxwell developed his theory of electromagnetic radiation and showed that light was the visible part of a vast spectrum of electromagnetic waves, the particle view of light became thoroughly discredited. By the end of the nineteenth century, scientists viewed the physical universe as roughly comprising two separate domains: matter composed of particles moving according to Newton's laws of motion, and electromagnetic radiation consisting of waves governed by Maxwell's equations. Today, these domains are referred to as classical mechanics and classical electrodynamics (or classical electromagnetism). Although there were a few physical phenomena that could not be explained within this framework, scientists at that time were so confident of the overall soundness of this framework that they viewed these aberrations as puzzling paradoxes that would ultimately be resolved somehow within this framework. As we shall see, these paradoxes led to a contemporary framework that intimately connects particles and waves at a fundamental level called wave-particle duality, which has superseded the classical view.

Visible light and other forms of electromagnetic radiation play important roles in chemistry, since they can be used to infer the energies of electrons within atoms and molecules. Much of modern technology is based on electromagnetic radiation. For example, radio waves from a mobile phone, X-rays used by dentists, the energy used to cook food in your microwave, the radiant heat from red-hot objects, and the light from your television screen are forms of electromagnetic radiation that all exhibit wavelike behavior.

Waves

A **wave** is an oscillation or periodic movement that can transport energy from one point in space to another. Common examples of waves are all around us. Shaking the end of a rope transfers energy from your hand to the other end of the rope, dropping a pebble into a pond causes waves to ripple outward along the water's surface, and the expansion of air that accompanies a lightning strike generates sound waves (thunder) that can travel outward for several miles. In each of these cases, kinetic energy is transferred through matter (the rope, water, or air) while the matter remains essentially in place. An insightful example of a wave occurs in sports stadiums when fans in a narrow region of seats rise simultaneously and stand with their arms raised up for a few seconds before sitting down again while the fans in neighboring sections likewise stand up and sit down in sequence. While this wave can quickly encircle a large stadium in a few seconds, none of the fans actually travel with the wave-they all stay in or above their seats.

Waves need not be restricted to travel through matter. As Maxwell showed, electromagnetic waves consist of an electric field oscillating in step with a perpendicular magnetic field, both of which are perpendicular to the direction of travel. These waves can travel through a vacuum at a constant speed of 2.998 108 m/s, the speed of light (denoted by c).

All waves, including forms of electromagnetic radiation, are characterized by, a **wavelength** (denoted by λ, the lowercase Greek letter lambda), a **frequency** (denoted by ν, the lowercase Greek letter nu), and an **amplitude**. As can be seen in Figure $3.2.1$, the wavelength is the distance between two consecutive peaks or troughs in a wave (measured in meters in the SI system).

Electromagnetic waves have wavelengths that fall within an enormous range-wavelengths of kilometers (10^3 m) to picometers (10^{-12} m) have been observed. The frequency is the number of wave cycles that pass a specified point in space in a specified amount of time (in the SI system, this is measured in seconds). A cycle corresponds to one complete wavelength. The unit for frequency, expressed as cycles per second [s^{-1}], is the **hertz** (Hz). Common multiples of this unit are megahertz, ($1\,\text{MHz} = 1 \times 10^6\,\text{Hz}$) and gigahertz ($1\,\text{GHz} = 1 \times 10^9\,\text{Hz}$). The amplitude corresponds to the magnitude of the wave's displacement and so, in Figure 3.2.1, this corresponds to one-half the height between the peaks and troughs. The amplitude is related to the intensity of the wave, which for light is the brightness, and for sound is the loudness.

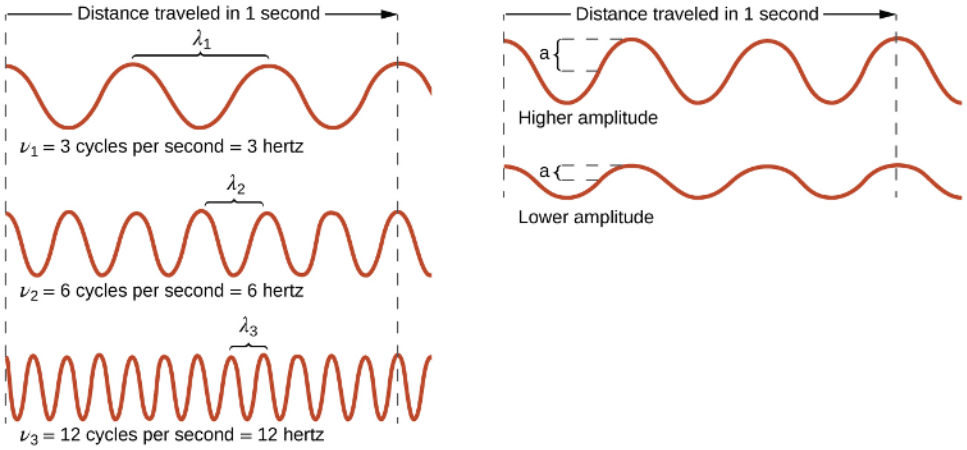


Figure 3.2.1: One-dimensional sinusoidal waves show the relationship among wavelength, frequency, and speed. The wave with the shortest wavelength has the highest frequency. Amplitude is one-half the height of the wave from peak to trough.

The product of a wave's wavelength (λ) and its frequency (ν), $\lambda\nu$, is the speed of the wave. Thus, for electromagnetic radiation in a vacuum, speed is equal to the fundamental constant, c:

$$c = 2.998 \times 10^8 ms^{-1} = \lambda\nu$$

Wavelength and frequency are inversely proportional: As the wavelength increases, the frequency decreases. The inverse proportionality is illustrated in Figure 3.2.2. This figure also shows the electromagnetic spectrum, the range of all types of electromagnetic radiation. Each of the various colors of visible light has specific frequencies and wavelengths associated with them, and you can see that visible light makes up only a small portion of the electromagnetic spectrum. Because the technologies developed to work in various parts of the electromagnetic spectrum are different, for reasons of convenience and historical legacies, different units are typically used for different parts of the spectrum. For example, radio waves are usually specified as frequencies (typically in units of MHz), while the visible region is usually specified in wavelengths (typically in units of nm or angstroms).

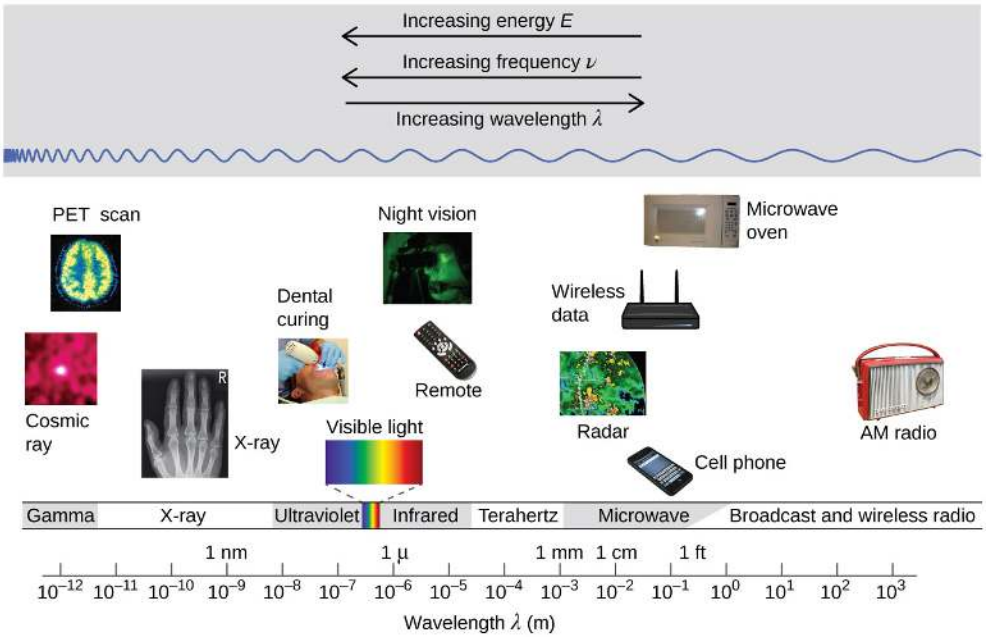


Figure 3.2.2: Portions of the electromagnetic spectrum are shown in order of decreasing frequency and increasing wavelength. (credit “Cosmic ray": modification of work by NASA; credit “PET scan": modification of work by the National Institute of Health; credit “X-ray": modification of work by Dr. Jochen Lengerke; credit “Dental curing": modification of work by the Department of the Navy; credit “Night vision": modification of work by the Department of the Army; credit “Remote": modification of work by Emilian Robert Vicol; credit “Cell phone": modification of work by Brett Jordan; credit “Microwave oven": modification of work by Billy Mabray; credit “Ultrasound": modification of work by Jane Whitney; credit “AM radio": modification of work by Dave Clausen)

Example 3.2.1: Determining the Frequency and Wavelength of Radiation

A sodium streetlight gives off yellow light that has a wavelength of 589 nm (1 nm = 1×10^{-9} m). What is the frequency of this light?

Solution

We can rearrange the Equation $c = \lambda\nu$ to solve for the frequency:

$$\nu = \frac{c}{\lambda}$$

Since c is expressed in meters per second, we must also convert 589 nm to meters.

$$\nu = \left(\frac{2.998 \times 10^{8} ms^{-1}}{589nm}\right)\left(\frac{1 \times 10^{9} nm}{1m}\right) = 5.09 \times 10^{14} s^{-1}$$

Exercise 3.2.1

One of the frequencies used to transmit and receive cellular telephone signals in the United States is 850 MHz. What is the wavelength in meters of these radio waves?

Answer

0.353 m = 35.3 cm

Chemistry in Everyday Life: Wireless Communication



Figure 3.2.3: Radio and cell towers are typically used to transmit long-wavelength electromagnetic radiation. Increasingly, cell towers are designed to blend in with the landscape, as with the Tucson, Arizona, cell tower (right) disguised as a palm tree. (credit left: modification of work by Sir Mildred Pierce; credit middle: modification of work by M.O. Stevens)

Many valuable technologies operate in the radio (3 kHz-300 GHz) frequency region of the electromagnetic spectrum. At the low frequency (low energy, long wavelength) end of this region are AM (amplitude modulation) radio signals (540-2830 kHz) that can travel long distances. FM (frequency modulation) radio signals are used at higher frequencies (87.5-108.0 MHz). In AM radio, the information is transmitted by varying the amplitude of the wave (Figure 3.2.4). In FM radio, by contrast, the amplitude is constant and the instantaneous frequency varies.

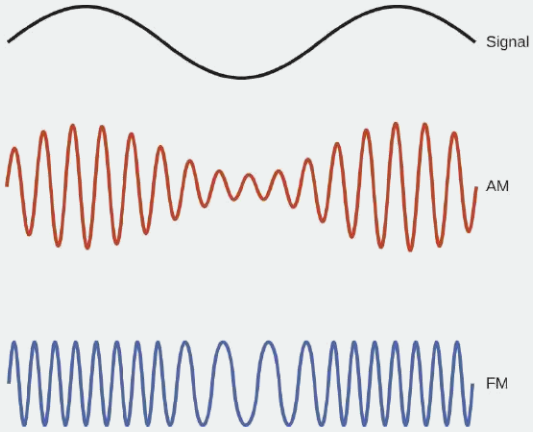


Figure 3.2.4: This schematic depicts how amplitude modulation (AM) and frequency modulation (FM) can be used to transmit a radio wave.

Other technologies also operate in the radio-wave portion of the electromagnetic spectrum. For example, 4G cellular telephone signals are approximately 880 MHz, while Global Positioning System (GPS) signals operate at 1.228 and 1.575 GHz, local area wireless technology (Wi-Fi) networks operate at 2.4 to 5 GHz, and highway toll sensors operate at 5.8 GHz. The frequencies associated with these applications are convenient because such waves tend not to be absorbed much by common building materials.

One particularly characteristic phenomenon of waves results when two or more waves come into contact: They interfere with each other. Figure 3.2.5 shows the interference patterns that arise when light passes through narrow slits closely spaced about a wavelength apart. The fringe patterns produced depend on the wavelength, with the fringes being more closely spaced for shorter wavelength light passing through a given set of slits. When the light passes through the two slits, each slit effectively acts as a new source, resulting in two closely spaced waves coming into contact at the detector (the camera in this case). The dark regions in Figure 3.2.5 correspond to regions where the peaks for the wave from one slit happen to coincide with the troughs for the wave from the other slit (destructive interference), while the brightest regions correspond to the regions where the peaks for the two waves (or their two troughs) happen to coincide (constructive interference). Likewise, when two stones are tossed close together into a pond, interference patterns are visible in the interactions between the waves produced by the stones. Such interference patterns cannot be explained by particles moving according to the laws of classical mechanics.

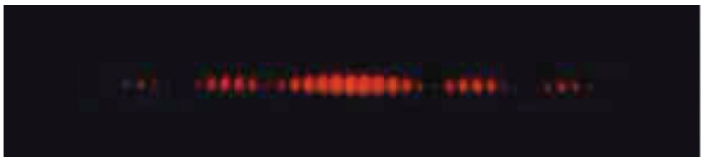


Figure 3.2.5: Interference fringe patterns are shown for light passing through two closely spaced, narrow slits. The spacing of the fringes depends on the wavelength, with the fringes being more closely spaced for the shorter-wavelength blue light. (credit: PASCO)

Portrait of a Chemist: Dorothy Crowfoot Hodgkin

X-rays exhibit wavelengths of approximately 0.01–10 nm. Since these wavelengths are comparable to the spaces between atoms in a crystalline solid, X-rays are scattered when they pass through crystals. The scattered rays undergo constructive and destructive interference that creates a specific diffraction pattern that may be measured and used to precisely determine the positions of atoms within the crystal. This phenomenon of X-ray diffraction is the basis for very powerful techniques enabling the determination of molecular structure. One of the pioneers who applied this powerful technology to important biochemical substances was Dorothy Crowfoot Hodgkin.

Born in Cairo, Egypt, in 1910 to British parents, Dorothy's fascination with chemistry was fostered early in her life. At age 11 she was enrolled in a prestigious English grammar school where she was one of just two girls allowed to study chemistry. On her 16th birthday, her mother, Molly, gifted her a book on X-ray crystallography, which had a profound impact on the trajectory of her career. She studied chemistry at Oxford University, graduating with first-class honors in 1932 and directly entering Cambridge University to pursue a doctoral degree. At Cambridge, Dorothy recognized the promise of X-ray crystallography for protein structure determinations, conducting research that earned her a PhD in 1937. Over the course of a very productive career, Dr. Hodgkin was credited with determining structures for several important biomolecules, including cholesterol iodide, penicillin, and vitamin B12. In recognition of her achievements in the use of X-ray techniques to elucidate the structures of biochemical substances, she was awarded the 1964 Nobel Prize in Chemistry. In 1969, she led a team of scientists who deduced the structure of insulin, facilitating the mass production of this hormone and greatly advancing the treatment of diabetic patients worldwide. Dr. Hodgkin continued working with the international scientific community, earning numerous distinctions and awards prior to her death in 1993.

Not all waves are travelling waves. Standing waves (also known as stationary waves) remain constrained within some region of space. As we shall see, standing waves play an important role in our understanding of the electronic structure of atoms and molecules. The simplest example of a standing wave is a one-dimensional wave associated with a vibrating string that is held fixed at its two end points. Figure 3.2.6 shows the four lowest-energy standing waves (the fundamental wave and the lowest three harmonics) for a vibrating string at a particular amplitude. Although the string's motion lies mostly within a plane, the wave itself is considered to be one dimensional, since it lies along the length of the string. The motion of string segments in a direction perpendicular to the string length generates the waves and so the amplitude of the waves is visible as the maximum displacement of the curves seen in Figure 3.2.6. The key observation from the figure is *that only those waves having an integer number, n, of half-wavelengths between the end points can form*. A system with fixed end points such as this restricts the number and type of the possible waveforms. This is an example of **quantization**, in which only discrete values from a more general set of continuous values of some property are observed. Another important observation is that the harmonic waves (those waves displaying more than one-half wavelength) all have one or more points between the two end points that are not in motion. These special points are **nodes**. The energies of the standing waves with a given amplitude in a vibrating string increase with the number of half-wavelengths n. Since the number of nodes is $n-1$, the energy can also be said to depend on the number of nodes, generally increasing as the number of nodes increases.

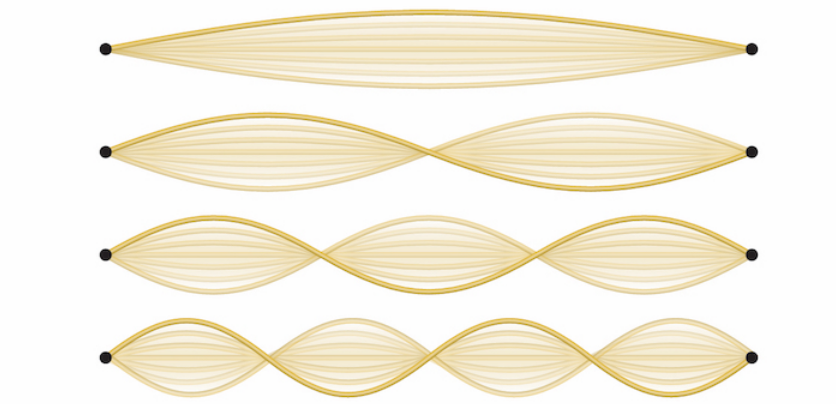


Figure 3.2.6: A vibrating string shows some one-dimensional standing waves. Since the two end points of the string are held fixed, only waves having an integer number of half-wavelengths can form. The points on the string between the end points that are not moving are called the nodes.

An example of two-dimensional standing waves is shown in Figure 3.2.7, which shows the vibrational patterns on a flat surface. Although the vibrational amplitudes cannot be seen like they could in the vibrating string, the nodes have been made visible by sprinkling the drum surface with a powder that collects on the areas of the surface that have minimal displacement. For one-dimensional standing waves, the nodes were points on the line, but for two-dimensional standing waves, the nodes are lines on the surface (for three-dimensional standing waves, the nodes are two-dimensional surfaces within the three-dimensional volume).

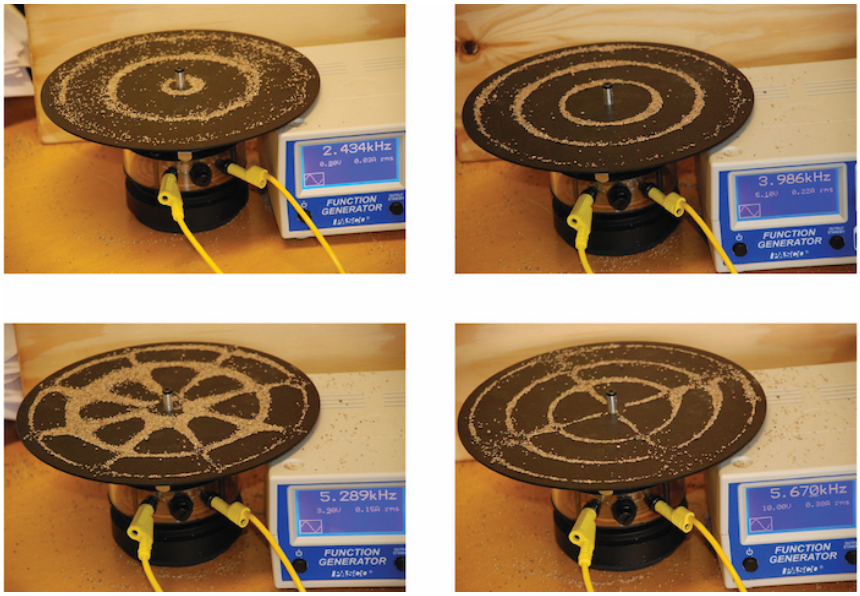


Figure 3.2.7: Two-dimensional standing waves can be visualized on a vibrating surface. The surface has been sprinkled with a powder that collects near the nodal lines. There are two types of nodes visible: radial nodes (circles) and angular nodes (radii).

Link to Learning

You can watch the formation of various radial nodes here as singer Imogen Heap projects her voice across a kettle drum.

Blackbody Radiation and the Ultraviolet Catastrophe

The last few decades of the nineteenth century witnessed intense research activity in commercializing newly discovered electric lighting. This required obtaining a better understanding of the distributions of light emitted from various sources being considered. Artificial lighting is usually designed to mimic natural sunlight within the limitations of the underlying technology. Such lighting consists of a range of broadly distributed frequencies that form a **continuous spectrum**. Figure 3.2.8 shows the wavelength distribution for sunlight. The most intense radiation is in the visible region, with the intensity dropping off rapidly for shorter wavelength ultraviolet (UV) light, and more slowly for longer wavelength infrared (IR) light.

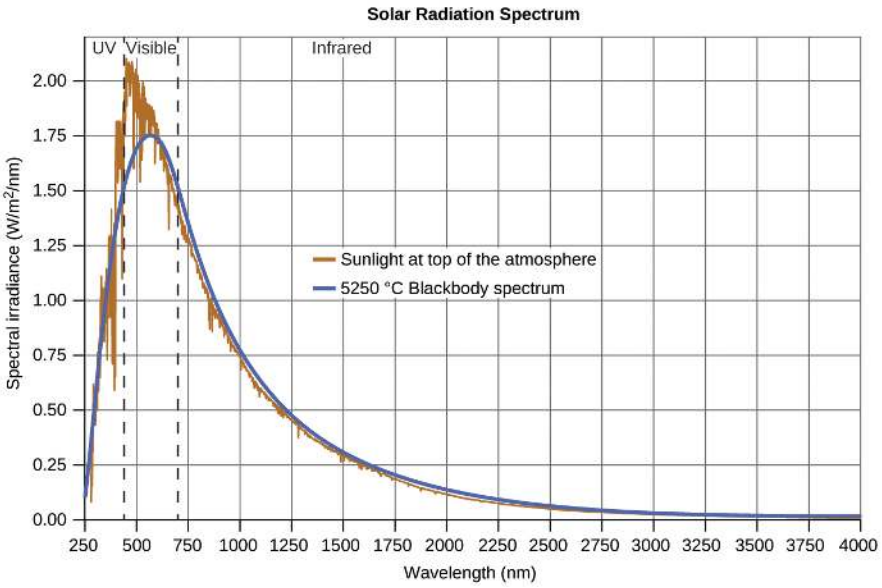


Figure 3.2.8: The spectral distribution (light intensity vs. wavelength) of sunlight reaches the Earth's atmosphere as UV light, visible light, and IR light. The unabsorbed sunlight at the top of the atmosphere has a distribution that approximately matches the theoretical distribution of a blackbody at 5250 °C, represented by the blue curve. (credit: modification of work by American Society for Testing and Materials (ASTM) Terrestrial Reference Spectra for Photovoltaic Performance Evaluation)

In Figure 3.2.8, the solar distribution is compared to a representative distribution, called a **blackbody spectrum**, that corresponds to a temperature of 5250 °C. The blackbody spectrum matches the solar spectrum quite well. A blackbody is a convenient, ideal emitter that approximates the behavior of many materials when heated. It is “ideal” in the same sense that an ideal gas is a convenient, simple representation of real gases that works well, provided that the pressure is not too high nor the temperature too low. A good approximation of a blackbody that can be used to observe blackbody radiation is a metal oven that can be heated to very high temperatures. The oven has a small hole allowing for the light being emitted within the oven to be observed with a spectrometer so that the wavelengths and their intensities can be measured. Figure 3.2.9 shows the resulting curves for some representative temperatures. Each distribution depends only on a single parameter: the temperature. The maxima in the blackbody curves, λ_{max}, shift to shorter wavelengths as the temperature increases, reflecting the observation that metals being heated to high temperatures begin to glow a darker red that becomes brighter as the temperature increases, eventually becoming white hot at very high temperatures as the intensities of all of the visible wavelengths become appreciable. This common observation was at the heart of the first paradox that showed the fundamental limitations of classical physics that we will examine.Physicists derived mathematical expressions for the blackbody curves using well-accepted concepts from the theories of classical mechanics and classical electromagnetism. The theoretical expressions as functions of temperature fit the observed experimental blackbody curves well at longer wavelengths, but showed significant discrepancies at shorter wavelengths. Not only did the theoretical curves not show a peak, they absurdly showed the intensity becoming infinitely large as the wavelength became smaller, which would imply that everyday objects at room temperature should be emitting large amounts of UV light. This became known as the “ultraviolet catastrophe” because no one could find any problems with the theoretical treatment that could lead to such unrealistic short-wavelength behavior. Finally, around 1900, Max Planck derived a theoretical expression for blackbody radiation that fit the experimental observations exactly (within experimental error). Planck developed his theoretical treatment by extending the earlier work that had been based on the premise that the atoms composing the oven vibrated at increasing frequencies (or decreasing wavelengths) as the temperature increased, with these vibrations being the source of the emitted electromagnetic radiation. But where the earlier treatments had allowed the vibrating atoms to have any energy values obtained from a continuous set of energies (perfectly reasonable, according to classical physics), Planck found that by restricting the vibrational energies to discrete values for each frequency, he could derive an expression for blackbody radiation that correctly had the intensity dropping rapidly for the short wavelengths in the UV region.

$$E = nh\nu, n = 1, 2, 3, \ldots \tag{3.2.1}$$

The quantity h is a constant now known as **Planck's constant**, in his honor. Although Planck was pleased he had resolved the blackbody radiation paradox, he was disturbed that to do so, he needed to assume the vibrating atoms required quantized energies,

which he was unable to explain. The value of Planck's constant is very small, 6.626 10^{-34} joule seconds (J s), which helps explain why energy quantization had not been observed previously in macroscopic phenomena.

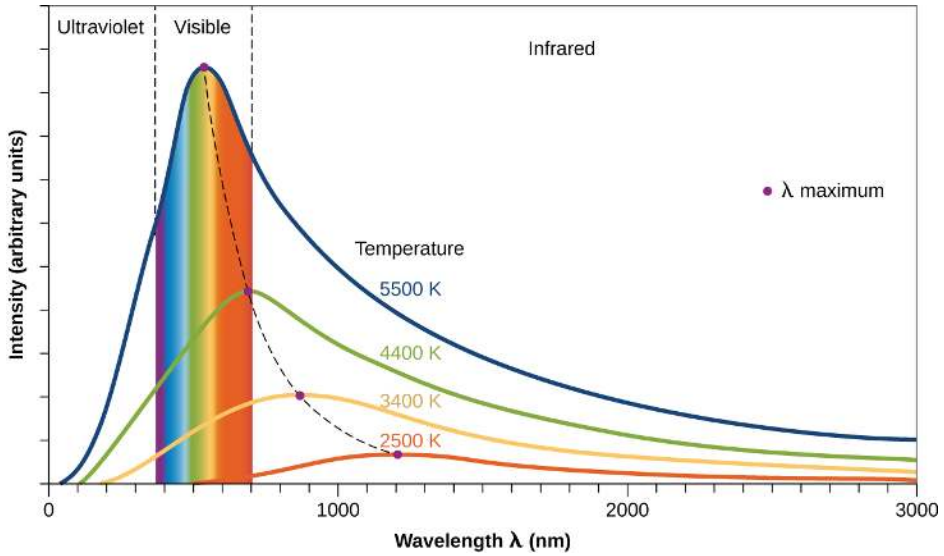


Figure 3.2.9: Blackbody spectral distribution curves are shown for some representative temperatures.

The Photoelectric Effect

The next paradox in the classical theory to be resolved concerned the photoelectric effect (Figure 3.2.10). It had been observed that electrons could be ejected from the clean surface of a metal when light having a frequency greater than some threshold frequency was shone on it. Surprisingly, the kinetic energy of the ejected electrons did not depend on the brightness of the light, but increased with increasing frequency of the light. Since the electrons in the metal had a certain amount of binding energy keeping them there, the incident light needed to have more energy to free the electrons. According to classical wave theory, a wave's energy depends on its intensity (which depends on its amplitude), not its frequency. One part of these observations was that the number of electrons ejected within in a given time period was seen to increase as the brightness increased. In 1905, Albert Einstein was able to resolve the paradox by incorporating Planck's quantization findings into the discredited particle view of light (Einstein actually won his Nobel prize for this work, and not for his theories of relativity for which he is most famous).

Einstein argued that the quantized energies that Planck had postulated in his treatment of blackbody radiation could be applied to the light in the photoelectric effect so that the light striking the metal surface should not be viewed as a wave, but instead as a stream of particles (later called **photons**) whose energy depended on their frequency, according to Planck's formula (Equation 3.2.1), $E = h\nu$ (or, in terms of wavelength using $c = \nu\lambda$, $E = \frac{hc}{\lambda}$). Electrons were ejected when hit by photons having sufficient energy (a frequency greater than the threshold). The greater the frequency, the greater the kinetic energy imparted to the escaping electrons by the collisions. Einstein also argued that the light intensity did not depend on the amplitude of the incoming wave, but instead corresponded to the number of photons striking the surface within a given time period. This explains why the number of ejected electrons increased with increasing brightness, since the greater the number of incoming photons, the greater the likelihood that they would collide with some of the electrons.

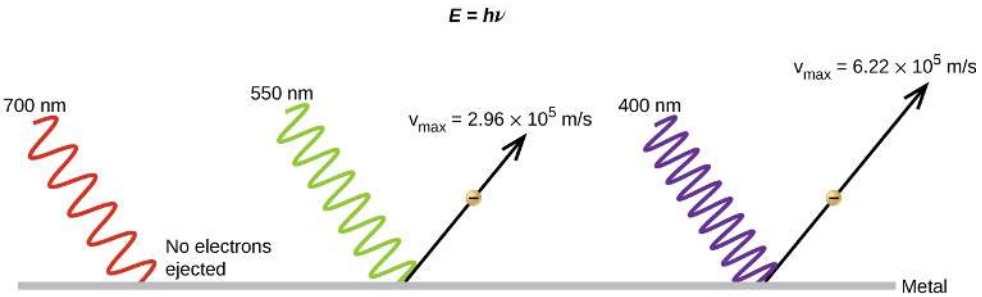


Figure 3.2.10: Photons with low frequencies do not have enough energy to cause electrons to be ejected via the photoelectric effect. For any frequency of light above the threshold frequency, the kinetic energy of an ejected electron will increase linearly with the energy of the incoming photon.

With Einstein's findings, the nature of light took on a new air of mystery. Although many light phenomena could be explained either in terms of waves or particles, certain phenomena, such as the interference patterns obtained when light passed through a double slit, were completely contrary to a particle view of light, while other phenomena, such as the photoelectric effect, were completely contrary to a wave view of light. Somehow, at a deep fundamental level still not fully understood, light is both wavelike and particle-like. This is known as **wave-particle duality**.

Example 3.2.2: Calculating the Energy of Radiation

When we see light from a neon sign, we are observing radiation from excited neon atoms. If this radiation has a wavelength of 640 nm, what is the energy of the photon being emitted?

Solution

We use the part of Planck's equation that includes the wavelength, λ, and convert units of nanometers to meters so that the units of λ and c are the same.

$$\begin{aligned} E &= \frac{hc}{\lambda} \\ &= \frac{\left(6.626 \times 10^{-34}\ \text{J}\cdot\cancel{\text{s}}\right)\left(2.998 \times 10^{8}\ \cancel{\text{m}}\ \cancel{\text{s}}^{-1}\right)}{\left(640\ \cancel{\text{nm}}\right)\left(\frac{1\ \cancel{\text{m}}}{10^{9}\ \cancel{\text{nm}}}\right)} \\ &= 3.10 \times 10^{-19} J \end{aligned}$$

Exercise 3.2.2

The microwaves in an oven are of a specific frequency that will heat the water molecules contained in food. (This is why most plastics and glass do not become hot in a microwave oven-they do not contain water molecules.) This frequency is about 3 10^9 Hz. What is the energy of one photon in these microwaves?

Answer

2 10^{-24} J

Link to Learning

Use this simulation program to experiment with the photoelectric effect to see how intensity, frequency, type of metal, and other factors influence the ejected photons.

Example 3.2.3: Photoelectric Effect

Identify which of the following statements are false and, where necessary, change the italicized word or phrase to make them true, consistent with Einstein's explanation of the photoelectric effect.

a. Increasing the brightness of incoming light *increases* the kinetic energy of the ejected electrons.
b. Increasing the *wavelength* of incoming light increases the kinetic energy of the ejected electrons.
c. Increasing the brightness of incoming light *increases* the number of ejected electrons.
d. Increasing the *frequency* of incoming light can increase the number of ejected electrons.

Solution

a. False. Increasing the brightness of incoming light *has no effect* on the kinetic energy of the ejected electrons. Only energy, not the number or amplitude, of the photons influences the kinetic energy of the electrons.
b. False. Increasing the *frequency* of incoming light increases the kinetic energy of the ejected electrons. Frequency is proportional to energy and inversely proportional to wavelength. Frequencies above the threshold value transfer the excess energy into the kinetic energy of the electrons.
c. True. Because the number of collisions with photons increases with brighter light, the number of ejected electrons increases.

d. True with regard to the threshold energy binding the electrons to the metal. Below this threshold, electrons are not emitted and above it they are. Once over the threshold value, further increasing the frequency does not increase the number of ejected electrons

? Exercise 3.2.3

Calculate the threshold energy in kJ/mol of electrons in aluminum, given that the lowest frequency photon for which the photoelectric effect is observed is 9.87 10^{14} Hz.

Answer

394 kJ/mol

Line Spectra

Another paradox within the classical electromagnetic theory that scientists in the late nineteenth century struggled with concerned the light emitted from atoms and molecules. When solids, liquids, or condensed gases are heated sufficiently, they radiate some of the excess energy as light. Photons produced in this manner have a range of energies, and thereby produce a continuous spectrum in which an unbroken series of wavelengths is present. Most of the light generated from stars (including our sun) is produced in this fashion. You can see all the visible wavelengths of light present in sunlight by using a prism to separate them. As can be seen in Figure 3.2.8: , sunlight also contains UV light (shorter wavelengths) and IR light (longer wavelengths) that can be detected using instruments but that are invisible to the human eye. Incandescent (glowing) solids such as tungsten filaments in incandescent lights also give off light that contains all wavelengths of visible light. These continuous spectra can often be approximated by blackbody radiation curves at some appropriate temperature, such as those shown in Figure 3.2.9.

In contrast to continuous spectra, light can also occur as discrete or **line spectra** having very narrow line widths interspersed throughout the spectral regions such as those shown in Figure 3.2.12 Exciting a gas at low partial pressure using an electrical current, or heating it, will produce line spectra. Fluorescent light bulbs and neon signs operate in this way (Figure 3.2.11). Each element displays its own characteristic set of lines, as do molecules, although their spectra are generally much more complicated.



Figure 3.2.11: Neon signs operate by exciting a gas at low partial pressure using an electrical current. This sign shows the elaborate artistic effects that can be achieved. (credit: Dave Shaver)

Each emission line consists of a single wavelength of light, which implies that the light emitted by a gas consists of a set of discrete energies. For example, when an electric discharge passes through a tube containing hydrogen gas at low pressure, the H_2 molecules are broken apart into separate H atoms and we see a blue-pink color. Passing the light through a prism produces a line spectrum, indicating that this light is composed of photons of four visible wavelengths, as shown in Figure 3.2.12

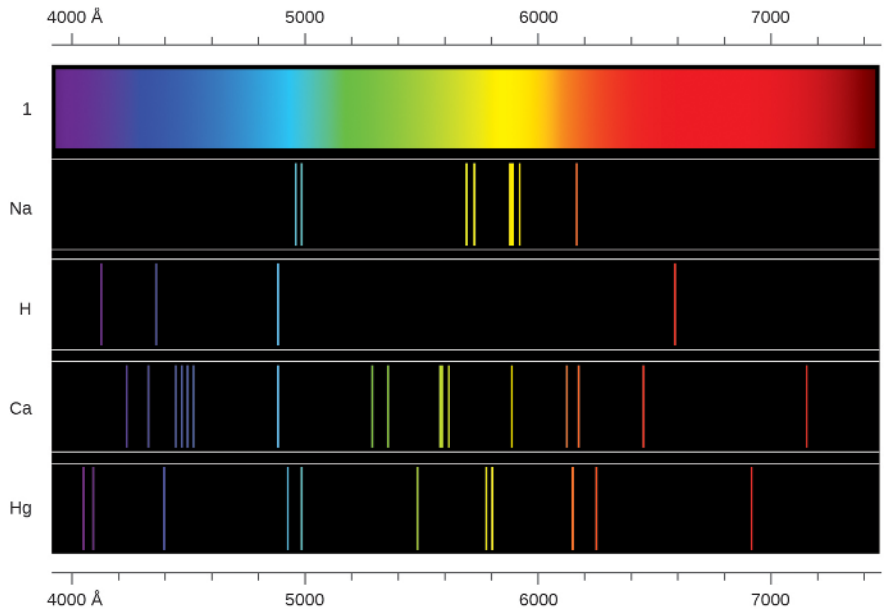


Figure 3.2.12: Compare the two types of emission spectra: continuous spectrum of white light (top) and the line spectra of the light from excited sodium, hydrogen, calcium, and mercury atoms.

The origin of discrete spectra in atoms and molecules was extremely puzzling to scientists in the late nineteenth century, since according to classical electromagnetic theory, only continuous spectra should be observed. Even more puzzling, in 1885, Johann Balmer was able to derive an empirical equation that related the four visible wavelengths of light emitted by hydrogen atoms to whole integers. That equation is the following one, in which *k* is a constant:

$$\frac{1}{\lambda} = k\left(\frac{1}{4} - \frac{1}{n^2}\right), \quad n = 3, 4, 5, 6$$

Other discrete lines for the hydrogen atom were found in the UV and IR regions. Johannes Rydberg generalized Balmer's work and developed an empirical formula that predicted all of hydrogen's emission lines, not just those restricted to the visible range, where n_1 and n_2 are integers, $n_1 < n_2$, and R_∞ is the **Rydberg constant** (1.097 10^7 m^{-1}).

$$\frac{1}{\lambda} = R_\infty\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Even in the late nineteenth century, spectroscopy was a very precise science, and so the wavelengths of hydrogen were measured to very high accuracy, which implied that the Rydberg constant could be determined very precisely as well. That such a simple formula as the Rydberg formula could account for such precise measurements seemed astounding at the time, but it was the eventual explanation for emission spectra by Neils Bohr in 1913 that ultimately convinced scientists to abandon classical physics and spurred the development of modern quantum mechanics.

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3.3: The Bohr Model

Learning Objectives

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

- Describe the Bohr model of the hydrogen atom
- Use the Rydberg equation to calculate energies of light emitted or absorbed by hydrogen atoms

Following the work of Ernest Rutherford and his colleagues in the early twentieth century, the picture of atoms consisting of tiny dense nuclei surrounded by lighter and even tinier electrons continually moving about the nucleus was well established. This picture was called the planetary model, since it pictured the atom as a miniature "solar system" with the electrons orbiting the nucleus like planets orbiting the sun. The simplest atom is hydrogen, consisting of a single proton as the nucleus about which a single electron moves. The electrostatic force attracting the electron to the proton depends only on the distance between the two particles. This classical mechanics description of the atom is incomplete, however, since an electron moving in an elliptical orbit would be accelerating (by changing direction) and, according to classical electromagnetism, it should continuously emit electromagnetic radiation. This loss in orbital energy should result in the electron's orbit getting continually smaller until it spirals into the nucleus, implying that atoms are inherently unstable.

In 1913, Niels Bohr attempted to resolve the atomic paradox by ignoring classical electromagnetism's prediction that the orbiting electron in hydrogen would continuously emit light. Instead, he incorporated into the classical mechanics description of the atom Planck's ideas of quantization and Einstein's finding that light consists of photons whose energy is proportional to their frequency. Bohr assumed that the electron orbiting the nucleus would not normally emit any radiation (the stationary state hypothesis), but it would emit or absorb a photon if it moved to a different orbit. The energy absorbed or emitted would reflect differences in the orbital energies according to this equation:

$$|\Delta E| = |E_f - E_i| = h\nu = \frac{hc}{\lambda}$$

In this equation, h is Planck's constant and E_i and E_f are the initial and final orbital energies, respectively. The absolute value of the energy difference is used, since frequencies and wavelengths are always positive. Instead of allowing for continuous values of energy, Bohr assumed the energies of these electron orbitals were quantized:

$$E_n = -\frac{k}{n^2}, n = 1, 2, 3, \ldots$$

In this expression, k is a constant comprising fundamental constants such as the electron mass and charge and Planck's constant. Inserting the expression for the orbit energies into the equation for ΔE gives

$$\Delta E = k\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) = \frac{hc}{\lambda}$$

or

$$\frac{1}{\lambda} = \frac{k}{hc}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

which is identical to the Rydberg equation in which

$$R_\infty = \frac{k}{hc}$$

When Bohr calculated his theoretical value for the Rydberg constant, R_∞, and compared it with the experimentally accepted value, he got excellent agreement. Since the Rydberg constant was one of the most precisely measured constants at that time, this level of agreement was astonishing and meant that Bohr's model was taken seriously, despite the many assumptions that Bohr needed to derive it.

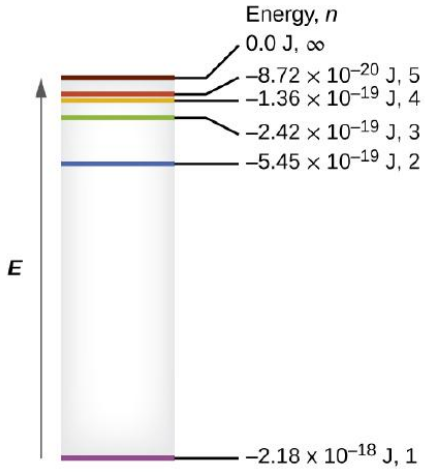


Figure 3.3.1: Quantum numbers and energy levels in a hydrogen atom. The more negative the calculated value, the lower the energy.

The lowest few energy levels are shown in Figure 3.3.1. One of the fundamental laws of physics is that matter is most stable with the lowest possible energy. Thus, the electron in a hydrogen atom usually moves in the $n = 1$ orbit, the orbit in which it has the lowest energy. When the electron is in this lowest energy orbit, the atom is said to be in its ground electronic state (or simply ground state). If the atom receives energy from an outside source, it is possible for the electron to move to an orbit with a higher n value and the atom is now in an excited electronic state (or simply an excited state) with a higher energy. When an electron transitions from an excited state (higher energy orbit) to a less excited state, or ground state, the difference in energy is emitted as a photon. Similarly, if a photon is absorbed by an atom, the energy of the photon moves an electron from a lower energy orbit up to a more excited one. We can relate the energy of electrons in atoms to what we learned previously about energy. The law of conservation of energy says that we can neither create nor destroy energy. Thus, if a certain amount of external energy is required to excite an electron from one energy level to another, that same amount of energy will be liberated when the electron returns to its initial state (Figure 3.3.2).

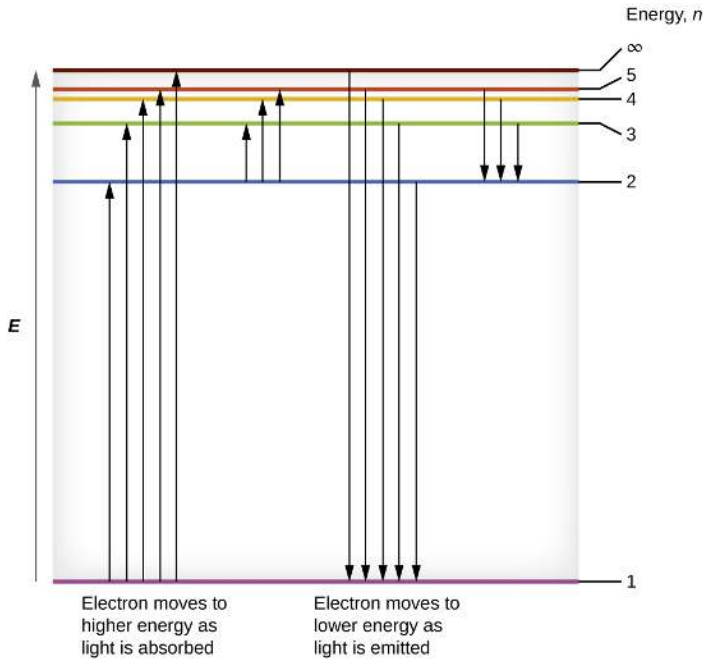


Figure 3.3.2: The horizontal lines show the relative energy of orbits in the Bohr model of the hydrogen atom, and the vertical arrows depict the energy of photons absorbed (left) or emitted (right) as electrons move between these orbits.

Since Bohr's model involved only a single electron, it could also be applied to the single electron ions He^+, Li^{2+}, Be^{3+}, and so forth, which differ from hydrogen only in their nuclear charges, and so one-electron atoms and ions are collectively referred to as

hydrogen-like atoms. The energy expression for hydrogen-like atoms is a generalization of the hydrogen atom energy, in which Z is the nuclear charge (+1 for hydrogen, +2 for He, +3 for Li, and so on) and k has a value of 2.179 10^{-18} J.

$$E_n = -\frac{kZ^2}{n^2} \tag{3.3.1}$$

The sizes of the circular orbits for hydrogen-like atoms are given in terms of their radii by the following expression, in which \ (a_0\) is a constant called the **Bohr radius**, with a value of 5.292 10^{-11} m:

$$r = \frac{n^2}{Z} a_0$$

The equation also shows us that as the electron's energy increases (as n increases), the electron is found at greater distances from the nucleus. This is implied by the inverse dependence of electrostatic attraction on distance, since, as the electron moves away from the nucleus, the electrostatic attraction between it and the nucleus decreases and it is held less tightly in the atom. Note that as n gets larger and the orbits get larger, their energies get closer to zero, and so the limits
$n \rightarrow \infty$ and $r \rightarrow \longrightarrow \infty$ imply that $E = 0$ corresponds to the ionization limit where the electron is completely removed from the nucleus. Thus, for hydrogen in the ground state n = 1, the ionization energy would be:

$$\Delta E = E_{n \longrightarrow \infty} - E_1 = 0 + k = k$$

With three extremely puzzling paradoxes now solved (blackbody radiation, the photoelectric effect, and the hydrogen atom), and all involving Planck's constant in a fundamental manner, it became clear to most physicists at that time that the classical theories that worked so well in the macroscopic world were fundamentally flawed and could not be extended down into the microscopic domain of atoms and molecules. Unfortunately, despite Bohr's remarkable achievement in deriving a theoretical expression for the Rydberg constant, he was unable to extend his theory to the next simplest atom, He, which only has two electrons. Bohr's model was severely flawed, since it was still based on the classical mechanics notion of precise orbits, a concept that was later found to be untenable in the microscopic domain, when a proper model of quantum mechanics was developed to supersede classical mechanics.

✔ Example 3.3.1: Calculating the Energy of an Electron in a Bohr Orbit

Early researchers were very excited when they were able to predict the energy of an electron at a particular distance from the nucleus in a hydrogen atom. If a spark promotes the electron in a hydrogen atom into an orbit with *n* = 3, what is the calculated energy, in joules, of the electron?

Solution

The energy of the electron is given by Equation 3.3.1:

$$E = \frac{-kZ^2}{n^2}$$

The atomic number, Z, of hydrogen is 1; *k* = 2.179 10^{-18} J; and the electron is characterized by an \(*n*\) value of 3. Thus,

$$E = \frac{-\left(2.179 \times 10^{-18} J\right) \times (1)^2}{(3)^2} = -2.421 \times 10^{-19} J$$

? Exercise 3.3.1

The electron in Figure 3.3.2 is promoted even further to an orbit with $n = 6$. What is its new energy?

Answer

−6.053 10^{-20} J

Example 3.3.2: Calculating the Energy and Wavelength of Electron Transitions in a One–electron (Bohr) System

What is the energy (in joules) and the wavelength (in meters) of the line in the spectrum of hydrogen that represents the movement of an electron from Bohr orbit with $n = 4$ to the orbit with $n = 6$? In what part of the electromagnetic spectrum do we find this radiation?

Solution

In this case, the electron starts out with $n = 4$, so $n_1 = 4$. It comes to rest in the $n = 6$ orbit, so $n_2 = 6$. The difference in energy between the two states is given by this expression:

$$\begin{aligned}\Delta E &= E_1 - E_2 \\ &= 2.179 \times 10^{-18}\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \\ &= 2.179 \times 10^{-18}\left(\frac{1}{4^2} - \frac{1}{6^2}\right) J \\ &= 2.179 \times 10^{-18}\left(\frac{1}{16} - \frac{1}{36}\right) J \\ &= 7.566 \times 10^{-20} J\end{aligned}$$

This energy difference is positive, indicating a photon enters the system (is absorbed) to excite the electron from the *n* = 4 orbit up to the *n* = 6 orbit. The wavelength of a photon with this energy is found by the expression $E = \frac{hc}{\lambda}$. Rearrangement gives:

$$\begin{aligned}\lambda &= \frac{hc}{E} \\ &= \left(6.626 \times 10^{-34}\ \text{J}\cdot\cancel{\text{s}}\right) \times \frac{2.998 \times 10^{8}\ \text{m}\ \cancel{\text{s}^{-1}}}{7.566 \times 10^{-20}\ \cancel{\text{J}}} \\ &= 2.626 \times 10^{-6}\ \text{m}\end{aligned}$$

From the illustration of the electromagnetic spectrum in Electromagnetic Energy, we can see that this wavelength is found in the infrared portion of the electromagnetic spectrum.

Exercise 3.3.2

What is the energy in joules and the wavelength in meters of the photon produced when an electron falls from the *n* = 5 to the *n* = 3 level in a He^+ ion (*Z* = 2 for He^+)?

Answer

6.198 10^{-19} J; 3.205 10^{-7} m

Bohr's model of the hydrogen atom provides insight into the behavior of matter at the microscopic level, but it does not account for electron–electron interactions in atoms with more than one electron. It does introduce several important features of all models used to describe the distribution of electrons in an atom. These features include the following:

- The energies of electrons (energy levels) in an atom are quantized, described by quantum numbers: integer numbers having only specific allowed value and used to characterize the arrangement of electrons in an atom.
- An electron's energy increases with increasing distance from the nucleus.
- The discrete energies (lines) in the spectra of the elements result from quantized electronic energies.

Of these features, the most important is the postulate of quantized energy levels for an electron in an atom. As a consequence, the model laid the foundation for the quantum mechanical model of the atom. Bohr won a Nobel Prize in Physics for his contributions to our understanding of the structure of atoms and how that is related to line spectra emissions.

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3.4: Development of Quantum Theory

Learning Objectives

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

- Extend the concept of wave–particle duality that was observed in electromagnetic radiation to matter as well
- Understand the general idea of the quantum mechanical description of electrons in an atom, and that it uses the notion of three-dimensional wave functions, or orbitals, that define the distribution of probability to find an electron in a particular part of space
- List and describe traits of the four quantum numbers that form the basis for completely specifying the state of an electron in an atom

Bohr's model explained the experimental data for the hydrogen atom and was widely accepted, but it also raised many questions. Why did electrons orbit at only fixed distances defined by a single quantum number n = 1, 2, 3, and so on, but never in between? Why did the model work so well describing hydrogen and one-electron ions, but could not correctly predict the emission spectrum for helium or any larger atoms? To answer these questions, scientists needed to completely revise the way they thought about matter.

Behavior in the Microscopic World

We know how matter behaves in the macroscopic world—objects that are large enough to be seen by the naked eye follow the rules of classical physics. A billiard ball moving on a table will behave like a particle: It will continue in a straight line unless it collides with another ball or the table cushion, or is acted on by some other force (such as friction). The ball has a well-defined position and velocity (or a well-defined momentum, $p = mv$, defined by mass m and velocity v) at any given moment. In other words, the ball is moving in a classical trajectory. This is the typical behavior of a classical object.

When waves interact with each other, they show interference patterns that are not displayed by macroscopic particles such as the billiard ball. For example, interacting waves on the surface of water can produce interference patterns similar to those shown on Figure 3.4.1. This is a case of wave behavior on the macroscopic scale, and it is clear that particles and waves are very different phenomena in the macroscopic realm.

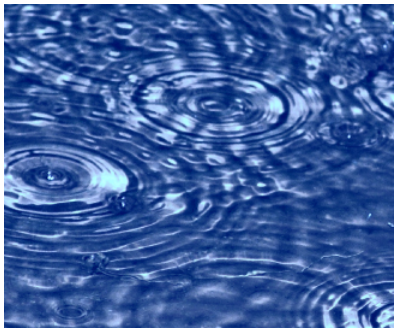


Figure 3.4.1: An interference pattern on the water surface is formed by interacting waves. The waves are caused by reflection of water from the rocks. (credit: modification of work by Sukanto Debnath)

As technological improvements allowed scientists to probe the microscopic world in greater detail, it became increasingly clear by the 1920s that very small pieces of matter follow a different set of rules from those we observe for large objects. The unquestionable separation of waves and particles was no longer the case for the microscopic world.

One of the first people to pay attention to the special behavior of the microscopic world was Louis de Broglie. He asked the question: If electromagnetic radiation can have particle-like character, can electrons and other submicroscopic particles exhibit wavelike character? In his 1925 doctoral dissertation, de Broglie extended the wave–particle duality of light that Einstein used to resolve the photoelectric-effect paradox to material particles. He predicted that a particle with mass m and velocity v (that is, with linear momentum p) should also exhibit the behavior of a wave with a wavelength value λ, given by this expression in which h is the familiar Planck's constant:

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

This is called the **de Broglie wavelength**. Unlike the other values of λ discussed in this chapter, the de Broglie wavelength is a characteristic of particles and other bodies, not electromagnetic radiation (note that this equation involves velocity [v, m/s], not frequency [ν, Hz]. Although these two symbols appear nearly identical, they mean very different things). Where Bohr had postulated the electron as being a particle orbiting the nucleus in quantized orbits, de Broglie argued that Bohr's assumption of quantization can be explained if the electron is considered not as a particle, but rather as a circular standing wave such that only an integer number of wavelengths could fit exactly within the orbit (Figure 3.4.2).

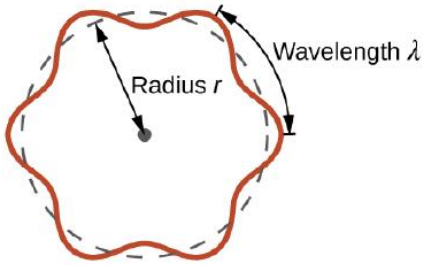


Figure 3.4.2: If an electron is viewed as a wave circling around the nucleus, an integer number of wavelengths must fit into the orbit for this standing wave behavior to be possible.

For a circular orbit of radius r, the circumference is $2\pi r$, and so de Broglie's condition is:

$$2\pi r = n\lambda, \qquad n = 1, 2, 3, \ldots$$

Shortly after de Broglie proposed the wave nature of matter, two scientists at Bell Laboratories, C. J. Davisson and L. H. Germer, demonstrated experimentally that electrons can exhibit wavelike behavior by showing an interference pattern for electrons travelling through a regular atomic pattern in a crystal. The regularly spaced atomic layers served as slits, as used in other interference experiments. Since the spacing between the layers serving as slits needs to be similar in size to the wavelength of the tested wave for an interference pattern to form, Davisson and Germer used a crystalline nickel target for their "slits," since the spacing of the atoms within the lattice was approximately the same as the de Broglie wavelengths of the electrons that they used. Figure 3.4.3 shows an interference pattern. It is strikingly similar to the interference patterns for light shown in Electromagnetic Energy for light passing through two closely spaced, narrow slits. The wave–particle duality of matter can be seen in Figure 3.4.3 by observing what happens if electron collisions are recorded over a long period of time. Initially, when only a few electrons have been recorded, they show clear particle-like behavior, having arrived in small localized packets that appear to be random. As more and more electrons arrived and were recorded, a clear interference pattern that is the hallmark of wavelike behavior emerged. Thus, it appears that while electrons are small localized particles, their motion does not follow the equations of motion implied by classical mechanics, but instead it is governed by some type of a wave equation. Thus the wave–particle duality first observed with photons is actually a fundamental behavior intrinsic to all quantum particles.

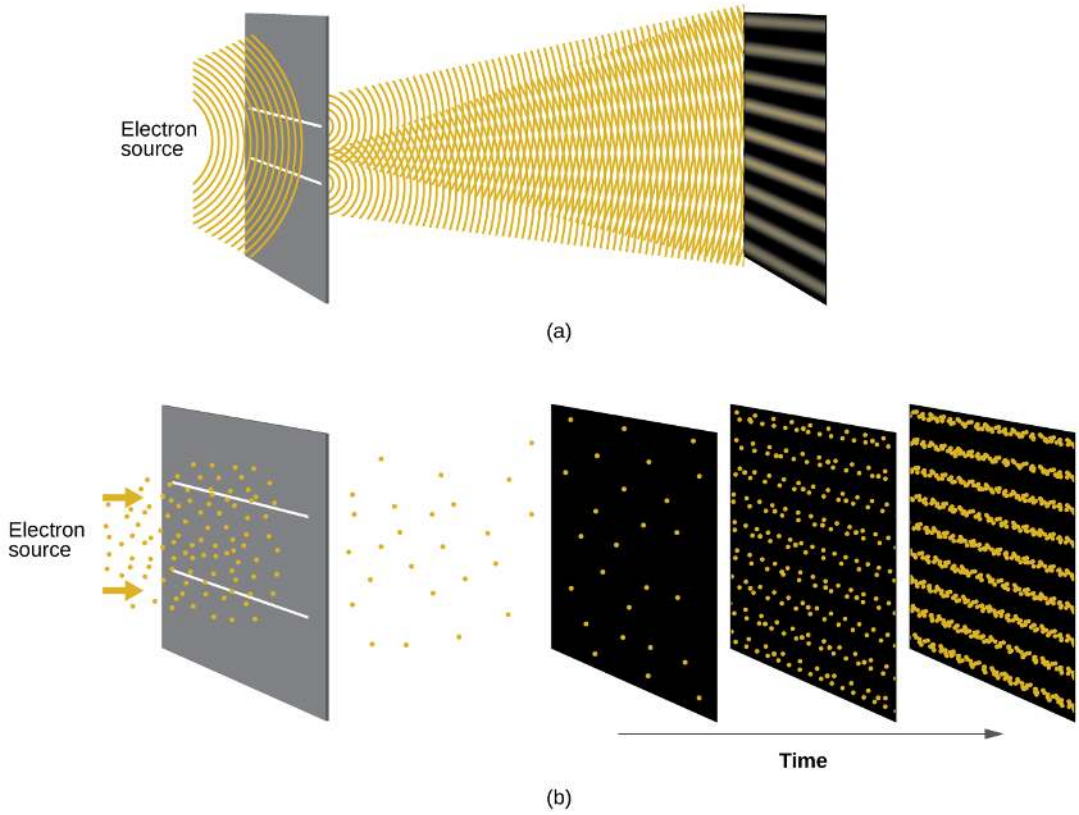


Figure 3.4.3: (a) The interference pattern for electrons passing through very closely spaced slits demonstrates that quantum particles such as electrons can exhibit wavelike behavior. (b) The experimental results illustrated here demonstrate the wave–particle duality in electrons.

Link to Learning

View the Dr. Quantum – Double Slit Experiment cartoon for an easy-to-understand description of wave–particle duality and the associated experiments.

Example 3.4.1: Calculating the Wavelength of a Particle

If an electron travels at a velocity of $1.000 \times 10^7\ \text{m s}^{-1}$ and has a mass of $9.109 \times 10^{-28}\ \text{g}$, what is its wavelength?

Solution

We can use de Broglie's equation to solve this problem, but we first must do a unit conversion of Planck's constant. You learned earlier that $1\ \text{J} = 1\text{kg m}^2/\text{s}^2$. Thus, we can write $h = 6.626 \times 10^{-34}\ \text{J} \cdot \text{s}$ as $6.626 \times 10^{-34}\ \text{kg m}^2/\text{s}$.

$$\begin{aligned} \lambda &= \frac{h}{mv} \\ &= \frac{6.626 \times 10^{-34}\ \text{kg m}^2/\text{s}}{\left(9.109 \times 10^{-31}\ \text{kg}\right)\left(1.000 \times 10^7\ \text{m/s}\right)} \\ &= 7.274 \times 10^{-11} m \end{aligned}$$

This is a small value, but it is significantly larger than the size of an electron in the classical (particle) view. This size is the same order of magnitude as the size of an atom. This means that electron wavelike behavior is going to be noticeable in an atom.

Exercise 3.4.1

Calculate the wavelength of a softball with a mass of 100 g traveling at a velocity of 35 m s^{-1}, assuming that it can be modeled as a single particle.

Answer

$1.9\ 10^{-34}$ m.

We never think of a thrown softball having a wavelength, since this wavelength is so small it is impossible for our senses or any known instrument to detect (strictly speaking, the wavelength of a real baseball would correspond to the wavelengths of its constituent atoms and molecules, which, while much larger than this value, would still be microscopically tiny). The de Broglie wavelength is only appreciable for matter that has a very small mass and/or a very high velocity.

Werner Heisenberg considered the limits of how accurately we can measure properties of an electron or other microscopic particles. He determined that there is a fundamental limit to how accurately one can measure both a particle's position and its momentum simultaneously. The more accurately we measure the momentum of a particle, the less accurately we can determine its position at that time, and vice versa. This is summed up in what we now call the Heisenberg uncertainty principle: *It is fundamentally impossible to determine simultaneously and exactly both the momentum and the position of a particle*. For a particle of mass \(m\) moving with velocity v_x in the x direction (or equivalently with momentum p_x), the product of the uncertainty in the position, Δx, and the uncertainty in the momentum, Δp_x, must be greater than or equal to $\hbar/2$ (where $\hbar = h/2$, the value of Planck's constant divided by 2π).

$$\Delta x \times \Delta p_x = (\Delta x)(m\Delta v) \geq \frac{\hbar}{2}$$

This equation allows us to calculate the limit to how precisely we can know both the simultaneous position of an object and its momentum. For example, if we improve our measurement of an electron's position so that the uncertainty in the position (Δx) has a value of, say, 1 pm (10^{-12} m, about 1% of the diameter of a hydrogen atom), then our determination of its momentum must have an uncertainty with a value of at least

$$\left[\Delta p = m\Delta v = \frac{\hbar}{(2\Delta x)}\right] = \frac{\left(1.055 \times 10^{-34}\ \mathrm{kg\,m^2/s}\right)}{\left(2 \times 1 \times 10^{-12}\ \mathrm{m}\right)} = 5 \times 10^{-23}\ \mathrm{kg\,m/s}$$

The value of $\hbar$ is not large, so the uncertainty in the position or momentum of a macroscopic object like a baseball is too insignificant to observe. However, the mass of a microscopic object such as an electron is small enough that the uncertainty can be large and significant.

It should be noted that Heisenberg's uncertainty principle is not just limited to uncertainties in position and momentum, but it also links other dynamical variables. For example, when an atom absorbs a photon and makes a transition from one energy state to another, the uncertainty in the energy and the uncertainty in the time required for the transition are similarly related, as:

$$\Delta E \Delta t \geq \frac{\hbar}{2}$$

Heisenberg's principle imposes ultimate limits on what is knowable in science. The uncertainty principle can be shown to be a consequence of wave–particle duality, which lies at the heart of what distinguishes modern quantum theory from classical mechanics.

Link to Learning

Read this article that describes a recent macroscopic demonstration of the uncertainty principle applied to microscopic objects.

The Quantum–Mechanical Model of an Atom

Shortly after de Broglie published his ideas that the electron in a hydrogen atom could be better thought of as being a circular standing wave instead of a particle moving in quantized circular orbits, Erwin Schrödinger extended de Broglie's work by deriving what is today known as the Schrödinger equation. When Schrödinger applied his equation to hydrogen-like atoms, he was able to

reproduce Bohr's expression for the energy and, thus, the Rydberg formula governing hydrogen spectra. Schrödinger described electrons as three-dimensional stationary waves, or **wavefunctions**, represented by the Greek letter psi, ψ. A few years later, Max Born proposed an interpretation of the wavefunction ψ that is still accepted today: Electrons are still particles, and so the waves represented by ψ are not physical waves but, instead, are complex probability amplitudes. The square of the magnitude of a wavefunction $|\psi|^2$ describes the probability of the quantum particle being present near a certain location in space. This means that wavefunctions can be used to determine the distribution of the electron's density with respect to the nucleus in an atom. In the most general form, the Schrödinger equation can be written as:

$$\hat{H}\psi = E\psi$$

$\hat{H}$ is the Hamiltonian operator, a set of mathematical operations representing the total energy of the quantum particle (such as an electron in an atom), ψ is the wavefunction of this particle that can be used to find the special distribution of the probability of finding the particle, and E is the actual value of the total energy of the particle.

Schrödinger's work, as well as that of Heisenberg and many other scientists following in their footsteps, is generally referred to as **quantum mechanics**.

Link to Learning

You may also have heard of Schrödinger because of his famous thought experiment. This story explains the concepts of superposition and entanglement as related to a cat in a box with poison.

Understanding Quantum Theory of Electrons in Atoms

The goal of this section is to understand the electron orbitals (location of electrons in atoms), their different energies, and other properties. The use of quantum theory provides the best understanding to these topics. This knowledge is a precursor to chemical bonding.

As was described previously, electrons in atoms can exist only on discrete energy levels but not between them. It is said that the energy of an electron in an atom is quantized, that is, it can be equal only to certain specific values and can jump from one energy level to another but not transition smoothly or stay between these levels.

The energy levels are labeled with an n value, where $n = 1, 2, 3, \ldots$. Generally speaking, the energy of an electron in an atom is greater for greater values of n. This number, n, is referred to as the principal quantum number. The principal quantum number defines the location of the energy level. It is essentially the same concept as the n in the Bohr atom description. Another name for the principal quantum number is the shell number. The shells of an atom can be thought of concentric circles radiating out from the nucleus. The electrons that belong to a specific shell are most likely to be found within the corresponding circular area. The further we proceed from the nucleus, the higher the shell number, and so the higher the energy level (Figure 3.4.4). The positively charged protons in the nucleus stabilize the electronic orbitals by electrostatic attraction between the positive charges of the protons and the negative charges of the electrons. So the further away the electron is from the nucleus, the greater the energy it has.

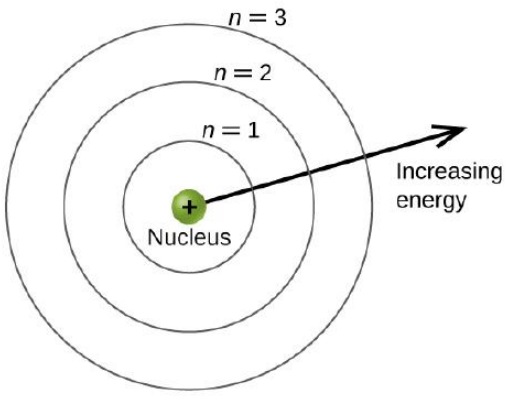


Figure 3.4.4: Different shells are numbered by principal quantum numbers.

This quantum mechanical model for where electrons reside in an atom can be used to look at electronic transitions, the events when an electron moves from one energy level to another. If the transition is to a higher energy level, energy is absorbed, and the energy

change has a positive value. To obtain the amount of energy necessary for the transition to a higher energy level, a photon is absorbed by the atom. A transition to a lower energy level involves a release of energy, and the energy change is negative. This process is accompanied by emission of a photon by the atom. The following equation summarizes these relationships and is based on the hydrogen atom:

$$\begin{aligned}\Delta E &= E_{\text{final}} - E_{\text{initial}} \\ &= -2.18 \times 10^{-18} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right) J\end{aligned}$$

The values n_f and n_i are the final and initial energy states of the electron.

The principal quantum number is one of three quantum numbers used to characterize an orbital. An **atomic orbital** is a general region in an atom within which an electron is most probable to reside. The quantum mechanical model specifies the probability of finding an electron in the three-dimensional space around the nucleus and is based on solutions of the Schrödinger equation. In addition, the principal quantum number defines the energy of an electron in a hydrogen or hydrogen-like atom or an ion (an atom or an ion with only one electron) and the general region in which discrete energy levels of electrons in a multi-electron atoms and ions are located.

Another quantum number is l, the **secondary (angular momentum) quantum number**. It is an integer that may take the values, $l = 0, 1, 2, \ldots, n-1$. This means that an orbital with $n = 1$ can have only one value of l, $l = 0$, whereas $n = 2$ permits $l = 0$ and $l = 1$, and so on. Whereas the principal quantum number, n, defines the general size and energy of the orbital, the secondary quantum number l specifies the shape of the orbital. Orbitals with the same value of l define a **subshell**.

Orbitals with $l = 0$ are called s orbitals and they make up the s subshells. The value $l = 1$ corresponds to the p orbitals. For a given n, p orbitals constitute a p subshell (e.g., 3p if n = 3). The orbitals with $l = 2$ are called the **d orbitals**, followed by the *f-*, *g-*, *and h-*orbitals for l = 3, 4, and 5.

There are certain distances from the nucleus at which the probability density of finding an electron located at a particular orbital is zero. In other words, the value of the wavefunction ψ is zero at this distance for this orbital. Such a value of radius r is called a **radial node**. The number of radial nodes in an orbital is $n-l-1$.

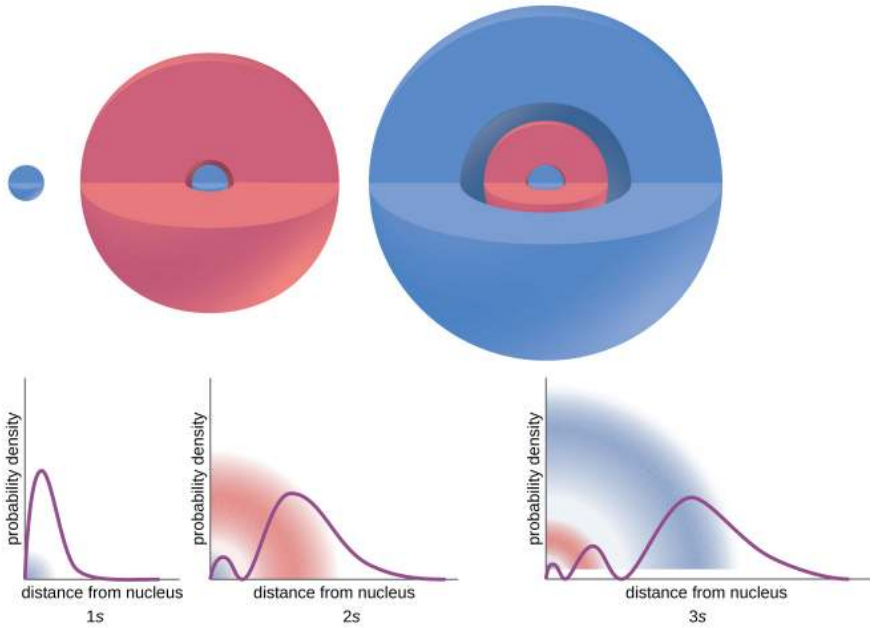


Figure 3.4.5: The graphs show the probability (*y* axis) of finding an electron for the 1*s*, 2*s*, 3*s* orbitals as a function of distance from the nucleus.

Consider the examples in Figure 3.4.5. The orbitals depicted are of the *s* type, thus $l = 0$ for all of them. It can be seen from the graphs of the probability densities that there are 1 – 0 – 1 = 0 places where the density is zero (nodes) for 1*s* (*n* = 1), 2 – 0 – 1 = 1 node for 2*s*, and 3 – 0 – 1 = 2 nodes for the 3*s* orbitals.

https://chem.libretexts.org/@go/page/532994

The *s* subshell electron density distribution is spherical and the *p* subshell has a dumbbell shape. The *d* and *f* orbitals are more complex. These shapes represent the three-dimensional regions within which the electron is likely to be found.

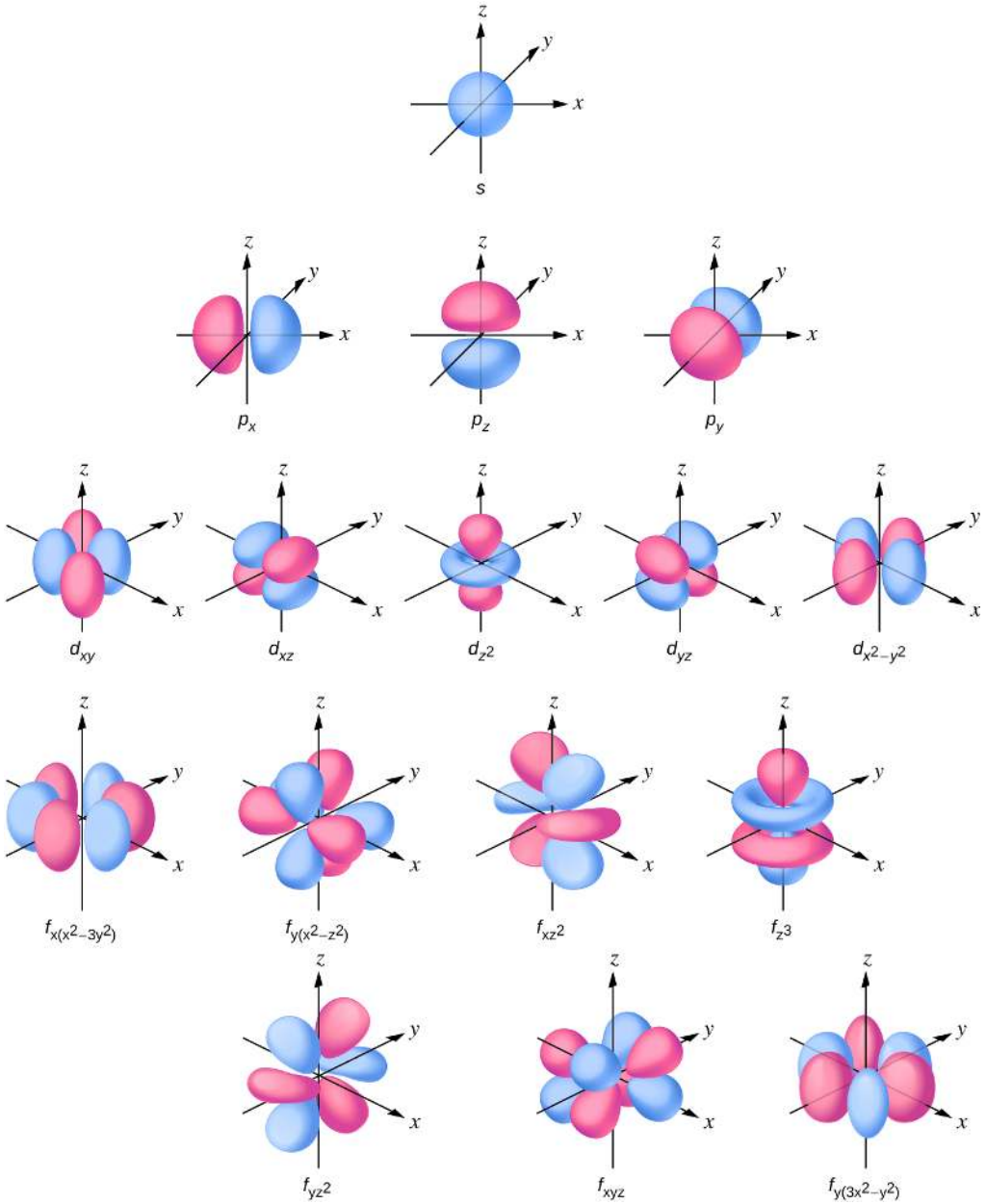


Figure 3.4.6: Shapes of *s*, *p*, *d*, and *f* orbitals.

The **magnetic quantum number**, m_l, specifies the relative spatial orientation of a particular orbital. Generally speaking, m_l are rangeed from $-l$ to $+l$, i.e.

$$-l, -(l-1), \ldots, 0, \ldots, (l-1), l\backslash).$$

The total number of possible orbitals with the same value of l (that is, in the same subshell) is $2l+1$. Thus, there is one s-orbital in an s subshell ($l=0$), there are three p-orbitals in a p subshell ($l=1$), five d-orbitals in a d subshell (l = 2), seven f-orbitals in an f subshell ($l=3$), and so forth. The principal quantum number defines the general value of the electronic energy. The angular momentum quantum number determines the shape of the orbital. And the magnetic quantum number specifies orientation of the orbital in space, as can be seen in Figure 3.4.6.

Figure 3.4.7 illustrates the energy levels for various orbitals. The number before the orbital name (such as 2*s*, 3*p*, and so forth) stands for the principal quantum number, n. The letter in the orbital name defines the subshell with a specific angular momentum

quantum number l = 0 for *s* orbitals, 1 for *p* orbitals, 2 for *d* orbitals. Finally, there are more than one possible orbitals for $l \geq 1$, each corresponding to a specific value of m_l. In the case of a hydrogen atom or a one-electron ion (such as He^+, Li^{2+}, and so on), energies of all the orbitals with the same n are the same. This is called a degeneracy, and the energy levels for the same principal quantum number, n, are called degenerate orbitals. However, in atoms with more than one electron, this degeneracy is eliminated by the electron–electron interactions, and orbitals that belong to different subshells have different energies, as shown on Figure 3.4.7. Orbitals within the same subshell are still degenerate and have the same energy.

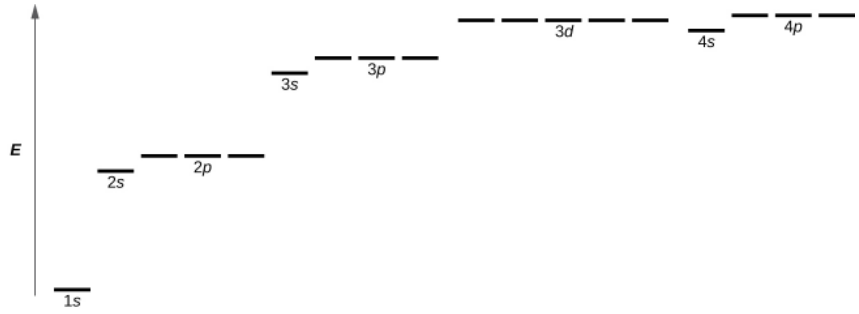


Figure 3.4.7: The chart shows the energies of electron orbitals in a multi-electron atom.

While the three quantum numbers discussed in the previous paragraphs work well for describing electron orbitals, some experiments showed that they were not sufficient to explain all observed results. It was demonstrated in the 1920s that when hydrogen-line spectra are examined at extremely high resolution, some lines are actually not single peaks but, rather, pairs of closely spaced lines. This is the so-called fine structure of the spectrum, and it implies that there are additional small differences in energies of electrons even when they are located in the same orbital. These observations led Samuel Goudsmit and George Uhlenbeck to propose that electrons have a fourth quantum number. They called this the **spin quantum number**, or m_s.

The other three quantum numbers, n, l, and m_l, are properties of specific atomic orbitals that also define in what part of the space an electron is most likely to be located. Orbitals are a result of solving the Schrödinger equation for electrons in atoms. The electron spin is a different kind of property. It is a completely quantum phenomenon with no analogues in the classical realm. In addition, it cannot be derived from solving the Schrödinger equation and is not related to the normal spatial coordinates (such as the Cartesian *x*, *y*, and *z*). Electron spin describes an intrinsic electron "rotation" or "spinning." Each electron acts as a tiny magnet or a tiny rotating object with an angular momentum, or as a loop with an electric current, even though this rotation or current cannot be observed in terms of spatial coordinates.

The magnitude of the overall electron spin can only have one value, and an electron can only “spin” in one of two quantized states. One is termed the α state, with the z component of the spin being in the positive direction of the z axis. This corresponds to the spin quantum number $m_s = +1/2$. The other is called the $\beta state$, with the z component of the spin being negative and $m_s = -1/2$. Any electron, regardless of the atomic orbital it is located in, can only have one of those two values of the spin quantum number. The energies of electrons having $m_s = +1/2$ and $m_s = -1/2$ are different if an external magnetic field is applied.

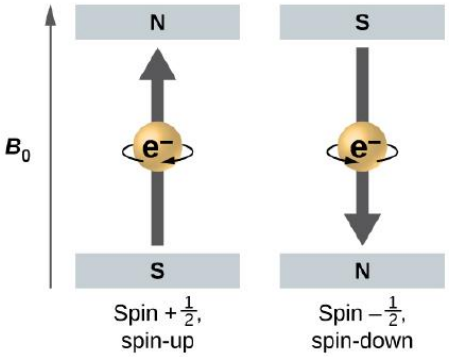


Figure 3.4.8: Electrons with spin values in an external magnetic field.

Figure 3.4.8 illustrates this phenomenon. An electron acts like a tiny magnet. Its moment is directed up (in the positive direction of the z axis) for the $m_s = +1/2$ spin quantum number and down (in the negative z direction) for the spin quantum number of $m_s = -1/2$. A magnet has a lower energy if its magnetic moment is aligned with the external magnetic field (the left electron on

Figure 3.4.8) and a higher energy for the magnetic moment being opposite to the applied field. This is why an electron with $m_s = +1/2$ has a slightly lower energy in an external field in the positive z direction, and an electron with $m_s = -1/2$ has a slightly higher energy in the same field. This is true even for an electron occupying the same orbital in an atom. A spectral line corresponding to a transition for electrons from the same orbital but with different spin quantum numbers has two possible values of energy; thus, the line in the spectrum will show a fine structure splitting.

The Pauli Exclusion Principle

An electron in an atom is completely described by four quantum numbers: n, l, m_l, and m_s. The first three quantum numbers define the orbital and the fourth quantum number describes the intrinsic electron property called spin. An Austrian physicist Wolfgang Pauli formulated a general principle that gives the last piece of information that we need to understand the general behavior of electrons in atoms. The **Pauli exclusion principle** can be formulated as follows: No two electrons in the same atom can have exactly the same set of all the four quantum numbers. What this means is that two electrons can share the same orbital (the same set of the quantum numbers n, l, and m_l) only if their spin quantum numbers m_s have different values. Since the spin quantum number can only have two values no more than two electrons can occupy the same orbital (and if two electrons are located in the same orbital, they must have opposite spins). Therefore, any atomic orbital can be populated by only zero, one, or two electrons. The properties and meaning of the quantum numbers of electrons in atoms are briefly summarized in Table 3.4.1.

Table 3.4.1: Quantum Numbers, Their Properties, and Significance

Name	Symbol	Allowed values	Physical meaning
principal quantum number	n	1, 2, 3, 4, ….	shell, the general region for the value of energy for an electron on the orbital
angular momentum or azimuthal quantum number	l	$0 \le l \le n-1$	subshell, the shape of the orbital
magnetic quantum number	m_l	$-l \le m_l \le l$	orientation of the orbital
spin quantum number	m_s	$+\frac{1}{2}$ and $-\frac{1}{2}$	direction of the intrinsic quantum “spinning” of the electron

Example 3.4.2: Working with Shells and Subshells

Indicate the number of subshells, the number of orbitals in each subshell, and the values of l and m_l for the orbitals in the $n = 4$ shell of an atom.

Solution

For $n = 4$, l can have values of 0, 1, 2, and 3. Thus, s, p, d, and f subshells are found in the n = 4 shell of an atom. For $l = 0$ (the s subshell), m_l can only be 0. Thus, there is only one 4s orbital. For $l = 1$ (p-type orbitals), m can have values of –1, 0, +1, so we find three 4p orbitals. For l = 2 (d-type orbitals), m_l can have values of –2, –1, 0, +1, +2, so we have five 4d orbitals. When $l = 3$ (f-type orbitals), m_l can have values of –3, –2, –1, 0, +1, +2, +3, and we can have seven 4f orbitals. Thus, we find a total of 16 orbitals in the $n = 4$ shell of an atom.

Exercise 3.4.2

Identify the subshell in which electrons with the following quantum numbers are found:

a. n = 3, l = 1
b. n = 5, l = 3
c. n = 2, l = 0

Answer

a. 3*p*
b. 5*f*

c. 2s

Example 3.4.3: Maximum Number of Electrons

Calculate the maximum number of electrons that can occupy a shell with (a) n = 2, (b) n = 5, and (c) n as a variable. Note you are only looking at the orbitals with the specified n value, not those at lower energies.

Solution

a. When $n = 2$, there are four orbitals (a single 2s orbital, and three orbitals labeled 2p). These four orbitals can contain eight electrons.

b. When $n = 2$, there are four orbitals (a single 2s orbital, and three orbitals labeled 2p). These four orbitals can contain eight electrons.
When $n = 5$, there are five subshells of orbitals that we need to sum:

$$\begin{aligned} &1 \text{ orbital labeled 5s} \\ &3 \text{ orbitals labeled 5p} \\ &5 \text{ orbitals labeled 5d} \\ &7 \text{ orbitals labeled 5f} \\ +\ &9 \text{ orbitals labeled 5g} \\ \hline &25 \text{ orbitals total} \end{aligned}$$

Again, each orbital holds two electrons, so 50 electrons can fit in this shell.

c. The number of orbitals in any shell n will equal n^2. There can be up to two electrons in each orbital, so the maximum number of electrons will be $2 \times n^2$.

Exercise 3.4.3

If a shell contains a maximum of 32 electrons, what is the principal quantum number, n?

Answer

$n = 4$

Example 3.4.4: Working with Quantum Numbers

Complete the following table for atomic orbitals:

Orbital	n	l	m_l degeneracy	Radial nodes (#)
4*f*				
	4	1		
	7		7	3
5*d*				

Solution

The table can be completed using the following rules:

- The orbital designation is nl, where $l = 0, 1, 2, 3, 4, 5, \ldots$ is mapped to the letter sequence s, p, d, f, g, h, …,
- The m_l degeneracy is the number of orbitals within an l subshell, and so is $2l + 1$ (there is one s orbital, three p orbitals, five d orbitals, seven f orbitals, and so forth).
- The number of radial nodes is equal to $n - l - 1$.

Orbital	n	l	m_l degeneracy	Radial nodes (#)
4*f*	4	3	7	0
4*p*	4	1	3	2
7*f*	7	3	7	3
5*d*	5	2	5	2

Exercise 3.4.1

How many orbitals have $l = 2$ and $n = 3$?

Answer

The five degenerate 3*d* orbitals

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3.5: Electronic Structure of Atoms (Electron Configurations)

Learning Objectives

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

- Derive the predicted ground-state electron configurations of atoms
- Identify and explain exceptions to predicted electron configurations for atoms and ions
- Relate electron configurations to element classifications in the periodic table

Having introduced the basics of atomic structure and quantum mechanics, we can use our understanding of quantum numbers to determine how atomic orbitals relate to one another. This allows us to determine which orbitals are occupied by electrons in each atom. The specific arrangement of electrons in orbitals of an atom determines many of the chemical properties of that atom.

Orbital Energies and Atomic Structure

The energy of atomic orbitals increases as the principal quantum number, n, increases. In any atom with two or more electrons, the repulsion between the electrons makes energies of subshells with different values of l differ so that the energy of the orbitals increases within a shell in the order $s < p < d < f$. Figure 3.5.1: depicts how these two trends in increasing energy relate. The 1s orbital at the bottom of the diagram is the orbital with electrons of lowest energy. The energy increases as we move up to the 2s and then 2p, 3s, and 3p orbitals, showing that the increasing n value has more influence on energy than the increasing l value for small atoms. However, this pattern does not hold for larger atoms. The 3d orbital is higher in energy than the 4s orbital. Such overlaps continue to occur frequently as we move up the chart.

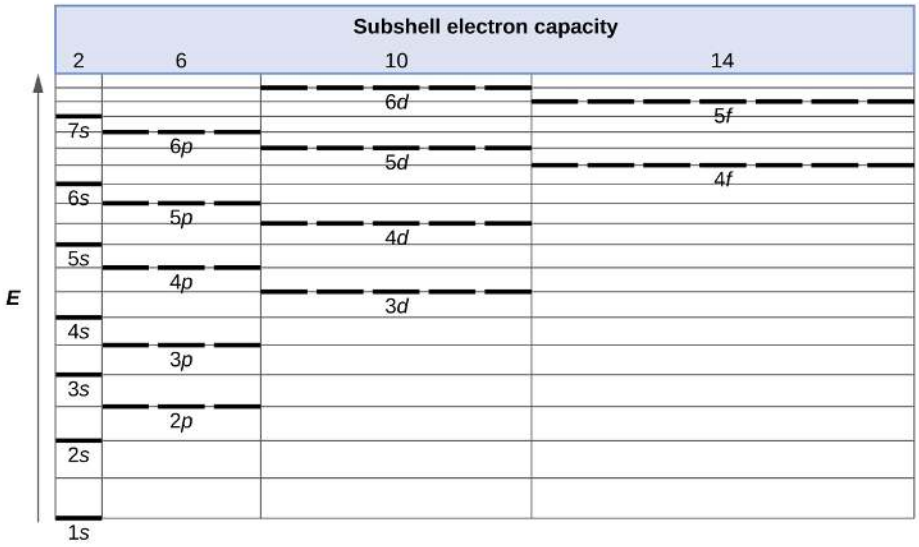


Figure 3.5.1: Generalized energy-level diagram for atomic orbitals in an atom with two or more electrons (not to scale).

Electrons in successive atoms on the periodic table tend to fill low-energy orbitals first. Thus, many students find it confusing that, for example, the 5p orbitals fill immediately after the 4d, and immediately before the 6s. The filling order is based on observed experimental results, and has been confirmed by theoretical calculations. As the principal quantum number, n, increases, the size of the orbital increases and the electrons spend more time farther from the nucleus. Thus, the attraction to the nucleus is weaker and the energy associated with the orbital is higher (less stabilized). But this is not the only effect we have to take into account. Within each shell, as the value of l increases, the electrons are less penetrating (meaning there is less electron density found close to the nucleus), in the order $s > p > d > f$. Electrons that are closer to the nucleus slightly repel electrons that are farther out, offsetting the more dominant electron–nucleus attractions slightly (recall that all electrons have −1 charges, but nuclei have +Z charges). This phenomenon is called shielding and will be discussed in more detail in the next section. Electrons in orbitals that experience more shielding are less stabilized and thus higher in energy. For small orbitals (1s through 3p), the increase in energy due to n is more significant than the increase due to l; however, for larger orbitals the two trends are comparable and cannot be simply predicted. We will discuss methods for remembering the observed order.

The arrangement of electrons in the orbitals of an atom is called the **electron configuration** of the atom. We describe an electron configuration with a symbol that contains three pieces of information (Figure 3.5.2):

1. The number of the principal quantum shell, n,
2. The letter that designates the orbital type (the subshell, l), and
3. A superscript number that designates the number of electrons in that particular subshell.

For example, the notation $2p^4$ (read "two–p–four") indicates four electrons in a *p* subshell ($l = 1$) with a principal quantum number (n) of 2. The notation $3d^8$ (read "three–d–eight") indicates eight electrons in the *d* subshell (i.e., $l = 2$) of the principal shell for which $n = 3$.

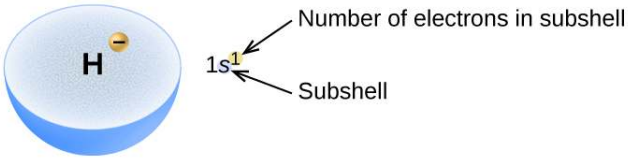


Figure 3.5.2: The diagram of an electron configuration specifies the subshell (*n* and *l* value, with letter symbol) and superscript number of electrons.

The Aufbau Principle

To determine the electron configuration for any particular atom, we can “build” the structures in the order of atomic numbers. Beginning with hydrogen, and continuing across the periods of the periodic table, we add one proton at a time to the nucleus and one electron to the proper subshell until we have described the electron configurations of all the elements. This procedure is called the **Aufbau principle**, from the German word *Aufbau* (“to build up”). Each added electron occupies the subshell of lowest energy available (in the order shown in Figure 3.5.1), subject to the limitations imposed by the allowed quantum numbers according to the Pauli exclusion principle. Electrons enter higher-energy subshells only after lower-energy subshells have been filled to capacity. Figure 3.5.3 illustrates the traditional way to remember the filling order for atomic orbitals.

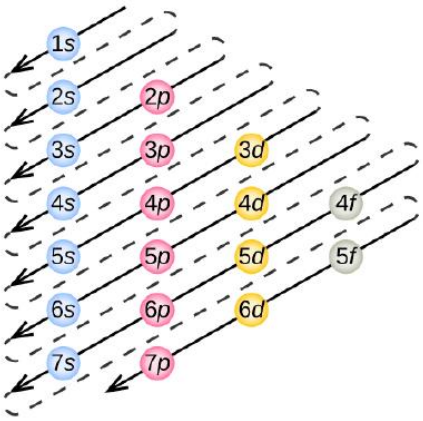


Figure 3.5.3: This diagram depicts the energy order for atomic orbitals and is useful for deriving ground-state electron configurations.

Since the arrangement of the periodic table is based on the electron configurations, Figure 3.5.4 provides an alternative method for determining the electron configuration. The filling order simply begins at hydrogen and includes each subshell as you proceed in increasing *Z* order. For example, after filling the 3*p* block up to Ar, we see the orbital will be 4s (K, Ca), followed by the 3*d* orbitals.

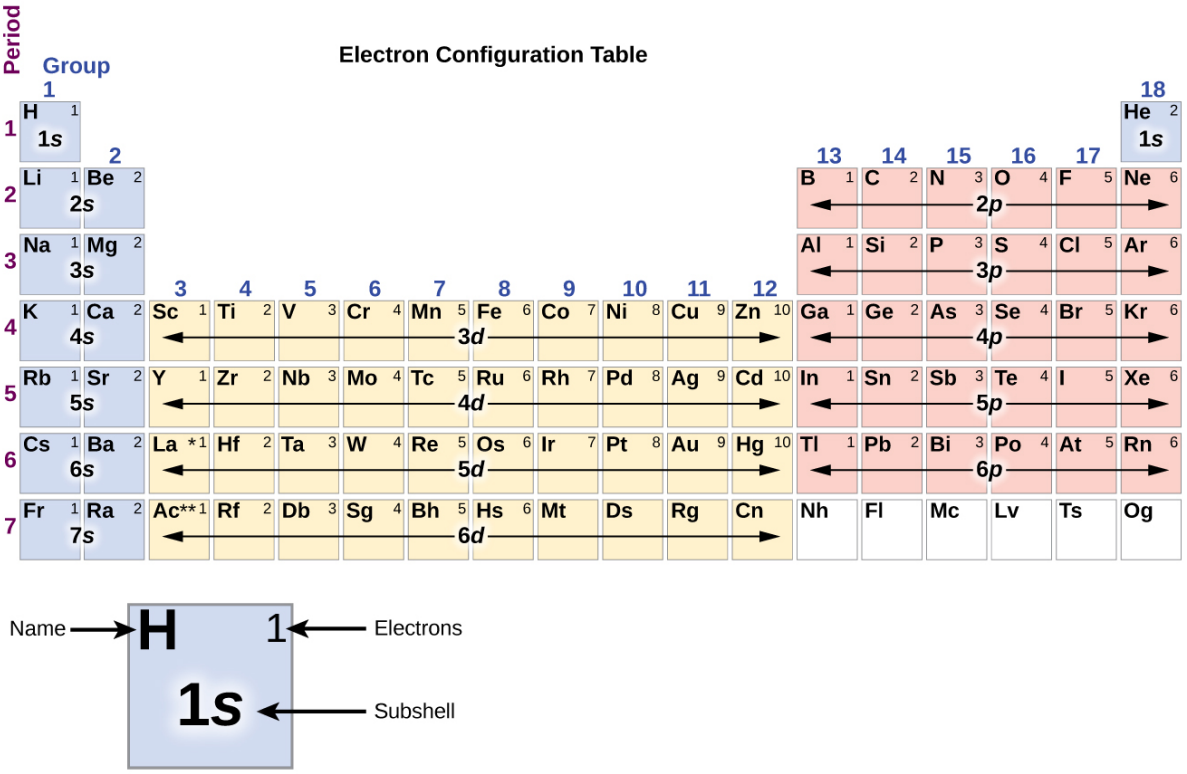
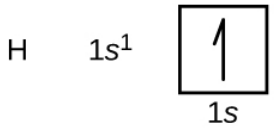
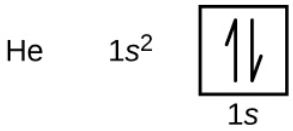


Figure 3.5.4: This partial periodic table shows electron configurations for the valence subshells of atoms. By "building up" from hydrogen, this table can be used to determine the electron configuration for atoms of most elements in the periodic table. (Electron configurations of the lanthanides and actinides are not accurately predicted by this simple approach. See Figure 3.5.6:

We will now construct the ground-state electron configuration and orbital diagram for a selection of atoms in the first and second periods of the periodic table. **Orbital diagrams** are pictorial representations of the electron configuration, showing the individual orbitals and the pairing arrangement of electrons. We start with a single hydrogen atom (atomic number 1), which consists of one proton and one electron. Referring to Figure 3.5.3 or Figure 3.5.4, we would expect to find the electron in the 1*s* orbital. By convention, the $m_s = +\frac{1}{2}$ value is usually filled first. The electron configuration and the orbital diagram are:



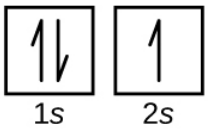
Following hydrogen is the noble gas helium, which has an atomic number of 2. The helium atom contains two protons and two electrons. The first electron has the same four quantum numbers as the hydrogen atom electron (\(n = 1\), \(l = 0\), \(m_l = 0\), $m_s = +\frac{1}{2}$. The second electron also goes into the 1*s* orbital and fills that orbital. The second electron has the same n, l, and ml quantum numbers, but must have the opposite spin quantum number, $m_s = -\frac{1}{2}$. This is in accord with the Pauli exclusion principle: No two electrons in the same atom can have the same set of four quantum numbers. For orbital diagrams, this means two arrows go in each box (representing two electrons in each orbital) and the arrows must point in opposite directions (representing paired spins). The electron configuration and orbital diagram of helium are:



The $n = 1$ shell is completely filled in a helium atom.

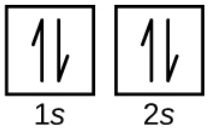
The next atom is the alkali metal lithium with an atomic number of 3. The first two electrons in lithium fill the 1*s* orbital and have the same sets of four quantum numbers as the two electrons in helium. The remaining electron must occupy the orbital of next lowest energy, the 2*s* orbital (Figure 3.5.3 or Figure 3.5.4). Thus, the electron configuration and orbital diagram of lithium are:

Li $1s^2 2s^1$



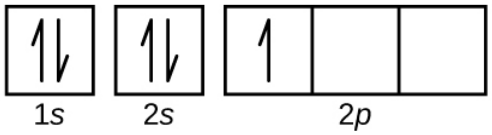
An atom of the alkaline earth metal beryllium, with an atomic number of 4, contains four protons in the nucleus and four electrons surrounding the nucleus. The fourth electron fills the remaining space in the 2*s* orbital.

Be $1s^2 2s^2$



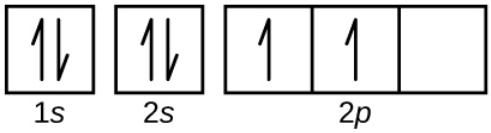
An atom of boron (atomic number 5) contains five electrons. The $n = 1$ shell is filled with two electrons and three electrons will occupy the $n = 2$ shell. Because any *s* subshell can contain only two electrons, the fifth electron must occupy the next energy level, which will be a 2*p* orbital. There are three degenerate 2*p* orbitals ($m_l = -1, 0, +1$) and the electron can occupy any one of these *p* orbitals. When drawing orbital diagrams, we include empty boxes to depict any empty orbitals in the same subshell that we are filling.

B $1s^2 2s^2 2p^1$

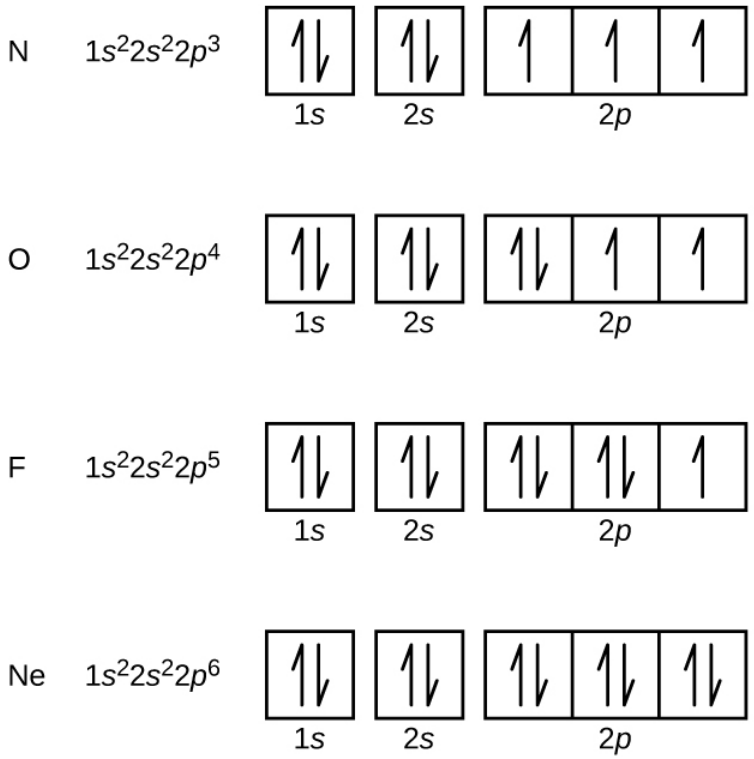


Carbon (atomic number 6) has six electrons. Four of them fill the 1*s* and 2*s* orbitals. The remaining two electrons occupy the 2*p* subshell. We now have a choice of filling one of the 2*p* orbitals and pairing the electrons or of leaving the electrons unpaired in two different, but degenerate, *p* orbitals. The orbitals are filled as described by **Hund's rule**: the lowest-energy configuration for an atom with electrons within a set of degenerate orbitals is that having the maximum number of unpaired electrons. Thus, the two electrons in the carbon 2*p* orbitals have identical n, l, and m_s quantum numbers and differ in their m_l quantum number (in accord with the Pauli exclusion principle). The electron configuration and orbital diagram for carbon are:

C $1s^2 2s^2 2p^2$



Nitrogen (atomic number 7) fills the 1*s* and 2*s* subshells and has one electron in each of the three 2*p* orbitals, in accordance with Hund's rule. These three electrons have unpaired spins. Oxygen (atomic number 8) has a pair of electrons in any one of the 2*p* orbitals (the electrons have opposite spins) and a single electron in each of the other two. Fluorine (atomic number 9) has only one 2*p* orbital containing an unpaired electron. All of the electrons in the noble gas neon (atomic number 10) are paired, and all of the orbitals in the $n = 1$ and the $n = 2$ shells are filled. The electron configurations and orbital diagrams of these four elements are:



The alkali metal sodium (atomic number 11) has one more electron than the neon atom. This electron must go into the lowest-energy subshell available, the 3*s* orbital, giving a $1s^2 2s^2 2p^6 3s^1$ configuration. The electrons occupying the outermost shell orbital(s) (highest value of *n*) are called **valence electrons**, and those occupying the inner shell orbitals are called **core electrons** (Figure 3.5.5). Since the core electron shells correspond to noble gas electron configurations, we can abbreviate electron configurations by writing the noble gas that matches the core electron configuration, along with the valence electrons in a condensed format. For our sodium example, the symbol [Ne] represents core electrons, ($1s^2 2s^2 2p^6$) and our abbreviated or condensed configuration is $[Ne]3s^1$.

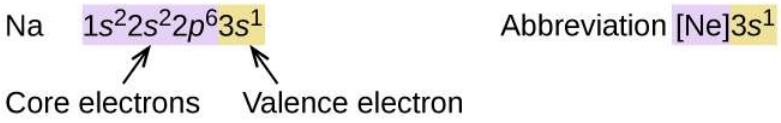


Figure 3.5.5: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.

Similarly, the abbreviated configuration of lithium can be represented as $[He]2s^1$, where [He] represents the configuration of the helium atom, which is identical to that of the filled inner shell of lithium. Writing the configurations in this way emphasizes the similarity of the configurations of lithium and sodium. Both atoms, which are in the alkali metal family, have only one electron in a valence *s* subshell outside a filled set of inner shells.

$$Li : [He]2s^1$$

$$Na : [Ne]3s^1$$

The alkaline earth metal magnesium (atomic number 12), with its 12 electrons in a $[Ne]3s^2$ configuration, is analogous to its family member beryllium, $[He]2s^2$. Both atoms have a filled *s* subshell outside their filled inner shells. Aluminum (atomic number 13), with 13 electrons and the electron configuration $[Ne]3s^2 3p^1$, is analogous to its family member boron, $[He]2s^2 2p^1$.

The electron configurations of silicon (14 electrons), phosphorus (15 electrons), sulfur (16 electrons), chlorine (17 electrons), and argon (18 electrons) are analogous in the electron configurations of their outer shells to their corresponding family members carbon, nitrogen, oxygen, fluorine, and neon, respectively, except that the principal quantum number of the outer shell of the heavier elements has increased by one to $n = 3$. Figure 3.5.6: shows the lowest energy, or ground-state, electron configuration for these elements as well as that for atoms of each of the known elements.

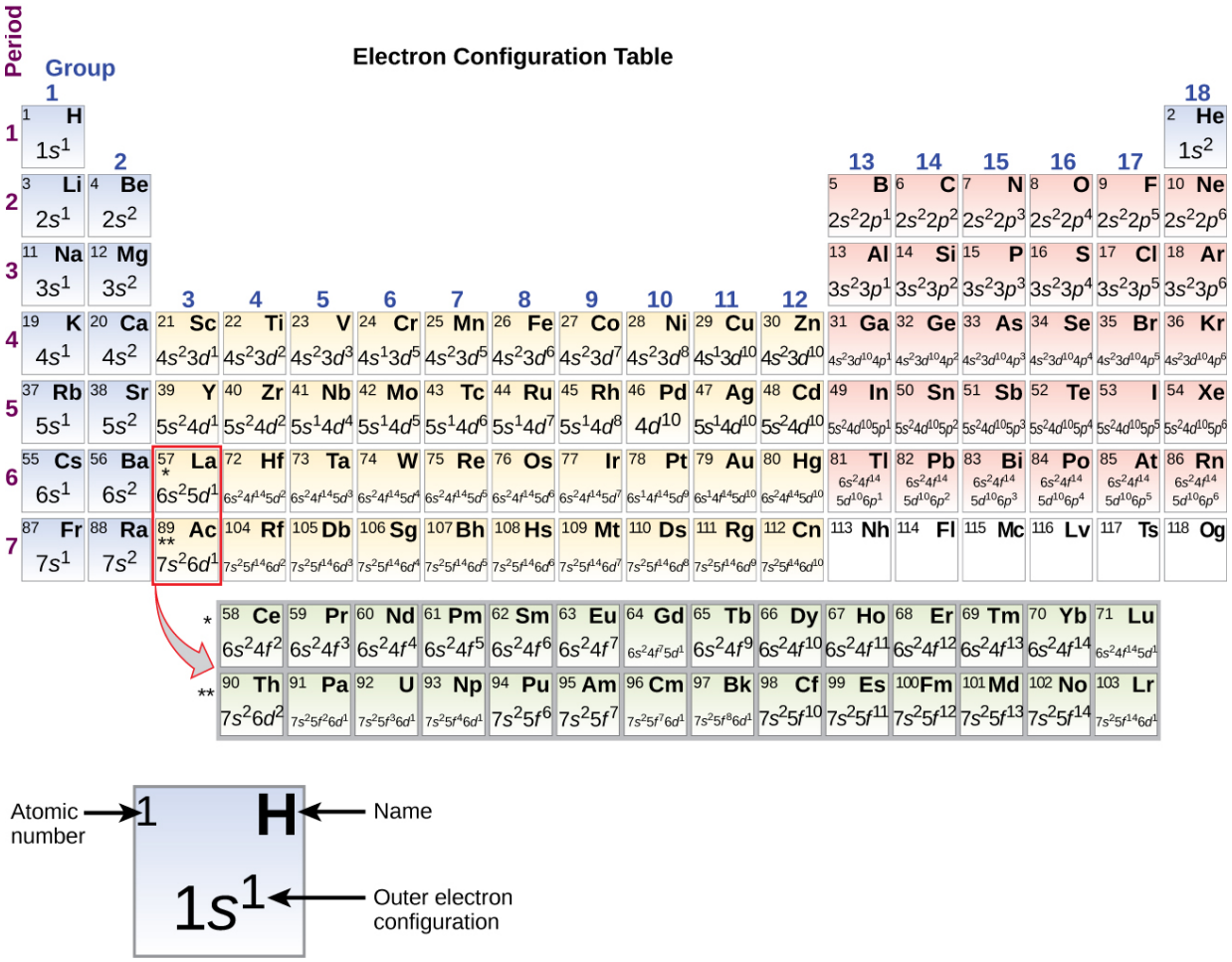


Figure 3.5.6: This version of the periodic table shows the outer-shell electron configuration of each element. Note that down each group, the configuration is often similar.

When we come to the next element in the periodic table, the alkali metal potassium (atomic number 19), we might expect that we would begin to add electrons to the 3*d* subshell. However, all available chemical and physical evidence indicates that potassium is like lithium and sodium, and that the next electron is not added to the 3*d* level but is, instead, added to the 4*s* level (Figure 3.5.6). As discussed previously, the 3*d* orbital with no radial nodes is higher in energy because it is less penetrating and more shielded from the nucleus than the 4*s*, which has three radial nodes. Thus, potassium has an electron configuration of $[Ar]4s^1$. Hence, potassium corresponds to Li and Na in its valence shell configuration. The next electron is added to complete the 4*s* subshell and calcium has an electron configuration of $[Ar]4s^2$. This gives calcium an outer-shell electron configuration corresponding to that of beryllium and magnesium.

Beginning with the transition metal scandium (atomic number 21), additional electrons are added successively to the 3*d* subshell. This subshell is filled to its capacity with 10 electrons (remember that for $l = 2$ [*d* orbitals], there are $2l + 1 = 5$ values of m_l, meaning that there are five *d* orbitals that have a combined capacity of 10 electrons). The 4*p* subshell fills next. Note that for three series of elements, scandium (Sc) through copper (Cu), yttrium (Y) through silver (Ag), and lutetium (Lu) through gold (Au), a total of 10 *d* electrons are successively added to the $(n - 1)$ shell next to the *n* shell to bring that $(n - 1)$ shell from 8 to 18 electrons. For two series, lanthanum (La) through lutetium (Lu) and actinium (Ac) through lawrencium (Lr), 14 *f* electrons ($l = 3$, $2l + 1 = 7$ m_l values; thus, seven orbitals with a combined capacity of 14 electrons) are successively added to the $(n - 2)$ shell to bring that shell from 18 electrons to a total of 32 electrons.

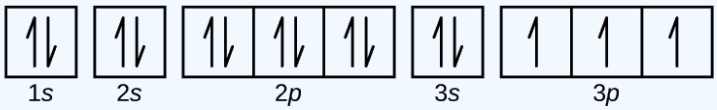
Example 3.5.1: Quantum Numbers and Electron Configurations

What is the electron configuration and orbital diagram for a phosphorus atom? What are the four quantum numbers for the last electron added?

Solution

The atomic number of phosphorus is 15. Thus, a phosphorus atom contains 15 electrons. The order of filling of the energy levels is 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, . . . The 15 electrons of the phosphorus atom will fill up to the 3*p* orbital, which will contain three electrons:

P $1s^2 2s^2 2p^6 3s^2 3p^3$



The last electron added is a 3*p* electron. Therefore, $n = 3$ and, for a *p*-type orbital, $l = 1$. The m_l value could be –1, 0, or +1. The three *p* orbitals are degenerate, so any of these m_l values is correct. For unpaired electrons, convention assigns the value of $+\frac{1}{2}$ for the spin quantum number; thus, $m_s = +\frac{1}{2}$

? Exercise 3.5.1

Identify the atoms from the electron configurations given:

a. $[Ar]4s^2 3d^5$
b. $[Kr]5s^2 4d^{10} 5p^6$

Answer

a. Mn
b. Xe

The periodic table can be a powerful tool in predicting the electron configuration of an element. However, we do find exceptions to the order of filling of orbitals that are shown in Figure 3.5.3: or Figure 3.5.4. For instance, the electron configurations (shown in Figure 3.5.6) of the transition metals chromium (Cr; atomic number 24) and copper (Cu; atomic number 29), among others, are not those we would expect. In general, such exceptions involve subshells with very similar energy, and small effects can lead to changes in the order of filling.

In the case of Cr and Cu, we find that half-filled and completely filled subshells apparently represent conditions of preferred stability. This stability is such that an electron shifts from the 4*s* into the 3*d* orbital to gain the extra stability of a half-filled 3*d* subshell (in Cr) or a filled 3*d* subshell (in Cu). Other exceptions also occur. For example, niobium (Nb, atomic number 41) is predicted to have the electron configuration $[Kr]5s^2 4d^3$. Experimentally, we observe that its ground-state electron configuration is actually $[Kr]5s^1 4d^4$. We can rationalize this observation by saying that the electron–electron repulsions experienced by pairing the electrons in the 5*s* orbital are larger than the gap in energy between the 5*s* and 4*d* orbitals. There is no simple method to predict the exceptions for atoms where the magnitude of the repulsions between electrons is greater than the small differences in energy between subshells.

Electron Configurations and the Periodic Table

As described earlier, the periodic table arranges atoms based on increasing atomic number so that elements with the same chemical properties recur periodically. When their electron configurations are added to the table (Figure 3.5.6), we also see a periodic recurrence of similar electron configurations in the outer shells of these elements. Because they are in the outer shells of an atom, valence electrons play the most important role in chemical reactions. The outer electrons have the highest energy of the electrons in an atom and are more easily lost or shared than the core electrons. Valence electrons are also the determining factor in some physical properties of the elements.

Elements in any one group (or column) have the same number of valence electrons; the alkali metals lithium and sodium each have only one valence electron, the alkaline earth metals beryllium and magnesium each have two, and the halogens fluorine and chlorine each have seven valence electrons. The similarity in chemical properties among elements of the same group occurs because they have the same number of valence electrons. It is the loss, gain, or sharing of valence electrons that defines how elements react.

It is important to remember that the periodic table was developed on the basis of the chemical behavior of the elements, well before any idea of their atomic structure was available. Now we can understand why the periodic table has the arrangement it has—the

arrangement puts elements whose atoms have the same number of valence electrons in the same group. This arrangement is emphasized in Figure 3.5.6: , which shows in periodic-table form the electron configuration of the last subshell to be filled by the Aufbau principle. The colored sections of Figure 3.5.6: show the three categories of elements classified by the orbitals being filled: main group, transition, and inner transition elements. These classifications determine which orbitals are counted in the **valence shell**, or highest energy level orbitals of an atom.

1. **Main group elements** (sometimes called **representative elements**) are those in which the last electron added enters an *s* or a *p* orbital in the outermost shell, shown in blue and red in Figure 3.5.6. This category includes all the nonmetallic elements, as well as many metals and the metalloids. The valence electrons for main group elements are those with the highest *n* level. For example, gallium (Ga, atomic number 31) has the electron configuration [Ar]$\underline{4s}^2 3d^{10} \underline{4p}^1$, which contains three valence electrons (underlined). The completely filled *d* orbitals count as core, not valence, electrons.
2. **Transition elements or transition metals**. These are metallic elements in which the last electron added enters a *d* orbital. The valence electrons (those added after the last noble gas configuration) in these elements include the *ns* and $(n - 1)$ *d* electrons. The official IUPAC definition of transition elements specifies those with partially filled *d* orbitals. Thus, the elements with completely filled orbitals (Zn, Cd, Hg, as well as Cu, Ag, and Au in Figure 3.5.6) are not technically transition elements. However, the term is frequently used to refer to the entire *d* block (colored yellow in Figure 3.5.6), and we will adopt this usage in this textbook.
3. **Inner transition elements** are metallic elements in which the last electron added occupies an *f* orbital. They are shown in green in Figure 3.5.6. The valence shells of the inner transition elements consist of the $(n - 2)f$, the $(n - 1)d$, and the *ns* subshells. There are two inner transition series:
    a. The lanthanide series: lanthanum (La) through lutetium (Lu)
    b. The actinide series: actinium (Ac) through lawrencium (Lr)

Lanthanum and actinium, because of their similarities to the other members of the series, are included and used to name the series, even though they are transition metals with no *f* electrons.

Electron Configurations of Ions

Ions are formed when atoms gain or lose electrons. A cation (positively charged ion) forms when one or more electrons are removed from a parent atom. For main group elements, the electrons that were added last are the first electrons removed. For transition metals and inner transition metals, however, electrons in the *s* orbital are easier to remove than the *d* or *f* electrons, and so the highest *ns* electrons are lost, and then the $(n - 1)d$ or $(n - 2)f$ electrons are removed. An anion (negatively charged ion) forms when one or more electrons are added to a parent atom. The added electrons fill in the order predicted by the Aufbau principle.

✔ Example 3.5.2: Predicting Electron Configurations of Ions

What is the electron configuration of:

a. Na^+
b. P^{3-}
c. Al^{2+}
d. Fe^{2+}
e. Sm^{3+}

Solution

First, write out the electron configuration for each parent atom. We have chosen to show the full, unabbreviated configurations to provide more practice for students who want it, but listing the core-abbreviated electron configurations is also acceptable.

Next, determine whether an electron is gained or lost. Remember electrons are negatively charged, so ions with a positive charge have *lost* an electron. For main group elements, the last orbital gains or loses the electron. For transition metals, the last *s* orbital loses an electron before the *d* orbitals.

a. Na: $1s^2 2s^2 2p^6 3s^1$. Sodium cation loses one electron, so Na^+: $1s^2 2s^2 2p^6 3s^1 = Na^+$: $1s^2 2s^2 2p^6$.
b. P: $1s^2 2s^2 2p^6 3s^2 3p^3$. Phosphorus trianion gains three electrons, so P^{3-}: $1s^2 2s^2 2p^6 3s^2 3p^6$.
c. Al: $1s^2 2s^2 2p^6 3s^2 3p^1$. Aluminum dication loses two electrons Al^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^1 = Al^{2+}$: $1s^2 2s^2 2p^6 3s^1$.
d. Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. Iron(II) loses two electrons and, since it is a transition metal, they are removed from the 4*s* orbital Fe^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6 = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$.

e. Sm: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^6$. Samarium trication loses three electrons. The first two will be lost from the 6*s* orbital, and the final one is removed from the 4*f* orbital. Sm^{3+}: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^6 = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 4f^5$.

? Exercise 3.5.21

Which ion with a +2 charge has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^5$? Which ion with a +3 charge has this configuration?

Answer

Tc^{2+}, Ru^{3+}

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3.6: Periodic Variations in Element Properties

Learning Objectives

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

- Describe and explain the observed trends in atomic size, ionization energy, and electron affinity of the elements

The elements in groups (vertical columns) of the periodic table exhibit similar chemical behavior. This similarity occurs because the members of a group have the same number and distribution of electrons in their valence shells. However, there are also other patterns in chemical properties on the periodic table. For example, as we move down a group, the metallic character of the atoms increases. Oxygen, at the top of group 16 (6A), is a colorless gas; in the middle of the group, selenium is a semiconducting solid; and, toward the bottom, polonium is a silver-grey solid that conducts electricity.

As we go across a period from left to right, we add a proton to the nucleus and an electron to the valence shell with each successive element. As we go down the elements in a group, the number of electrons in the valence shell remains constant, but the principal quantum number increases by one each time. An understanding of the electronic structure of the elements allows us to examine some of the properties that govern their chemical behavior. These properties vary periodically as the electronic structure of the elements changes. They are (1) size (radius) of atoms and ions, (2) ionization energies, and (3) electron affinities.

Link to Learning

Explore visualizations of the periodic trends discussed in this section (and many more trends). With just a few clicks, you can create three-dimensional versions of the periodic table showing atomic size or graphs of ionization energies from all measured elements.

Variation in Covalent Radius

The quantum mechanical picture makes it difficult to establish a definite size of an atom. However, there are several practical ways to define the radius of atoms and, thus, to determine their relative sizes that give roughly similar values. We will use the covalent radius (Figure 3.6.1), which is defined as one-half the distance between the nuclei of two identical atoms when they are joined by a covalent bond (this measurement is possible because atoms within molecules still retain much of their atomic identity). We know that as we scan down a group, the principal quantum number, *n*, increases by one for each element. Thus, the electrons are being added to a region of space that is increasingly distant from the nucleus. Consequently, the size of the atom (and its covalent radius) must increase as we increase the distance of the outermost electrons from the nucleus. This trend is illustrated for the covalent radii of the halogens in Table 3.6.1 and Figure 3.6.1. The trends for the entire periodic table can be seen in Figure 3.6.2.

Table 3.6.1: Covalent Radii of the Halogen Group Elements

Atom	Covalent radius (pm)	Nuclear charge
F	64	+9
Cl	99	+17
Br	114	+35
I	133	+53
At	148	+85

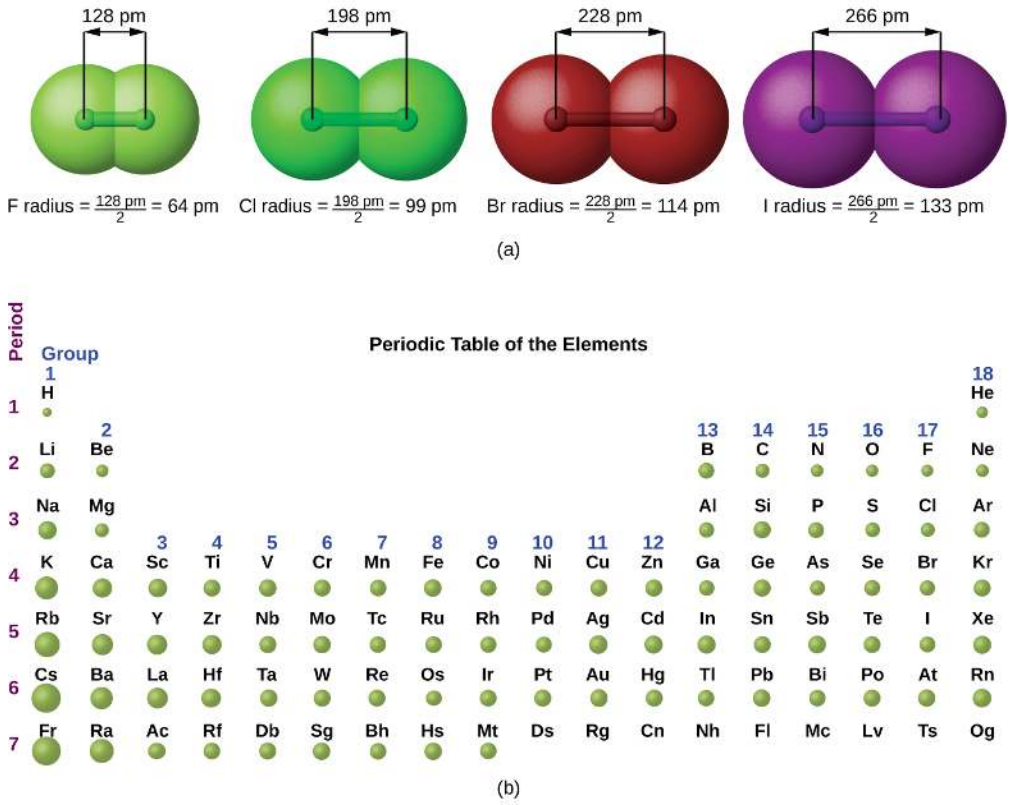


Figure 3.6.1: (a) The radius of an atom is defined as one-half the distance between the nuclei in a molecule consisting of two identical atoms joined by a covalent bond. The atomic radius for the halogens increases down the group as *n* increases. (b) Covalent radii of the elements are shown to scale. The general trend is that radii increase down a group and decrease across a period.

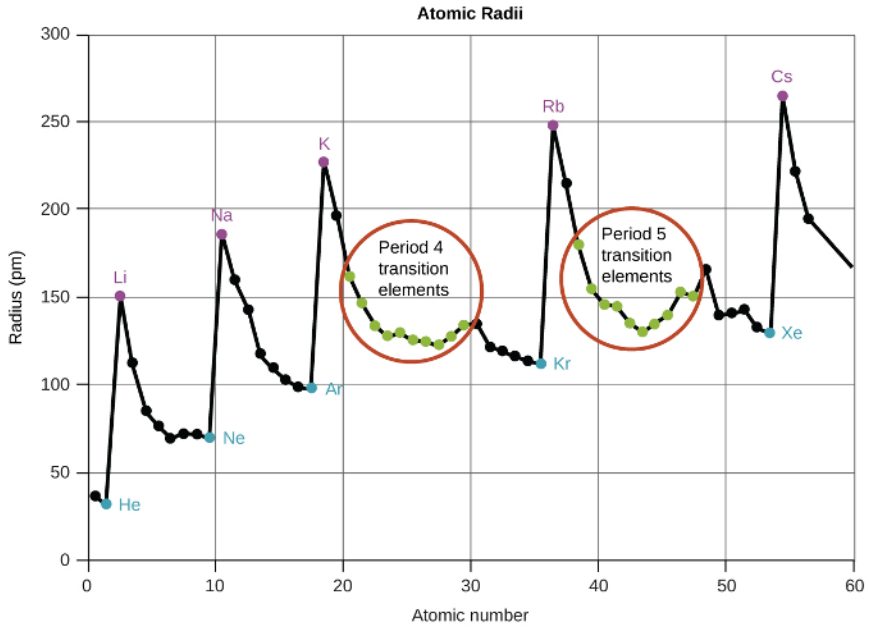


Figure 3.6.2: Within each period, the trend in atomic radius decreases as *Z* increases; for example, from K to Kr. Within each group (e.g., the alkali metals shown in purple), the trend is that atomic radius increases as *Z* increases.

As shown in Figure 3.6.2, as we move across a period from left to right, we generally find that each element has a smaller covalent radius than the element preceding it. This might seem counterintuitive because it implies that atoms with more electrons have a smaller atomic radius. This can be explained with the concept of **effective nuclear charge**, Z_{eff}. This is the pull exerted on a

specific electron by the nucleus, taking into account any electron–electron repulsions. For hydrogen, there is only one electron and so the nuclear charge (Z) and the effective nuclear charge (Z_{eff}) are equal. For all other atoms, the inner electrons partially *shield* the outer electrons from the pull of the nucleus, and thus:

$$Z_{\text{eff}} = Z - \text{ shielding}$$

Shielding is determined by the probability of another electron being between the electron of interest and the nucleus, as well as by the electron–electron repulsions the electron of interest encounters. Core electrons are adept at shielding, while electrons in the same valence shell do not block the nuclear attraction experienced by each other as efficiently. Thus, each time we move from one element to the next across a period, Z increases by one, but the shielding increases only slightly. Thus, Z_{eff} increases as we move from left to right across a period. The stronger pull (higher effective nuclear charge) experienced by electrons on the right side of the periodic table draws them closer to the nucleus, making the covalent radii smaller.

Thus, as we would expect, the outermost or valence electrons are easiest to remove because they have the highest energies, are shielded more, and are farthest from the nucleus. As a general rule, when the representative elements form cations, they do so by the loss of the *ns* or *np* electrons that were added last in the Aufbau process. The transition elements, on the other hand, lose the *ns* electrons before they begin to lose the $(n - 1)d$ electrons, even though the *ns* electrons are added first, according to the Aufbau principle.

Example 3.6.1: Sorting Atomic Radii

Predict the order of increasing covalent radius for Ge, Fl, Br, Kr.

Solution

Radius increases as we move down a group, so Ge < Fl (Note: Fl is the symbol for flerovium, element 114, NOT fluorine). Radius decreases as we move across a period, so Kr < Br < Ge. Putting the trends together, we obtain Kr < Br < Ge < Fl.

Exercise 3.6.1

Give an example of an atom whose size is smaller than fluorine.

Answer

Ne or He

Variation in Ionic Radii

Ionic radius is the measure used to describe the size of an ion. A cation always has fewer electrons and the same number of protons as the parent atom; it is smaller than the atom from which it is derived (Figure 3.6.3). For example, the covalent radius of an aluminum atom ($1s^2 2s^2 2p^6 3s^2 3p^1$) is 118 pm, whereas the ionic radius of an Al^{3+} ($1s^2 2s^2 2p^6$) is 68 pm. As electrons are removed from the outer valence shell, the remaining core electrons occupying smaller shells experience a greater effective nuclear charge Z_{eff} (as discussed) and are drawn even closer to the nucleus.



Figure 3.6.3: The radius for a cation is smaller than the parent atom (Al), due to the lost electrons; the radius for an anion is larger than the parent (S), due to the gained electrons.

Cations with larger charges are smaller than cations with smaller charges (e.g., V^{2+} has an ionic radius of 79 pm, while that of V^{3+} is 64 pm). Proceeding down the groups of the periodic table, we find that cations of successive elements with the same charge generally have larger radii, corresponding to an increase in the principal quantum number, n.

An anion (negative ion) is formed by the addition of one or more electrons to the valence shell of an atom. This results in a greater repulsion among the electrons and a decrease in Z_{eff} per electron. Both effects (the increased number of electrons and the decreased Z_{eff}) cause the radius of an anion to be larger than that of the parent atom (Figure 3.6.3). For example, a sulfur atom ($[Ne]3s^2 3p^4$) has a covalent radius of 104 pm, whereas the ionic radius of the sulfide anion ($[Ne]3s^2 3p^6$) is 170 pm. For consecutive elements proceeding down any group, anions have larger principal quantum numbers and, thus, larger radii.

Atoms and ions that have the same electron configuration are said to be **isoelectronic**. Examples of isoelectronic species are N^{3-}, O^{2-}, F^-, Ne, Na^+, Mg^{2+}, and Al^{3+} ($1s^22s^22p^6$). Another isoelectronic series is P^{3-}, S^{2-}, Cl^-, Ar, K^+, Ca^{2+}, and Sc^{3+} ([Ne]$3s^23p^6$). For atoms or ions that are isoelectronic, the number of protons determines the size. The greater the nuclear charge, the smaller the radius in a series of isoelectronic ions and atoms.

Variation in Ionization Energies

The amount of energy required to remove the most loosely bound electron from a gaseous atom in its ground state is called its first **ionization energy** (IE_1). The first ionization energy for an element, X, is the energy required to form a cation with +1 charge:

$$\mathrm{X(g)} \longrightarrow \mathrm{X^+(g) + e^-} \qquad IE_1$$

The energy required to remove the second most loosely bound electron is called the second ionization energy (IE_2).

$$\mathrm{X^+(g)} \longrightarrow \mathrm{X^{2+}(g) + e^-} \qquad IE_2$$

The energy required to remove the third electron is the third ionization energy, and so on. Energy is always required to remove electrons from atoms or ions, so ionization processes are endothermic and IE values are always positive. For larger atoms, the most loosely bound electron is located farther from the nucleus and so is easier to remove. Thus, as size (atomic radius) increases, the ionization energy should decrease. Relating this logic to what we have just learned about radii, we would expect first ionization energies to decrease down a group and to increase across a period.

Figure 3.6.4 graphs the relationship between the first ionization energy and the atomic number of several elements. The values of first ionization energy for the elements are given in Figure 3.6.5. Within a period, the IE_1 generally increases with increasing *Z*. Down a group, the IE_1 value generally decreases with increasing *Z*. There are some systematic deviations from this trend, however. Note that the ionization energy of boron (atomic number 5) is less than that of beryllium (atomic number 4) even though the nuclear charge of boron is greater by one proton. This can be explained because the energy of the subshells increases as *l* increases, due to penetration and shielding (as discussed previously in this chapter). Within any one shell, the *s* electrons are lower in energy than the *p* electrons. This means that an *s* electron is harder to remove from an atom than a *p* electron in the same shell. The electron removed during the ionization of beryllium ([He]$2s^2$) is an *s* electron, whereas the electron removed during the ionization of boron ([He]$2s^22p^1$) is a *p* electron; this results in a lower first ionization energy for boron, even though its nuclear charge is greater by one proton. Thus, we see a small deviation from the predicted trend occurring each time a new subshell begins.

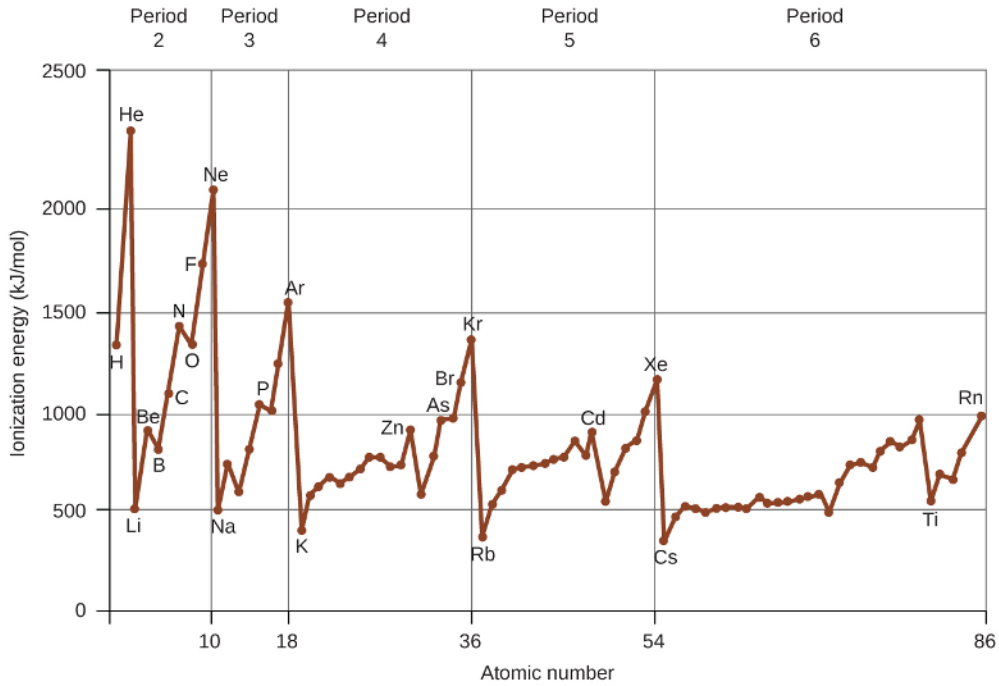


Figure 3.6.4: The first ionization energy of the elements in the first five periods are plotted against their atomic number.

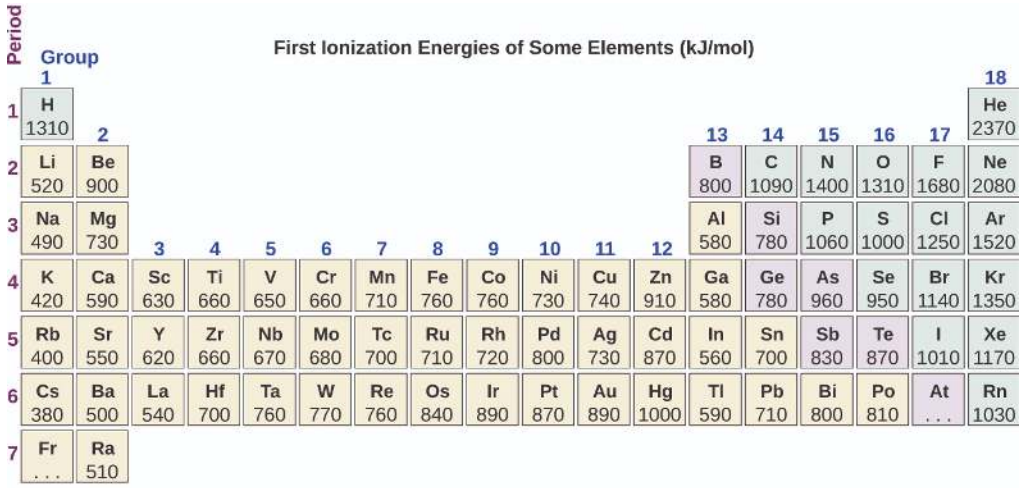
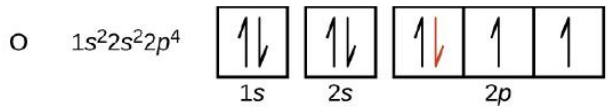


Figure 3.6.5: This version of the periodic table shows the first ionization energy (IE_1), in kJ/mol, of selected elements.

Another deviation occurs as orbitals become more than one-half filled. The first ionization energy for oxygen is slightly less than that for nitrogen, despite the trend in increasing IE_1 values across a period. Looking at the orbital diagram of oxygen, we can see that removing one electron will eliminate the electron–electron repulsion caused by pairing the electrons in the 2*p* orbital and will result in a half-filled orbital (which is energetically favorable). Analogous changes occur in succeeding periods (note the dip for sulfur after phosphorus in Figure 3.6.5).



Removing an electron from a cation is more difficult than removing an electron from a neutral atom because of the greater electrostatic attraction to the cation. Likewise, removing an electron from a cation with a higher positive charge is more difficult than removing an electron from an ion with a lower charge. Thus, successive ionization energies for one element always increase. As seen in Table 3.6.2, there is a large increase in the ionization energies for each element. This jump corresponds to removal of the core electrons, which are harder to remove than the valence electrons. For example, Sc and Ga both have three valence electrons, so the rapid increase in ionization energy occurs after the third ionization.

Table 3.6.2: Successive Ionization Energies for Selected Elements (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
K	418.8	3051.8	4419.6	5876.9	7975.5	9590.6	11343
Ca	589.8	1145.4	4912.4	6490.6	8153.0	10495.7	12272.9
Sc	633.1	1235.0	2388.7	7090.6	8842.9	10679.0	13315.0
Ga	578.8	1979.4	2964.6	6180	8298.7	10873.9	13594.8
Ge	762.2	1537.5	3302.1	4410.6	9021.4	Not available	Not available
As	944.5	1793.6	2735.5	4836.8	6042.9	12311.5	Not available

Example 3.6.2: Ranking Ionization Energies

Predict the order of increasing energy for the following processes: IE_1 for Al, IE_1 for Tl, IE_2 for Na, IE_3 for Al.

Solution

Removing the $6p^1$ electron from Tl is easier than removing the $3p^1$ electron from Al because the higher *n* orbital is farther from the nucleus, so $IE_1(Tl) < IE_1(Al)$. Ionizing the third electron from Al ($Al^{2+} \longrightarrow Al^{3+} + e^-$) requires more energy because the cation Al^{2+} exerts a stronger pull on the electron than the neutral Al atom, so $IE_1(Al) < IE_3(Al)$. The second ionization

energy for sodium removes a core electron, which is a much higher energy process than removing valence electrons. Putting this all together, we obtain: $IE_1(Tl) < IE_1(Al) < IE_3(Al) < IE_2(Na)$.

? Exercise 3.6.2

Which has the lowest value for IE_1: O, Po, Pb, or Ba?

Answer

Ba

Variation in Electron Affinities

The **electron affinity** (EA) is the energy change for the process of adding an electron to a gaseous atom to form an anion (negative ion).

$$\mathrm{X(g) + e^- \longrightarrow X^-(g)} \quad EA_1$$

This process can be either endothermic or exothermic, depending on the element. The EA of some of the elements is given in Figure 3.6.6. You can see that many of these elements have negative values of EA, which means that energy is released when the gaseous atom accepts an electron. However, for some elements, energy is required for the atom to become negatively charged and the value of their EA is positive. Just as with ionization energy, subsequent EA values are associated with forming ions with more charge. The second EA is the energy associated with adding an electron to an anion to form a –2 ion, and so on.

As we might predict, it becomes easier to add an electron across a series of atoms as the effective nuclear charge of the atoms increases. We find, as we go from left to right across a period, EAs tend to become more negative. The exceptions found among the elements of group 2 (2A), group 15 (5A), and group 18 (8A) can be understood based on the electronic structure of these groups. The noble gases, group 18 (8A), have a completely filled shell and the incoming electron must be added to a higher *n* level, which is more difficult to do. Group 2 (2A) has a filled *ns* subshell, and so the next electron added goes into the higher energy *np*, so, again, the observed EA value is not as the trend would predict. Finally, group 15 (5A) has a half-filled *np* subshell and the next electron must be paired with an existing *np* electron. In all of these cases, the initial relative stability of the electron configuration disrupts the trend in EA.

We also might expect the atom at the top of each group to have the most negative EA; their first ionization potentials suggest that these atoms have the largest effective nuclear charges. However, as we move down a group, we see that the *second* element in the group most often has the most negative EA. This can be attributed to the small size of the $n = 2$ shell and the resulting large electron–electron repulsions. For example, chlorine, with an EA value of –348 kJ/mol, has the highest value of any element in the periodic table. The EA of fluorine is –322 kJ/mol. When we add an electron to a fluorine atom to form a fluoride anion (F^-), we add an electron to the $n = 2$ shell. The electron is attracted to the nucleus, but there is also significant repulsion from the other electrons already present in this small valence shell. The chlorine atom has the same electron configuration in the valence shell, but because the entering electron is going into the $n = 3$ shell, it occupies a considerably larger region of space and the electron–electron repulsions are reduced. The entering electron does not experience as much repulsion and the chlorine atom accepts an additional electron more readily, resulting in a more negative EA.

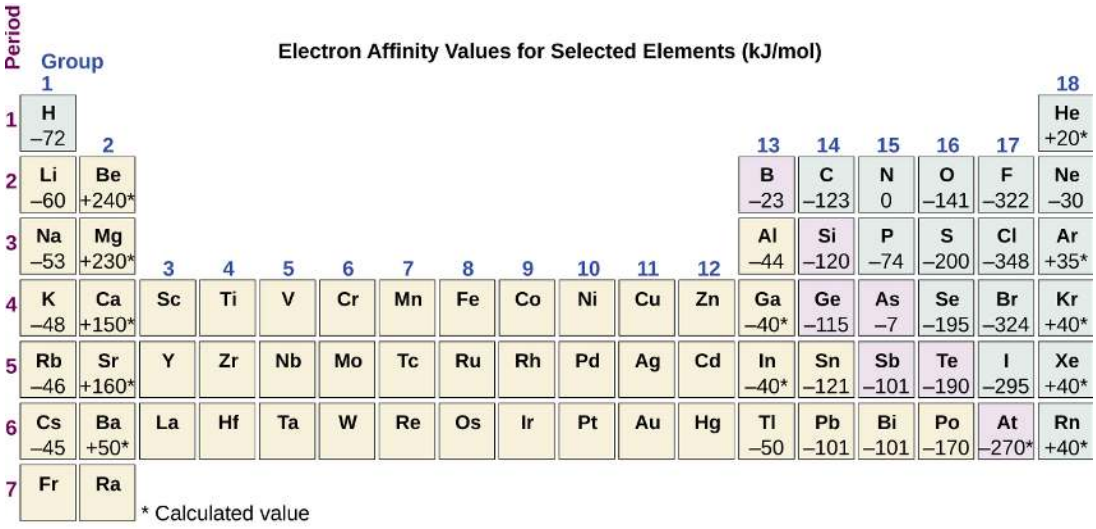


Figure 3.6.6: This version of the periodic table displays the electron affinity values (in kJ/mol) for selected elements.

The properties discussed in this section (size of atoms and ions, effective nuclear charge, ionization energies, and electron affinities) are central to understanding chemical reactivity. For example, because fluorine has an energetically favorable EA and a large energy barrier to ionization (IE), it is much easier to form fluorine anions than cations. Metallic properties including conductivity and malleability (the ability to be formed into sheets) depend on having electrons that can be removed easily. Thus, metallic character increases as we move down a group and decreases across a period in the same trend observed for atomic size because it is easier to remove an electron that is farther away from the nucleus.

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3.7: The Periodic Table

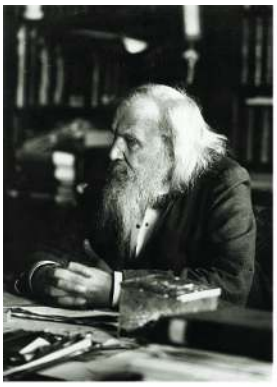
Learning Objectives

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

- State the periodic law and explain the organization of elements in the periodic table
- Predict the general properties of elements based on their location within the periodic table
- Identify metals, nonmetals, and metalloids by their properties and/or location on the periodic table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time. Both published tables with the elements arranged according to increasing atomic mass. But Mendeleev went one step further than Meyer: He used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work. Although Mendeleev and Meyer had a long dispute over priority, Mendeleev's contributions to the development of the periodic table are now more widely recognized (Figure 3.7.1).



(a)

Reihen	Gruppe I. — R^1O	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	—	— — — —

(b)

Figure 3.7.1: (a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)

By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses. The modern statement of this relationship, the **periodic law**, is as follows: *the properties of the elements are periodic functions of their atomic numbers*. A modern **periodic table** arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column (Figure 3.7.2). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called **periods** or **series**, and 18 vertical columns, called **groups**. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written below the main body of the table.

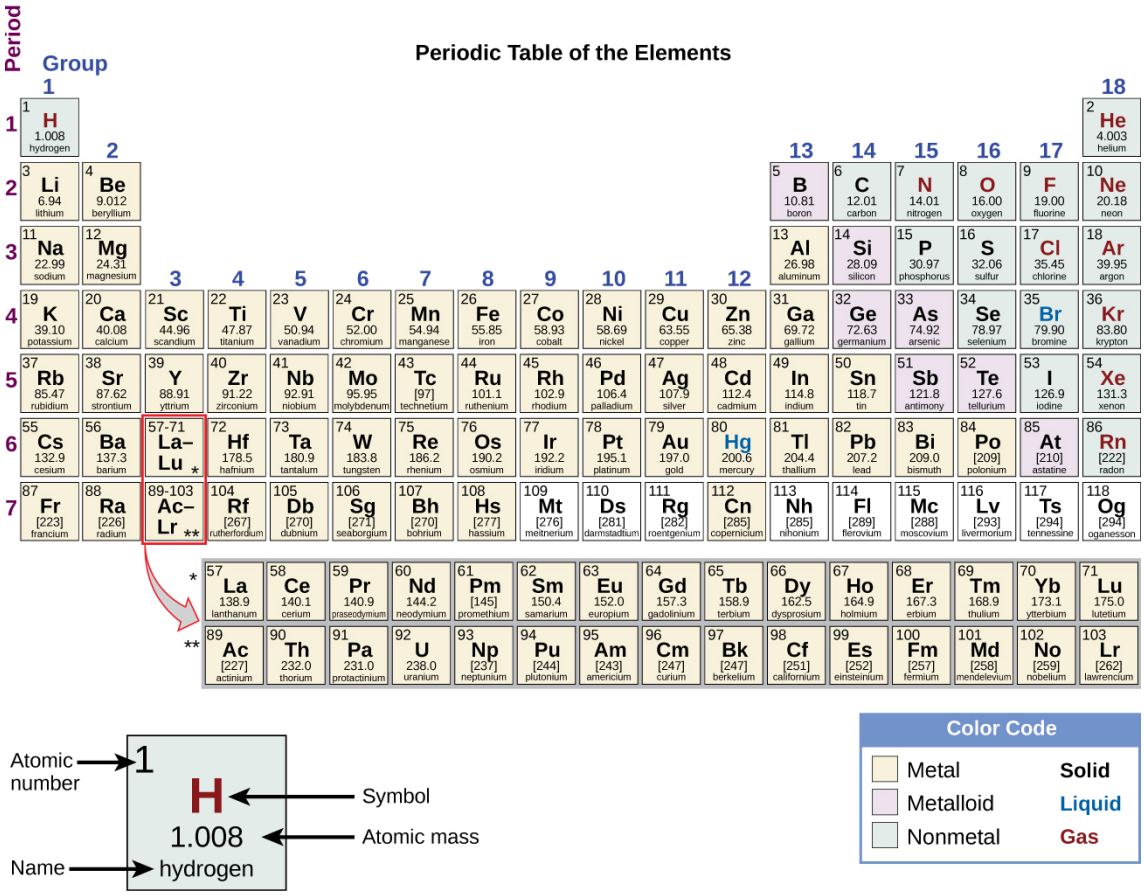


Figure 3.7.2: Elements in the periodic table are organized according to their properties.

Even after the periodic nature of elements and the table itself were widely accepted, gaps remained. Mendeleev had predicted, and others including Henry Moseley had later confirmed, that there should be elements below Manganese in Group 7. German chemists Ida Tacke and Walter Noddack set out to find the elements, a quest being pursued by scientists around the world. Their method was unique in that they did not only consider the properties of manganese, but also the elements horizontally adjacent to the missing elements 43 and 75 on the table. Thus, by investigating ores containing minerals of ruthenium (Ru), tungsten (W), osmium (Os), and so on, they were able to identify naturally occurring elements that helped complete the table. Rhenium, one of their discoveries, was one of the last natural elements to be discovered and is the last stable element to be discovered. (Francium, the last natural element to be discovered, was identified by Marguerite Perey in 1939.)

Many elements differ dramatically in their chemical and physical properties, but some elements are similar in their behaviors. For example, many elements appear shiny, are malleable (able to be deformed without breaking) and ductile (can be drawn into wires), and conduct heat and electricity well. Other elements are not shiny, malleable, or ductile, and are poor conductors of heat and electricity. We can sort the elements into large classes with common properties: **metals** (elements that are shiny, malleable, good conductors of heat and electricity—shaded yellow); **nonmetals** (elements that appear dull, poor conductors of heat and electricity—shaded green); and **metalloids** (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals—shaded purple).

The elements can also be classified into the main-group elements (or representative elements) in the columns labeled 1, 2, and 13–18; the transition metals in the columns labeled 3–12[3]; and inner transition metals in the two rows at the bottom of the table (the top-row elements are called lanthanides and the bottom-row elements are actinides; Figure 3.7.3). The elements can be subdivided further by more specific properties, such as the composition of the compounds they form. For example, the elements in group 1 (the first column) form compounds that consist of one atom of the element and one atom of hydrogen. These elements (except hydrogen) are known as **alkali metals**, and they all have similar chemical properties. The elements in group 2 (the second column) form compounds consisting of one atom of the element and two atoms of hydrogen: These are called **alkaline earth metals**, with similar properties among members of that group. Other groups with specific names are the **pnictogens** (group 15), **chalcogens** (group 16), halogens (group 17), and the **noble gases** (group 18, also known as inert gases). The groups can also be referred to by

the first element of the group: For example, the chalcogens can be called the oxygen group or oxygen family. Hydrogen is a unique, nonmetallic element with properties similar to both group 1 and group 17 elements. For that reason, hydrogen may be shown at the top of both groups, or by itself.

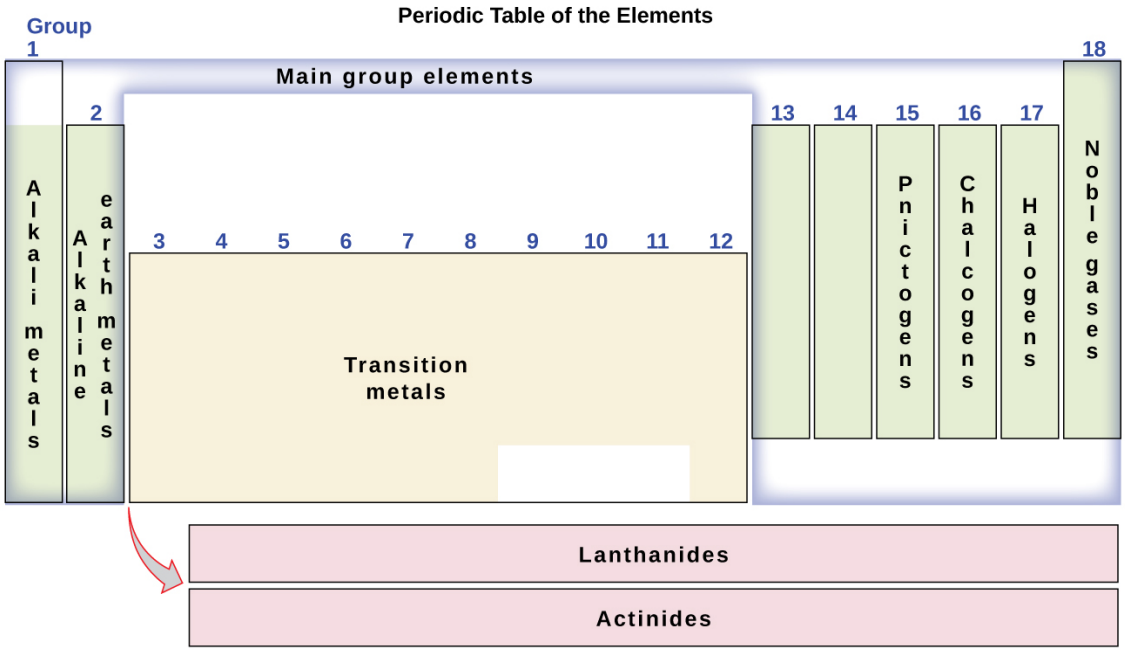


Figure 3.7.3: The periodic table organizes elements with similar properties into groups.

Link to Learning

Click on this link for an interactive periodic table, which you can use to explore the properties of the elements (includes podcasts and videos of each element). You may also want to try this one that shows photos of all the elements.

Example 3.7.1: Naming Groups of Elements

Atoms of each of the following elements are essential for life. Give the group name for the following elements:

a. chlorine
b. calcium
c. sodium
d. sulfur

Solution

The family names are as follows:

a. halogen
b. alkaline earth metal
c. alkali metal
d. chalcogen

Exercise 3.7.1

Give the group name for each of the following elements:

a. krypton
b. selenium
c. barium
d. lithium

Answer

(a) noble gas; (b) chalcogen; (c) alkaline earth metal; (d) alkali metal

As you will learn in your further study of chemistry, elements in groups often behave in a somewhat similar manner. This is partly due to the number of electrons in their outer shell and their similar readiness to bond. These shared properties can have far-ranging implications in nature, science, and medicine. For example, when Gertrude Elion and George Hitchens were investigating ways to interrupt cell and virus replication to fight diseases, they utilized the similarity between sulfur and oxygen (both in Group 16) and their capacity to bond in similar ways. Elion focused on purines, which are key components of DNA and which contain oxygen. She found that by introducing sulfur-based compounds (called purine analogues) that mimic the structure of purines, molecules within DNA would bond to the analogues rather than the "regular" DNA purine. With the normal DNA bonding and structure altered, Elion successfully interrupted cell replication. At its core, the strategy worked because of the similarity between sulfur and oxygen. Her discovery led directly to important treatments for leukemia. Overall, Elion's work with George Hitchens not only led to more treatments, but also changed the entire methodology of drug development. By using specific elements and compounds to target specific aspects of tumor cells, viruses, and bacteria, they laid the groundwork for many of today's most common and important medicines, used to help millions of people each year. They were awarded the Nobel Prize in 1988.

In studying the periodic table, you might have noticed something about the atomic masses of some of the elements. Element 43 (technetium), element 61 (promethium), and most of the elements with atomic number 84 (polonium) and higher have their atomic mass given in square brackets. This is done for elements that consist entirely of unstable, radioactive isotopes (you will learn more about radioactivity in the nuclear chemistry chapter). An average atomic weight cannot be determined for these elements because their radioisotopes may vary significantly in relative abundance, depending on the source, or may not even exist in nature. The number in square brackets is the atomic mass number (an approximate atomic mass) of the most stable isotope of that element.

Footnotes

- 3Per the IUPAC definition, group 12 elements are not transition metals, though they are often referred to as such. Additional details on this group’s elements are provided in a chapter on transition metals and coordination chemistry.

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3.8: Ionic and Molecular Compounds

Learning Objectives

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

- Define ionic and molecular (covalent) compounds
- Predict the type of compound formed from elements based on their location within the periodic table
- Determine formulas for simple ionic compounds

In ordinary chemical reactions, the nucleus of each atom (and thus the identity of the element) remains unchanged. Electrons, however, can be added to atoms by transfer from other atoms, lost by transfer to other atoms, or shared with other atoms. The transfer and sharing of electrons among atoms govern the chemistry of the elements. During the formation of some compounds, atoms gain or lose electrons, and form electrically charged particles called ions (Figure 3.8.1).

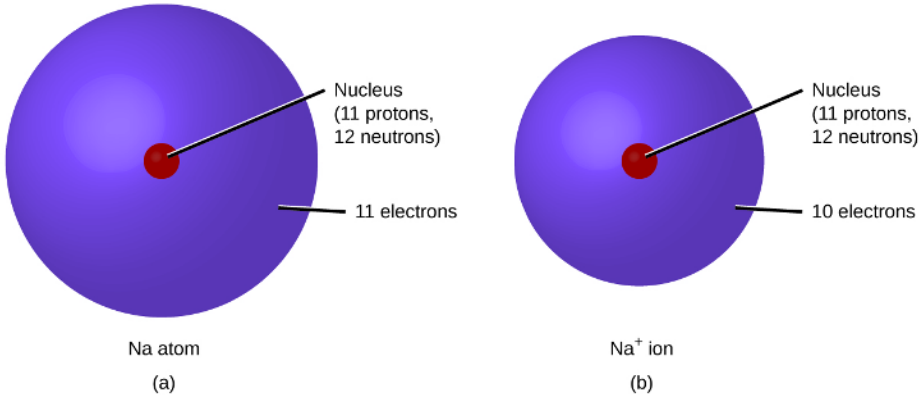


Table 3.8.1: (a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na^+) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.

You can use the periodic table to predict whether an atom will form an anion or a cation, and you can often predict the charge of the resulting ion. Atoms of many main-group metals lose enough electrons to leave them with the same number of electrons as an atom of the preceding noble gas. To illustrate, an atom of an alkali metal (group 1) loses one electron and forms a cation with a 1+ charge; an alkaline earth metal (group 2) loses two electrons and forms a cation with a 2+ charge, and so on. For example, a neutral calcium atom, with 20 protons and 20 electrons, readily loses two electrons. This results in a cation with 20 protons, 18 electrons, and a 2+ charge. It has the same number of electrons as atoms of the preceding noble gas, argon, and is symbolized Ca^{2+}. The name of a metal ion is the same as the name of the metal atom from which it forms, so Ca^{2+} is called a calcium ion.

When atoms of nonmetal elements form ions, they generally gain enough electrons to give them the same number of electrons as an atom of the next noble gas in the periodic table. Atoms of group 17 gain one electron and form anions with a 1− charge; atoms of group 16 gain two electrons and form ions with a 2− charge, and so on. For example, the neutral bromine atom, with 35 protons and 35 electrons, can gain one electron to provide it with 36 electrons. This results in an anion with 35 protons, 36 electrons, and a 1− charge. It has the same number of electrons as atoms of the next noble gas, krypton, and is symbolized Br^-. (A discussion of the theory supporting the favored status of noble gas electron numbers reflected in these predictive rules for ion formation is provided in a later chapter of this text.)

Note the usefulness of the periodic table in predicting likely ion formation and charge (Figure 3.8.2). Moving from the far left to the right on the periodic table, main-group elements tend to form cations with a charge equal to the group number. That is, group 1 elements form 1+ ions; group 2 elements form 2+ ions, and so on. Moving from the far right to the left on the periodic table, elements often form anions with a negative charge equal to the number of groups moved left from the noble gases. For example, group 17 elements (one group left of the noble gases) form 1− ions; group 16 elements (two groups left) form 2− ions, and so on. This trend can be used as a guide in many cases, but its predictive value decreases when moving toward the center of the periodic table. In fact, transition metals and some other metals often exhibit variable charges that are not predictable by their location in the table. For example, copper can form ions with a 1+ or 2+ charge, and iron can form ions with a 2+ or 3+ charge.

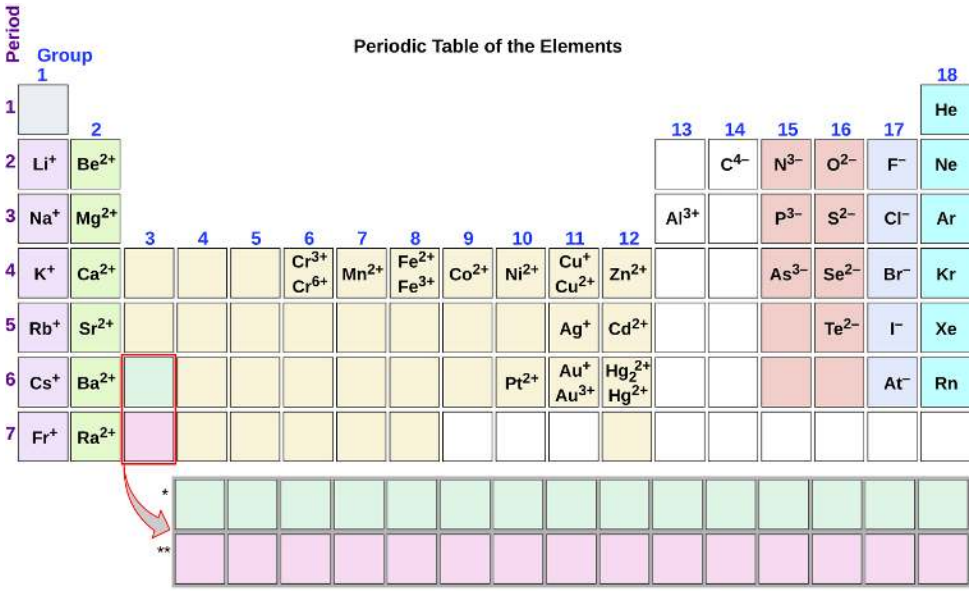


Table 3.8.2: Some elements exhibit a regular pattern of ionic charge when they form ions.

Example 3.8.1: Composition of Ions

An ion found in some compounds used as antiperspirants contains 13 protons and 10 electrons. What is its symbol?

Solution

Because the number of protons remains unchanged when an atom forms an ion, the atomic number of the element must be 13. Knowing this lets us use the periodic table to identify the element as Al (aluminum). The Al atom has lost three electrons and thus has three more positive charges (13) than it has electrons (10). This is the aluminum cation, Al^{3+}.

Exercise 3.8.1

Give the symbol and name for the ion with 34 protons and 36 electrons.

Answer

Se^{2-}, the selenide ion

Example 3.8.2: Formation of Ions

Magnesium and nitrogen react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Solution

Magnesium's position in the periodic table (group 2) tells us that it is a metal. Metals form positive ions (cations). A magnesium atom must lose two electrons to have the same number electrons as an atom of the previous noble gas, neon. Thus, a magnesium atom will form a cation with two fewer electrons than protons and a charge of 2+. The symbol for the ion is Mg^{2+}, and it is called a magnesium ion.

Nitrogen's position in the periodic table (group 15) reveals that it is a nonmetal. Nonmetals form negative ions (anions). A nitrogen atom must gain three electrons to have the same number of electrons as an atom of the following noble gas, neon. Thus, a nitrogen atom will form an anion with three more electrons than protons and a charge of 3−. The symbol for the ion is N^{3-}, and it is called a nitride ion.

Exercise 3.8.2

Aluminum and carbon react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Answer

Al will form a cation with a charge of 3+: Al^{3+}, an aluminum ion. Carbon will form an anion with a charge of 4−: C^{4-}, a carbide ion.

The ions that we have discussed so far are called monatomic ions, that is, they are ions formed from only one atom. We also find many polyatomic ions. These ions, which act as discrete units, are electrically charged molecules (a group of bonded atoms with an overall charge). Some of the more important polyatomic ions are listed in Table 3.8.1. Oxyanions are polyatomic ions that contain one or more oxygen atoms. At this point in your study of chemistry, you should memorize the names, formulas, and charges of the most common polyatomic ions. Because you will use them repeatedly, they will soon become familiar.

Table 3.8.1: Common Polyatomic Ions

Name	Formula	Related Acid	Formula
ammonium	NH_4^+		
hydronium	H_3O		
peroxide	O_2^{2-}		
hydroxide	OH^-		
acetate	CH_3COO^-	acetic acid	CH_3COOH
cyanide	CN^-	hydrocyanic acid	HCN
azide	N_3^-	hydrazoic acid	HN_3
carbonate	CO_3^{2-}	carbonic acid	H_2CO_3
bicarbonate	HCO_3^-		
nitrate	NO_3^-	nitric acid	HNO_3
nitrite	NO_2^-	nitrous acid	HNO_2
sulfate	SO_{42}^-	sulfuric acid	H_2SO_4
hydrogen sulfate	HSO_4^-		
sulfite	SO_3^{2-}	sulfurous acid	H_2SO_3
hydrogen sulfite	HSO_3^-		
phosphate	PO_4^{3-}	phosphoric acid	H_3PO_4
hydrogen phosphate	HPO_4^{2-}		
dihydrogen phosphate	$H_2PO_4^-$		
perchlorate	ClO_4^-	perchloric acid	$HClO_4$
chlorate	ClO_3^-	chloric acid	$HClO_3$
chlorite	ClO_2^-	chlorous acid	$HClO_2$
hypochlorite	ClO^-	hypochlorous acid	$HClO$
chromate	CrO_4^{2-}	chromic acid	H_2CrO_4
dichromate	$Cr_2O_7^{2-}$	dichromic acid	$H_2Cr_2O_7$

Name	Formula	Related Acid	Formula
permanganate	MnO_4^-	permanganic acid	$HMnO_4$

Note that there is a system for naming some polyatomic ions; -ate and -ite are suffixes designating polyatomic ions containing more or fewer oxygen atoms. Per- (short for “hyper”) and hypo- (meaning “under”) are prefixes meaning more oxygen atoms than -ate and fewer oxygen atoms than -ite, respectively. For example, perchlorate is ClO_4^-, chlorate is ClO_3^-, chlorite is ClO_2^- and hypochlorite is ClO^-. Unfortunately, the number of oxygen atoms corresponding to a given suffix or prefix is not consistent; for example, nitrate is NO_3^- while sulfate is SO_4^{2-}. This will be covered in more detail in the next module on nomenclature.

The nature of the attractive forces that hold atoms or ions together within a compound is the basis for classifying chemical bonding. When electrons are transferred and ions form, ionic bonds result. Ionic bonds are electrostatic forces of attraction, that is, the attractive forces experienced between objects of opposite electrical charge (in this case, cations and anions). When electrons are “shared” and molecules form, covalent bonds result. Covalent bonds are the attractive forces between the positively charged nuclei of the bonded atoms and one or more pairs of electrons that are located between the atoms. Compounds are classified as ionic or molecular (covalent) on the basis of the bonds present in them.

Ionic Compounds

When an element composed of atoms that readily lose electrons (a metal) reacts with an element composed of atoms that readily gain electrons (a nonmetal), a transfer of electrons usually occurs, producing ions. The compound formed by this transfer is stabilized by the electrostatic attractions (ionic bonds) between the ions of opposite charge present in the compound. For example, when each sodium atom in a sample of sodium metal (group 1) gives up one electron to form a sodium cation, Na^+, and each chlorine atom in a sample of chlorine gas (group 17) accepts one electron to form a chloride anion, Cl^-, the resulting compound, NaCl, is composed of sodium ions and chloride ions in the ratio of one Na^+ ion for each Cl^- ion. Similarly, each calcium atom (group 2) can give up two electrons and transfer one to each of two chlorine atoms to form $CaCl_2$, which is composed of Ca^{2+} and Cl^- ions in the ratio of one Ca^{2+} ion to two Cl^- ions.

A compound that contains ions and is held together by ionic bonds is called an ionic compound. The periodic table can help us recognize many of the compounds that are ionic: When a metal is combined with one or more nonmetals, the compound is usually ionic. This guideline works well for predicting ionic compound formation for most of the compounds typically encountered in an introductory chemistry course. However, it is not always true (for example, aluminum chloride, $AlCl_3$, is not ionic).

You can often recognize ionic compounds because of their properties. Ionic compounds are solids that typically melt at high temperatures and boil at even higher temperatures. For example, sodium chloride melts at 801 °C and boils at 1413 °C. (As a comparison, the molecular compound water melts at 0 °C and boils at 100 °C.) In solid form, an ionic compound is not electrically conductive because its ions are unable to flow (“electricity” is the flow of charged particles). When molten, however, it can conduct electricity because its ions are able to move freely through the liquid (Figure 3.8.3).

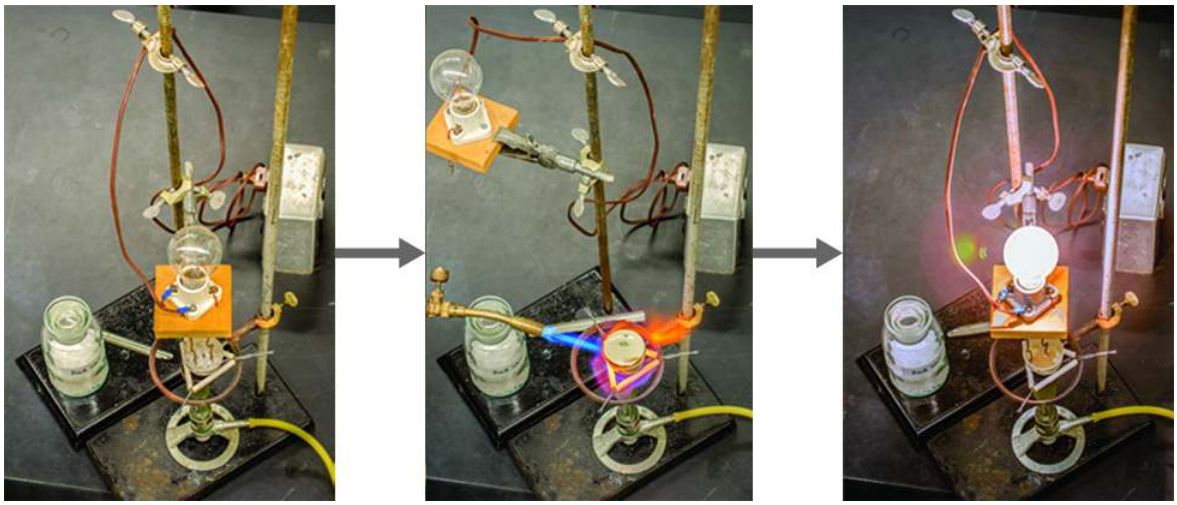


Table 3.8.3: Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

Link to Learning

Watch this video to see a mixture of salts melt and conduct electricity.

In every ionic compound, the total number of positive charges of the cations equals the total number of negative charges of the anions. Thus, ionic compounds are electrically neutral overall, even though they contain positive and negative ions. We can use this observation to help us write the formula of an ionic compound. The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.

Example 3.8.3: Predicting the Formula of an Ionic Compound

The gemstone sapphire (Figure 3.8.4) is mostly a compound of aluminum and oxygen that contains aluminum cations, Al^{3+}, and oxygen anions, O^{2-}. What is the formula of this compound?



Table 3.8.4: Although pure aluminum oxide is colorless, trace amounts of iron and titanium give blue sapphire its characteristic color. (credit: modification of work by Stanislav Doronenko via Wikipedia)

Solution

Because the ionic compound must be electrically neutral, it must have the same number of positive and negative charges. Two aluminum ions, each with a charge of 3+, would give us six positive charges, and three oxide ions, each with a charge of 2−, would give us six negative charges. The formula would be Al_2O_3.

Exercise 3.8.3

Predict the formula of the ionic compound formed between the sodium cation, Na^+, and the sulfide anion, S^{2-}.

Answer

Na_2S

Many ionic compounds contain polyatomic ions (Table 3.8.1) as the cation, the anion, or both. As with simple ionic compounds, these compounds must also be electrically neutral, so their formulas can be predicted by treating the polyatomic ions as discrete units. We use parentheses in a formula to indicate a group of atoms that behave as a unit. For example, the formula for calcium phosphate, one of the minerals in our bones, is $Ca_3(PO_4)_2$. This formula indicates that there are three calcium ions (Ca^{2+}) for every two phosphate (PO_4^{3-}) groups. The PO_4^{3-} groups are discrete units, each consisting of one phosphorus atom and four oxygen atoms, and having an overall charge of 3−. The compound is electrically neutral, and its formula shows a total count of three Ca, two P, and eight O atoms.

Example 3.8.4: Predicting the Formula of a Compound with a Polyatomic Anion

Baking powder contains calcium dihydrogen phosphate, an ionic compound composed of the ions Ca^{2+} and $H_2PO_4^-$. What is the formula of this compound?

Solution

The positive and negative charges must balance, and this ionic compound must be electrically neutral. Thus, we must have two negative charges to balance the 2+ charge of the calcium ion. This requires a ratio of one Ca^{2+} ion to two $H_2PO_4^-$ ions. We

designate this by enclosing the formula for the dihydrogen phosphate ion in parentheses and adding a subscript 2. The formula is $Ca(H_2PO_4)_2$.

? Exercise 3.8.4

Predict the formula of the ionic compound formed between the lithium ion and the peroxide ion, O_2^{2-} (Hint: Use the periodic table to predict the sign and the charge on the lithium ion.)

Answer

Li_2O_2

Because an ionic compound is not made up of single, discrete molecules, it may not be properly symbolized using a *molecular* formula. Instead, ionic compounds must be symbolized by a formula indicating the relative numbers of its constituent ions. For compounds containing only monatomic ions (such as $NaCl$) and for many compounds containing polyatomic ions (such as $CaSO_4$), these formulas are just the empirical formulas introduced earlier in this chapter. However, the formulas for some ionic compounds containing polyatomic ions are not empirical formulas. For example, the ionic compound sodium oxalate is comprised of Na^+ and $C_2O_4^{2-}$ ions combined in a 2:1 ratio, and its formula is written as $Na_2C_2O_4$. The subscripts in this formula are not the smallest-possible whole numbers, as each can be divided by 2 to yield the empirical formula, $NaCO_2$. This is not the accepted formula for sodium oxalate, however, as it does not accurately represent the compound's polyatomic anion, $C_2O_4^{2-}$.

Molecular Compounds

Many compounds do not contain ions but instead consist solely of discrete, neutral molecules. These molecular compounds (covalent compounds) result when atoms share, rather than transfer (gain or lose), electrons. Covalent bonding is an important and extensive concept in chemistry, and it will be treated in considerable detail in a later chapter of this text. We can often identify molecular compounds on the basis of their physical properties. Under normal conditions, molecular compounds often exist as gases, low-boiling liquids, and low-melting solids, although many important exceptions exist.

Whereas ionic compounds are usually formed when a metal and a nonmetal combine, covalent compounds are usually formed by a combination of nonmetals. Thus, the periodic table can help us recognize many of the compounds that are covalent. While we can use the positions of a compound's elements in the periodic table to predict whether it is ionic or covalent at this point in our study of chemistry, you should be aware that this is a very simplistic approach that does not account for a number of interesting exceptions. Shades of gray exist between ionic and molecular compounds, and you'll learn more about those later.

✓ Example 3.8.5: Predicting the Type of Bonding in Compounds

Predict whether the following compounds are ionic or molecular:

a. KI, the compound used as a source of iodine in table salt
b. H_2O_2, the bleach and disinfectant hydrogen peroxide
c. $CHCl_3$, the anesthetic chloroform
d. Li_2CO_3, a source of lithium in antidepressants

Solution

a. Potassium (group 1) is a metal, and iodine (group 17) is a nonmetal; KI is predicted to be ionic.
b. Hydrogen (group 1) is a nonmetal, and oxygen (group 16) is a nonmetal; H_2O_2 is predicted to be molecular.
c. Carbon (group 14) is a nonmetal, hydrogen (group 1) is a nonmetal, and chlorine (group 17) is a nonmetal; $CHCl_3$ is predicted to be molecular.
d. Lithium (group 1) is a metal, and carbonate is a polyatomic ion; Li_2CO_3 is predicted to be ionic.

? Exercise 3.8.5

Using the periodic table, predict whether the following compounds are ionic or covalent:

a. SO_2
b. CaF_2

c. N_2H_4
d. $Al_2(SO_4)_3$

Answer

a. molecular
b. ionic
c. molecular
d. ionic

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3.9: 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

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3.10: Key Equations

$c = \lambda\nu$
$E = h\nu = \frac{hc}{\lambda}$, where $h = 6.626 \times 10^{-34}$ J s
$\frac{1}{\lambda} = R_\infty \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$

$E_n = -\frac{kZ^2}{n^2}$, $n = 1,\ 2,\ 3,\ \ldots$
$\Delta E = kZ^2 \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$
$r =$

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3.11: 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.

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CHAPTER OVERVIEW

4: 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.

4.1: Introduction
4.2: Ionic Bonding
4.3: Covalent Bonding
4.4: Chemical Nomenclature
4.5: Lewis Symbols and Structures
4.6: Formal Charges and Resonance
4.7: Molecular Structure and Polarity
4.8: Key Terms
4.9: Key Equations
4.10: Summary

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4.1: 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 4.1.1: 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)

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4.2: Ionic Bonding

Learning Objectives

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

- Explain the formation of cations, anions, and ionic compounds
- Predict the charge of common metallic and nonmetallic elements, and write their electron configurations

As you have learned, ions are atoms or molecules bearing an electrical charge. A cation (a positive ion) forms when a neutral atom loses one or more electrons from its valence shell, and an anion (a negative ion) forms when a neutral atom gains one or more electrons in its valence shell.

Compounds composed of ions are called ionic compounds (or salts), and their constituent ions are held together by ionic bonds: electrostatic forces of attraction between oppositely charged cations and anions. The properties of ionic compounds shed some light on the nature of ionic bonds. Ionic solids exhibit a crystalline structure and tend to be rigid and brittle; they also tend to have high melting and boiling points, which suggests that ionic bonds are very strong. Ionic solids are also poor conductors of electricity for the same reason—the strength of ionic bonds prevents ions from moving freely in the solid state. Most ionic solids, however, dissolve readily in water. Once dissolved or melted, ionic compounds are excellent conductors of electricity and heat because the ions can move about freely.

Neutral atoms and their associated ions have very different physical and chemical properties. Sodium *atoms* form sodium metal, a soft, silvery-white metal that burns vigorously in air and reacts explosively with water. Chlorine *atoms* form chlorine gas, Cl_2, a yellow-green gas that is extremely corrosive to most metals and very poisonous to animals and plants. The vigorous reaction between the elements sodium and chlorine forms the white, crystalline compound sodium chloride, common table salt, which contains sodium *cations* and chloride *anions* (Figure 4.2.1). The compound composed of these ions exhibits properties entirely different from the properties of the elements sodium and chlorine. Chlorine is poisonous, but sodium chloride is essential to life; sodium atoms react vigorously with water, but sodium chloride simply dissolves in water.



Figure 4.2.1: (a) Sodium is a soft metal that must be stored in mineral oil to prevent reaction with air or water. (b) Chlorine is a pale yellow-green gas. (c) When combined, they form white crystals of sodium chloride (table salt). (credit a: modification of work by “Jurii”/Wikimedia Commons)

The Formation of Ionic Compounds

Binary ionic compounds are composed of just two elements: a metal (which forms the cations) and a nonmetal (which forms the anions). For example, NaCl is a binary ionic compound. We can think about the formation of such compounds in terms of the periodic properties of the elements. Many metallic elements have relatively low ionization potentials and lose electrons easily. These elements lie to the left in a period or near the bottom of a group on the periodic table. Nonmetal atoms have relatively high electron affinities and thus readily gain electrons lost by metal atoms, thereby filling their valence shells. Nonmetallic elements are found in the upper-right corner of the periodic table.

As all substances must be electrically neutral, the total number of positive charges on the cations of an ionic compound must equal the total number of negative charges on its anions. The formula of an ionic compound represents the simplest ratio of the numbers of ions necessary to give identical numbers of positive and negative charges. For example, the formula for aluminum oxide, Al_2O_3,

indicates that this ionic compound contains two aluminum cations, Al^{3+}, for every three oxide anions, O^{2-} [thus, (2 +3) + (3 –2) = 0].

It is important to note, however, that the formula for an ionic compound does *not* represent the physical arrangement of its ions. It is incorrect to refer to a sodium chloride (NaCl) "molecule" because there is not a single ionic bond, per se, between any specific pair of sodium and chloride ions. The attractive forces between ions are isotropic—the same in all directions—meaning that any particular ion is equally attracted to all of the nearby ions of opposite charge. This results in the ions arranging themselves into a tightly bound, three-dimensional lattice structure. Sodium chloride, for example, consists of a regular arrangement of equal numbers of Na^+ cations and Cl^- anions (Figure 4.2.2).

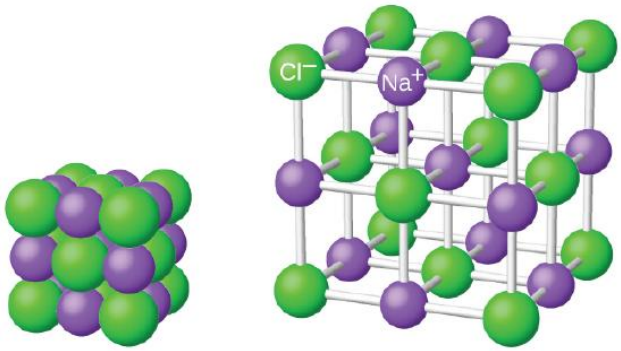


Figure 4.2.2: The atoms in sodium chloride (common table salt) are arranged to (a) maximize opposite charges interacting. The smaller spheres represent sodium ions, the larger ones represent chloride ions. In the expanded view (b), the geometry can be seen more clearly. Note that each ion is "bonded" to all of the surrounding ions—six in this case.

The strong electrostatic attraction between Na^+ and Cl^- ions holds them tightly together in solid NaCl. It requires 769 kJ of energy to dissociate one mole of solid NaCl into separate gaseous Na^+ and Cl^- ions:

$$\mathrm{NaCl(s)} \longrightarrow \mathrm{Na^+(g) + Cl^-(g)} \qquad \Delta H = 769\ \mathrm{kJ}$$

Electronic Structures of Cations

When forming a cation, an atom of a main group element tends to lose all of its valence electrons, thus assuming the electronic structure of the noble gas that precedes it in the periodic table. For groups 1 (the alkali metals) and 2 (the alkaline earth metals), the group numbers are equal to the numbers of valence shell electrons and, consequently, to the charges of the cations formed from atoms of these elements when all valence shell electrons are removed. For example, calcium is a group 2 element whose neutral atoms have 20 electrons and a ground state electron configuration of $1s^22s^22p^63s^23p^64s^2$. When a Ca atom loses both of its valence electrons, the result is a cation with 18 electrons, a 2+ charge, and an electron configuration of $1s^22s^22p^63s^23p^6$. The Ca^{2+} ion is therefore isoelectronic with the noble gas Ar.

For groups 13–17, the group numbers exceed the number of valence electrons by 10 (accounting for the possibility of full *d* subshells in atoms of elements in the fourth and greater periods). Thus, the charge of a cation formed by the loss of all valence electrons is equal to the group number minus 10. For example, aluminum (in group 13) forms 3+ ions (Al^{3+}).

Exceptions to the expected behavior involve elements toward the bottom of the groups. In addition to the expected ions Tl^{3+}, Sn^{4+}, Pb^{4+}, and Bi^{5+}, a partial loss of these atoms' valence shell electrons can also lead to the formation of Tl^+, Sn^{2+}, Pb^{2+}, and Bi^{3+} ions. The formation of these 1+, 2+, and 3+ cations is ascribed to the inert pair effect, which reflects the relatively low energy of the valence *s*-electron pair for atoms of the heavy elements of groups 13, 14, and 15. Mercury (group 12) also exhibits an unexpected behavior: it forms a diatomic ion, (an ion formed from two mercury atoms, with an Hg-Hg bond), in addition to the expected monatomic ion Hg^{2+} (formed from only one mercury atom).

Transition and inner transition metal elements behave differently than main group elements. Most transition metal cations have 2+ or 3+ charges that result from the loss of their outermost *s* electron(s) first, sometimes followed by the loss of one or two *d* electrons from the next-to-outermost shell. For example, iron ($1s^22s^22p^63s^23p^63d^64s^2$) forms the ion Fe^{2+} ($1s^22s^22p^63s^23p^63d^6$) by the loss of the 4*s* electrons and the ion Fe^{3+} ($1s^22s^22p^63s^23p^63d^5$) by the loss of the 4*s* electrons and one of the 3*d* electrons. Although the *d* orbitals of the transition elements are—according to the Aufbau principle—the last to fill when building up electron

configurations, the outermost *s* electrons are the first to be lost when these atoms ionize. When the inner transition metals form ions, they usually have a 3+ charge, resulting from the loss of their outermost *s* electrons and a *d* or *f* electron.

Example 4.2.1: Determining the Electronic Structures of Cations

There are at least 14 elements categorized as "essential trace elements" for the human body. They are called "essential" because they are required for healthy bodily functions, "trace" because they are required only in small amounts, and "elements" in spite of the fact that they are really ions. Two of these essential trace elements, chromium and zinc, are required as Cr^{3+} and Zn^{2+}. Write the electron configurations of these cations.

Solution

First, write the electron configuration for the neutral atoms:

Zn: $[Ar]3d^{10}4s^2$

Cr: $[Ar]3d^5 4s^1$

Next, remove electrons from the highest energy orbital. For the transition metals, electrons are removed from the *s* orbital first and then from the *d* orbital. For the *p*-block elements, electrons are removed from the *p* orbitals and then from the *s* orbital. Zinc is a member of group 12, so it should have a charge of 2+, and thus loses only the two electrons in its *s* orbital. Chromium is a transition element and should lose its *s* electrons and then its *d* electrons when forming a cation. Thus, we find the following electron configurations of the ions:

Zn^{2+}: $[Ar]3d^{10}$

Cr^{3+}: $[Ar]3d^3$

Exercise 4.2.1

Potassium and magnesium are required in our diet. Write the electron configurations of the ions expected from these elements.

Answer

K^+: [Ar], Mg^{2+}: [Ne]

Electronic Structures of Anions

Most monatomic anions form when a neutral nonmetal atom gains enough electrons to completely fill its outer *s* and *p* orbitals, thereby reaching the electron configuration of the next noble gas. Thus, it is simple to determine the charge on such a negative ion: The charge is equal to the number of electrons that must be gained to fill the *s* and *p* orbitals of the parent atom. Oxygen, for example, has the electron configuration $1s^2 2s^2 2p^4$, whereas the oxygen anion has the electron configuration of the noble gas neon (Ne), $1s^2 2s^2 2p^6$. The two additional electrons required to fill the valence orbitals give the oxide ion the charge of 2– (O^{2-}).

Example 4.2.1: Determining the Electronic Structure of Anions

Selenium and iodine are two essential trace elements that form anions. Write the electron configurations of the anions.

Solution

Se^{2-}: $[Ar]3d^{10}4s^2 4p^6$

I^-: $[Kr]4d^{10}5s^2 5p^6$

Exercise 4.2.2

Write the electron configurations of a phosphorus atom and its negative ion. Give the charge on the anion.

Answer

P: $[Ne]3s^2 3p^3$; P^{3-}: $[Ne]3s^2 3p^6$

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4.3: 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 $4.3.1$ 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.

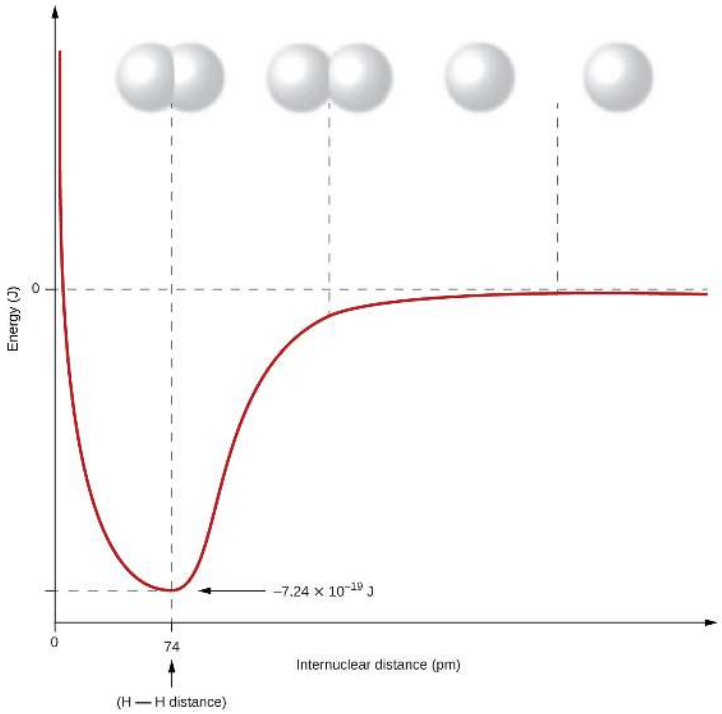


Figure 4.3.1: The potential energy of two separate hydrogen atoms (right) decreases as they approach each other, and the single electrons on each atom are shared to form a covalent bond. The bond length is the internuclear distance at which the lowest potential energy is achieved.

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 4.3.2 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 4.3.1, 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 4.3.2.

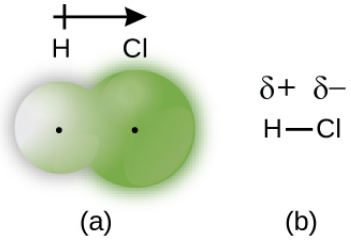


Figure 4.3.2: (a) The distribution of electron density in the HCl molecule is uneven. The electron density is greater around the chlorine nucleus. The small, black dots indicate the location of the hydrogen and chlorine nuclei in the molecule. (b) Symbols δ+ and δ– indicate the polarity of the H–Cl bond.

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 4.3.3 shows the electronegativity values of the elements as proposed by one of the most famous chemists of the twentieth century: Linus Pauling (Figure 4.3.4). 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.)

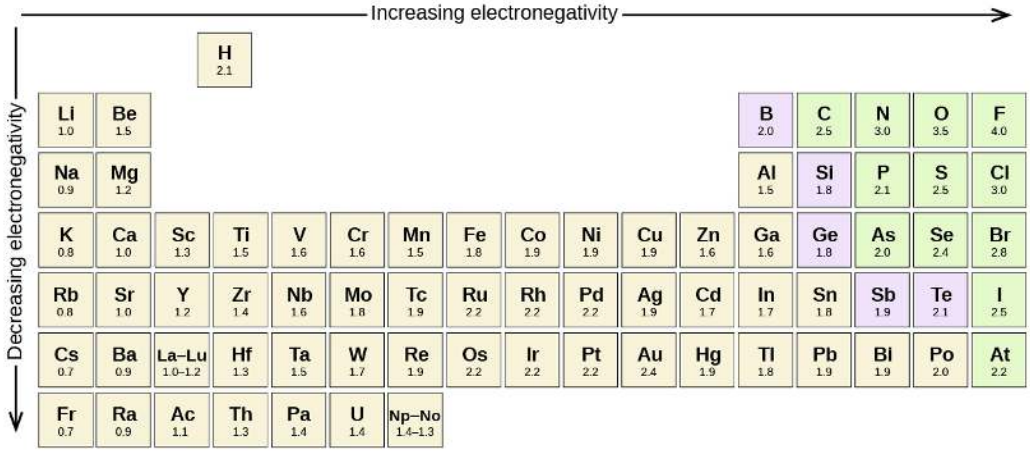


Figure 4.3.3: The electronegativity values derived by Pauling follow predictable periodic trends, with the higher electronegativities toward the upper right of the periodic table.

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 4.3.4, 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.

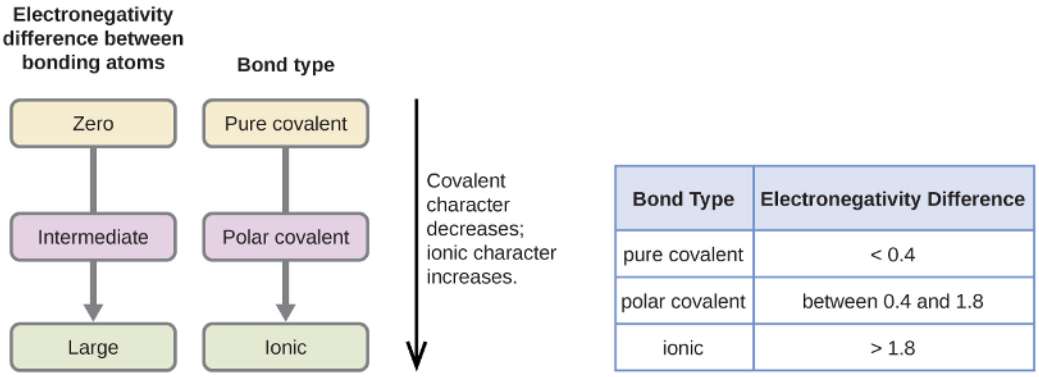


Figure 4.3.4: Linus Pauling (1901–1994) made many important contributions to the field of chemistry. He was also a prominent activist, publicizing issues related to health and nuclear weapons.

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 4.3.5 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 4.3.5: 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 4.3.5. 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 4.3.1: Electronegativity and Bond Polarity

Bond polarities play an important role in determining the structure of proteins. Using the electronegativity values in Figure 4.3.3, 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 4.3.1 shows these bonds in order of increasing polarity.

Table 4.3.1: 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 4.3.1

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 4.3.3, 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	

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4.4: 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 4.4.1.

Table 4.4.1: 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 4.4.2.

Table 4.4.2: 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 4.4.3. 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 4.4.3: 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 4.4.4.

Table 4.4.4: 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 4.4.5) 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 4.4.5: 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 4.4.1: 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 4.4.1

Write the formulas of the following ionic compounds:

a. chromium(III) phosphide
b. mercury(II) sulfide
c. manganese(II) phosphate

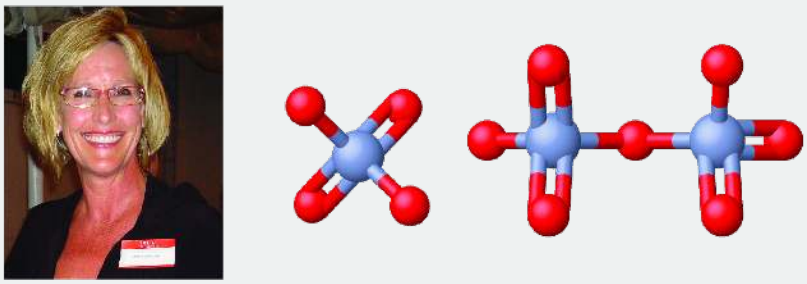
d. copper(I) oxide
e. iron(III) chloride dihydrate

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 4.4.1) 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 4.4.1: (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 4.4.5.

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 4.4.6.

Table 4.4.6: 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 4.4.2: 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 4.4.2

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 4.4.7.

Table 4.4.7: 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 4.4.8. 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 4.4.8: 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

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4.5: 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 4.5.1: 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 4.5.1: 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 4.5.2 demonstrates the use of Lewis symbols to show the transfer of electrons during the formation of ionic compounds.

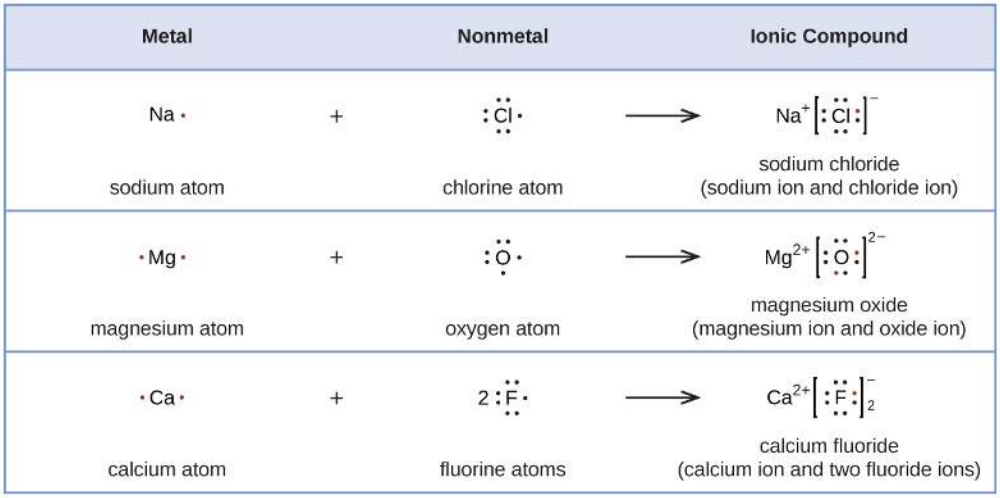
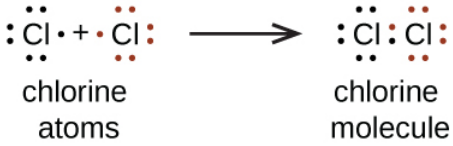


Figure 4.5.2: Cations are formed when atoms lose electrons, represented by fewer Lewis dots, whereas anions are formed by atoms gaining electrons. The total number of electrons does not change.

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:

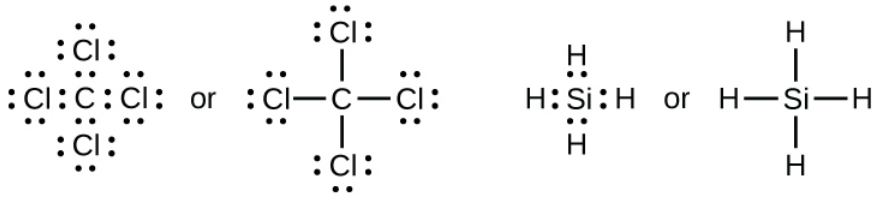


A single shared pair of electrons is called a single bond. Each Cl atom interacts with eight valence electrons: the six in the lone pairs and the two in the single bond.

The Octet Rule

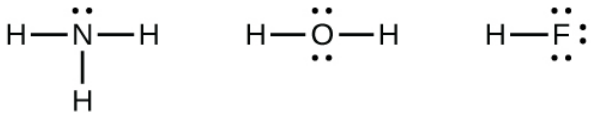
The other halogen molecules (F_2, Br_2, I_2, and At_2) form bonds like those in the chlorine molecule: one single bond between atoms and three lone pairs of electrons per atom. This allows each halogen atom to have a noble gas electron configuration. The tendency of main group atoms to form enough bonds 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

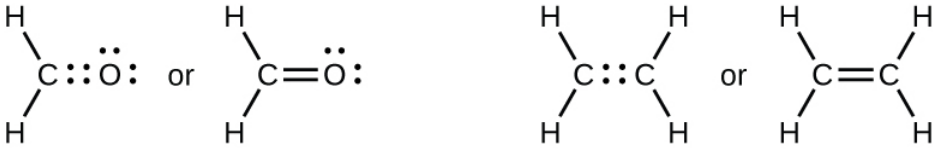
Group 15 elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To obtain an octet, these atoms form three covalent bonds, as in NH_3 (ammonia). Oxygen and other atoms in group 16 obtain an octet by forming two covalent bonds:



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

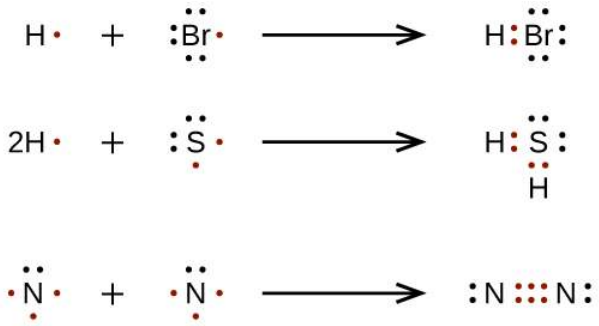
A triple bond forms when three electron pairs are shared by a pair of atoms, as in carbon monoxide (CO) and the cyanide ion (CN^-):



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:

$$\begin{array}{rr} & SiH_4 \\ \text{Si: 4 valence electrons/atom} \times \text{ 1 atom} & = 4 \\ + \quad \text{H : 1 valence electron/atom} \times \text{ 4 atoms} & = 4 \\ \hline & \text{total valence electrons} = 8 \end{array}$$

   - 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):

$$\begin{array}{rr} & CHO_2^- \\ \text{C: 4 valence electrons/atom} \times \text{ 1 atom} & = 4 \\ \text{H: 1 valence electron/atom} \times \text{ 1 atoms} & = 1 \\ \text{O: 6 valence electron/atom} \times \text{ 2 atoms} & = 12 \\ + \quad \text{1 electron} & = 1 \\ \hline & \text{total valence electrons} = 18 \end{array}$$

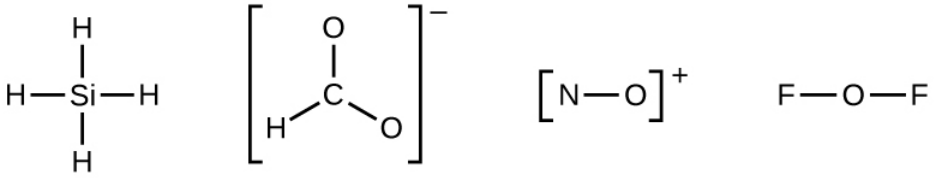
   - For a *positive ion*, such as NO^+, we add the number of valence electrons on the atoms in the ion and then subtract the number of positive charges on the ion (one electron is lost for each single positive charge) from the total number of valence electrons:

$$\begin{array}{rr} & NO^+ \\ \text{N: 5 valence electrons/atom} \times \text{ 1 atom} & = 5 \\ \text{O: 6 valence electron/atom} \times \text{ 1 atoms} & = 6 \\ + \quad \text{e: -1 electron (positive charge)} \times \text{ 1 electron} & = -1 \\ \hline & \text{total valence electrons} = 10 \end{array}$$

   - Since OF_2 is a neutral molecule, we simply add the number of valence electrons:

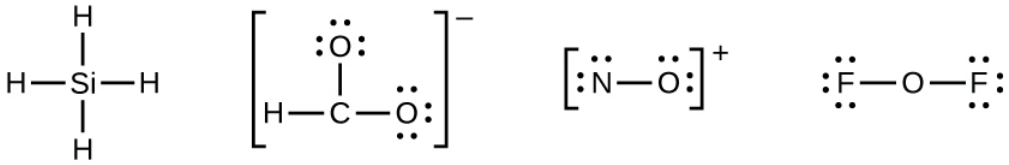
$$\begin{array}{rr} & \mathrm{OF_2} \\ \mathrm{O:\ 6\ valence\ electrons/atom} \times\ 1\ \mathrm{atom} & = 6 \\ +\quad \mathrm{F:\ 7\ valence\ electron/atom} \times\ 2\ \mathrm{atoms} & = 14 \\ \hline \mathrm{total\ valence\ electrons} & = 20 \end{array}$$

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:)

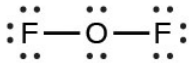


When several arrangements of atoms are possible, as for $\mathrm{CHO_2^-}$, we must use experimental evidence to choose the correct one. In general, the less electronegative elements are more likely to be central atoms. In $\mathrm{CHO_2^-}$, the less electronegative carbon atom occupies the central 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.

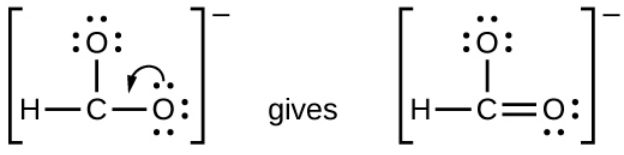
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons.
   - There are no remaining electrons on $\mathrm{SiH_4}$, so it is unchanged:



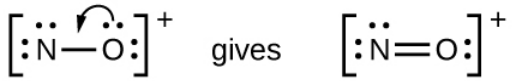
4. Place all remaining electrons on the central atom.
   - For $\mathrm{SiH_4}$, $\mathrm{CHO_2^-}$, and $\mathrm{NO^+}$, there are no remaining electrons; we already placed all of the electrons determined in Step 1.
   - For $\mathrm{OF_2}$, we had 16 electrons remaining in Step 3, and we placed 12, leaving 4 to be placed on the central atom:



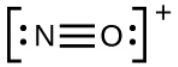
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.
   - $\mathrm{SiH_4}$: Si already has an octet, so nothing needs to be done.
   - $\mathrm{CHO2^-}$: We have distributed the valence electrons as lone pairs on the oxygen 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 4.5.1: Writing Lewis Structures

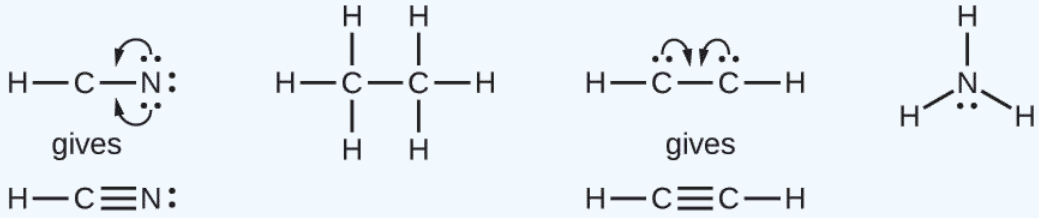
NASA's Cassini-Huygens mission detected a large cloud of toxic hydrogen cyanide (HCN) on Titan, one of Saturn's moons. Titan also contains ethane (H_3CCH_3), acetylene (HCCH), and ammonia (NH_3). What are the Lewis structures of these molecules?

Solution

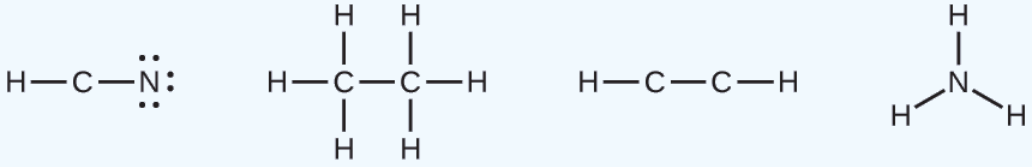
Step 1. *Calculate the number of valence electrons.*

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:

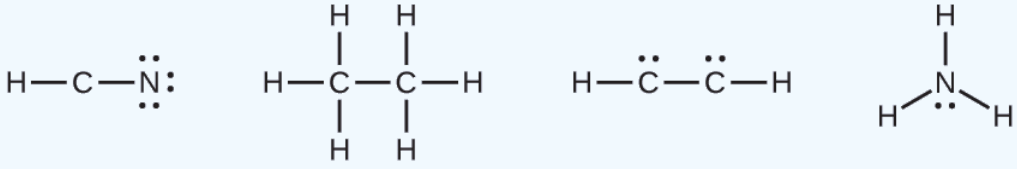


Step 3. *Where needed, distribute electrons to the terminal atoms:*



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

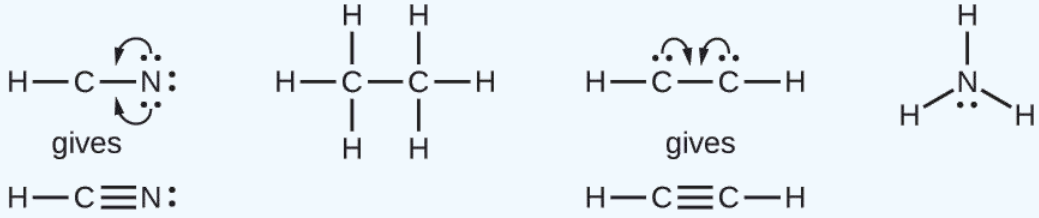
Step 4. *Where needed, place remaining electrons on the central atom:*



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 4.5.1

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 4.5.3), 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.

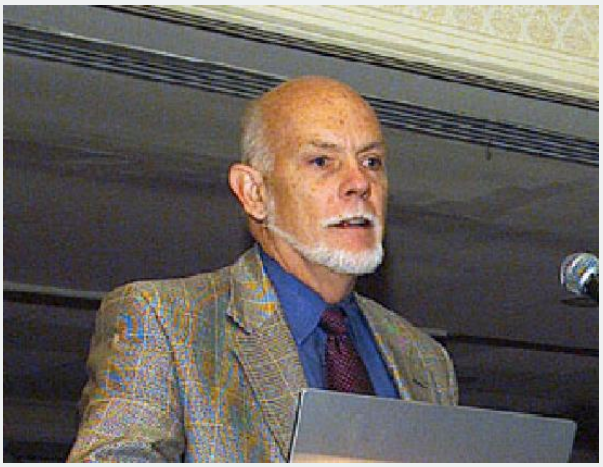


Figure 4.5.3: Richard Smalley (1943–2005), a professor of physics, chemistry, and astronomy at Rice University, was one of the leading advocates for fullerene chemistry. Upon his death in 2005, the US Senate honored him as the "Father of Nanotechnology." (credit: United States Department of Energy)

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:

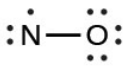
- Odd-electron molecules have an odd number of valence electrons, and therefore have an unpaired electron.
- Electron-deficient molecules have a central atom that has fewer electrons than needed for a noble gas configuration.
- Hypervalent molecules have a central atom that has more electrons than needed for a noble gas configuration.

Odd-electron Molecules

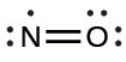
We call molecules that contain an odd number of electrons free radicals. Nitric oxide, NO, is an example of an odd-electron molecule; it is produced in internal combustion engines when oxygen and nitrogen react at high temperatures.

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:

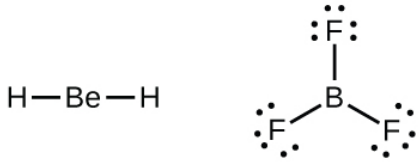


4. *Place all remaining electrons on the central atom.* Since there are no remaining electrons, this step does not apply.
5. *Rearrange the electrons to make multiple bonds with the central atom in order to obtain octets wherever possible.* We know that an odd-electron molecule cannot have an octet for 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:)

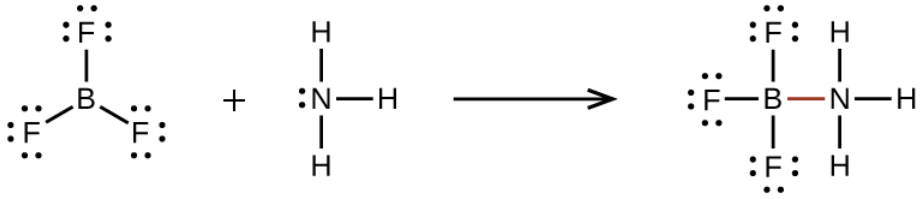


Electron-deficient Molecules

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 4.5.4: shows the Lewis structures for two hypervalent molecules, PCl_5 and SF_6.

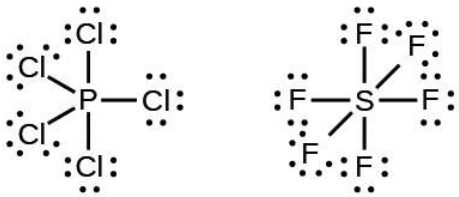
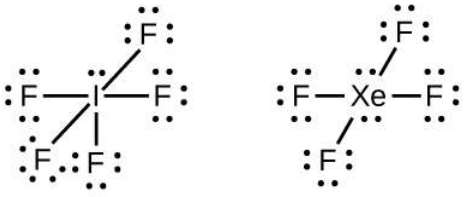


Figure 4.5.4: In PCl_5, the central atom phosphorus shares five pairs of electrons. In SF_6, sulfur shares six pairs of electrons.

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 4.5.2: 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?

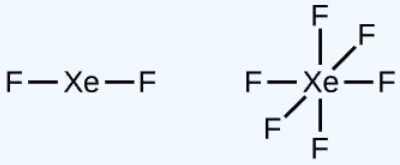
Solution

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:

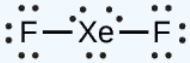
XeF_2: 8 + (2 × 7) = 22
XeF_6: 8 + (6 × 7) = 50

Step 2. Draw a skeleton joining the atoms by single bonds. Xenon will be the central atom because fluorine cannot be a central atom:

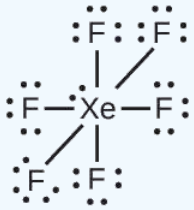


Step 3. Distribute the remaining electrons.

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:



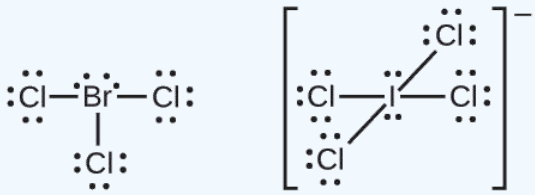
XeF_6: We place three lone pairs of electrons around each F atom, accounting for 36 electrons. Two electrons remain, and this lone pair is placed on the Xe atom:



Exercise 4.5.2: 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



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4.6: 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.

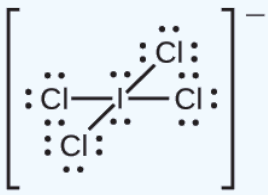
We must remember that the formal charge calculated for an atom is not the *actual* charge of the atom in the molecule. Formal charge is only a useful bookkeeping procedure; it does not indicate the presence of actual charges.

Example 4.6.1: 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 4.6.1

Calculate the formal charge for each atom in the carbon monoxide molecule:



Answer

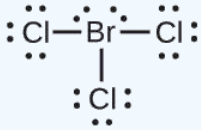
C −1, O +1

✓ Example 4.6.2: 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

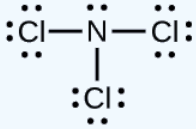
All atoms in $BrCl_3$ have a formal charge of zero, and the sum of the formal charges totals zero, as it must in a neutral molecule.

? Exercise 4.6.2

Determine the formal charge for each atom in NCl_3.

Answer

N: 0; all three Cl atoms: 0

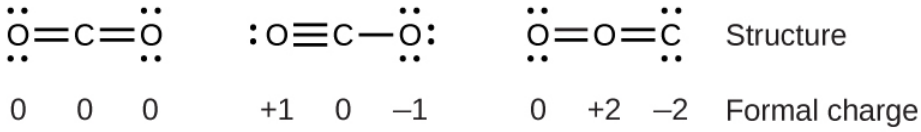


Using Formal Charge to Predict Molecular Structure

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:

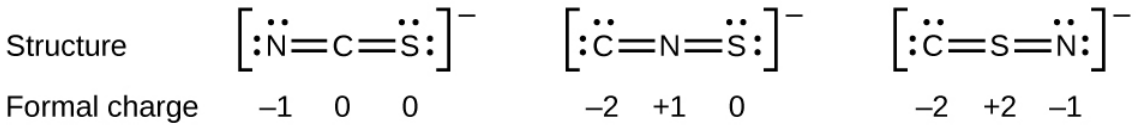
1. A molecular structure in which all formal charges are zero is preferable to one in which some formal charges are not zero.
2. If the Lewis structure must have nonzero formal charges, the arrangement with the smallest nonzero formal charges is preferable.
3. Lewis structures are preferable when adjacent 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).

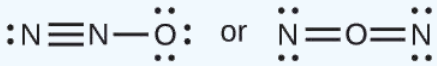
As another example, the thiocyanate ion, an ion formed from a carbon atom, a nitrogen atom, and a sulfur atom, could have three different molecular structures: NCS^-, CNS^-, or CSN^-. The formal charges present in each of these molecular structures can help us pick the most likely arrangement of atoms. 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).

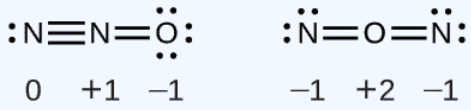
Example 4.6.3: Using Formal Charge to Determine Molecular Structure

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?

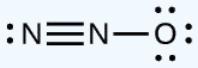


Solution

Determining formal charge yields the following:



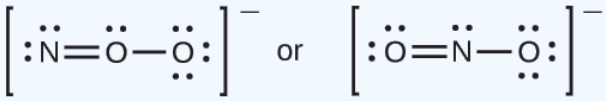
The structure with a terminal oxygen atom best satisfies the criteria for the most stable distribution of formal charge:



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 4.6.3

Which is the most likely molecular structure for the nitrite ion?



Answer

ONO^-

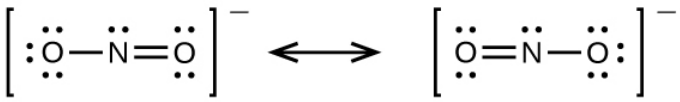
Resonance

Notice that the more likely structure for the nitrite anion in Example 4.6.3 may actually be drawn in two different ways, distinguished by the locations of the N-O and N=O bonds:



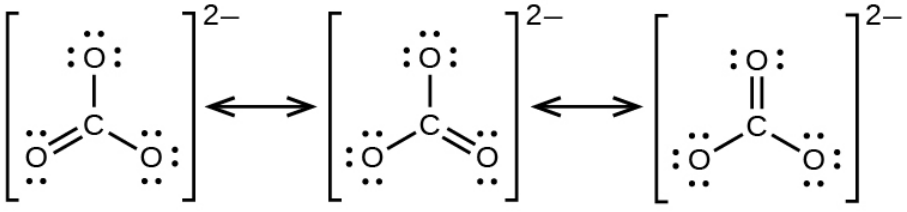
If nitrite ions do indeed contain a single and a double bond, we would expect for the two bond lengths to be different. A double bond between two atoms is shorter (and stronger) than a single bond between the same two atoms. Experiments show, however, that both N–O bonds in have the same strength and 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.

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4.7: 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 4.7.1). 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).

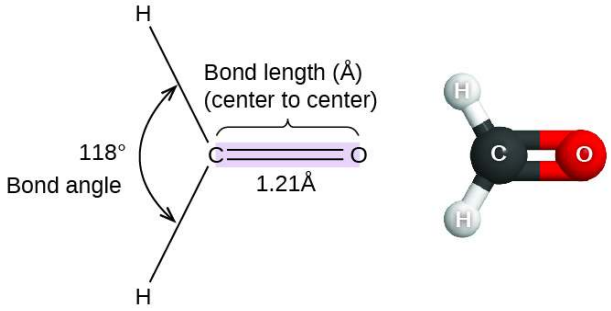


Figure 4.7.1: Bond distances (lengths) and angles are shown for the formaldehyde molecule, H_2CO.

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 4.7.2) 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 4.7.2).

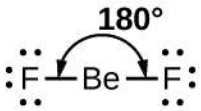


Figure 4.7.2: The BeF_2 molecule adopts a linear structure in which the two bonds are as far apart as possible, on opposite sides of the Be atom.

Figure 4.7.3 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.

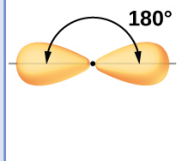
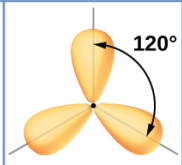
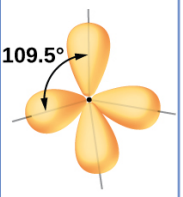
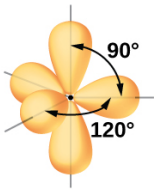
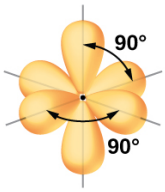
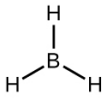
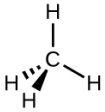
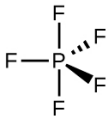
Number of regions	Two regions of high electron density (bonds and/or unshared pairs)	Three regions of high electron density (bonds and/or unshared pairs)	Four regions of high electron density (bonds and/or unshared pairs)	Five regions of high electron density (bonds and/or unshared pairs)	Six regions 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 4.7.3: 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 4.7.3 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.

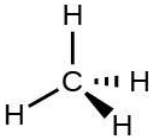


Figure 4.7.4: The molecular structure of the methane molecule, CH_4, is shown with a tetrahedral arrangement of the hydrogen atoms. VSEPR structures like this one are often drawn using the wedge and dash notation, in which solid lines represent bonds in the plane of the page, solid wedges represent bonds 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 4.7.4). 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 4.7.5).

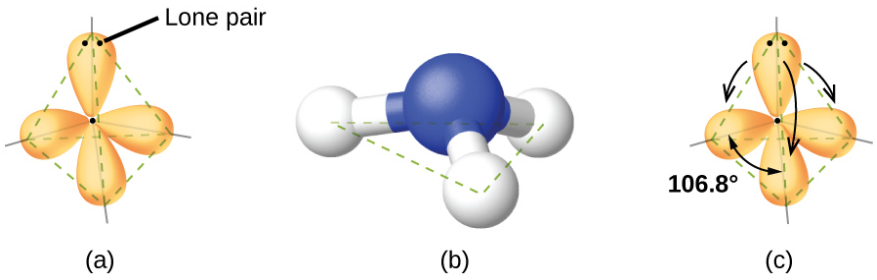


Figure 4.7.5: (a) The electron-pair geometry for the ammonia molecule is tetrahedral with one lone pair and three single bonds. (b) The trigonal pyramidal molecular structure is determined from the electron-pair geometry. (c) The actual bond angles deviate slightly from the idealized angles because the 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 4.7.5, small distortions from the ideal angles in Figure 4.7.3 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 4.7.1). 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 4.7.5) 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 4.7.3) because the lone pair-bonding pair repulsion is greater than the bonding pair-bonding pair repulsion (Figure 4.7.5). Figure 4.7.6 illustrates the ideal molecular structures, which are predicted based on the electron-pair geometries for various combinations of lone pairs and bonding pairs.

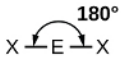
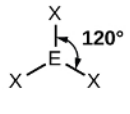
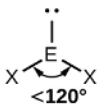
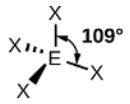
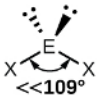
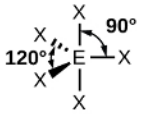
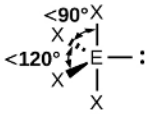
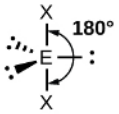
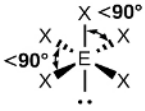
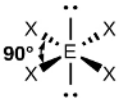
Number of electron regions	Electron region geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramid	 Square planar	 T-shape	 Linear

Figure 4.7.6: 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 4.7.6) 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 4.7.7 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 4.7.6, 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 4.7.7). The stable structure is the one that puts the lone pairs in equatorial locations, giving a T-shaped molecular structure.

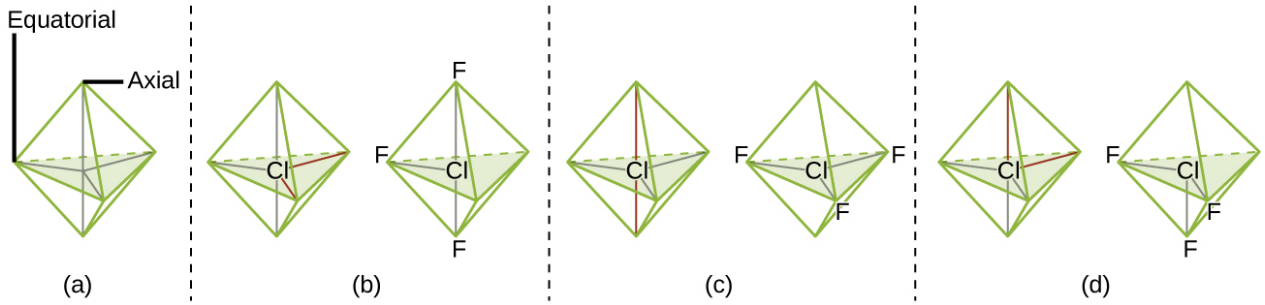


Figure 4.7.7: (a) In a trigonal bipyramid, the two axial positions are located directly across from one another, whereas the three equatorial positions are located in a triangular arrangement. (b–d) The two lone pairs (red lines) in ClF_3 have several possible arrangements, but the T-shaped molecular 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 4.7.6).

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 4.7.6, first column).
4. Use the number of lone pairs to determine the molecular structure (Figure 4.7.6). 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.

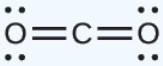
Example 4.7.1: Predicting Electron-pair Geometry and Molecular Structure: CO_2 and BCl_3

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

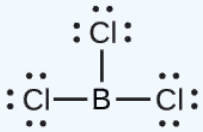
(a) We write the Lewis structure of CO_2 as:



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:

https://chem.libretexts.org/@go/page/533010



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 4.7.8).

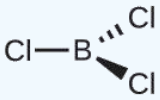


Figure 4.7.8:

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 4.7.1

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

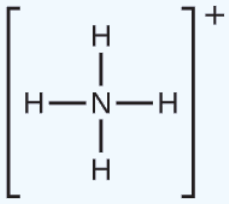
The electron-pair geometry is trigonal planar and the molecular structure is trigonal planar. Due to resonance, all three C–O bonds are identical. Whether they are single, double, or an average of the two, each bond counts as one region of electron density.

✓ Example 4.7.2: 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 4.7.6). Therefore, the electron pair geometry of Figure 4.7.9.

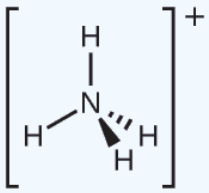


Figure 4.7.9: The ammonium ion displays a tetrahedral electron-pair geometry as well as a tetrahedral molecular structure.

Exercise 4.7.2

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.

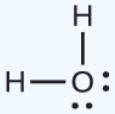
The next several examples illustrate the effect of lone pairs of electrons on molecular structure.

Example 4.7.3: 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 4.7.10), as indicated in Figure 4.7.6. 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°.

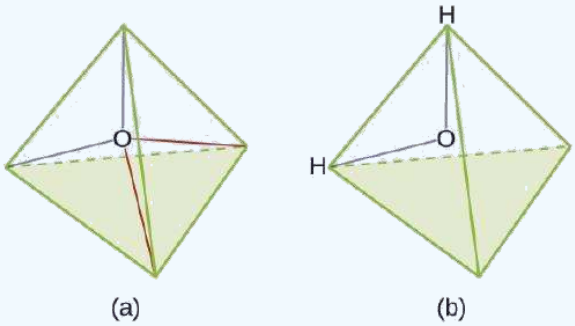


Figure 4.7.10: (a) H_2O has four regions of electron density around the central atom, so it has a tetrahedral electron-pair geometry. (b) Two of the electron regions are lone pairs, so the molecular structure is bent.

Exercise 4.7.3

The hydronium ion, H_3O^+, forms when acids are dissolved in water. Predict the electron-pair geometry and molecular structure of this cation.

Answer

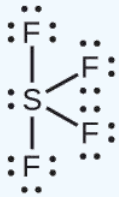
electron pair geometry: tetrahedral; molecular structure: trigonal pyramidal

Example 4.7.4: 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 4.7.11) is that of a seesaw (Figure 4.7.6).

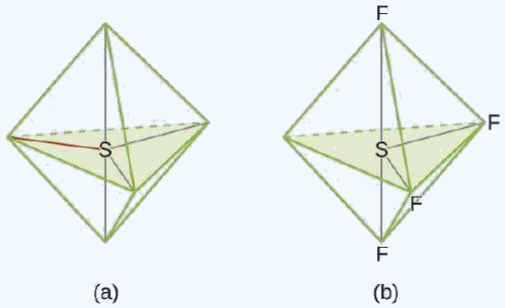


Figure 4.7.11: (a) SF4 has a trigonal bipyramidal arrangement of the five regions of electron density. (b) One of the regions is a lone pair, which results in a seesaw-shaped molecular structure.

Exercise 4.7.4

Predict the electron pair geometry and molecular structure for molecules of XeF_2.

Answer

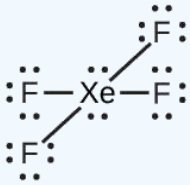
The electron-pair geometry is trigonal bipyramidal. The molecular structure is linear.

Example 4.7.5: 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 4.7.6), which is the electron-pair geometry. To minimize repulsions, the lone pairs should be on opposite sides of the central atom (Figure 4.7.12). The five atoms are all in the same plane and have a square planar molecular structure.

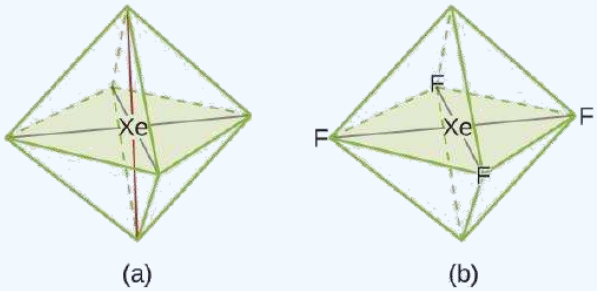


Figure 4.7.12: (a) XeF_4 adopts an octahedral arrangement with two lone pairs (red lines) and four bonds in the electron-pair geometry. (b) The molecular structure is square planar with the lone pairs directly across from one another.

Exercise 4.7.5

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

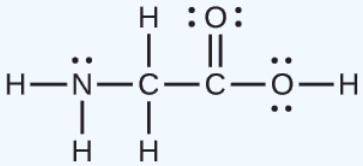
electron pair geometry: trigonal bipyramidal; molecular structure: linear

Molecular Structure for Multicenter Molecules

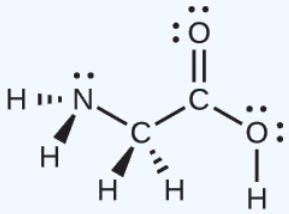
When a molecule or polyatomic ion has only one central atom, the molecular structure completely describes the shape of the molecule. Larger molecules do not have a single central atom, but are connected by a chain of interior atoms that each possess a "local" geometry. The way these local structures 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 4.7.6: 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:

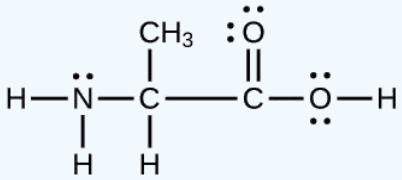
- nitrogen—four regions of electron density; tetrahedral
- carbon ($\underline{C}H_2$)—four regions of electron density; tetrahedral
- carbon ($\underline{C}O_2$)—three regions of electron density; trigonal planar
- oxygen ($\underline{O}H$)—four regions of electron density; tetrahedral

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 4.7.6

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 4.7.7: 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 4.7.7

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 4.7.13). 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.

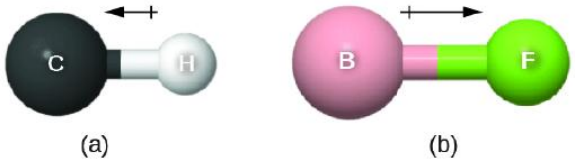


Figure 4.7.13: (a) There is a small difference in electronegativity between C and H, represented as a short vector. (b) The electronegativity difference between B and F is much larger, so the vector representing the bond moment is much longer.

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 4.7.14). 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 4.7.14), 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).

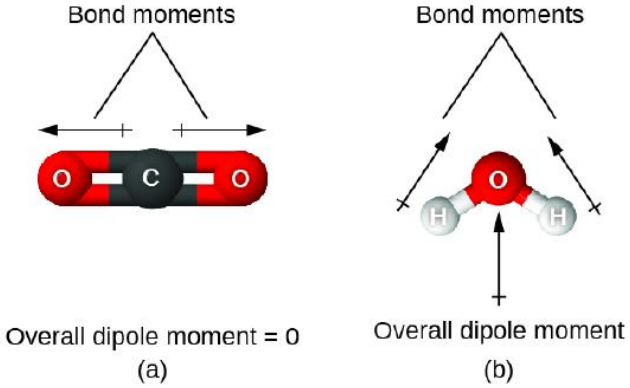
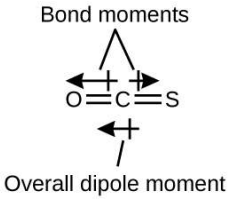


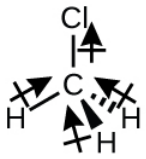
Figure 4.7.14: The overall dipole moment of a molecule depends on the individual bond dipole moments and how they are arranged. (a) Each CO bond has a bond dipole moment, but they point in opposite directions so that the net CO_2 molecule is nonpolar. (b) In contrast, water is polar because the OH bond 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:

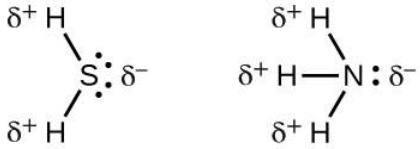


The C-O bond is considerably polar. Although C and S have very similar electronegativity values, S is slightly more electronegative than C, and so the C-S bond is just slightly polar. Because oxygen is more electronegative than sulfur, the oxygen end of the molecule is the negative end.

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 4.7.15). 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.

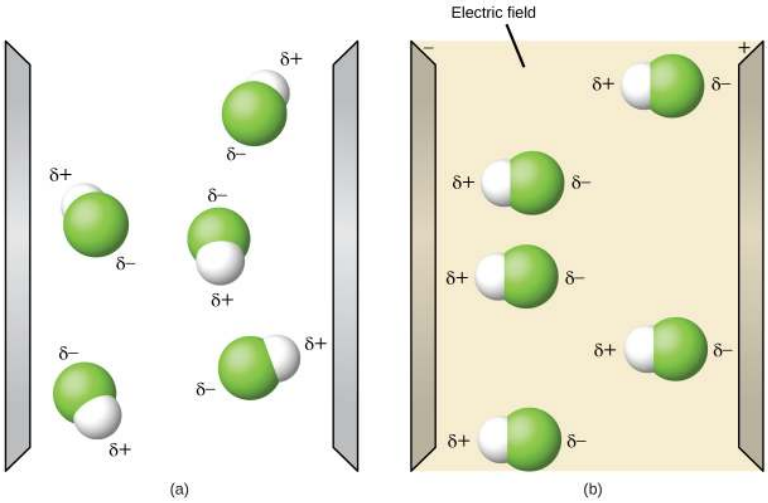


Figure 4.7.15: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction.

Link to Learning

The molecule polarity simulation provides many ways to explore dipole moments of bonds and molecules.

Example 4.7.8: 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 4.7.15

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 4.7.8

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.

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4.8: 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

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4.9: Key Equations

$$\text{formal charge} = \#\text{ valence shell electrons (free atom)} - \#\text{ lone pair electrons} - \tfrac{1}{2}\ \#\text{ bonding}$$

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4.10: 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.

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CHAPTER OVERVIEW

5: 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.

5.1: Introduction
5.2: Valence Bond Theory
5.3: Hybrid Atomic Orbitals
5.4: Multiple Bonds
5.5: Molecular Orbital Theory
5.6: Key Terms
5.7: Key Equations
5.8: Summary

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5.1: 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.

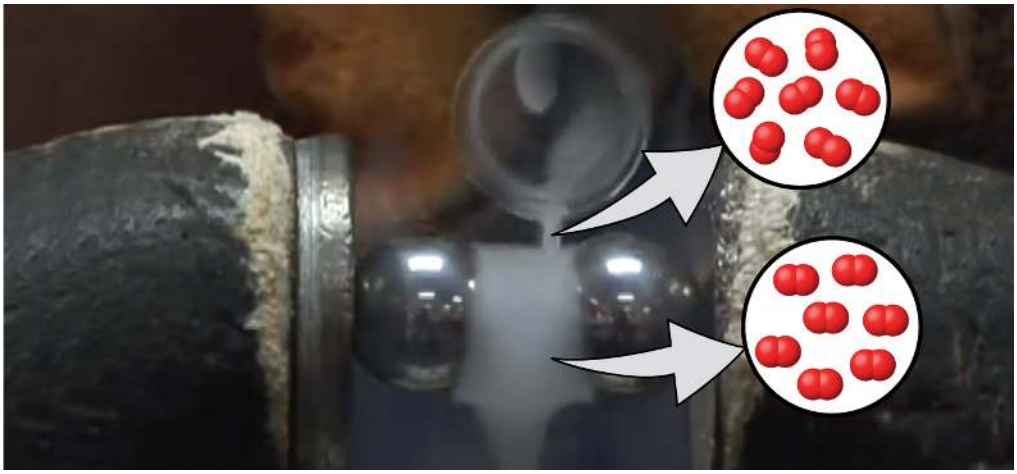


Figure 5.1.1: Oxygen molecules orient randomly most of the time, as shown in the top magnified view. However, when we pour liquid oxygen through a magnet, the molecules line up with the magnetic field, and the attraction allows them to stay suspended between the poles of the magnet where the magnetic 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 5.1.1) 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.



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5.2: 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.

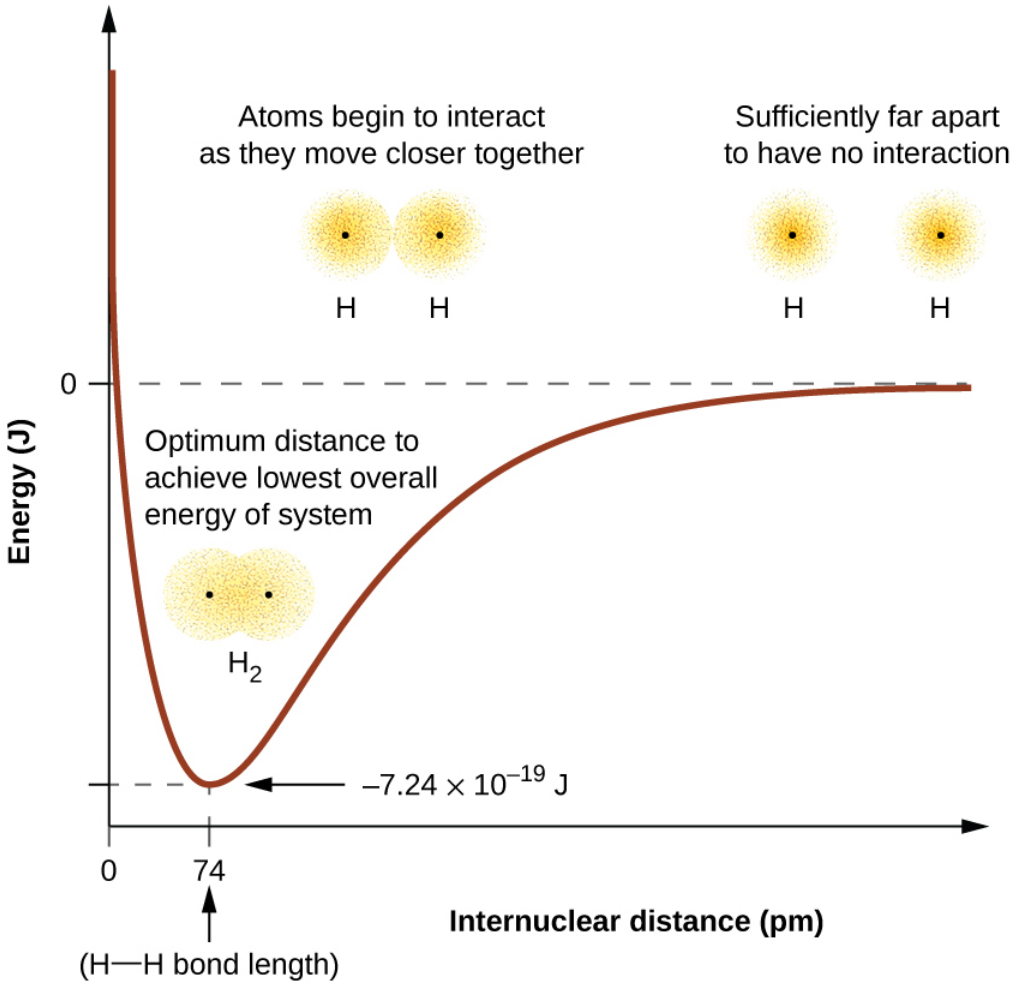


Figure 5.2 (a) The interaction of two hydrogen atoms changes as a function of distance. (b) The energy of the system changes as the atoms interact. The lowest (most stable) energy occurs at a distance of 74 pm, which is the bond length observed for the H_2 molecule.

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.

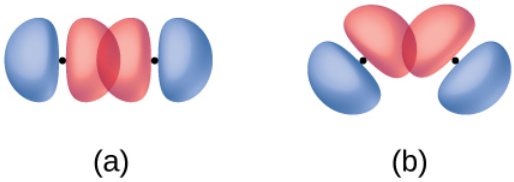


Figure 5.3 (a) The overlap of two *p* orbitals is greatest when the orbitals are directed end to end. (b) Any other arrangement results in less overlap. The dots indicate the locations of the nuclei.

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.

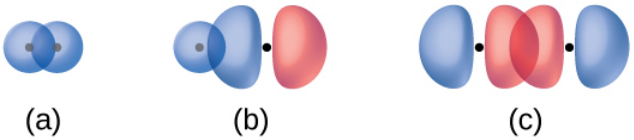


Figure 5.4 Sigma (σ) bonds form from the overlap of the following: (a) two *s* orbitals, (b) an *s* orbital and a *p* orbital, and (c) two *p* orbitals. The dots indicate the locations of the nuclei.

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.

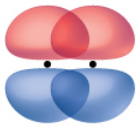
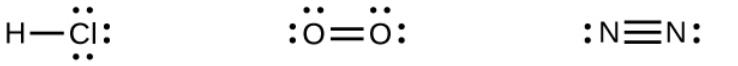


Figure 5.5 Pi (π) bonds form from the side-by-side overlap of two *p* orbitals. The dots indicate the location of the nuclei.

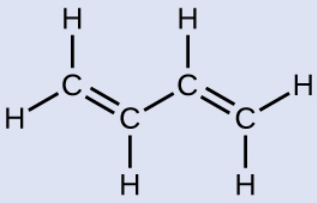
While all single bonds are σ bonds, multiple bonds consist of both σ and π bonds. As the Lewis structures below suggest, O_2 contains a double bond, and N_2 contains a triple bond. The double bond consists of one σ bond and one π bond, and the triple bond consists of one σ bond and two π bonds. Between 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	One σ bond	One σ bond
No π bonds	One π 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.

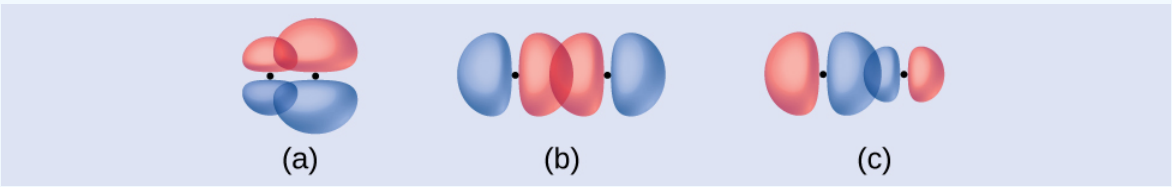
Check Your Learning

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.

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5.3: 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 5.3.1, 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.

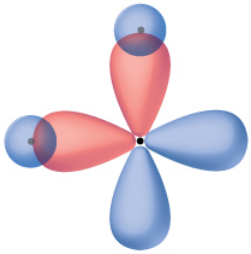


Figure 5.3.1: The hypothetical overlap of two of the 2*p* orbitals on an oxygen atom (red) with the 1*s* orbitals of two hydrogen atoms (blue) would produce a bond angle of 90°. This is not consistent with experimental evidence.[1]

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 5.3.2). 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.

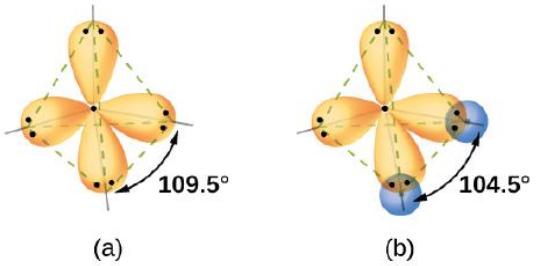


Figure 5.3.2: (a) A water molecule has four regions of electron density, so VSEPR theory predicts a tetrahedral arrangement of hybrid orbitals. (b) Two of the hybrid orbitals on oxygen contain lone pairs, and the other two overlap with the 1*s* orbitals of hydrogen atoms to form the O–H bonds in H_2O. 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 5.3.3). 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.

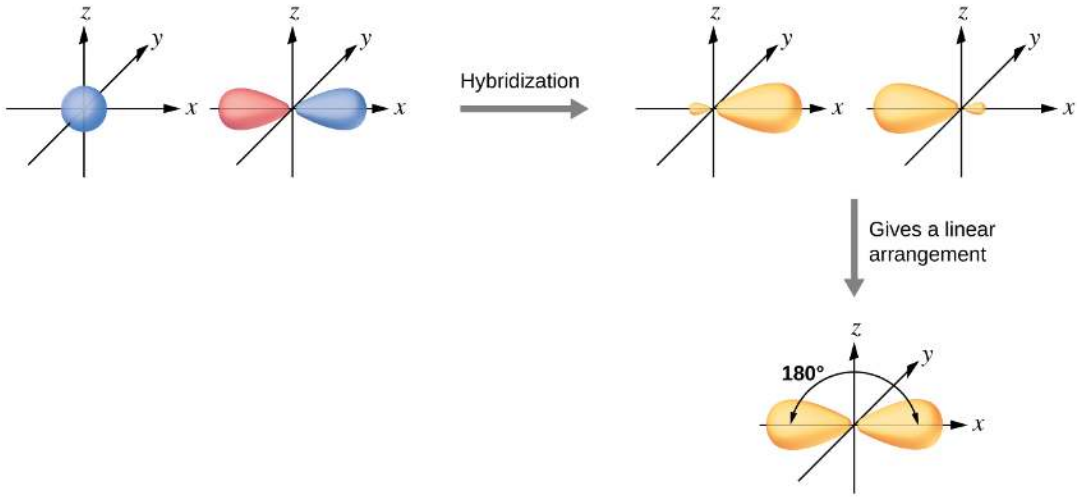


Figure 5.3.3: Hybridization of an *s* orbital (blue) and a *p* orbital (red) of the same atom produces two *sp* hybrid orbitals (yellow). Each hybrid orbital is oriented primarily in just one direction. Note that each *sp* orbital contains one lobe that is significantly larger than the other. The set 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 5.3.4. 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.

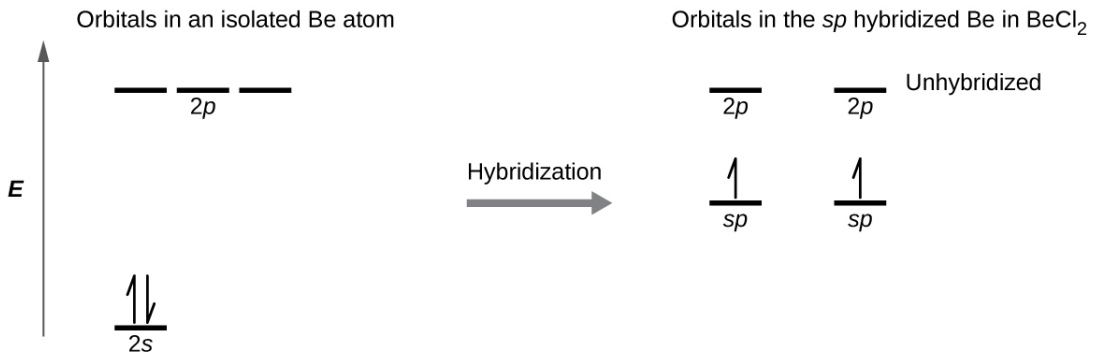


Figure 5.3.4: This orbital energy-level diagram shows the *sp* hybridized orbitals on Be in the linear $BeCl_2$ molecule. Each of the two *sp* hybrid orbitals holds one electron and is thus half filled and available for bonding via overlap with a Cl 3*p* orbital.

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 5.3.5).

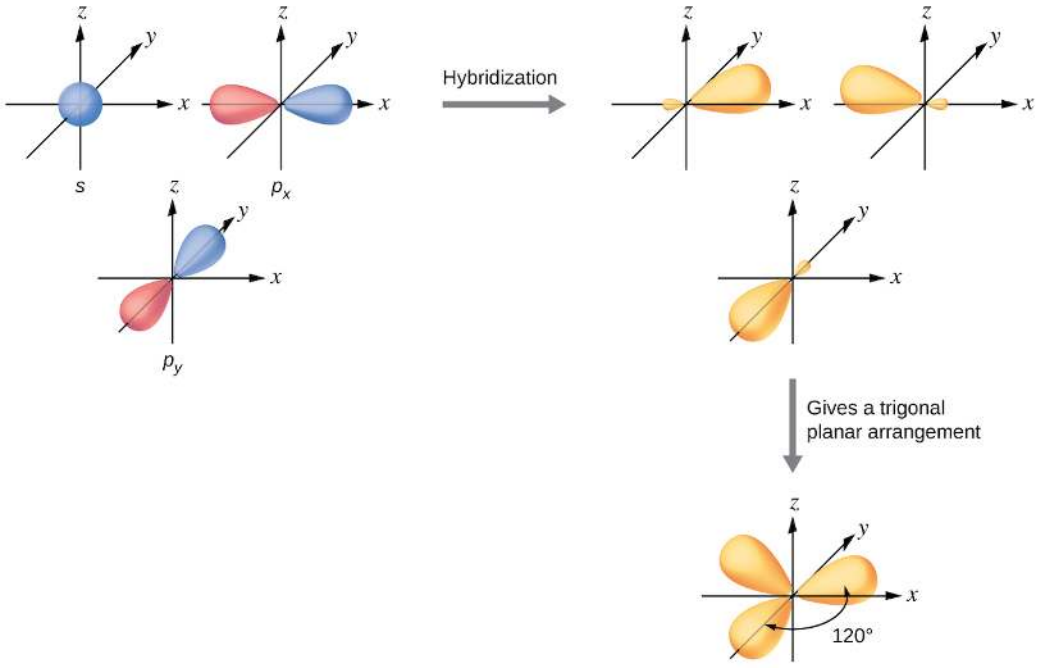


Figure 5.3.5: The hybridization of an *s* orbital (blue) and two *p* orbitals (red) produces three equivalent sp^2 hybridized orbitals (yellow) oriented at 120° with respect to each other. The remaining unhybridized *p* orbital is not shown here, but is located along the z axis.

Although quantum mechanics yields the "plump" orbital lobes as depicted in Figure 5.3.5, sometimes for clarity these orbitals are drawn thinner and without the minor lobes, as in Figure 5.3.6, 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.

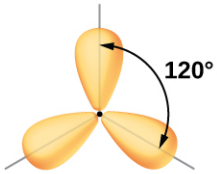


Figure 5.3.6: This alternate way of drawing the trigonal planar sp^2 hybrid orbitals is sometimes used in more crowded figures.

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 5.3.7). 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 5.3.8. 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.

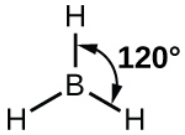


Figure 5.3.7: BH_3 is an electron-deficient molecule with a trigonal planar structure.

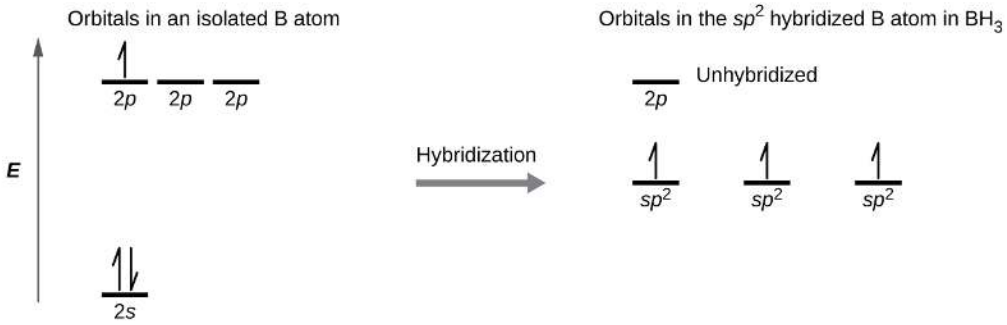


Figure 5.3.8: In an isolated B atom, there are one 2*s* and three 2*p* valence orbitals. When boron is in a molecule with three regions of electron density, three of the orbitals hybridize and create a set of three sp^2 orbitals and one unhybridized 2*p* orbital. The three half-filled hybrid orbitals 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 5.3.9), 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.

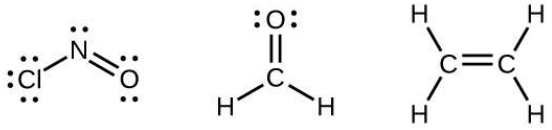


Figure 5.3.9: The central atom(s) in each of the structures shown contain three regions of electron density and are sp^2 hybridized. As we know from the discussion of VSEPR theory, a region of electron density contains all of the electrons that point in one direction. A lone pair, an unpaired electron, 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 5.3.10). Each of these hybrid orbitals points toward a different corner of a tetrahedron.

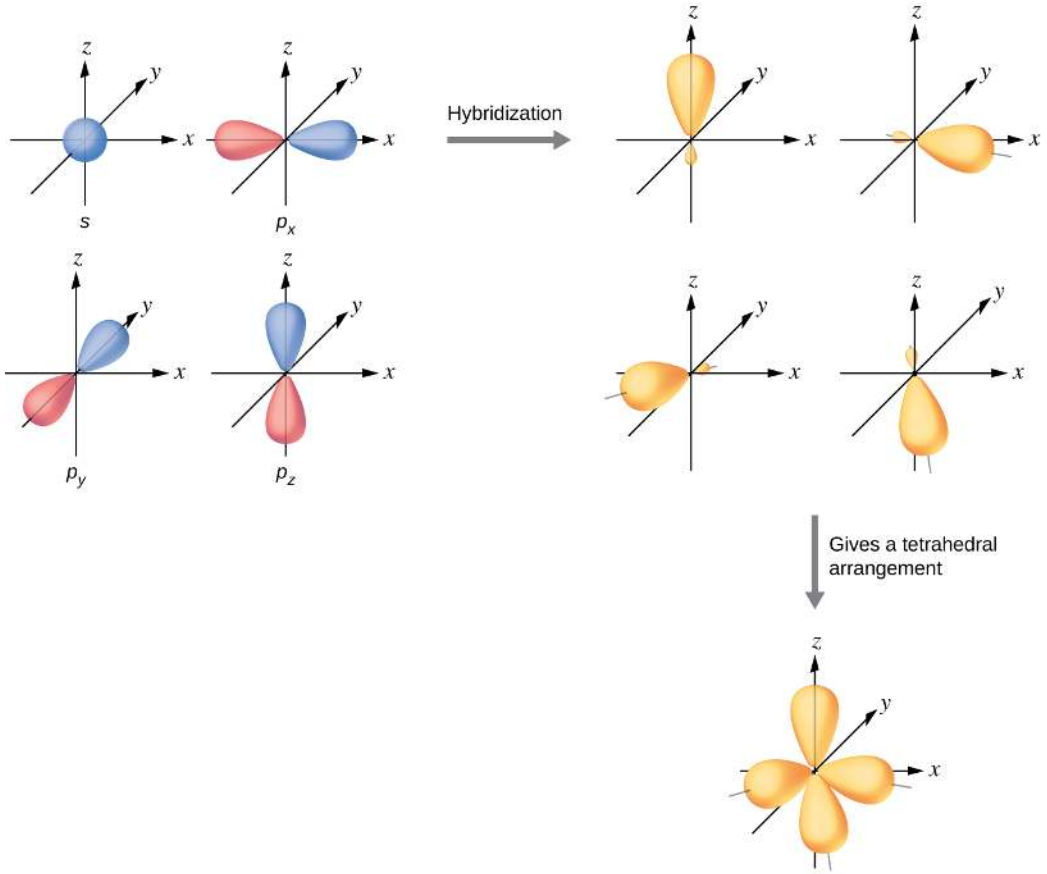


Figure 5.3.10: The hybridization of an *s* orbital (blue) and three *p* orbitals (red) produces four equivalent sp^3 hybridized orbitals (yellow) oriented at 109.5° with respect to each other.

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 5.3.11. 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.

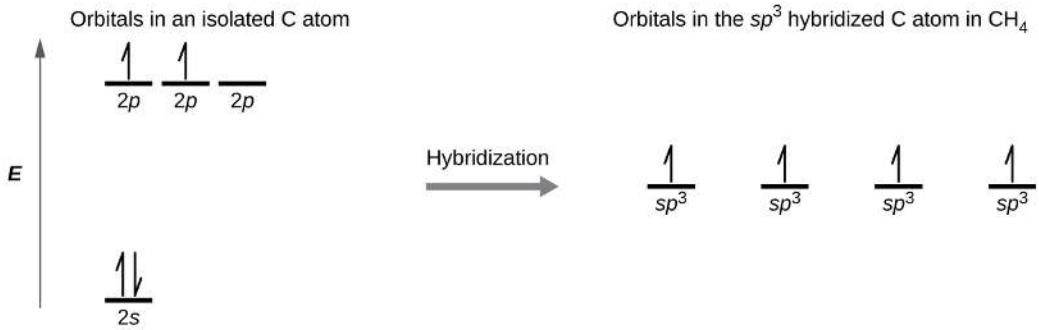


Figure 5.3.11: The four valence atomic orbitals from an isolated carbon atom all hybridize when the carbon bonds in a molecule like CH_4 with four regions of electron density. This creates four equivalent sp^3 hybridized orbitals. Overlap of each of the hybrid orbitals with a hydrogen orbital creates 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 5.3.12). 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 5.3.12. 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.

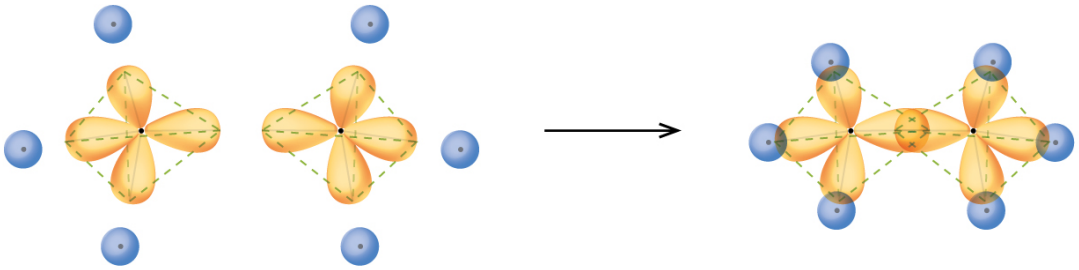


Figure 5.3.12: (a) In the ethane molecule, C_2H_6, each carbon has four sp^3 orbitals. (b) These four orbitals overlap to form seven σ bonds.

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).

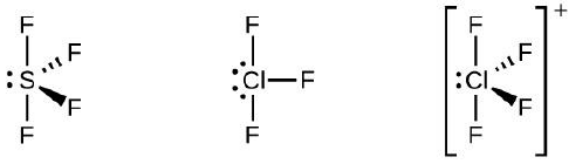


Figure 5.3.13: The three compounds pictured exhibit sp^3d hybridization in the central atom and a trigonal bipyramid form. SF_4 and have one lone pair of electrons on the central atom, and ClF_3 has two lone pairs giving it the T-shape shown.

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 5.3.14) 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.)

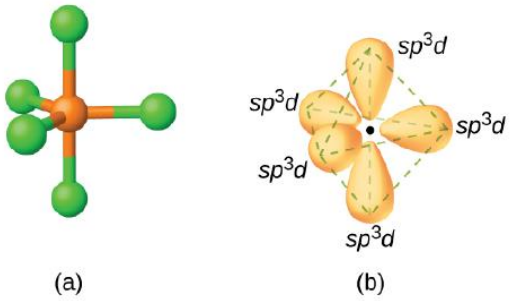


Figure 5.3.14: (a) The five regions of electron density around phosphorus in PCl_5 require five hybrid sp^3d orbitals. (b) These orbitals combine to form a trigonal bipyramidal structure with each large lobe of the hybrid orbital pointing at a vertex. As before, there are also small lobes pointing in 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.

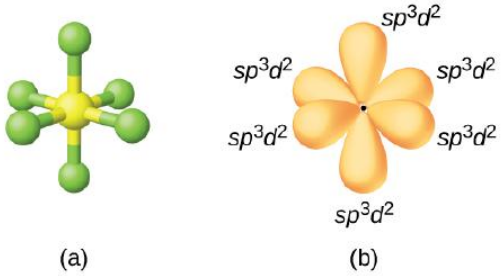


Figure 5.3.15: (a) Sulfur hexafluoride, SF_6, has an octahedral structure that requires sp^3d^2 hybridization. (b) The six sp^3d^2 orbitals form an octahedral structure around sulfur. Again, the minor lobe of each orbital is not shown for clarity.

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 5.3.16. 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 5.3.16 that corresponds to this geometry.

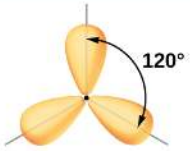
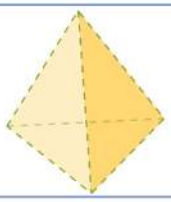
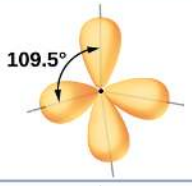
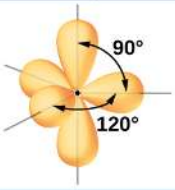
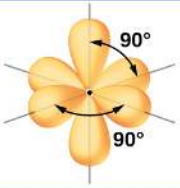
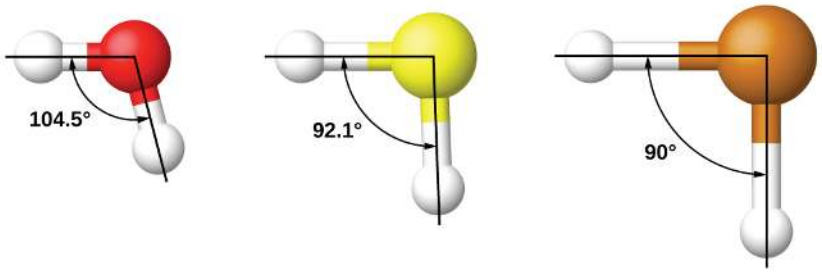
Regions of Electron Density	Arrangement		Hybridization	
2		linear	*sp*	
3		trigonal planar	sp^2	
4		tetrahedral	sp^3	
5		trigonal bipyramidal	sp^3d	
6		octahedral	sp^3d^2	

Figure 5.3.16: 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.

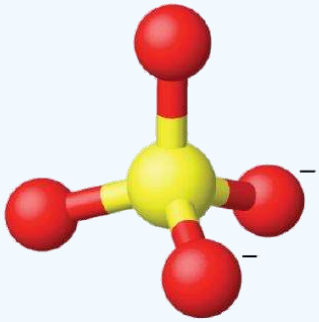


Example 5.3.1: Assigning Hybridization

Ammonium sulfate is important as a fertilizer. What is the hybridization of the sulfur atom in the sulfate ion,

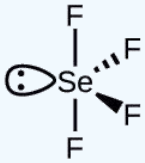
Solution

The Lewis structure of sulfate shows there are four regions of electron density. The hybridization is sp^3.



Exercise 5.3.1

What is the hybridization of the selenium atom in SeF_4?



Answer

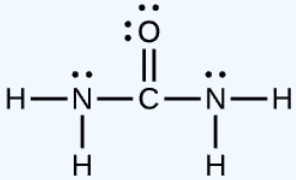
The selenium atom is sp^3d hybridized.

Example 5.3.2: 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

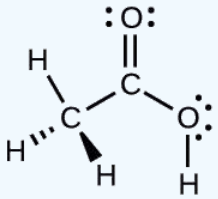
The Lewis structure of urea is



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 5.3.16), which is the hybridization of the carbon atom in urea.

Exercise 5.3.1

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.

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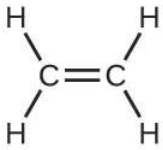
5.4: Multiple Bonds

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 5.4.1).

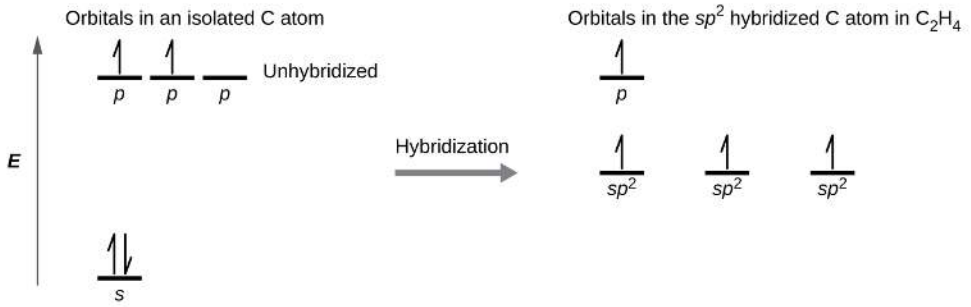


Figure 5.4.1: In ethene, each carbon atom is sp^2 hybridized, and the sp^2 orbitals and the *p* orbital are singly occupied. The hybrid orbitals overlap to form σ bonds, while the *p* orbitals on each carbon atom overlap to form a π bond.

These orbitals form the C–H single bonds and the σ bond in the Figure 5.4.2. 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 5.4.2) 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 5.4.2) and form a π bond.

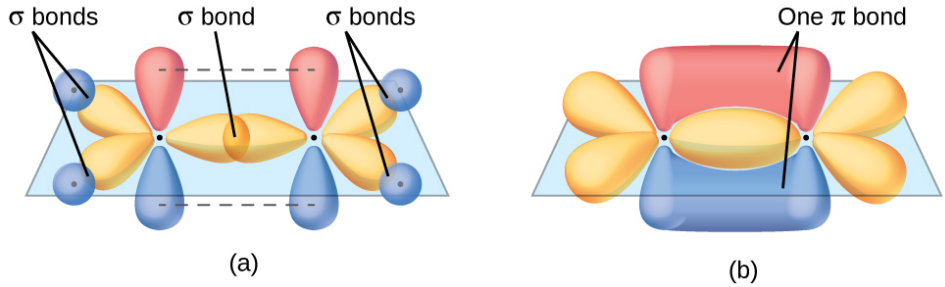


Figure 5.4.2: In the ethene molecule, C_2H_4, there are (a) five σ bonds. One C–C σ bond results from overlap of sp^2 hybrid orbitals on the carbon atom with one sp^2 hybrid orbital on the other carbon atom. Four C–H bonds result from the overlap between the C atoms' sp^2 orbitals with *s* orbitals on 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 5.4.3). We find this situation in acetylene, $\mathrm{H-C\equiv C-H}$ (Figure 5.4.4), 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.

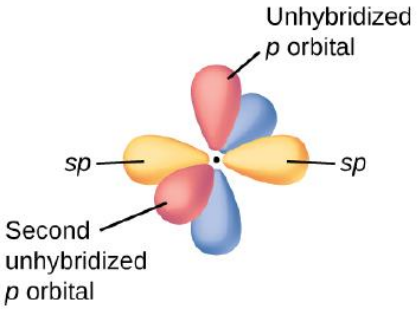


Figure 5.4.3: Diagram of the two linear *sp* hybrid orbitals of a carbon atom, which lie in a straight line, and the two unhybridized *p* orbitals at perpendicular angles.

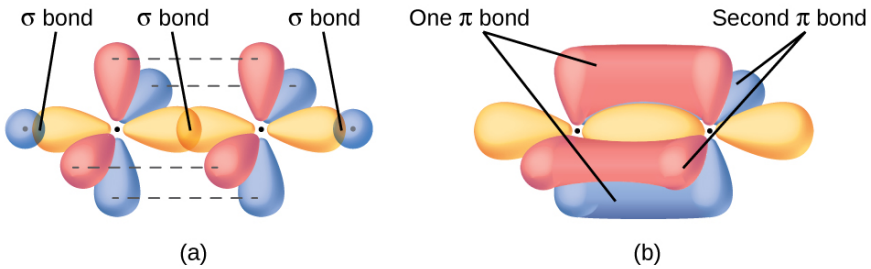


Figure 5.4.4: (a) In the acetylene molecule, C_2H_2, there are two C–H σ bonds and a triple bond involving one C–C σ bond and two C–C π bonds. The dashed lines, each connecting two lobes, indicate the side-by-side overlap of the four unhybridized *p* orbitals. (b) This shows the overall outline of the 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 5.4.5). 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.)

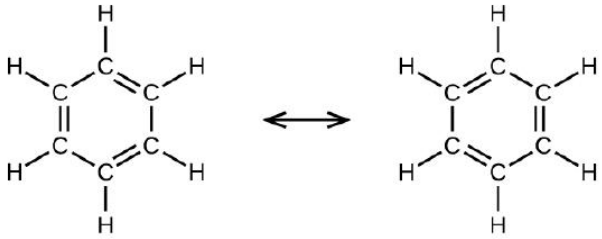


Figure 5.4.5: Each carbon atom in benzene, C_6H_6, is sp^2 hybridized, independently of which resonance form is considered. The electrons in the π bonds are not located in one set of *p* orbitals or the other, but rather delocalized throughout the molecule.

Example 5.4.1: 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 5.4.1

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

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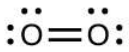
5.5: Molecular Orbital Theory

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 5.5.1), 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.

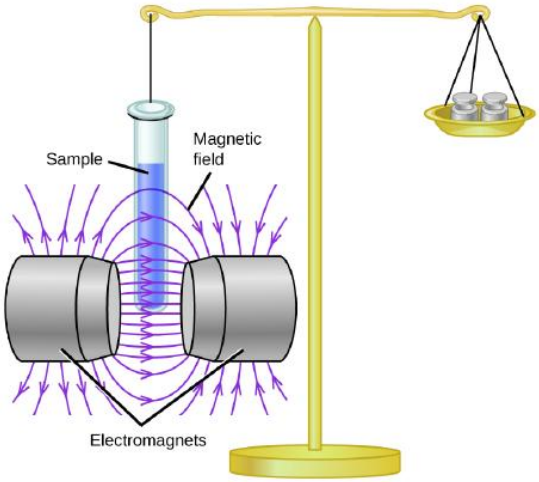


Figure 5.5.1: A Gouy balance compares the mass of a sample in the presence of a magnetic field with the mass with the electromagnet turned off to determine the number of unpaired electrons in a sample.

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 5.5.1 summarizes the main points of the two complementary bonding theories. Both theories provide different, useful ways of describing molecular structure.

Table 5.5.1: 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 5.5.2). 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.

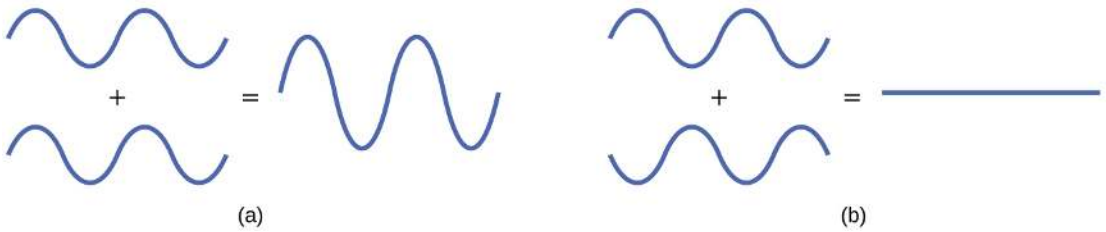


Figure 5.5.2: (a) When in-phase waves combine, constructive interference produces a wave with greater amplitude. (b) When out-of-phase waves combine, destructive interference produces a wave with less (or no) amplitude.

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 5.5.3. 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.

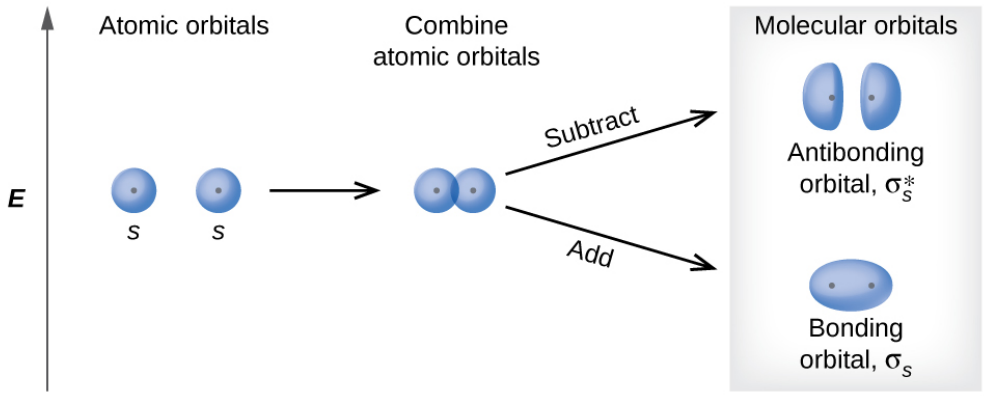


Figure 5.5.3: Sigma (σ) and sigma-star (σ*) molecular orbitals are formed by the combination of two *s* atomic orbitals. The dots (·) indicate the locations of nuclei.

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 5.5.4). 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.

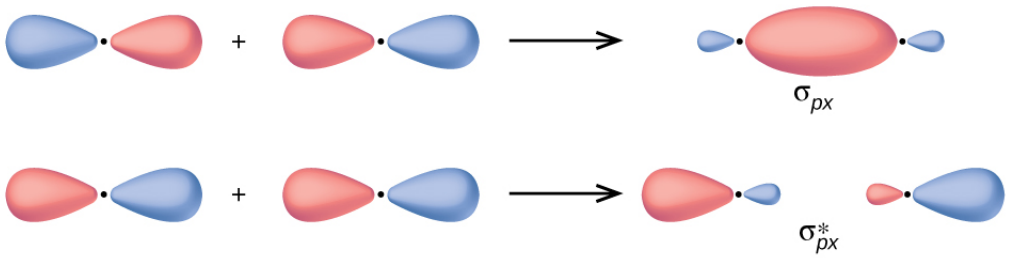


Figure 5.5.4: Combining wave functions of two *p* atomic orbitals along the internuclear axis creates two molecular orbitals, σ_p and σ_p^*.

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 5.5.5. 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.

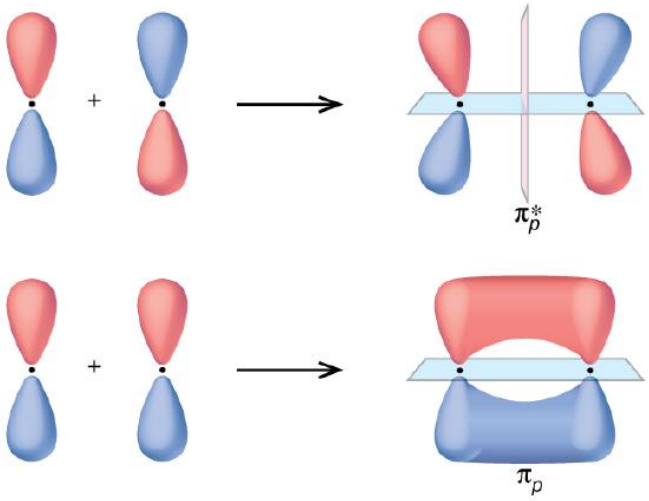
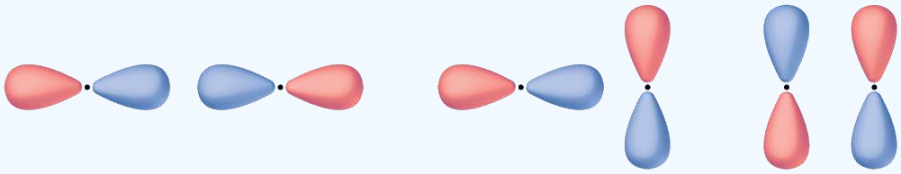


Figure 5.5.5: Side-by-side overlap of each two *p* orbitals results in the formation of two π molecular orbitals. Combining the out-of-phase orbitals results in an antibonding molecular orbital with two nodes. One contains the internuclear axis, and one is perpendicular to the axis. Combining the in-phase 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 5.5.1: 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.



$3p_x$ and $3p_x$ (a) $3p_x$ and $3p_y$ (b) $3p_y$ and $3p_y$ (c)

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 5.5.1

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 5.5.6) 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.

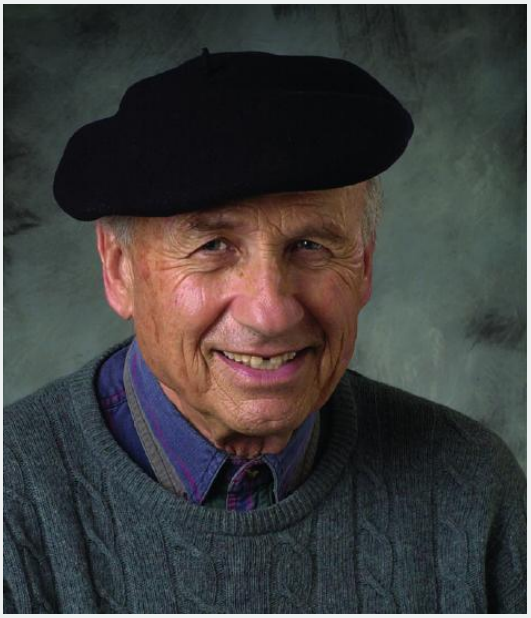


Figure 5.5.6: Walter Kohn developed methods to describe molecular orbitals. (credit: image courtesy of Walter Kohn)

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 5.5.7). 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.

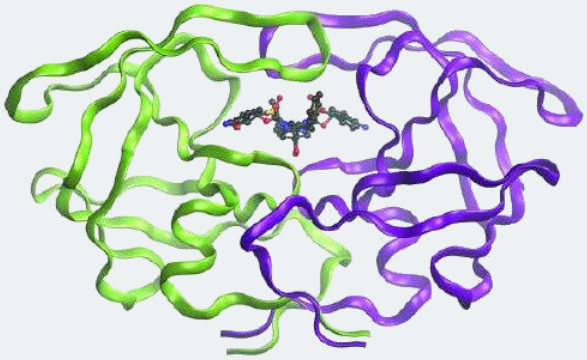


Figure 5.5.7: The molecule shown, HIV-1 protease, is an important target for pharmaceutical research. By designing molecules that bind to this protein, scientists are able to drastically inhibit the progress of the disease.

Molecular Orbital Energy Diagrams

The relative energy levels of atomic and molecular orbitals are typically shown in a molecular orbital diagram (Figure 5.5.8). 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 5.5.8). 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.

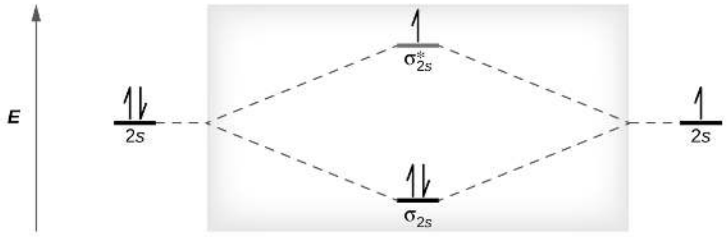


Figure 5.5.8: This is the molecular orbital diagram for the homonuclear diatomic showing the molecular orbitals of the valence shell only. The molecular orbitals are filled in the same manner as atomic orbitals, using the Aufbau principle and Hund's rule.

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 5.5.9) in which a single upward arrow indicates one electron in an orbital, and two (upward and downward) arrows indicate two electrons of opposite spin.

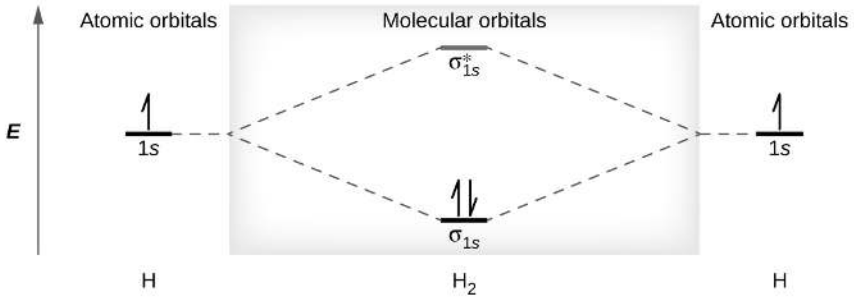


Figure 5.5.9: The molecular orbital energy diagram predicts that H_2 will be a stable molecule with lower energy than the separated atoms.

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.

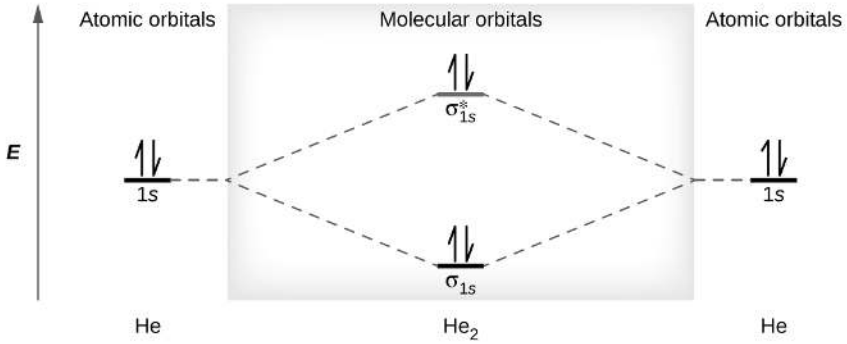


Figure 5.5.10: The molecular orbital energy diagram predicts that He_2 will not be a stable molecule, since it has equal numbers of bonding and antibonding electrons.

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 5.5.2).

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 5.5.11. 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.

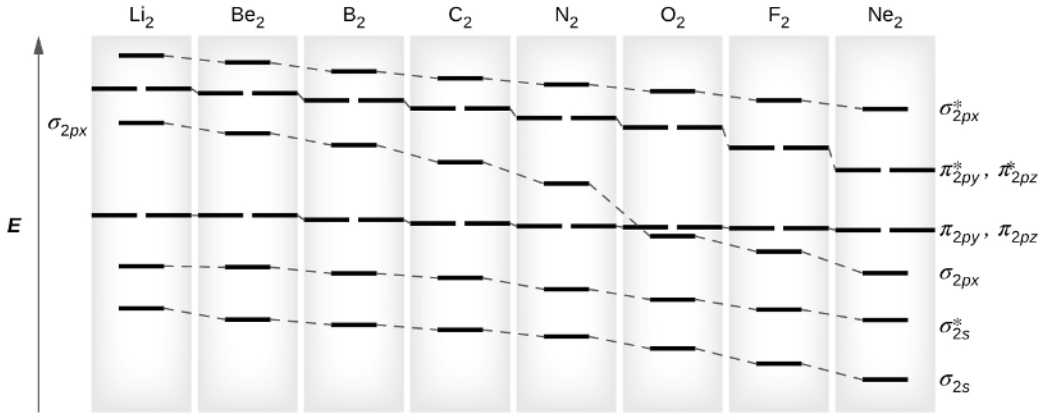


Figure 5.5.11: This shows the MO diagrams for each homonuclear diatomic molecule in the second period. The orbital energies decrease across the period as the effective nuclear charge increases and atomic radius decreases. Between N_2 and O_2, the order of the orbitals changes.

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 5.5.12). Similarly, the antibonding orbitals also undergo s-p mixing, with the σ_{s*} becoming more stable and the σ_{p*} becoming less stable.

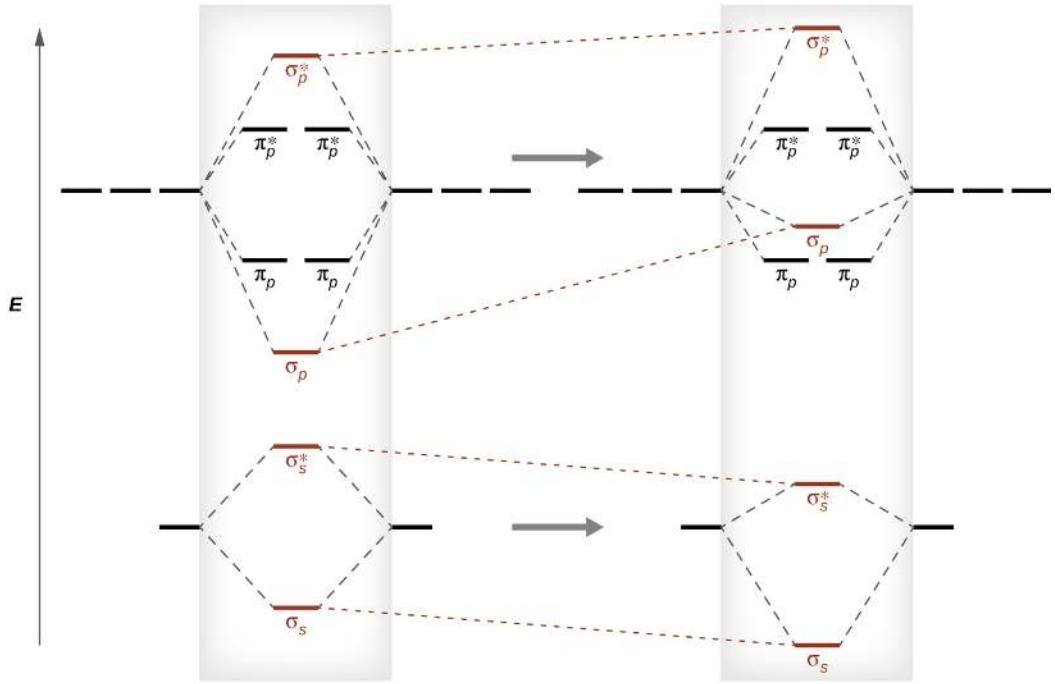


Figure 5.5.12: Without mixing, the MO pattern occurs as expected, with the σ_p orbital lower in energy than the π_p orbitals. When s-p mixing occurs, the orbitals shift as shown, with the σ_p orbital higher in energy than the π_p orbitals.

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 5.5.11. 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 5.5.11: , 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 5.5.2, Be_2 and Ne_2 molecules would have a bond order of 0, and these molecules do not exist.

Table 5.5.2: 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 5.5.2 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 5.5.3). 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 5.5.13 shows the bands for three important classes of materials: insulators, semiconductors, and conductors.

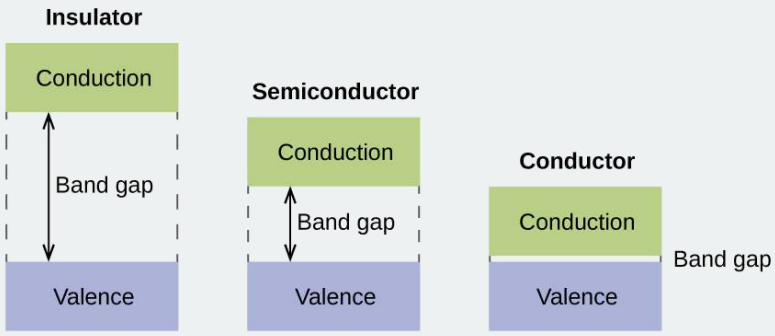


Figure 5.5.13: Molecular orbitals in solids are so closely spaced that they are described as bands. The valence band is lower in energy and the conduction band is higher in energy. The type of solid is determined by the size of the "band gap" between the valence and conduction bands. Only a very small 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 5.5.2: 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 5.5.11. Each oxygen atom contributes six electrons, so the diagram appears as shown in Figure 5.5.14

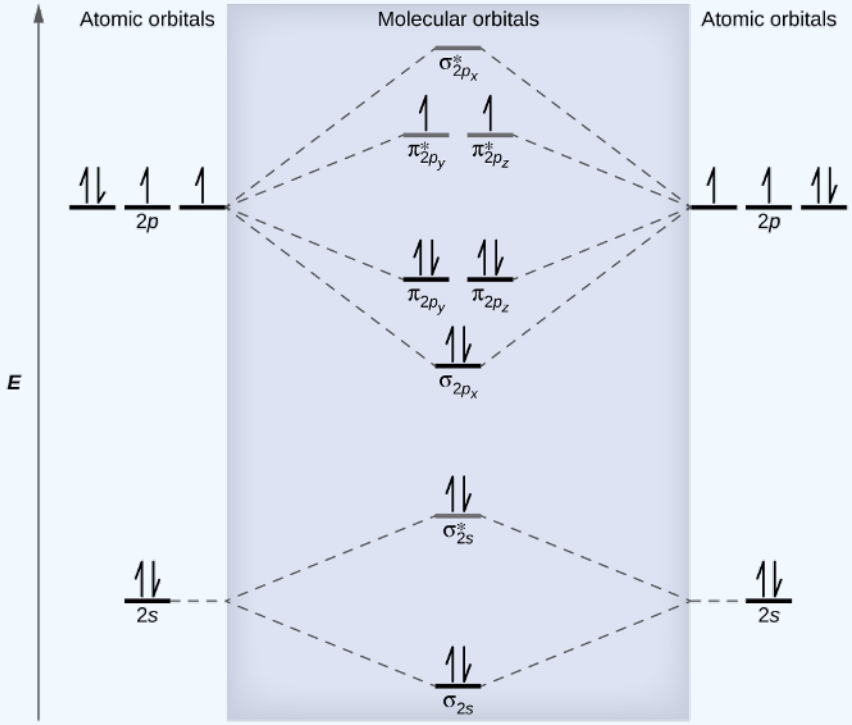


Figure 5.5.14: The molecular orbital energy diagram for O_2 predicts two unpaired electrons.

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 5.5.2

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 5.5.3: 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 5.5.3

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.

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5.6: 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

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5.7: Key Equations

$$\text{bond order} = \frac{(\text{number of bonding electron}) - (\text{number of antibonding electrons})}{}$$

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5.8: 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.

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CHAPTER OVERVIEW

6: Composition of Substances and Solutions

Quantitative aspects of the composition of substances and mixtures are the subject of this chapter.

6.1: Introduction
6.2: Formula Mass
6.3: Determining Empirical and Molecular Formulas
6.4: Molarity
6.5: Other Units for Solution Concentrations
6.6: Key Terms
6.7: Key Equations
6.8: Summary

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6.1: Introduction

Swimming pools have long been a popular means of recreation, exercise, and physical therapy. Since it is impractical to refill large pools with fresh water on a frequent basis, pool water is regularly treated with chemicals to prevent the growth of harmful bacteria and algae. Proper pool maintenance requires regular additions of various chemical compounds in carefully measured amounts. For example, the relative amount of calcium ion, Ca^{2+}, in the water should be maintained within certain limits to prevent eye irritation and avoid damage to the pool bed and plumbing. To maintain proper calcium levels, calcium cations are added to the water in the form of an ionic compound that also contains anions; thus, it is necessary to know both the relative amount of Ca^{2+} in the compound and the volume of water in the pool in order to achieve the proper calcium level. Quantitative aspects of the composition of substances (such as the calcium-containing compound) and mixtures (such as the pool water) are the subject of this chapter.



Figure 6.1.1: The water in a swimming pool is a complex mixture of substances whose relative amounts must be carefully maintained to ensure the health and comfort of people using the pool. (credit: modification of work by Vic Brincat)

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6.2: 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 6.2.1 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

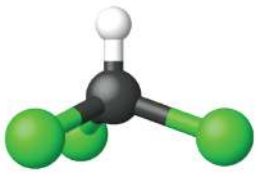


Figure 6.2.1: The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

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 6.2.2).

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

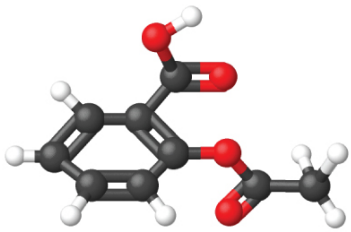


Figure 6.2.2: The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $C_9H_8O_4$.

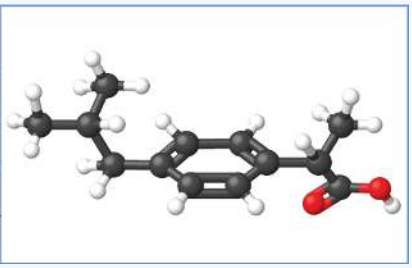
Example 6.2.1: Computing Molecular Mass for a Covalent Compound

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 6.2.1

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 6.2.3).

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

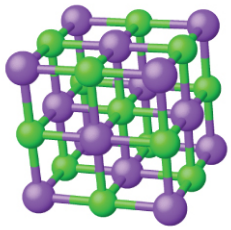


Figure 6.2.3: Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

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.

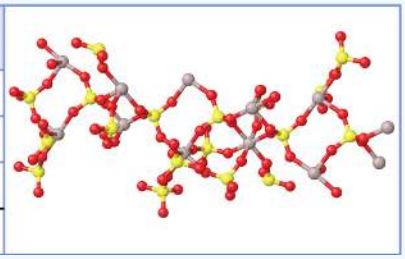
Example 6.2.2: Computing Formula Mass for an Ionic Compound

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 6.2.2

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 $6.02214076 \times 10^{23}$.

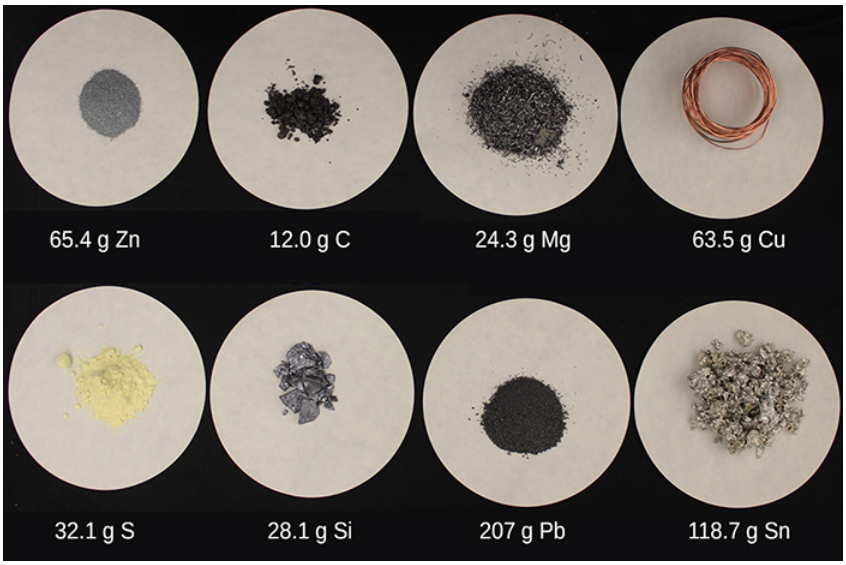


Figure 6.2.4: Each sample contains 6.022 10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark 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 6.2.5).



Figure 6.2.5: 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 6.2.6). 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 6.2.6: 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.

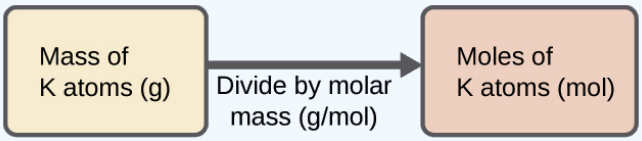
Example 6.2.3: Deriving Moles from Grams for an Element

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 6.2.3

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

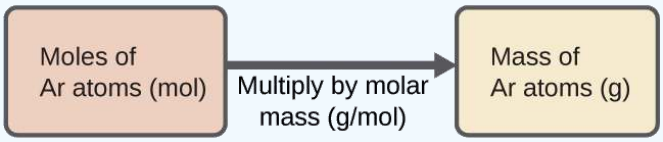
0.360 mol

Example 6.2.4: 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):

$$9.2 \times 10^{-4}\ \cancel{\text{mol Ar}} \left(\frac{39.95\ \text{g Ar}}{\cancel{\text{mol Ar}}} \right) = 0.037\ \text{g Ar}$$

The result is in agreement with our expectations, around 0.04 g Ar.

Exercise 6.2.1

What is the mass of 2.561 mol of gold?

Answer

504.4 g

Example 6.2.5: Deriving Number of Atoms from Mass for an Element

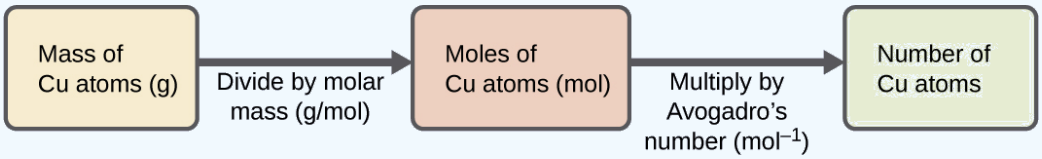
Copper is commonly used to fabricate electrical wire (Figure 6.2.7). How many copper atoms are in 5.00 g of copper wire?



Figure 6.2.7: 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:

$$5.00\ \text{g}\ \cancel{\text{Cu}} \left(\frac{\text{mol}\ \cancel{\text{Cu}}}{63.55\ \text{g}\ \cancel{\text{Cu}}}\right)\left(\frac{6.022 \times 10^{23}\ \text{Cu atoms}}{\text{mol}\ \cancel{\text{Cu}}}\right) = 4.74 \times 10^{22}\ \text{Cu atoms}$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 10^{22} as expected.

? Exercise 6.2.5

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

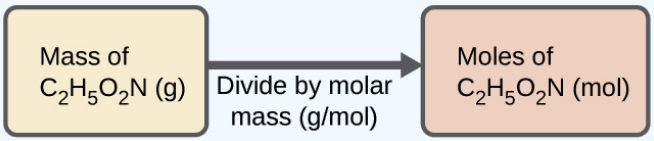
4.586×10^{22} Au atoms

✓ Example 6.2.6: 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?

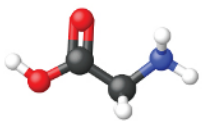
Solution

Derive the number of moles of a compound from its mass following the same procedure used for an element in Example 6.2.3:



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 6.2.6

How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a 25-g sample of sucrose?

Answer

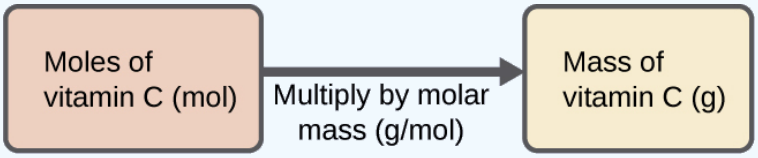
0.073 mol

✓ Example 6.2.7: 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 6.2.7

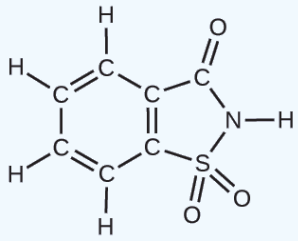
What is the mass of 0.443 mol of hydrazine, N_2H_4?

Answer

14.2 g

✓ Example 6.2.8: Deriving the Number of Atoms and Molecules from the Mass of a Compound

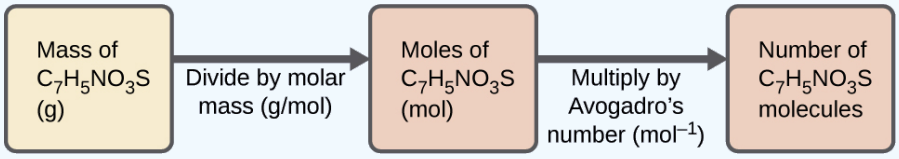
A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



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 6.2.6, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400\ \text{g}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}} \left(\frac{\text{mol}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}}}{183.18\ \text{g}\ \cancel{\text{CH}_5\text{NO}_3\text{S}}}\right)\left(\frac{6.022 \times 10^{23}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}}{1\ \text{mol}\ \cancel{\text{C}_7\text{H}_5\text{NO}_3\text{S}}}\right)$$
$$= 1.31 \times 10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is:

$$1.31 \times 10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}\left(\frac{7\ \text{C atoms}}{1\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecule}}\right) = 9.17 \times 10^{20}\ \text{C atoms}$$

Exercise 6.2.8

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 6.2.8). 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).

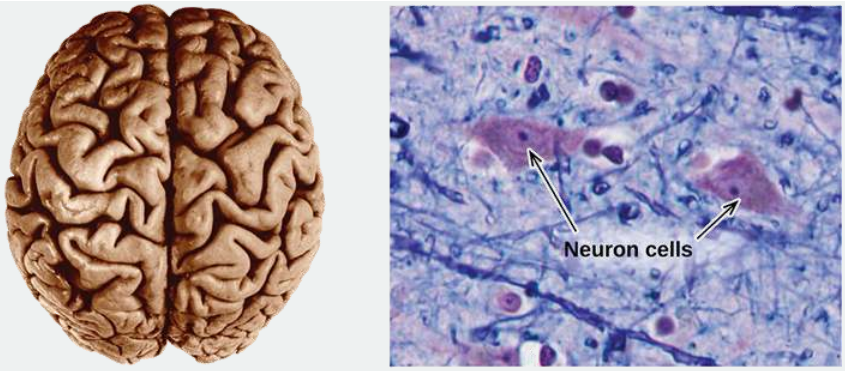


Figure 6.2.8: (a) A typical human brain weighs about 1.5 kg and occupies a volume of roughly 1.1 L. (b) Information is transmitted in brain tissue and throughout the central nervous system by specialized cells called neurons (micrograph shows cells at 1600× magnification).

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 6.2.9). 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.

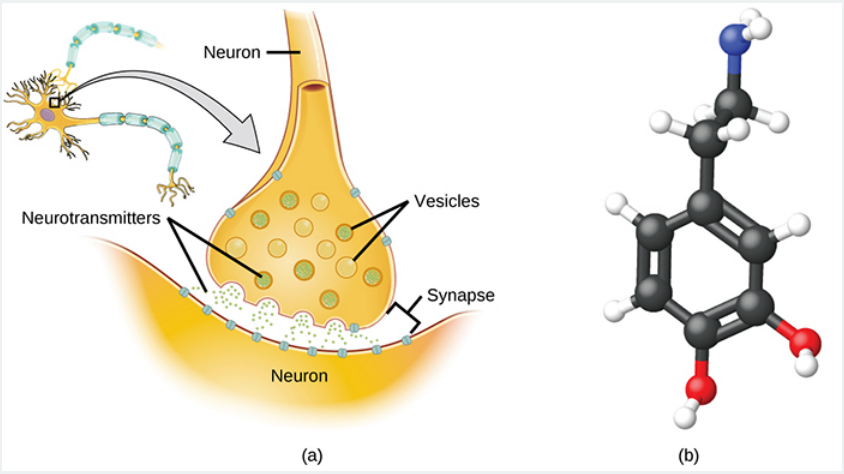


Figure 6.2.9: (a) Chemical signals are transmitted from neurons to other cells by the release of neurotransmitter molecules into the small gaps (synapses) between the cells. (b) Dopamine, $C_8H_{11}NO_2$, is a neurotransmitter involved in a number of neurological processes.

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.

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6.3: Determining Empirical and Molecular Formulas

Learning Objectives

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

- Compute the percent composition of a compound
- Determine the empirical formula of a compound
- Determine the molecular formula of a compound

The previous section discussed the relationship between the bulk mass of a substance and the number of atoms or molecules it contains (moles). Given the chemical formula of the substance, one may determine the amount of the substance (moles) from its mass, and vice versa. But what if the chemical formula of a substance is unknown? In this section, these same principles will be applied to derive the chemical formulas of unknown substances from experimental mass measurements.

Percent Composition

The elemental makeup of a compound defines its chemical identity, and chemical formulas are the most succinct way of representing this elemental makeup. When a compound's formula is unknown, measuring the mass of each of its constituent elements is often the first step in the process of determining the formula experimentally. The results of these measurements permit the calculation of the compound's **percent composition**, defined as the percentage by mass of each element in the compound. For example, consider a gaseous compound composed solely of carbon and hydrogen. The percent composition of this compound could be represented as follows:

$$\%H = \frac{\text{mass } H}{\text{mass compound}} \times 100\%$$
$$\%C = \frac{\text{mass } C}{\text{mass compound}} \times 100\%$$

If analysis of a 10.0-g sample of this gas showed it to contain 2.5 g H and 7.5 g C, the percent composition would be calculated to be 25% H and 75% C:

$$\%H = \frac{2.5gH}{10.0g\,\text{compound}} \times 100\% = 25\%$$
$$\%C = \frac{7.5gC}{10.0g\,\text{compound}} \times 100\% = 75\%$$

Example 6.3.1: Calculation of Percent Composition

Analysis of a 12.04-g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N. What is the percent composition of this compound?

Solution

To calculate percent composition, divide the experimentally derived mass of each element by the overall mass of the compound, and then convert to a percentage:

$$\%C = \frac{7.34gC}{12.04g\,\text{compound}} \times 100\% = 61.0\%$$
$$\%H = \frac{1.85gH}{12.04g\,\text{compound}} \times 100\% = 15.4\%$$
$$\%N = \frac{2.85gN}{12.04g\,\text{compound}} \times 100\% = 23.7\%$$

The analysis results indicate that the compound is 61.0% C, 15.4% H, and 23.7% N by mass.

? Exercise 6.3.1

A 24.81-g sample of a gaseous compound containing only carbon, oxygen, and chlorine is determined to contain 3.01 g C, 4.00 g O, and 17.81 g Cl. What is this compound's percent composition?

Answer

12.1% C, 16.1% O, 71.79% Cl

Determining Percent Composition from Molecular or Empirical Formulas

Percent composition is also useful for evaluating the relative abundance of a given element in different compounds of known formulas. As one example, consider the common nitrogen-containing fertilizers ammonia (NH_3), ammonium nitrate (NH_4NO_3), and urea (CH_4N_2O). The element nitrogen is the active ingredient for agricultural purposes, so the mass percentage of nitrogen in the compound is a practical and economic concern for consumers choosing among these fertilizers. For these sorts of applications, the percent composition of a compound is easily derived from its formula mass and the atomic masses of its constituent elements. A molecule of NH_3 contains one N atom weighing 14.01 amu and three H atoms weighing a total of (3 1.008 amu) = 3.024 amu. The formula mass of ammonia is therefore (14.01 amu + 3.024 amu) = 17.03 amu, and its percent composition is:

$$\%N = \frac{14.01 amu N}{17.03 amu HH} \times 100\% = 82.27\%$$
$$\%H = \frac{3.024 amu H}{17.03 amu H} \times 100\% = 17.76\%$$

This same approach may be taken considering a pair of molecules, a dozen molecules, or a mole of molecules, etc. The latter amount is most convenient and would simply involve the use of molar masses instead of atomic and formula masses, as demonstrated Example 6.3.2. As long as the molecular or empirical formula of the compound in question is known, the percent composition may be derived from the atomic or molar masses of the compound's elements.

✓ Example 6.3.2: Determining Percent Composition from a Molecular Formula

Aspirin is a compound with the molecular formula $C_9H_8O_4$. What is its percent composition?

Solution

To calculate the percent composition, the masses of C, H, and O in a known mass of $C_9H_8O_4$ are needed. It is convenient to consider 1 mol of $C_9H_8O_4$ and use its molar mass (180.159 g/mole, determined from the chemical formula) to calculate the percentages of each of its elements:

$$\%C = \frac{9 mol C \times \text{ molar mass } C}{\text{molar mass } C_9H_8O_4} \times 100 = \frac{9 \times 12.01 g/mol}{180.159 g/mol} \times 100 = \frac{108.09 g/mol}{180.159 g/mol} \times 100$$
$$\%C = 60.00\%C$$

$$\%H = \frac{8 mol H \times \text{ molar mass } H}{\text{molar mass } C_9H_8O_4} \times 100 = \frac{8 \times 1.008 g/mol}{180.159 g/mol} \times 100 = \frac{8.064 g/mol}{180.159 g/mol} \times 100$$
$$\%H = 4.476\%H$$

$$\%O = \frac{4 mol O \times \text{ molar mass}}{\text{molar mass } C_9H_8O_4} \times 100 = \frac{4 \times 16.00 g/mol}{180.159 g/mol} \times 100 = \frac{64.00 g/mol}{180.159 g/mol} \times 100$$
$$\%O = 35.52\%$$

Note that these percentages sum to equal 100.00% when appropriately rounded.

? Exercise 6.3.2

To three significant digits, what is the mass percentage of iron in the compound Fe_2O_3?

Answer

69.9% Fe

Determination of Empirical Formulas

As previously mentioned, the most common approach to determining a compound's chemical formula is to first measure the masses of its constituent elements. However, keep in mind that chemical formulas represent the relative *numbers*, not masses, of atoms in the substance. Therefore, any experimentally derived data involving mass must be used to derive the corresponding numbers of atoms in the compound. This is accomplished using molar masses to convert the mass of each element to a number of moles. These molar amounts are used to compute whole-number ratios that can be used to derive the empirical formula of the substance. Consider a sample of compound determined to contain 1.71 g C and 0.287 g H. The corresponding numbers of atoms (in moles) are:

$$1.71 gC \times \frac{1 molC}{12.01 gC} = 0.142 molC$$

$$0.287 gH \times \frac{1 molH}{1.008 gH} = 0.284 molH$$

Thus, this compound may be represented by the formula $C_{0.142}H_{0.284}$. Per convention, formulas contain whole-number subscripts, which can be achieved by dividing each subscript by the smaller subscript:

$$C_{\frac{0.142}{0.142}} H_{\frac{0.234}{0.142}} \text{ or } CH_2$$

(Recall that subscripts of "1" are not written but rather assumed if no other number is present.)

The empirical formula for this compound is thus CH_2. This may or may not be the compound's *molecular formula* as well; however, additional information is needed to make that determination (as discussed later in this section).

Consider as another example a sample of compound determined to contain 5.31 g Cl and 8.40 g O. Following the same approach yields a tentative empirical formula of:

$$Cl_{0.150}O_{0.525} = Cl_{\frac{0.150}{0.150}} O_{\frac{0.525}{0.150}} = ClO_{3.5}$$

In this case, dividing by the smallest subscript still leaves us with a decimal subscript in the empirical formula. To convert this into a whole number, multiply each of the subscripts by two, retaining the same atom ratio and yielding Cl_2O_7 as the final empirical formula.

In summary, empirical formulas are derived from experimentally measured element masses by:

1. Deriving the number of moles of each element from its mass
2. Dividing each element's molar amount by the smallest molar amount to yield subscripts for a tentative empirical formula
3. Multiplying all coefficients by an integer, if necessary, to ensure that the smallest whole-number ratio of subscripts is obtained

Figure 6.3.1 outlines this procedure in flow chart fashion for a substance containing elements A and X.

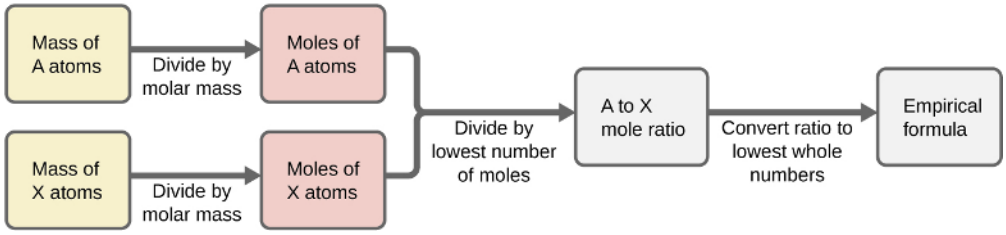


Figure 6.3.1: The empirical formula of a compound can be derived from the masses of all elements in the sample.

Example 6.3.3: Determining a Compound's Empirical Formula from the Masses of Its Elements

A sample of the black mineral hematite (Figure 6.3.2), an oxide of iron found in many iron ores, contains 34.97 g of iron and 15.03 g of oxygen. What is the empirical formula of hematite?



Figure 6.3.2: Hematite is an iron oxide that is used in jewelry. (credit: Mauro Cateb)

Solution

This problem provides the mass in grams of each element. Begin by finding the moles of each:

$$34.97 gFe \left(\frac{molFe}{55.85 g} \right) = 0.6261 molFe$$

$$15.03 gO \left(\frac{molO}{16.00 g} \right) = 0.9394 molO$$

Next, derive the iron-to-oxygen molar ratio by dividing by the lesser number of moles:

$$\frac{0.6261}{0.6261} = 1.000 molFe$$

$$\frac{0.9394}{0.6261} = 1.500 molO$$

The ratio is 1.000 mol of iron to 1.500 mol of oxygen ($Fe_1O_{1.5}$). Finally, multiply the ratio by two to get the smallest possible whole number subscripts while still maintaining the correct iron-to-oxygen ratio:

$$2 \left(Fe_1O_{1.5} \right) = Fe_2O_3$$

The empirical formula is Fe_2O_3.

? Exercise 6.3.3

What is the empirical formula of a compound if a sample contains 0.130 g of nitrogen and 0.370 g of oxygen?

Answer

N_2O_5

Link to Learning

For additional worked examples illustrating the derivation of empirical formulas, watch the brief video clip.

Deriving Empirical Formulas from Percent Composition

Finally, with regard to deriving empirical formulas, consider instances in which a compound's percent composition is available rather than the absolute masses of the compound's constituent elements. In such cases, the percent composition can be used to calculate the masses of elements present in any convenient mass of compound; these masses can then be used to derive the empirical formula in the usual fashion.

Example 6.3.4: Determining an Empirical Formula from Percent Composition

The bacterial fermentation of grain to produce ethanol forms a gas with a percent composition of 27.29% C and 72.71% O (Figure 6.3.3). What is the empirical formula for this gas?



Figure 6.3.3: An oxide of carbon is removed from these fermentation tanks through the large copper pipes at the top. (credit: "Dual Freq"/Wikimedia Commons)

Solution

Since the scale for percentages is 100, it is most convenient to calculate the mass of elements present in a sample weighing 100 g. The calculation is "most convenient" because, per the definition for percent composition, the mass of a given element in grams is numerically equivalent to the element's mass percentage. This numerical equivalence results from the definition of the "percentage" unit, whose name is derived from the Latin phrase *per centum* meaning "by the hundred." Considering this definition, the mass percentages provided may be more conveniently expressed as fractions:

$$27.29\%C = \frac{27.29gC}{100g\text{ compound}}$$
$$72.71\%O = \frac{72.71gO}{100g\text{ compound}}$$

The molar amounts of carbon and oxygen in a 100-g sample are calculated by dividing each element's mass by its molar mass:

$$27.29gC\left(\frac{molC}{12.01g}\right) = 2.272molC$$
$$72.71gO\left(\frac{molo}{16.00g}\right) = 4.544molO$$

Coefficients for the tentative empirical formula are derived by dividing each molar amount by the lesser of the two:

$$\frac{2.272molC}{2.272} = 1$$
$$\frac{4.544molO}{2.272} = 2$$

Since the resulting ratio is one carbon to two oxygen atoms, the empirical formula is CO_2.

Exercise 6.3.4

What is the empirical formula of a compound containing 40.0% C, 6.71% H, and 53.28% O?

Answer

CH_2O

Derivation of Molecular Formulas

Recall that empirical formulas are symbols representing the *relative* numbers of a compound's elements. Determining the *absolute* numbers of atoms that compose a single molecule of a covalent compound requires knowledge of both its empirical formula and its molecular mass or molar mass. These quantities may be determined experimentally by various measurement techniques. Molecular mass, for example, is often derived from the mass spectrum of the compound (see discussion of this technique in the previous chapter on atoms and molecules). Molar mass can be measured by a number of experimental methods, many of which will be introduced in later chapters of this text.

Molecular formulas are derived by comparing the compound's molecular or molar mass to its **empirical formula mass**. As the name suggests, an empirical formula mass is the sum of the average atomic masses of all the atoms represented in an empirical formula. If the molecular (or molar) mass of the substance is known, it may be divided by the empirical formula mass to yield the number of empirical formula units per molecule (*n*):

$$\frac{\text{molecular or molar mass (amu or } \frac{\text{g}}{\text{mol}})}{\text{empirical formula mass } (\text{amu or } \frac{\text{g}}{\text{mol}})} = n \text{ formula units/molecule}$$

The molecular formula is then obtained by multiplying each subscript in the empirical formula by *n*, as shown by the generic empirical formula A_xB_y:

$$(A_xB_y)_n = A_{nx}B_{ny}$$

For example, consider a covalent compound whose empirical formula is determined to be CH_2O. The empirical formula mass for this compound is approximately 30 amu (the sum of 12 amu for one C atom, 2 amu for two H atoms, and 16 amu for one O atom). If the compound's molecular mass is determined to be 180 amu, this indicates that molecules of this compound contain six times the number of atoms represented in the empirical formula:

$$\frac{180 \text{ amu / molecule}}{30 \frac{\text{amu}}{\text{formula unit}}} = 6 \text{ formula units / molecule}$$

Molecules of this compound are then represented by molecular formulas whose subscripts are six times greater than those in the empirical formula:

$$(CH_2O)_6 = C_6H_{12}O_6$$

Note that this same approach may be used when the molar mass (g/mol) instead of the molecular mass (amu) is used. In this case, *one mole* of empirical formula units and molecules is considered, as opposed to single units and molecules.

Example 6.3.5: Determination of the Molecular Formula for Nicotine

Nicotine, an alkaloid in the nightshade family of plants that is mainly responsible for the addictive nature of cigarettes, contains 74.02% C, 8.710% H, and 17.27% N. If 40.57 g of nicotine contains 0.2500 mol nicotine, what is the molecular formula?

Solution

Determining the molecular formula from the provided data will require comparison of the compound's empirical formula mass to its molar mass. As the first step, use the percent composition to derive the compound's empirical formula. Assuming a convenient, a 100-g sample of nicotine yields the following molar amounts of its elements:

$$(74.02gC)\left(\frac{1molC}{12.01gC}\right) = 6.163molC$$

$$(8.710gH)\left(\frac{1molH}{1.008gH}\right) = 8.641molH$$

$$(17.27gN)\left(\frac{1molN}{14.01gN}\right) = 1.233molN$$

Next, calculate the molar ratios of these elements relative to the least abundant element, N.

$$6.163\ \text{mol C}/1.233\ \text{mol N} = 58.641\ \text{mol H}/1.233\ \text{mol N} = 71.233\ \text{mol N}/1.233\ \text{mol N} = 1$$

$$\frac{1.233}{1.233} = 1.000 molN$$
$$\frac{6.163}{1.233} = 4.998 molC$$
$$\frac{8.624}{1.233} = 6.994 molH$$

The C-to-N and H-to-N molar ratios are adequately close to whole numbers, and so the empirical formula is C_5H_7N. The empirical formula mass for this compound is therefore 81.13 amu/formula unit, or 81.13 g/mol formula unit.

Calculate the molar mass for nicotine from the given mass and molar amount of compound:

$$\frac{40.57 g\ \text{nicotine}}{0.2500 mol\ \text{nicotine}} = \frac{162.3 g}{mol}$$

Comparing the molar mass and empirical formula mass indicates that each nicotine molecule contains two formula units:

$$\frac{162.3 g/mol}{81.13 \frac{g}{\text{formula unit}}} = 2\ \text{formula units / molecule}$$

Finally, derive the molecular formula for nicotine from the empirical formula by multiplying each subscript by two:

$$(\text{C}_5\text{H}_7\text{N})_2 = \text{C}_{10}\text{H}_{14}\text{N}_2$$

Exercise 6.3.1

What is the molecular formula of a compound with a percent composition of 49.47% C, 5.201% H, 28.84% N, and 16.48% O, and a molecular mass of 194.2 amu?

Answer

$\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$

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6.4: 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 6.4.1). This section will describe one of the most common ways in which the relative compositions of mixtures may be quantified.



Figure 6.4.1: 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 6.4.1: 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 6.4.1

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 6.4.2: 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 6.4.1?

Solution

Rearrange the definition of molarity to isolate the quantity sought, moles of sugar, then substitute the value for molarity derived in Example 6.4.1, 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 6.4.2

What volume (mL) of the sweetened tea described in Example 6.4.1 contains the same amount of sugar (mol) as 10 mL of the soft drink in this example?

Answer

80 mL

Example 6.4.3: Calculating Molar Concentrations from the Mass of Solute

Distilled white vinegar (Figure 6.4.2) 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 6.4.2: 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 6.4.3

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 6.4.4: 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 6.4.2:

$$\begin{aligned} M &= \frac{\text{mol solute}}{L\ \text{solution}} \\ \text{mol solute} &= M \times L\ \text{solution} \\ &= 5.30 \frac{\text{mol NaCl}}{\text{L}} \times 0.250 L \\ &= 1.325\ \text{mol NaCl} \end{aligned}$$

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 6.4.4

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 6.4.4, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In Example 6.4.4, 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 6.4.5). This eliminates intermediate steps so that only the final result is rounded.

Example 6.4.5: Determining the Volume of Solution Containing a Given Mass of Solute

In Example 6.4.3, 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 6.4.5

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 6.4.3).

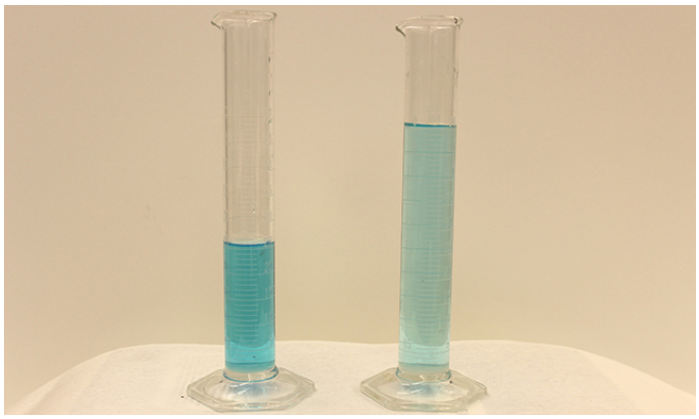


Figure 6.4.3: Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott)

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 6.4.6: 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 6.4.6

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 6.4.7: 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 6.4.7

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 6.4.8: 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 6.4.8

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

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6.5: Other Units for Solution Concentrations

Learning Objectives

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

- Define the concentration units of mass percentage, volume percentage, mass-volume percentage, parts-per-million (ppm), and parts-per-billion (ppb)
- Perform computations relating a solution's concentration and its components' volumes and/or masses using these units

The previous section introduced molarity, a very useful measurement unit for evaluating the concentration of solutions. However, molarity is only one measure of concentration. This section will describe some other units of concentration that are commonly used in various applications, either for convenience or by convention.

Mass Percentage

Earlier in this chapter, percent composition was introduced as a measure of the relative amount of a given element in a compound. Percentages are also commonly used to express the composition of mixtures, including solutions. The mass percentage of a solution component is defined as the ratio of the component's mass to the solution's mass, expressed as a percentage:

$$\text{mass percentage} = \frac{\text{mass of component}}{\text{mass of solution}} \times 100\%$$

Mass percentage is also referred to by similar names such as *percent mass*, *percent weight*, *weight/weight percent*, and other variations on this theme. The most common symbol for mass percentage is simply the percent sign, %, although more detailed symbols are often used including %mass, %weight, and (w/w)%. Use of these more detailed symbols can prevent confusion of mass percentages with other types of percentages, such as volume percentages (to be discussed later in this section).

Mass percentages are popular concentration units for consumer products. The label of a typical liquid bleach bottle (Figure 6.5.1) cites the concentration of its active ingredient, sodium hypochlorite (NaOCl), as being 7.4%. A 100.0-g sample of bleach would therefore contain 7.4 g of NaOCl.



Figure 6.5.1: Liquid bleach is an aqueous solution of sodium hypochlorite (NaOCl). This brand has a concentration of 7.4% NaOCl by mass.

Example 6.5.1: Calculation of Percent by Mass

A 5.0-g sample of spinal fluid contains 3.75 mg (0.00375 g) of glucose. What is the percent by mass of glucose in spinal fluid?

Solution

The spinal fluid sample contains roughly 4 mg of glucose in 5000 mg of fluid, so the mass fraction of glucose should be a bit less than one part in 1000, or about 0.1%. Substituting the given masses into the equation defining mass percentage yields:

$$\%\ \text{glucose} = \frac{3.75\ \text{mg glucose} \times \frac{1\ \text{g}}{1000\ \text{mg}}}{5.0\ \text{g spinal fluid}} = 0.075\%$$

The computed mass percentage agrees with our rough estimate (it's a bit less than 0.1%).

Note that while any mass unit may be used to compute a mass percentage (mg, g, kg, oz, and so on), the same unit must be used for both the solute and the solution so that the mass units cancel, yielding a dimensionless ratio. In this case, the solute mass unit in the numerator was converted from mg to g to match the units in the denominator. Alternatively, the spinal fluid mass unit in the denominator could have been converted from g to mg instead. As long as identical mass units are used for both solute and solution, the computed mass percentage will be correct.

? Exercise 6.5.1

A bottle of a tile cleanser contains 135 g of HCl and 775 g of water. What is the percent by mass of HCl in this cleanser?

Answer

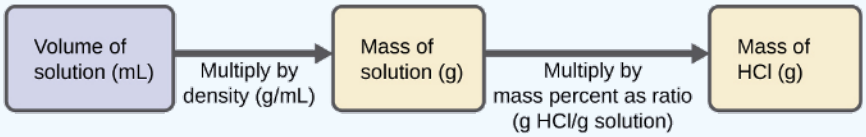
14.8%

✓ Example 6.5.2: Calculations using Mass Percentage

"Concentrated" hydrochloric acid is an aqueous solution of 37.2% HCl that is commonly used as a laboratory reagent. The density of this solution is 1.19 g/mL. What mass of HCl is contained in 0.500 L of this solution?

Solution

The HCl concentration is near 40%, so a 100-g portion of this solution would contain about 40 g of HCl. Since the solution density isn't greatly different from that of water (1 g/mL), a reasonable estimate of the HCl mass in 500 g (0.5 L) of the solution is about five times greater than that in a 100 g portion, or 5 40 = 200 g. In order to derive the mass of solute in a solution from its mass percentage, the mass of the solution must be known. Using the solution density given, convert the solution's volume to mass, and then use the given mass percentage to calculate the solute mass. This mathematical approach is outlined in this flowchart:



For proper unit cancellation, the 0.500-L volume is converted into 500 mL, and the mass percentage is expressed as a ratio, 37.2 g HCl/g solution:

$$500 mL\ \text{solution} \left(\frac{1.19\ \text{g solution}}{mL\ \text{solution}}\right)\left(\frac{37.2\ \text{g HCl}}{100\ \text{g solution}}\right) = 221\ \text{g HCl}$$

This mass of HCl is consistent with our rough estimate of approximately 200 g.

? Exercise 6.5.2

What volume of concentrated HCl solution contains 125 g of HCl?

Answer

282 mL

Volume Percentage

Liquid volumes over a wide range of magnitudes are conveniently measured using common and relatively inexpensive laboratory equipment. The concentration of a solution formed by dissolving a liquid solute in a liquid solvent is therefore often expressed as a

volume percentage, %vol or (v/v)%:

$$\text{volume percentage} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\%$$

Example 6.5.3: Calculations using Volume Percentage

Rubbing alcohol (isopropanol) is usually sold as a 70%vol aqueous solution. If the density of isopropyl alcohol is 0.785 g/mL, how many grams of isopropyl alcohol are present in a 355 mL bottle of rubbing alcohol?

Solution

Per the definition of volume percentage, the isopropanol volume is 70% of the total solution volume. Multiplying the isopropanol volume by its density yields the requested mass:

$$(355\,\text{mL solution}) \left(\frac{70\,\text{mL isopropyl alcohol}}{100\,\text{mL solution}}\right) \left(\frac{0.785\,\text{g isopropyl alcohol}}{1\,\text{mL isopropyl alcohol}}\right) = 195\,\text{g isopropyl}$$

Exercise 6.5.3

Wine is approximately 12% ethanol (CH_3CH_2OH) by volume. Ethanol has a molar mass of 46.06 g/mol and a density 0.789 g/mL. How many moles of ethanol are present in a 750-mL bottle of wine?

Answer

1.5 mol ethanol

Mass-Volume Percentage

"Mixed" percentage units, derived from the mass of solute and the volume of solution, are popular for certain biochemical and medical applications. A **mass-volume percent** is a ratio of a solute's mass to the solution's volume expressed as a percentage. The specific units used for solute mass and solution volume may vary, depending on the solution. For example, physiological saline solution, used to prepare intravenous fluids, has a concentration of 0.9% mass/volume (m/v), indicating that the composition is 0.9 g of solute per 100 mL of solution. The concentration of glucose in blood (commonly referred to as "blood sugar") is also typically expressed in terms of a mass-volume ratio. Though not expressed explicitly as a percentage, its concentration is usually given in milligrams of glucose per deciliter (100 mL) of blood (Figure 6.5.2).



(a) (b)

Figure 6.5.2: "Mixed" mass-volume units are commonly encountered in medical settings. (a) The NaCl concentration of physiological saline is 0.9% (m/v). (b) This device measures glucose levels in a sample of blood. The normal range for glucose concentration in blood (fasting) is around 70–100 mg/dL. (credit a: modification of work by "The National Guard"/Flickr; credit b: modification of work by Biswarup Ganguly)

Parts per Million and Parts per Billion

Very low solute concentrations are often expressed using appropriately small units such as parts per million (ppm) or parts per billion (ppb). Like percentage ("part per hundred") units, ppm and ppb may be defined in terms of masses, volumes, or mixed mass-volume units. There are also ppm and ppb units defined with respect to numbers of atoms and molecules.

The mass-based definitions of ppm and ppb are given here:

$$\text{ppm} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^6 \text{ ppm}$$

$$\text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^9 \text{ ppb}$$

Both ppm and ppb are convenient units for reporting the concentrations of pollutants and other trace contaminants in water. Concentrations of these contaminants are typically very low in treated and natural waters, and their levels cannot exceed relatively low concentration thresholds without causing adverse effects on health and wildlife. For example, the EPA has identified the maximum safe level of fluoride ion in tap water to be 4 ppm. Inline water filters are designed to reduce the concentration of fluoride and several other trace-level contaminants in tap water (Figure 6.5.3).



(a) (b)

Figure 6.5.3: (a) In some areas, trace-level concentrations of contaminants can render unfiltered tap water unsafe for drinking and cooking. (b) Inline water filters reduce the concentration of solutes in tap water. (credit a: modification of work by Jenn Durfey; credit b: modification of work by "vastateparkstaff"/Wikimedia commons)

Example 6.5.4: Calculation of Parts per Million and Parts per Billion Concentrations

According to the EPA, when the concentration of lead in tap water reaches 15 ppb, certain remedial actions must be taken. What is this concentration in ppm? At this concentration, what mass of lead (μg) would be contained in a typical glass of water (300 mL)?

Solution

The definitions of the ppm and ppb units may be used to convert the given concentration from ppb to ppm. Comparing these two unit definitions shows that ppm is 1000 times greater than ppb (1 ppm = 10^3 ppb). Thus:

$$15 \text{ ppb} \times \frac{1 \text{ ppm}}{10^3 \text{ ppb}} = 0.015 \text{ ppm}$$

The definition of the ppb unit may be used to calculate the requested mass if the mass of the solution is provided. Since the volume of solution (300 mL) is given, its density must be used to derive the corresponding mass. Assume the density of tap water to be roughly the same as that of pure water (~1.00 g/mL), since the concentrations of any dissolved substances should not be very large. Rearranging the equation defining the ppb unit and substituting the given quantities yields:

$$\text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^9 \text{ ppb}$$

$$\text{mass solute} = \frac{\text{ppb} \times \text{mass solution}}{10^9 \text{ ppb}}$$

$$= \frac{15 \text{ ppb} \times 300 mL \times \frac{1.00 \text{ g}}{\text{mL}}}{10^9 \text{ ppb}} = 4.5 \times 10^{-6} \text{ g}$$

Finally, convert this mass to the requested unit of micrograms:

$$4.5 \times 10^{-6} \text{ g} \times \frac{1 \, \mu\text{g}}{10^{-6} \text{ g}} = 4.5 \, \mu\text{g}$$

? Exercise 6.5.4

A 50.0-g sample of industrial wastewater was determined to contain 0.48 mg of mercury. Express the mercury concentration of the wastewater in ppm and ppb units.

Answer

9.6 ppm, 9600 ppb

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6.6: Key Terms

aqueous solution
solution for which water is the solvent

Avogadro's number (N_A)
experimentally determined value of the number of entities comprising 1 mole of substance, equal to 6.022×10^{23} mol^{-1}

concentrated
qualitative term for a solution containing solute at a relatively high concentration

concentration
quantitative measure of the relative amounts of solute and solvent present in a solution

dilute
qualitative term for a solution containing solute at a relatively low concentration

dilution
process of adding solvent to a solution in order to lower the concentration of solutes

dissolved
describes the process by which solute components are dispersed in a solvent

empirical formula mass
sum of average atomic masses for all atoms represented in an empirical formula

formula mass
sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass

mass percentage
ratio of solute-to-solution mass expressed as a percentage

mass-volume percent
ratio of solute mass to solution volume, expressed as a percentage

molar mass
mass in grams of 1 mole of a substance

molarity (M)
unit of concentration, defined as the number of moles of solute dissolved in 1 liter of solution

mole
amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

parts per billion (ppb)
ratio of solute-to-solution mass multiplied by 10^9

parts per million (ppm)
ratio of solute-to-solution mass multiplied by 10^6

percent composition
percentage by mass of the various elements in a compound

solute

solution component present in a concentration less than that of the solvent

solvent

solution component present in a concentration that is higher relative to other components

volume percentage

ratio of solute-to-solution volume expressed as a percentage

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6.7: Key Equations

$$\%\,\mathrm{X} = \frac{\text{mass}}{\mathrm{X}} \times 100\,\%$$

$$\frac{\text{molecular or molar mass } \left(\text{amuor } \frac{\text{g}}{\text{mol}}\right)}{\text{empirical formula mass } \left(\text{amuor } \frac{\text{g}}{\text{mol}}\right)} = n \text{ formulaunits/}$$

$$(\mathrm{A_xB_y})_\mathrm{n} = \mathrm{A_{nx}B_{ny}}$$

$$M = \frac{\text{mol}}{\text{solute}}$$

$$C_1V_1 = C_2V_2$$

$$\text{Percent by mass} = \frac{\text{mass of}}{\text{solute}}$$

$$\text{ppm} = \frac{\text{mass}}{\text{solute}}$$

$$\text{ppb} = \frac{\text{mass}}{\text{solute}}$$

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6.8: Summary

The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H_2O molecule weighs approximately18 amu and 1 mole of H_2O molecules weighs approximately 18 g).

The chemical identity of a substance is defined by the types and relative numbers of atoms composing its fundamental entities (molecules in the case of covalent compounds, ions in the case of ionic compounds). A compound's percent composition provides the mass percentage of each element in the compound, and it is often experimentally determined and used to derive the compound's empirical formula. The empirical formula mass of a covalent compound may be compared to the compound's molecular or molar mass to derive a molecular formula.

Solutions are homogeneous mixtures. Many solutions contain one component, called the solvent, in which other components, called solutes, are dissolved. An aqueous solution is one for which the solvent is water. The concentration of a solution is a measure of the relative amount of solute in a given amount of solution. Concentrations may be measured using various units, with one very useful unit being molarity, defined as the number of moles of solute per liter of solution. The solute concentration of a solution may be decreased by adding solvent, a process referred to as dilution. The dilution equation is a simple relation between concentrations and volumes of a solution before and after dilution.

In addition to molarity, a number of other solution concentration units are used in various applications. Percentage concentrations based on the solution components' masses, volumes, or both are useful for expressing relatively high concentrations, whereas lower concentrations are conveniently expressed using ppm or ppb units. These units are popular in environmental, medical, and other fields where mole-based units such as molarity are not as commonly used.

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CHAPTER OVERVIEW

7: Stoichiometry of Chemical Reactions

This chapter will describe how to symbolize chemical reactions using chemical equations, how to classify some common chemical reactions by identifying patterns of reactivity, and how to determine the quantitative relations between the amounts of substances involved in chemical reactions—that is, the reaction *stoichiometry*.

7.1: Introduction
7.2: Writing and Balancing Chemical Equations
7.3: Classifying Chemical Reactions
7.4: Reaction Stoichiometry
7.5: Reaction Yields
7.6: Quantitative Chemical Analysis
7.7: Key Terms
7.8: Key Equations
7.9: Summary

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7.1: Introduction

Solid-fuel rockets are a central feature in the world's space exploration programs, including the new Space Launch System being developed by the National Aeronautics and Space Administration (NASA) to replace the retired Space Shuttle fleet (Figure 7.1.1). The engines of these rockets rely on carefully prepared solid mixtures of chemicals combined in precisely measured amounts. Igniting the mixture initiates a vigorous chemical reaction that rapidly generates large amounts of gaseous products. These gases are ejected from the rocket engine through its nozzle, providing the thrust needed to propel heavy payloads into space. Both the nature of this chemical reaction and the relationships between the amounts of the substances being consumed and produced by the reaction are critically important considerations that determine the success of the technology.



Figure 7.1.1: Many modern rocket fuels are solid mixtures of substances combined in carefully measured amounts and ignited to yield a thrust-generating chemical reaction. (credit: modification of work by NASA)

This chapter will describe how to symbolize chemical reactions using chemical equations, how to classify some common chemical reactions by identifying patterns of reactivity, and how to determine the quantitative relations between the amounts of substances involved in chemical reactions—that is, the reaction *stoichiometry*.

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7.2: 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 7.2.1: , with space-filling molecular models shown in the lower half of the figure.

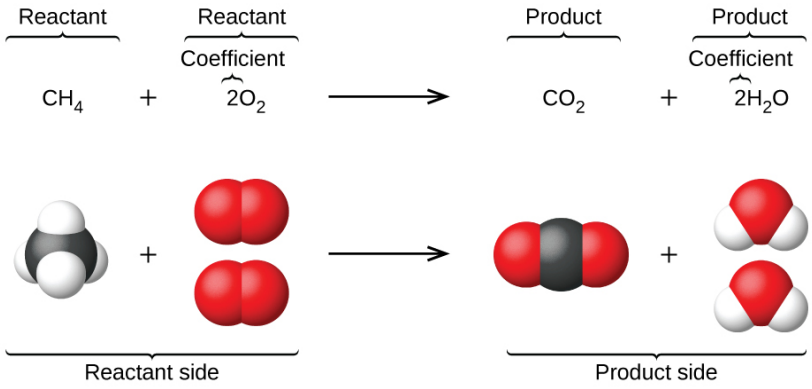


Figure 7.2.1: The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be represented by a chemical equation using formulas (top).

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 7.2.2). 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.

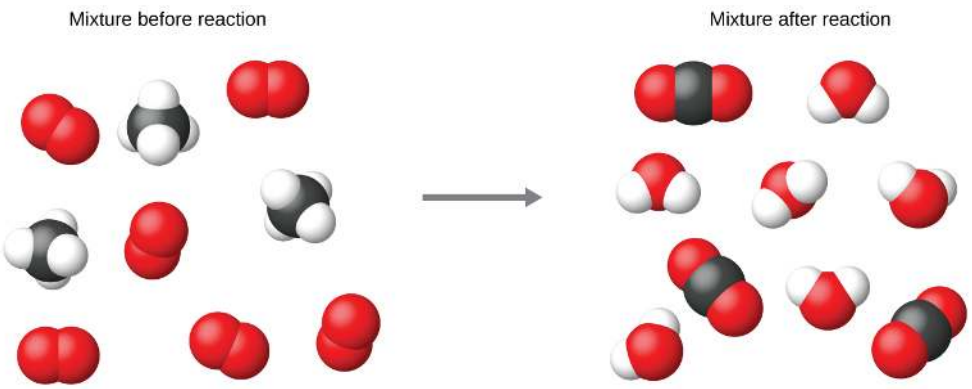


Figure 7.2.2: Regardless of the absolute numbers of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given by the chemical reaction equation.

Balancing Equations

The chemical equation described above in Figure 7.2.1 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 7.2.1: 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 7.2.1

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 7.2.2: 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 7.2.2

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

$$2\,NaCl(aq) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g) + Cl_2(g) \quad \text{(molecular)}$$

$$2\,Na^+(aq) + 2\,Cl^-(aq) + 2\,H_2O(l) \longrightarrow 2\,Na^+(aq) + 2\,OH^-(aq) + H_2(g) + Cl_2(g) \quad \text{(complete ionic)}$$

$$2\,Cl^-(aq) + 2\,H_2O(l) \longrightarrow 2\,OH^-(aq) + H_2(g) + Cl_2(g) \quad \text{(net ionic)}$$

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7.3: Classifying Chemical Reactions

Learning Objectives

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

- Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction)
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Identify common acids and bases
- Predict the solubility of common inorganic compounds by using solubility rules
- Compute the oxidation states for elements in compounds

Humans interact with one another in various and complex ways, and we classify these interactions according to common patterns of behavior. When two humans exchange information, we say they are communicating. When they exchange blows with their fists or feet, we say they are fighting. Faced with a wide range of varied interactions between chemical substances, scientists have likewise found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This module will provide an introduction to three of the most prevalent types of chemical reactions: precipitation, acid-base, and oxidation-reduction.

Precipitation Reactions and Solubility Rules

A precipitation reaction is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as *double displacement*, *double replacement*, or *metathesis* reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and *gravimetric methods* for determining the composition of matter (see the last module of this chapter).

The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its solubility, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be soluble. A substance will precipitate when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be insoluble, and these are the substances that readily precipitate from solution. More information on these important concepts is provided in a later chapter on solutions. For purposes of predicting the identities of solids formed by precipitation reactions, one may simply refer to patterns of solubility that have been observed for many ionic compounds (Table 7.3.1).

Table 7.3.1

Soluble Ionic Compounds	contain these ions	exceptions
	NH_4^+ group I cations: Li^+ Na^+ K^+ Rb^+ Cs^+	none
	Cl^- Br^- I^-	compounds with Ag^+, Hg_2^{2+}, and Pb^{2+}
	F^-	compounds with group 2 metal cations, Pb^{2+}, and Fe^{3+}

	$C_2H_3O_2^-$ HCO_3^- NO_3^- ClO_3^-	none
	SO_4^{2-}	compounds with Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+} and Sr^{2+}
Insoluble Ionic Compounds	contain these ions	exceptions
	CO_3^{2-} CrO_4^{2-} PO_4^{3-} S^{2-}	compounds with group 1 cations and NH_4^+
	OH^-	compounds with group 1 cations and Ba^{2+}

A vivid example of precipitation is observed when solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide:

$$2\,\mathrm{KI(aq)} + \mathrm{Pb(NO_3)_2(aq)} \longrightarrow \mathrm{PbI_2(s)} + 2\,\mathrm{KNO_3(aq)}$$

This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts.

The net ionic equation representing this reaction is:

$$\mathrm{Pb^{2+}(aq)} + 2\,\mathrm{I^-(aq)} \longrightarrow \mathrm{PbI_2(s)}$$

Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (Figure 7.3.1). The properties of pure PbI_2 crystals make them useful for fabrication of X-ray and gamma ray detectors.

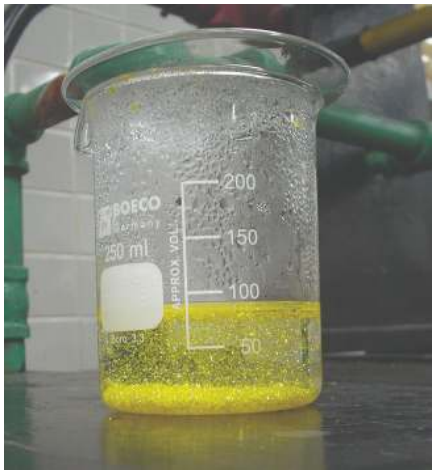


Figure 7.3.1: A precipitate of PbI_2 forms when solutions containing Pb^{2+} and I^- are mixed. (credit: Der Kreole/Wikimedia Commons)

The solubility guidelines in Table 7.3.1 may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound. For example, mixing solutions of silver nitrate and sodium chloride will yield a solution containing Ag^+, NO_3^-, Na^+, and Cl^- ions. Aside from the two ionic compounds originally present in the solutions, $AgNO_3$ and NaCl, two additional ionic compounds may be derived from this collection of ions: $NaNO_3$ and AgCl. The solubility guidelines indicate all nitrate salts are soluble but that AgCl is one of insoluble. A precipitation reaction, therefore, is predicted to occur, as described by the following equations:

$$\mathrm{NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)} \quad \text{(molecular)}$$

$$\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)} \quad \text{(net ionic)}$$

Example 7.3.1: Predicting Precipitation Reactions

Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction.

a. potassium sulfate and barium nitrate
b. lithium chloride and silver acetate
c. lead nitrate and ammonium carbonate

Solution

(a) The two possible products for this combination are KNO_3 and $BaSO_4$. The solubility guidelines indicate $BaSO_4$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$\mathrm{Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)}$$

(b) The two possible products for this combination are $LiC_2H_3O_2$ and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$$

(c) The two possible products for this combination are $PbCO_3$ and NH_4NO_3. The solubility guidelines indicate $PbCO_3$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$\mathrm{Pb^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow PbCO_3(s)}$$

Exercise 7.3.1

Which solution could be used to precipitate the barium ion, Ba^{2+}, in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate?

Answer

sodium sulfate, $\mathrm{BaSO_4}$

Acid-Base Reactions

An acid-base reaction is one in which a hydrogen ion, H^+, is transferred from one chemical species to another. Such reactions are of central importance to numerous natural and technological processes, ranging from the chemical transformations that take place within cells and the lakes and oceans, to the industrial-scale production of fertilizers, pharmaceuticals, and other substances essential to society. The subject of acid-base chemistry, therefore, is worthy of thorough discussion, and a full chapter is devoted to this topic later in the text.

For purposes of this brief introduction, we will consider only the more common types of acid-base reactions that take place in aqueous solutions. In this context, an **acid** is a substance that will dissolve in water to yield hydronium ions, H_3O^+. As an example, consider the equation shown here:

$$\mathrm{HCl(aq) + H_2O(aq) \longrightarrow Cl^-(aq) + H_3O^+(aq)}$$

The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water, H_3O^+ ions are produced by a chemical reaction in which H^+ ions are transferred from HCl molecules to H_2O molecules (Figure 7.3.2).

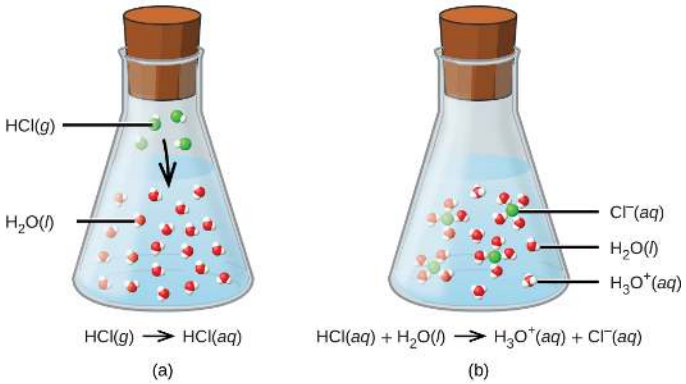


Figure 7.3.2: When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions (and solvated chloride ions).

The nature of HCl is such that its reaction with water as just described is essentially 100% efficient: Virtually every HCl molecule that dissolves in water will undergo this reaction. Acids that completely react in this fashion are called **strong acids**, and HCl is one among just a handful of common acid compounds that are classified as strong (Table 7.3.2). A far greater number of compounds behave as **weak acids** and only partially react with water, leaving a large majority of dissolved molecules in their original form and generating a relatively small amount of hydronium ions. Weak acids are commonly encountered in nature, being the substances partly responsible for the tangy taste of citrus fruits, the stinging sensation of insect bites, and the unpleasant smells associated with body odor. A familiar example of a weak acid is acetic acid, the main ingredient in food vinegars:

$$\mathrm{CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons CH_3CO_2^-(aq) + H_3O^+(aq)}$$

When dissolved in water under typical conditions, only about 1% of acetic acid molecules are present in the ionized form,

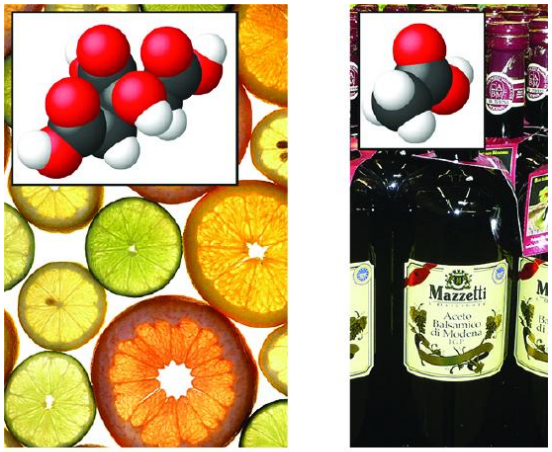


Figure 7.3.3: (a) Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid. (credit a: modification of work by Scott Bauer; credit b: modification of work by Brücke-Osteuropa/Wikimedia Commons)

Table 7.3.2: Common Strong Acids

Compound Formula	Name in Aqueous Solution
HBr	hydrobromic acid
HCl	hydrochloric acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid

A base is a substance that will dissolve in water to yield hydroxide ions, OH^-. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion—for example, NaOH and $Ca(OH)_2$. Unlike the acid compounds discussed previously, these compounds do not react chemically with water; instead they dissolve and dissociate, releasing hydroxide ions directly into the solution. For example, KOH and $Ba(OH)_2$ dissolve in water and dissociate completely to produce cations (K^+ and Ba^{2+}, respectively) and hydroxide ions, OH^-. These bases, along with other hydroxides that completely dissociate in water, are considered **strong bases**.

Consider as an example the dissolution of lye (sodium hydroxide) in water:

$$\mathrm{NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)}$$

This equation confirms that sodium hydroxide is a base. When dissolved in water, NaOH dissociates to yield Na^+ and OH^- ions. This is also true for any other ionic compound containing hydroxide ions. Since the dissociation process is essentially complete when ionic compounds dissolve in water under typical conditions, NaOH and other ionic hydroxides are all classified as strong bases.

Unlike ionic hydroxides, some compounds produce hydroxide ions when dissolved by chemically reacting with water molecules. In all cases, these compounds react only partially and so are classified as **weak bases**. These types of compounds are also abundant in nature and important commodities in various technologies. For example, global production of the weak base ammonia is typically well over 100 metric tons annually, being widely used as an agricultural fertilizer, a raw material for chemical synthesis of other compounds, and an active ingredient in household cleaners (Figure 7.3.4). When dissolved in water, ammonia reacts partially to yield hydroxide ions, as shown here:

$$\mathrm{NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)}$$

This is, by definition, an acid-base reaction, in this case involving the transfer of H^+ ions from water molecules to ammonia molecules. Under typical conditions, only about 1% of the dissolved ammonia is present as ions.



Figure 7.3.4: Ammonia is a weak base used in a variety of applications. (a) Pure ammonia is commonly applied as an agricultural fertilizer. (b) Dilute solutions of ammonia are effective household cleansers. (credit a: modification of work by National Resources Conservation Service; credit b: modification of work by pat00139)

A **neutralization reaction** is a specific type of acid-base reaction in which the reactants are an acid and a base (but not water), and the products are often a salt and water

$$\mathrm{acid + base \longrightarrow salt + water}$$

To illustrate a neutralization reaction, consider what happens when a typical antacid such as milk of magnesia (an aqueous suspension of solid $Mg(OH)_2$) is ingested to ease symptoms associated with excess stomach acid (HCl):

$$\mathrm{Mg(OH)_2(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + 2\,H_2O(l)}$$

Note that in addition to water, this reaction produces a salt, magnesium chloride.

Example 7.3.2: Writing Equations for Acid-Base Reactions

Write balanced chemical equations for the acid-base reactions described here:

a. the weak acid hydrogen hypochlorite reacts with water
b. a solution of barium hydroxide is neutralized with a solution of nitric acid

Solution

(a) The two reactants are provided, HOCl and H_2O. Since the substance is reported to be an acid, its reaction with water will involve the transfer of H^+ from HOCl to H_2O to generate hydronium ions, H_3O^+ and hypochlorite ions, OCl^-.

$$\mathrm{HOCl(aq) + H_2O(l) \rightleftharpoons OCl^-(aq) + H_3O^+(aq)}$$

A double-arrow is appropriate in this equation because it indicates the HOCl is a weak acid that has not reacted completely.

(b) The two reactants are provided, $Ba(OH)_2$ and HNO_3. Since this is a neutralization reaction, the two products will be water and a salt composed of the cation of the ionic hydroxide (Ba^{2+}) and the anion generated when the acid transfers its hydrogen ion

$$\mathrm{Ba(OH)_2(aq) + 2\,HNO_3(aq) \longrightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)}$$

? Exercise 7.3.2

Write the net ionic equation representing the neutralization of any strong acid with an ionic hydroxide. (Hint: Consider the ions produced when a strong acid is dissolved in water.)

Answer

$$\mathrm{H_3O^+(aq) + OH^-(aq) \longrightarrow 2\,H_2O(l)}$$

Chemistry in Everyday Life: Stomach Antacids

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$\mathrm{CaCO_3(s) + 2\,HCl(aq) \rightleftharpoons CaCl_2(aq) + H_2O(l) + CO_2(g)}$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

$$\mathrm{Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\,OH^-(aq)}$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that:

$$\mathrm{H_3O^+ + OH^- \rightleftharpoons 2\,H_2O(l)}$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

Chemistry in Everyday Life: Culinary Aspects of Chemistry

Examples of acid-base chemistry are abundant in the culinary world. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter "rises." Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure 7.3.5). It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This

reduces the odor of the fish, and also adds a “sour” taste that we seem to enjoy.



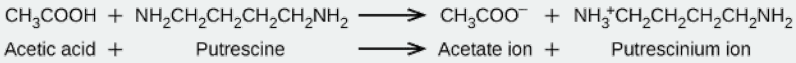


Figure 7.3.5: A neutralization reaction takes place between citric acid in lemons or acetic acid in vinegar, and the bases in the flesh of fish.

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favors the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavor of the vegetables with the acid making them taste sour.

Link to Learning

Explore the microscopic view of strong and weak acids and bases.

Oxidation-Reduction Reactions

Earth’s atmosphere contains about 20% molecular oxygen, O_2, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape the world. The term **oxidation** was originally used to describe chemical reactions involving O_2, but its meaning has evolved to refer to a broad and important reaction class known as *oxidation-reduction (redox) reactions*. A few examples of such reactions will be used to develop a clear picture of this classification.

Some redox reactions involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride:

$$2\,\mathrm{Na(s)} + \mathrm{Cl_2(g)} \longrightarrow 2\,\mathrm{NaCl(s)}$$

It is helpful to view the process with regard to each individual reactant, that is, to represent the fate of each reactant in the form of an equation called a **half-reaction**:

$$2\,\mathrm{Na(s)} \longrightarrow 2\,\mathrm{Na^+(s)} + 2\,\mathrm{e^-}$$

$$\mathrm{Cl_2(g)} + 2\,\mathrm{e^-} \longrightarrow 2\,\mathrm{Cl^-(s)}$$

These equations show that Na atoms *lose electrons* while Cl atoms (in the Cl_2 molecule) *gain electrons*, the “*s*” subscripts for the resulting ions signifying they are present in the form of a solid ionic compound. For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur:

$$\text{oxidation} = \text{loss of electrons}$$
$$\text{reduction} = \text{gain of electrons}$$

In this reaction, then, sodium is *oxidized* and chlorine undergoes **reduction**. Viewed from a more active perspective, sodium functions as a **reducing agent** (reductant), since it provides electrons to (or reduces) chlorine. Likewise, chlorine functions as an **oxidizing agent** (oxidant), as it effectively removes electrons from (oxidizes) sodium.

$$\text{reducing agent} = \text{species that is oxidized}$$
$$\text{oxidizing agent} = \text{species that is reduced}$$

Some redox processes, however, do not involve the transfer of electrons. Consider, for example, a reaction similar to the one yielding NaCl:

$$\mathrm{H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)}$$

The product of this reaction is a covalent compound, so transfer of electrons in the explicit sense is not involved. To clarify the similarity of this reaction to the previous one and permit an unambiguous definition of redox reactions, a property called *oxidation number* has been defined. The **oxidation number** (or **oxidation state**) of an element in a compound is the charge its atoms would possess *if the compound were ionic*. The following guidelines are used to assign oxidation numbers to each element in a molecule or ion.

1. The oxidation number of an atom in an elemental substance is zero.
2. The oxidation number of a monatomic ion is equal to the ion's charge.
3. Oxidation numbers for common nonmetals are usually assigned as follows:
   - Hydrogen: +1 when combined with nonmetals, −1 when combined with metals
   - Oxygen: −2 in most compounds, sometimes −1 (so-called peroxides, very rarely (so-called superoxides, positive values when combined with F (values vary)
   - Halogens: −1 for F always, −1 for other halogens except when combined with oxygen or other halogens (positive oxidation numbers in these cases, varying values)
4. The sum of oxidation numbers for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion.

Note: The proper convention for reporting charge is to write the number first, followed by the sign (e.g., 2+), while oxidation number is written with the reversed sequence, sign followed by number (e.g., +2). This convention aims to emphasize the distinction between these two related properties.

Example 7.3.3: Assigning Oxidation Numbers

Follow the guidelines in this section of the text to assign oxidation numbers to all the elements in the following species:

a. $\mathrm{H_2S}$
b. $\mathrm{SO_3^{2-}}$
c. $\mathrm{Na_2SO_4}$

Solution

(a) According to guideline 3, the oxidation number for H is +1.

Using this oxidation number and the compound's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$\text{charge on } H_2S = 0 = (2 \times +1) + (1 \times x)$$
$$x = 0 - (2 \times +1) = -2$$

(b) Guideline 3 suggests the oxidation number for oxygen is −2.

Using this oxidation number and the ion's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$\text{charge on } {SO_3}^{2-} = -2 = (3 \times -2) + (1 \times x)$$
$$x = -2 - (3 \times -2) = +4$$

(c) For ionic compounds, it's convenient to assign oxidation numbers for the cation and anion separately.

According to guideline 2, the oxidation number for sodium is +1.

Assuming the usual oxidation number for oxygen (−2 per guideline 3), the oxidation number for sulfur is calculated as directed by guideline 4:

$$\text{charge on } {SO_4}^{2-} = -2 = (4 \times -2) + (1 \times x)$$
$$x = -2 - (4 \times -2) = +6$$

? Exercise 7.3.3

Assign oxidation states to the elements whose atoms are underlined in each of the following compounds or ions:

a. $\mathrm{K\underline{N}O_3}$
b. $\mathrm{\underline{Al}H_3}$
c. $\mathrm{\underline{N}H_4^+}$
d. $\mathrm{H_2\underline{P}O_4^-}$

Answer

a. N, +5;
b. Al, +3;
c. N, −3;
d. P, +5

Using the oxidation number concept, an all-inclusive definition of redox reaction has been established. **Oxidation-reduction (redox) reactions** are those in which one or more elements involved undergo a change in oxidation number. (While the vast majority of redox reactions involve changes in oxidation number for two or more elements, a few interesting exceptions to this rule do exist Example 4.6.) Definitions for the complementary processes of this reaction class are correspondingly revised as shown here:

$$\begin{aligned} \text{oxidation} &= \text{increase in oxidation number} \\ \text{reduction} &= \text{decrease in oxidation number} \end{aligned}$$

Returning to the reactions used to introduce this topic, they may now both be identified as redox processes. In the reaction between sodium and chlorine to yield sodium chloride, sodium is oxidized (its oxidation number increases from 0 in Na to +1 in NaCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in NaCl). In the reaction between molecular hydrogen and chlorine, hydrogen is oxidized (its oxidation number increases from 0 in H_2 to +1 in HCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in HCl).

Several subclasses of redox reactions are recognized, including **combustion reactions** in which the reductant (also called a *fuel*) and oxidant (often, but not necessarily, molecular oxygen) react vigorously and produce significant amounts of heat, and often light, in the form of a flame. Solid rocket-fuel reactions are combustion processes. A typical propellant reaction in which solid aluminum is oxidized by ammonium perchlorate is represented by this equation:

$$\begin{aligned} \text{oxidation} &= \text{increase in oxidation number} \\ \text{reduction} &= \text{decrease in oxidation number} \end{aligned}$$

Link to Learning

Watch a brief video showing the test firing of a small-scale, prototype, hybrid rocket engine planned for use in the new Space Launch System being developed by NASA. The first engines firing at 3 s (green flame) use a liquid fuel/oxidant mixture, and the second, more powerful engines firing at 4 s (yellow flame) use a solid mixture.

Single-displacement (replacement) reactions are redox reactions in which an ion in solution is displaced (or replaced) via the oxidation of a metallic element. One common example of this type of reaction is the acid oxidation of certain metals:

$$\mathrm{Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)}$$

Metallic elements may also be oxidized by solutions of other metal salts; for example:

$$\mathrm{Cu(s) + 2\,AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2\,Ag(s)}$$

This reaction may be observed by placing copper wire in a solution containing a dissolved silver salt. Silver ions in solution are reduced to elemental silver at the surface of the copper wire, and the resulting Cu^{2+} ions dissolve in the solution to yield a characteristic blue color (Figure 7.3.6).

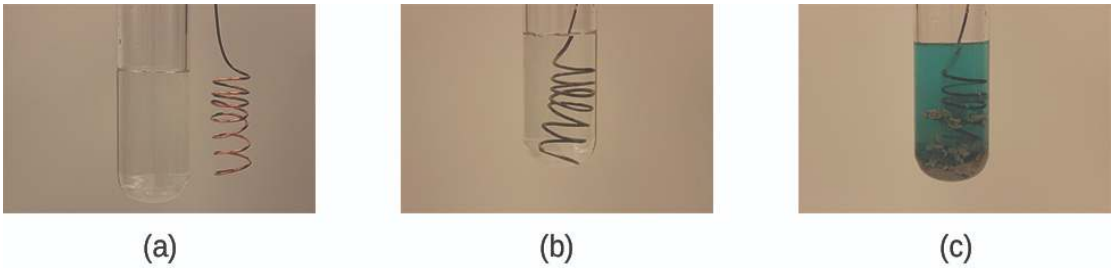


Figure 7.3.6: (a) A copper wire is shown next to a solution containing silver(I) ions. (b) Displacement of dissolved silver ions by copper ions results in (c) accumulation of gray-colored silver metal on the wire and development of a blue color in the solution, due to dissolved copper ions. (credit: modification of work by Mark Ott)

Example 7.3.4: Describing Redox Reactions

Identify which equations represent redox reactions, providing a name for the reaction if appropriate. For those reactions identified as redox, name the oxidant and reductant.

a. $\mathrm{ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)}$
b. $\mathrm{2\,Ga(l) + 3\,Br_2(l) \longrightarrow 2\,GaBr_3(s)}$
c. $\mathrm{2\,H_2O_2(aq) \longrightarrow 2\,H_2O(l) + O_2(g)}$
d. $\mathrm{BaCl_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,KCl(aq)}$
e. $\mathrm{C_2H_4(g) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 2\,H_2O(l)}$

Solution

Redox reactions are identified per definition if one or more elements undergo a change in oxidation number.

a. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.
b. This is a redox reaction. Gallium is oxidized, its oxidation number increasing from 0 in Ga(l) to +3 in $GaBr_3$(s). The reducing agent is Ga(l). Bromine is reduced, its oxidation number decreasing from 0 in Br_2(l) to −1 in $GaBr_3$(s). The oxidizing agent is Br_2(l).
c. This is a redox reaction. It is a particularly interesting process, as it involves the same element, oxygen, undergoing both oxidation and reduction (a so-called *disproportionation reaction)*. Oxygen is oxidized, its oxidation number increasing from −1 in H_2O_2(aq) to 0 in O_2(g). Oxygen is also reduced, its oxidation number decreasing from −1 in H_2O_2(aq) to −2 in H_2O(l). For disproportionation reactions, the same substance functions as an oxidant and a reductant.
d. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.
e. This is a redox reaction (combustion). Carbon is oxidized, its oxidation number increasing from −2 in C_2H_4(g) to +4 in CO_2(g). The reducing agent (fuel) is C_2H_4(g). Oxygen is reduced, its oxidation number decreasing from 0 in O_2(g) to −2 in H_2O(l). The oxidizing agent is O_2(g).

Exercise 7.3.4

This equation describes the production of tin(II) chloride:

$$\mathrm{Sn(s) + 2\,HCl(g) \longrightarrow SnCl_2(s) + H_2(g)}$$

Is this a redox reaction? If so, provide a more specific name for the reaction if appropriate, and identify the oxidant and reductant.

Answer

Yes, a single-replacement reaction. $\mathrm{Sn(s)}$ is the reductant and $\mathrm{HCl(g)}$ is the oxidant.

Balancing Redox Reactions via the Half-Reaction Method

Redox reactions that take place in aqueous media often involve water, hydronium ions, and hydroxide ions as reactants or products. Although these species are not oxidized or reduced, they do participate in chemical change in other ways (e.g., by providing the elements required to form oxyanions). Equations representing these reactions are sometimes very difficult to balance by inspection, so systematic approaches have been developed to assist in the process. One very useful approach is to use the method of half-reactions, which involves the following steps:

1. Write the two half-reactions representing the redox process.
2. Balance all elements except oxygen and hydrogen.
3. Balance oxygen atoms by adding H_2O molecules.
4. Balance hydrogen atoms by adding H^+ ions.
5. Balance charge by adding electrons.
6. If necessary, multiply each half-reaction's coefficients by the smallest possible integers to yield equal numbers of electrons in each.
7. Add the balanced half-reactions together and simplify by removing species that appear on both sides of the equation.
8. For reactions occurring in basic media (excess hydroxide ions), carry out these additional steps:
   a. Add OH^- ions to both sides of the equation in numbers equal to the number of H^+ ions.
   b. On the side of the equation containing both H^+ and OH^- ions, combine these ions to yield water molecules.
   c. Simplify the equation by removing any redundant water molecules.
9. Finally, check to see that both the number of atoms and the total charges[1] are balanced.

Example 7.3.5: Balancing Redox Reactions in Acidic Solution

Write a balanced equation for the reaction between dichromate ion and iron(II) to yield iron(III) and chromium(III) in acidic solution.

$$\mathrm{Cr_2O_7^{2-} + Fe^{2+} \longrightarrow Cr^{3+} + Fe^{3+}}$$

Solution

Step 1.Write the two half-reactions.

Each half-reaction will contain one reactant and one product with one element in common.

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+}}$$

$$\mathrm{Cr_2O_7^{2-} \longrightarrow Cr^{3+}}$$

Step 2. Balance oxygen atoms by adding $\mathrm{H_2O}$ molecules.

The iron half-reaction is already balanced, but the chromium half-reaction shows two Cr atoms on the left and one Cr atom on the right. Changing the coefficient on the right side of the equation to 2 achieves balance with regard to Cr atoms.

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+}}$$

$$\mathrm{Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+}}$$

Step 3. Balance charge by adding electrons.

The iron half-reaction does not contain O atoms. The chromium half-reaction shows seven O atoms on the left and none on the right, so seven water molecules are added to the right side.

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+}}$$

$$\mathrm{Cr_2O_7^{2-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O}$$

Step 4. Balance hydrogen atoms by adding $\mathrm{H^+}$ ions.

The iron half-reaction does not contain H atoms. The chromium half-reaction shows 14 H atoms on the right and none on the left, so 14 hydrogen ions are added to the left side.

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+}}$$

$$\mathrm{Cr_2O_7^{2-} + 14\,H^+ \longrightarrow 2\,Cr^{3+} + 7\,H_2O}$$

Step 5. Balance charge by adding electrons.

The iron half-reaction shows a total charge of 2+ on the left side (1 $\mathrm{Fe^{2+}}$ ion) and 3+ on the right side (1 $\mathrm{Fe^{3+}}$ ion). Adding one electron to the right side brings that side's total charge to (3+) + (1−) = 2+, and charge balance is achieved.

The chromium half-reaction shows a total charge of (1 × 2−) + (14 × 1+) = 12+ on the left side (1 $\mathrm{Cr_2O_7^{2-}}$ ion and 14 $\mathrm{H^+}$ ions). The total charge on the right side is (2 × 3+) = 6 + (2 $\mathrm{Cr^{3+}}$ ions). Adding six electrons to the left side will bring that side's total charge to (12+ + 6−) = 6+, and charge balance is achieved.

$$\mathrm{Fe^{2+} \longrightarrow Fe^{3+} + e^-}$$

$$\mathrm{Cr_2O_7^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O}$$

Step 6: Multiply the two half-reactions so the number of electrons in one reaction equals the number of electrons in the other reaction.

To be consistent with mass conservation, and the idea that redox reactions involve the transfer (not creation or destruction) of electrons, the iron half-reaction's coefficient must be multiplied by 6.

$$\mathrm{6\,Fe^{2+} \longrightarrow 6\,Fe^{3+} + 6\,e^-}$$

$$\mathrm{Cr_2O_7^{2-} + 6\,e^- + 14\,H^+ \longrightarrow 2\,Cr^{3+} + 7\,H_2O}$$

Step 7. Add the balanced half-reactions and cancel species that appear on both sides of the equation.

$$\mathrm{6\,Fe^{2+} + Cr_2O_7^{2-} + 6\,e^- + 14\,H^+ \longrightarrow 6\,Fe^{3+} + 6\,e^- + 2\,Cr^{3+} + 7\,H_2O}$$

Only the six electrons are redundant species. Removing them from each side of the equation yields the simplified, balanced equation here:

$$\mathrm{6\,Fe^{2+} + Cr_2O_7^{2-} + 14\,H^+ \longrightarrow 6\,Fe^{3+} + 2\,Cr^{3+} + 7\,H_2O}$$

A final check of atom and charge balance confirms the equation is balanced.

	Reactants	Products
Fe	6	6
Cr	2	2
O	7	7
H	14	14
charge	24+	24+

? Exercise 7.3.5

In basic solution, molecular chlorine, $\mathrm{Cl_2}$, reacts with hydroxide ions, $\mathrm{OH^-}$, to yield chloride ions, $\mathrm{Cl^-}$. and chlorate ions, $\mathrm{ClO_3^-}$. HINT: This is a *disproportionation reaction* in which the element chlorine is both oxidized and reduced. Write a balanced equation for this reaction.

Answer

$$\mathrm{3\,Cl_2(aq) + 6\,OH^-(aq) \longrightarrow 5\,Cl^-(aq) + ClO_3^-(aq) + 3\,H_2O(l)}$$

Footnotes

- 1The requirement of “charge balance” is just a specific type of “mass balance” in which the species in question are electrons. An equation must represent equal numbers of electrons on the reactant and product sides, and so both atoms and charges must be balanced.

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7.4: 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 7.4.1: Moles of Reactant Required in a Reaction

How many moles of I_2 are required to react with 0.429 mol of Al according to the following equation (see Figure 7.4.1)?

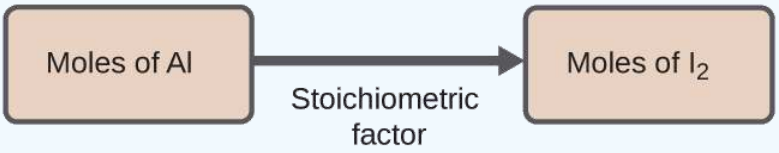
$$\mathrm{2\,Al + 3\,I_2 \longrightarrow 2\,AlI_3}$$



Figure 7.4.1: 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 7.4.1

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 7.4.2: 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?

$$\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$$

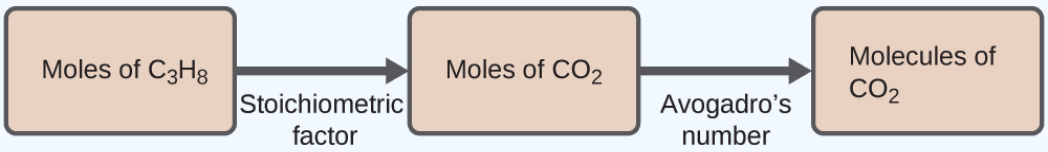
Solution

The approach here is the same as for Example 7.4.1, 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 7.4.2

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.

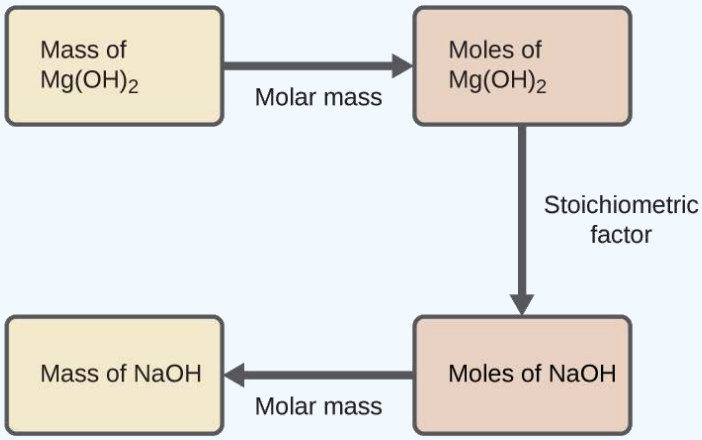
Example 7.4.3: Relating Masses of Reactants and Products

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 7.4.1 and Example 7.4.2 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 7.4.3

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

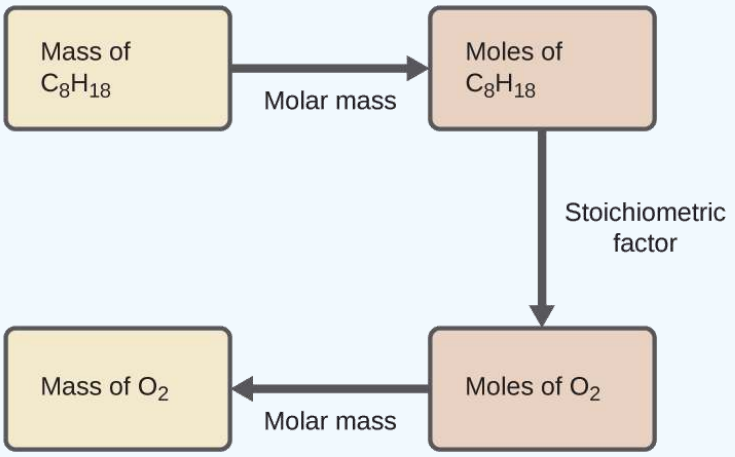
Example 7.4.4: Relating Masses of Reactants

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 7.4.3, 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 7.4.4

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 7.4.2: provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

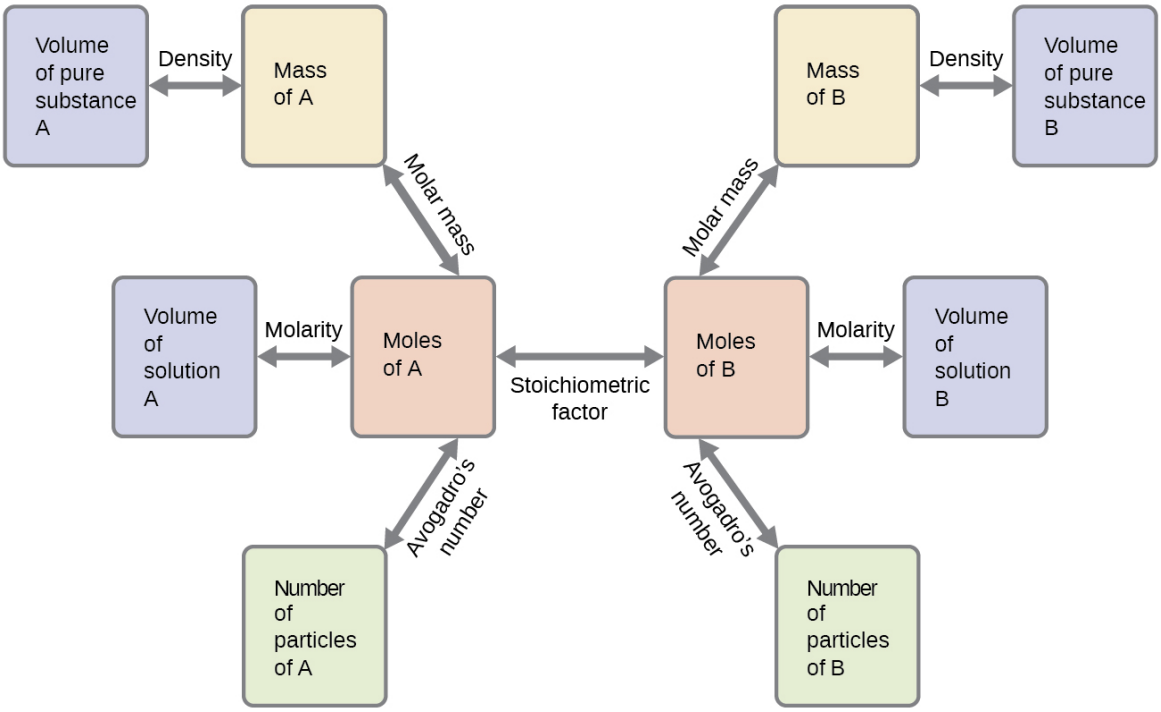


Figure 7.4.2: The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations.

Chemistry in Everyday Life: Airbags

Airbags (Figure 7.4.3) 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 7.4.3: Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

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7.5: 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 7.5.1):

$$1 \text{ slice of cheese} + 2 \text{ slices of bread} \longrightarrow 1 \text{ sandwich}$$

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*.

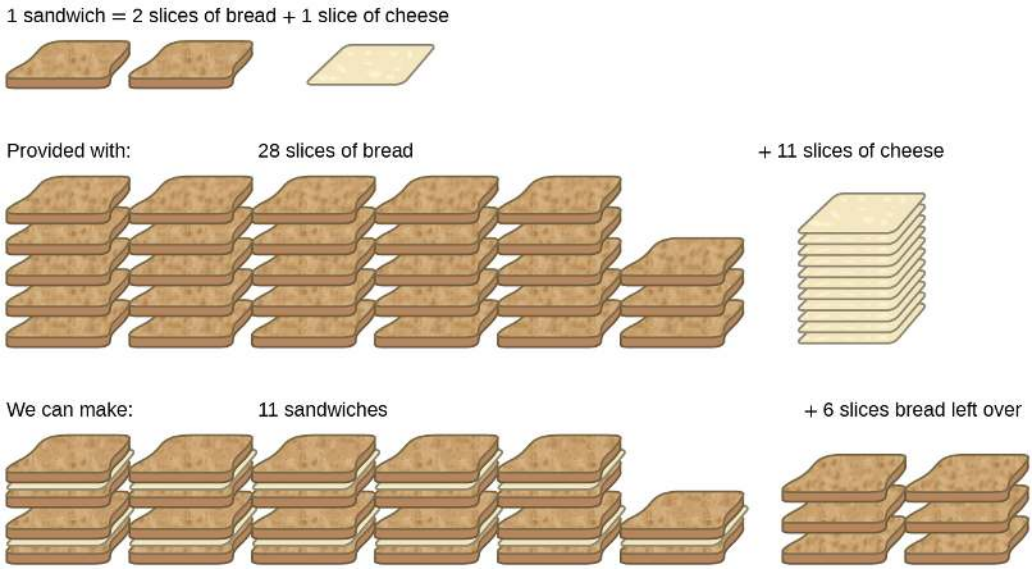


Figure 7.5.1: Sandwich making can illustrate the concepts of limiting and excess reactants.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)$$

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 7.5.2).

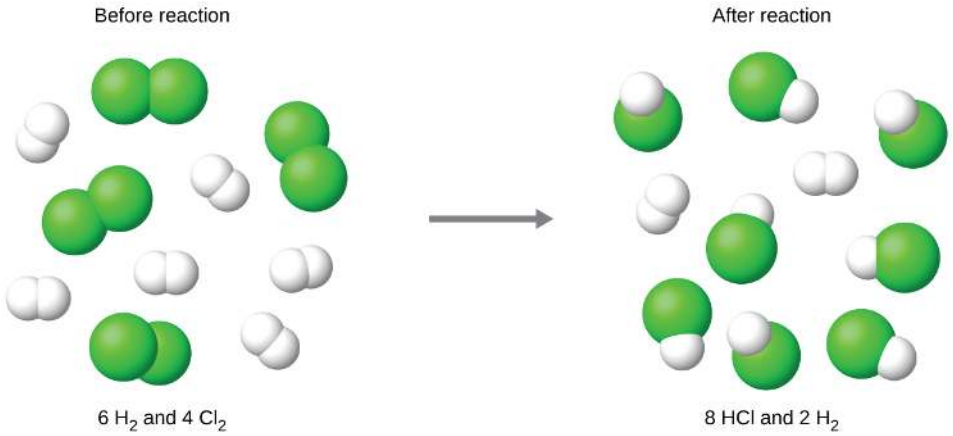


Figure 7.5.2: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

Link to Learning

View this interactive simulation illustrating the concepts of limiting and excess reactants.

Example 7.5.1: 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 7.5.1

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**:

$$\text{percent yield } = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{7.5.1}$$

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 7.5.2: 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 7.5.1 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 7.5.2

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:

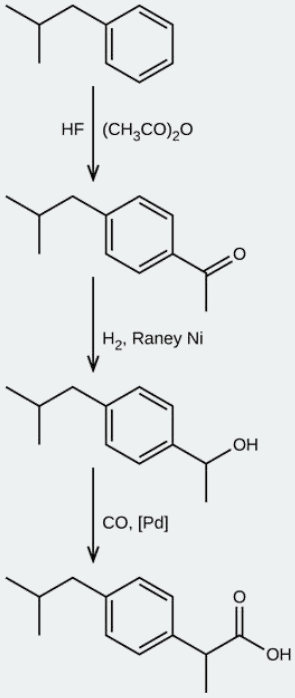
$$\text{atom economy} = \frac{\text{mass of product}}{\text{mass of reactants}} \times 100\%$$

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 7.5.3). 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 7.5.3: (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)

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7.6: Quantitative Chemical Analysis

Learning Objectives

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

- Describe the fundamental aspects of titrations and gravimetric analysis.
- Perform stoichiometric calculations using typical titration and gravimetric data.

In the 18th century, the strength (actually the concentration) of vinegar samples was determined by noting the amount of potassium carbonate, K_2CO_3, which had to be added, a little at a time, before bubbling ceased. The greater the weight of potassium carbonate added to reach the point where the bubbling ended, the more concentrated the vinegar.

We now know that the effervescence that occurred during this process was due to reaction with acetic acid, CH_3CO_2H, the compound primarily responsible for the odor and taste of vinegar. Acetic acid reacts with potassium carbonate according to the following equation:

The bubbling was due to the production of CO_2.

The test of vinegar with potassium carbonate is one type of quantitative analysis—the determination of the amount or concentration of a substance in a sample. In the analysis of vinegar, the concentration of the solute (acetic acid) was determined from the amount of reactant that combined with the solute present in a known volume of the solution. In other types of chemical analyses, the amount of a substance present in a sample is determined by measuring the amount of product that results.

Titration

The described approach to measuring vinegar strength was an early version of the analytical technique known as titration analysis. A typical titration analysis involves the use of a buret (Figure 7.6.1) to make incremental additions of a solution containing a known concentration of some substance (the titrant) to a sample solution containing the substance whose concentration is to be measured (the analyte). The titrant and analyte undergo a chemical reaction of known stoichiometry, and so measuring the volume of titrant solution required for complete reaction with the analyte (the equivalence point of the titration) allows calculation of the analyte concentration. The equivalence point of a titration may be detected visually if a distinct change in the appearance of the sample solution accompanies the completion of the reaction. The halt of bubble formation in the classic vinegar analysis is one such example, though, more commonly, special dyes called indicators are added to the sample solutions to impart a change in color at or very near the equivalence point of the titration. Equivalence points may also be detected by measuring some solution property that changes in a predictable way during the course of the titration. Regardless of the approach taken to detect a titration's equivalence point, the volume of titrant actually measured is called the end point. Properly designed titration methods typically ensure that the difference between the equivalence and end points is negligible. Though any type of chemical reaction may serve as the basis for a titration analysis, the three described in this chapter (precipitation, acid-base, and redox) are most common. Additional details regarding titration analysis are provided in the chapter on acid-base equilibria.

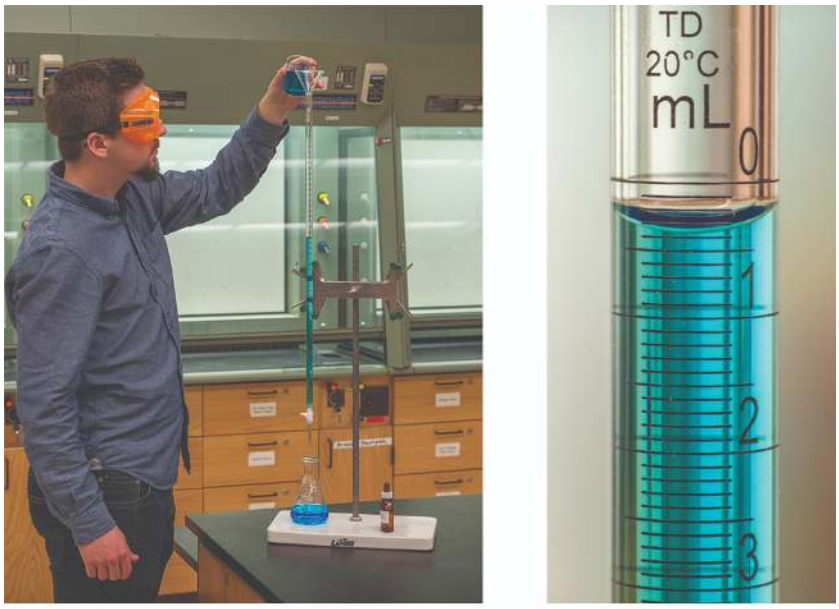


Figure 7.6.1: (a) A student fills a buret in preparation for a titration analysis. (b) A typical buret permits volume measurements to the nearest 0.01 mL. (credit a: modification of work by Mark Blaser and Matt Evans; credit b: modification of work by Mark Blaser and Matt Evans)

Example 4.14: Titration Analysis

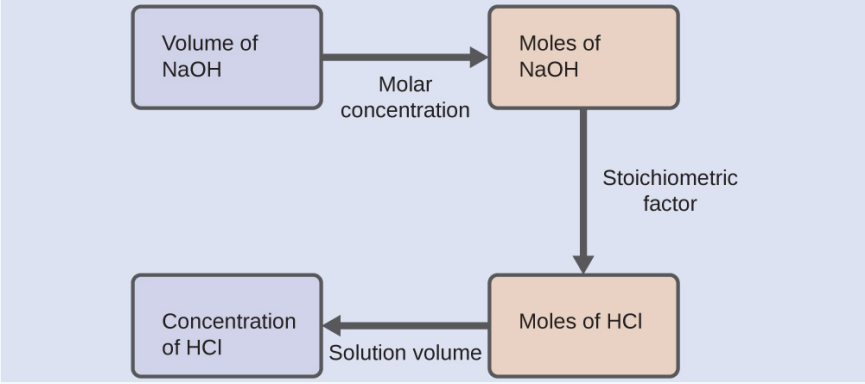
The end point in a titration of a 50.00-mL sample of aqueous HCl was reached by addition of 35.23 mL of 0.250 M NaOH titrant. The titration reaction is:

What is the molarity of the HCl?

Solution

As for all reaction stoichiometry calculations, the key issue is the relation between the molar amounts of the chemical species of interest as depicted in the balanced chemical equation. The approach outlined in previous modules of this chapter is followed, with additional considerations required, since the amounts of reactants provided and requested are expressed as solution concentrations.

For this exercise, the calculation will follow the following outlined steps:



The molar amount of HCl is calculated to be:

Using the provided volume of HCl solution and the definition of molarity, the HCl concentration is:

Note: For these types of titration calculations, it is convenient to recognize that solution molarity is also equal to the number of *milli*moles of solute per *milli*liter of solution:

Using this version of the molarity unit will shorten the calculation by eliminating two conversion factors:

? Exercise 7.6.1

A 20.00-mL sample of aqueous oxalic acid, $H_2C_2O_4$, was titrated with a 0.09113-*M* solution of potassium permanganate, $KMnO_4$ (see net ionic equation below).

A volume of 23.24 mL was required to reach the end point. What is the oxalic acid molarity?

Answer

0.2648 *M*

Gravimetric Analysis

A gravimetric analysis is one in which a sample is subjected to some treatment that causes a change in the physical state of the analyte that permits its separation from the other components of the sample. Mass measurements of the sample, the isolated analyte, or some other component of the analysis system, used along with the known stoichiometry of the compounds involved, permit calculation of the analyte concentration. Gravimetric methods were the first techniques used for quantitative chemical analysis, and they remain important tools in the modern chemistry laboratory.

The required change of state in a gravimetric analysis may be achieved by various physical and chemical processes. For example, the moisture (water) content of a sample is routinely determined by measuring the mass of a sample before and after it is subjected to a controlled heating process that evaporates the water. Also common are gravimetric techniques in which the analyte is subjected to a precipitation reaction of the sort described earlier in this chapter. The precipitate is typically isolated from the reaction mixture by filtration, carefully dried, and then weighed (Figure 7.6.2). The mass of the precipitate may then be used, along with relevant stoichiometric relationships, to calculate analyte concentration.



Figure 7.6.2: Precipitate may be removed from a reaction mixture by filtration.

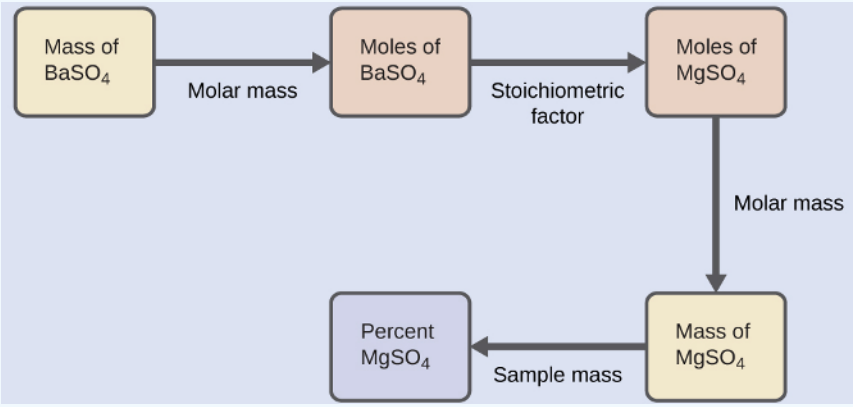
Example 4.15: Gravimetric Analysis

A 0.4550-g solid mixture containing $MgSO_4$ is dissolved in water and treated with an excess of $Ba(NO_3)_2$, resulting in the precipitation of 0.6168 g of $BaSO_4$.

What is the concentration (mass percent) of $MgSO_4$ in the mixture?

Solution

The plan for this calculation is similar to others used in stoichiometric calculations, the central step being the connection between the moles of $BaSO_4$ and $MgSO_4$ through their stoichiometric factor. Once the mass of $MgSO_4$ is computed, it may be used along with the mass of the sample mixture to calculate the requested percentage concentration.



The mass of $MgSO_4$ that would yield the provided precipitate mass is

The concentration of $MgSO_4$ in the sample mixture is then calculated to be

? Exercise 7.6.1

What is the percent of chloride ion in a sample if 1.1324 g of the sample produces 1.0881 g of AgCl when treated with excess Ag^+?

Answer

23.76%

The elemental composition of hydrocarbons and related compounds may be determined via a gravimetric method known as combustion analysis. In a combustion analysis, a weighed sample of the compound is heated to a high temperature under a stream of oxygen gas, resulting in its complete combustion to yield gaseous products of known identities. The complete combustion of hydrocarbons, for example, will yield carbon dioxide and water as the only products. The gaseous combustion products are swept through separate, preweighed collection devices containing compounds that selectively absorb each product (Figure 7.6.3). The mass increase of each device corresponds to the mass of the absorbed product and may be used in an appropriate stoichiometric calculation to derive the mass of the relevant element.

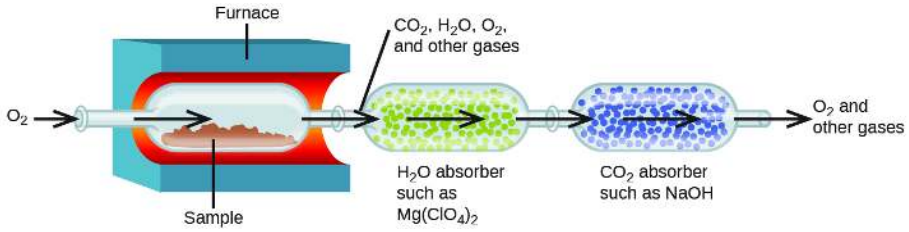


Figure 7.6.3: This schematic diagram illustrates the basic components of a combustion analysis device for determining the carbon and hydrogen content of a sample.

✓ Example 4.16: Combustion Analysis

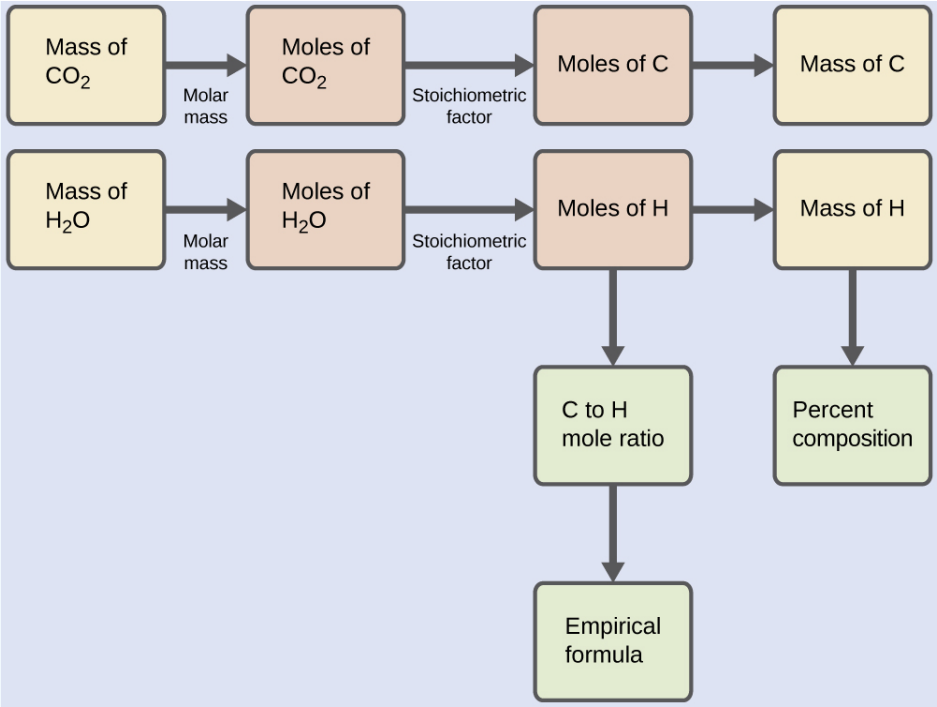
Polyethylene is a hydrocarbon polymer used to produce food-storage bags and many other flexible plastic items. A combustion analysis of a 0.00126-g sample of polyethylene yields 0.00394 g of CO_2 and 0.00161 g of H_2O. What is the empirical formula of polyethylene?

Solution

The primary assumption in this exercise is that all the carbon in the sample combusted is converted to carbon dioxide, and all the hydrogen in the sample is converted to water:

Note that a balanced equation is not necessary for the task at hand. To derive the empirical formula of the compound, only the subscripts *x* and *y* are needed.

First, calculate the molar amounts of carbon and hydrogen in the sample, using the provided masses of the carbon dioxide and water, respectively. With these molar amounts, the empirical formula for the compound may be written as described in the previous chapter of this text. An outline of this approach is given in the following flow chart:



The empirical formula for the compound is then derived by identifying the smallest whole-number multiples for these molar amounts. The H-to-C molar ratio is

and the empirical formula for polyethylene is CH_2.

Exercise 7.6.1

A 0.00215-g sample of polystyrene, a polymer composed of carbon and hydrogen, produced 0.00726 g of CO_2 and 0.00148 g of H_2O in a combustion analysis. What is the empirical formula for polystyrene?

Answer

CH

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7.7: Key Terms

acid | substance that produces H_3O^+ when dissolved in water

acid-base reaction | reaction involving the transfer of a hydrogen ion between reactant species

actual yield | amount of product formed in a reaction

analyte | chemical species of interest

balanced equation | chemical equation with equal numbers of atoms for each element in the reactant and product

base | substance that produces OH^- when dissolved in water

buret | device used for the precise delivery of variable liquid volumes, such as in a titration analysis

chemical equation | symbolic representation of a chemical reaction

coefficient | number placed in front of symbols or formulas in a chemical equation to indicate their relative amount

combustion analysis | gravimetric technique used to determine the elemental composition of a compound via the collection and weighing of its gaseous combustion products

combustion reaction | vigorous redox reaction producing significant amounts of energy in the form of heat and, sometimes, light

complete ionic equation | chemical equation in which all dissolved ionic reactants and products, including spectator ions, are explicitly represented by formulas for their dissociated ions

end point | measured volume of titrant solution that yields the change in sample solution appearance or other property expected for stoichiometric equivalence (see *equivalence point*)

equivalence point | volume of titrant solution required to react completely with the analyte in a titration analysis; provides a stoichiometric amount of titrant for the sample's analyte according to the titration reaction

excess reactant | reactant present in an amount greater than required by the reaction stoichiometry

gravimetric analysis | quantitative chemical analysis method involving the separation of an analyte from a sample by a physical or chemical process and subsequent mass measurements of the analyte, reaction product, and/or sample

half-reaction | an equation that shows whether each reactant loses or gains electrons in a reaction.

indicator | substance added to the sample in a titration analysis to permit visual detection of the end point

insoluble | of relatively low solubility; dissolving only to a slight extent

limiting reactant | reactant present in an amount lower than required by the reaction stoichiometry, thus limiting the amount of product generated

molecular equation | chemical equation in which all reactants and products are represented as neutral substances

net ionic equation | chemical equation in which only those dissolved ionic reactants and products that undergo a chemical or physical change are represented (excludes spectator ions)

neutralization reaction | reaction between an acid and a base to produce salt and water

oxidation | process in which an element's oxidation number is increased by loss of electrons

oxidation number | (also, oxidation state) the charge each atom of an element would have in a compound if the compound were ionic

oxidation-reduction reaction | (also, redox reaction) reaction involving a change in oxidation number for one or more reactant elements

oxidizing agent | (also, oxidant) substance that brings about the oxidation of another substance, and in the process becomes reduced

percent yield | measure of the efficiency of a reaction, expressed as a percentage of the theoretical yield

precipitate | insoluble product that forms from reaction of soluble reactants

precipitation reaction | reaction that produces one or more insoluble products; when reactants are ionic compounds, sometimes called double-displacement or metathesis

product | substance formed by a chemical or physical change; shown on the right side of the arrow in a chemical equation

quantitative analysis | the determination of the amount or concentration of a substance in a sample

reactant | substance undergoing a chemical or physical change; shown on the left side of the arrow in a chemical equation

reducing agent | (also, reductant) substance that brings about the reduction of another substance, and in the process becomes oxidized

reduction | process in which an element's oxidation number is decreased by gain of electrons

salt | ionic compound that can be formed by the reaction of an acid with a base that contains a cation and an anion other than hydroxide or oxide

single-displacement reaction | (also, replacement) redox reaction involving the oxidation of an elemental substance by an ionic species

solubility | the extent to which a substance may be dissolved in water, or any solvent

soluble | of relatively high solubility; dissolving to a relatively large extent

spectator ion | ion that does not undergo a chemical or physical change during a reaction, but its presence is required to maintain charge neutrality

stoichiometric factor | ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products

stoichiometry | relationships between the amounts of reactants and products of a chemical reaction

strong acid | acid that reacts completely when dissolved in water to yield hydronium ions

strong base | base that reacts completely when dissolved in water to yield hydroxide ions

theoretical yield | amount of product that may be produced from a given amount of reactant(s) according to the reaction stoichiometry

titrant | solution containing a known concentration of substance that will react with the analyte in a titration analysis

titration analysis | quantitative chemical analysis method that involves measuring the volume of a reactant solution required to completely react with the analyte in a sample

weak acid | acid that reacts only to a slight extent when dissolved in water to yield hydronium ions

weak base | base that reacts only to a slight extent when dissolved in water to yield hydroxide ions

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7.8: Key Equations

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7.9: Summary

Chemical equations are symbolic representations of chemical and physical changes. Formulas for the substances undergoing the change (reactants) and substances generated by the change (products) are separated by an arrow and preceded by integer coefficients indicating their relative numbers. Balanced equations are those whose coefficients result in equal numbers of atoms for each element in the reactants and products. Chemical reactions in aqueous solution that involve ionic reactants or products may be represented more realistically by complete ionic equations and, more succinctly, by net ionic equations.

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties.

When reactions are carried out using less-than-stoichiometric quantities of reactants, the amount of product generated will be determined by the limiting reactant. The amount of product generated by a chemical reaction is its actual yield. This yield is often less than the amount of product predicted by the stoichiometry of the balanced chemical equation representing the reaction (its theoretical yield). The extent to which a reaction generates the theoretical amount of product is expressed as its percent yield.

The stoichiometry of chemical reactions may serve as the basis for quantitative chemical analysis methods. Titrations involve measuring the volume of a titrant solution required to completely react with a sample solution. This volume is then used to calculate the concentration of analyte in the sample using the stoichiometry of the titration reaction. Gravimetric analysis involves separating the analyte from the sample by a physical or chemical process, determining its mass, and then calculating its concentration in the sample based on the stoichiometry of the relevant process. Combustion analysis is a gravimetric method used to determine the elemental composition of a compound by collecting and weighing the gaseous products of its combustion.

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CHAPTER OVERVIEW

8: 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.

8.1: Introduction
8.2: Gas Pressure
8.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law
8.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions
8.5: Effusion and Diffusion of Gases
8.6: The Kinetic-Molecular Theory
8.7: Non-Ideal Gas Behavior
8.8: Key Terms
8.9: Key Equations
8.10: Summary

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8.1: 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 8.1.1) 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 8.1.1: 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.

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8.2: 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 8.2.1). 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.

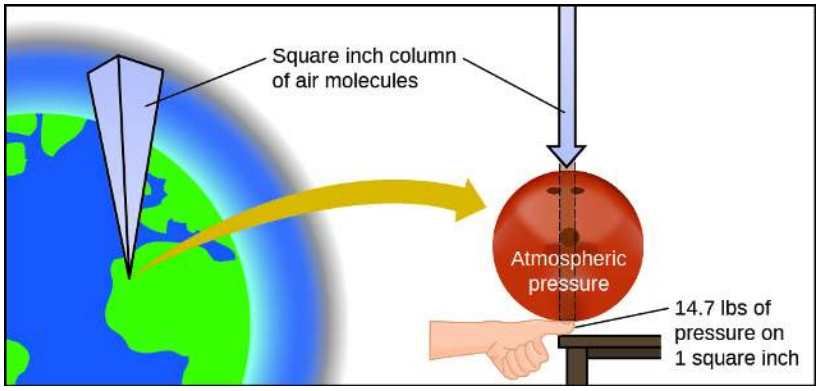


Figure 8.2.1: The atmosphere above us exerts a large pressure on objects at the surface of the earth, roughly equal to the weight of a bowling ball pressing on an area the size of a human thumbnail.

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 8.2.2—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 8.2.2: 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 8.2.1 provides some information on these and a few other common units for pressure measurements

Table 8.2.1: 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 8.2.1: 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 8.2.1.

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 8.2.1

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 8.2.3). 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.

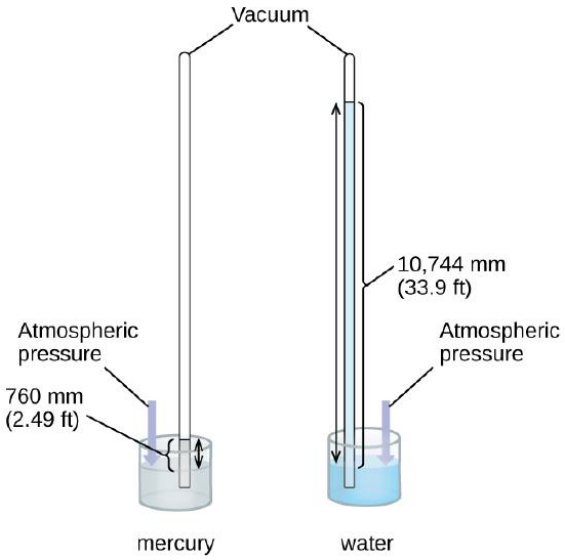


Figure 8.2.3: In a barometer, the height, *h*, of the column of liquid is used as a measurement of the air pressure. Using very dense liquid mercury (left) permits the construction of reasonably sized barometers, whereas using water (right) would require a barometer more than 30 feet tall.

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 8.2.2: 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 8.2.2

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 8.2.4) 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.

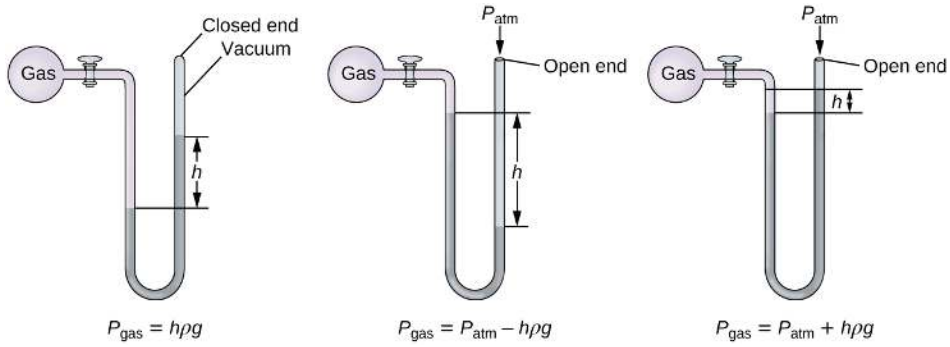
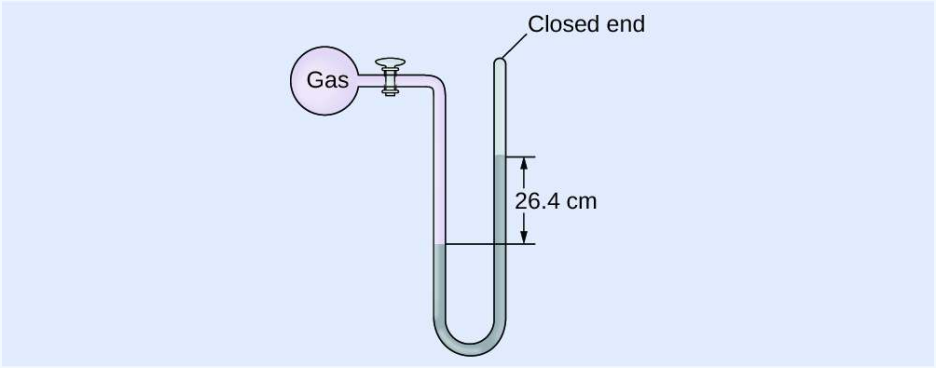


Figure 8.2.4: A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.

Example 8.2.3: 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 8.2.2, but it is simpler to just convert between units using Table 8.2.1.

a. $$26.4 em Hg_g \times \frac{10 mm Hg_g}{1 em Hg} \times \frac{1 \text{ torr}}{1 mm \Pi Hg} = 264 \text{ torr}$$

b. $$264 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \times \frac{101,325 Pa}{1 \text{ atm}} = 35,200 Pa$$

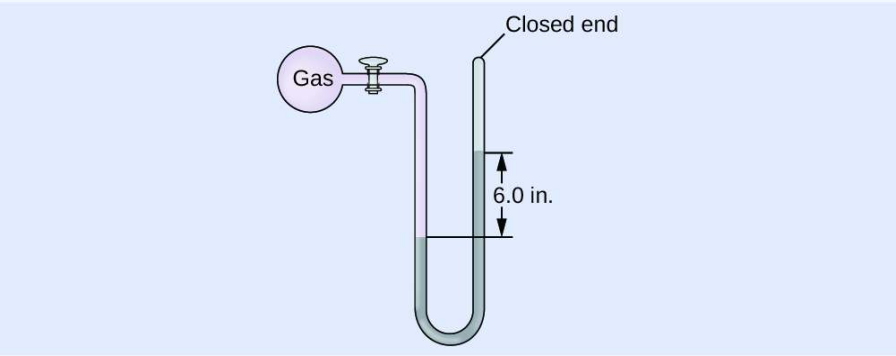
c. $$35,200 Pa \times \frac{1 bar}{100,000 Pa} = 0.352 bar$$

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Exercise 8.2.3

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



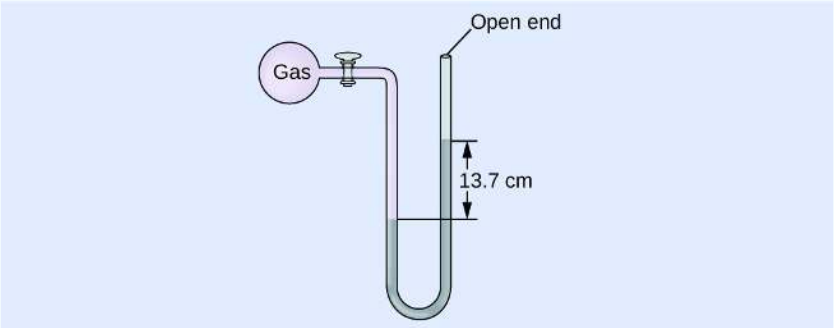
Answer

(a) ~150 torr; (b) ~20,000 Pa; (c) ~0.20 bar

Example 8.2.4: 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

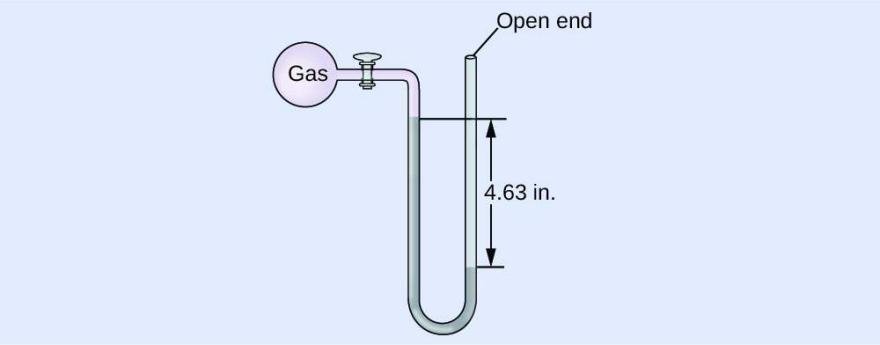
b. $$897 mm Hg_g \times \frac{1 atm}{760 mm Hg_g} = 1.18 atm$$

c. $$1.18 \text{ atm} \times \frac{101.325 kPa}{1 - 20 \times 10^2 kPa} = 1.20$$

Exercise 8.2.4

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 8.2.5). 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 8.2.5: (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 8.2.6) are the result of thousands of measurements of air pressure, temperature, and the like, which are compiled, modeled, and analyzed in weather centers worldwide.

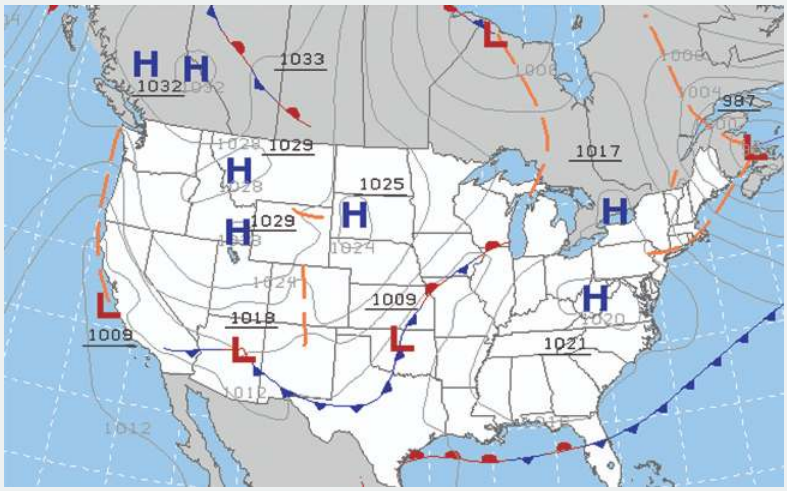


Figure 8.2.6: Meteorologists use weather maps to describe and predict weather. Regions of high (H) and low (L) pressure have large effects on weather conditions. The gray lines represent locations of constant pressure known as isobars. (credit: modification of work by National Oceanic and Atmospheric 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 8.2.7 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.

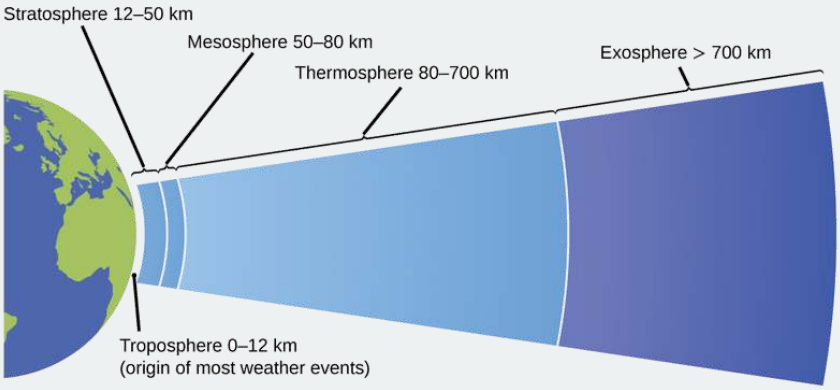


Figure 8.2.7: Earth's atmosphere has five layers: the troposphere, the stratosphere, the mesosphere, the thermosphere, and the exosphere.

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.

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8.3: 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 8.3.1), 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.

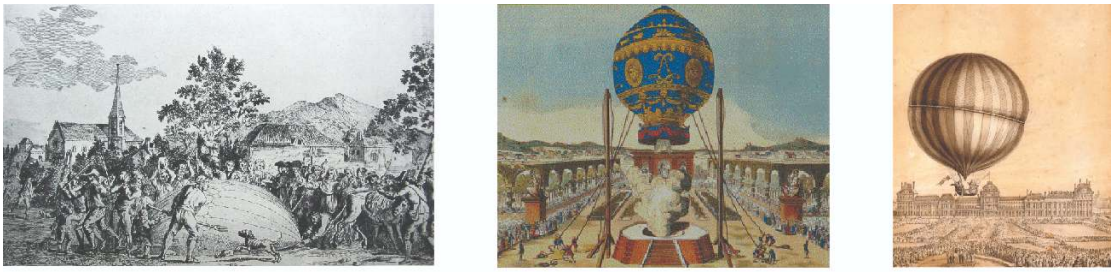


Figure 8.3.1: In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. 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 8.3.2) and the pressure increases.

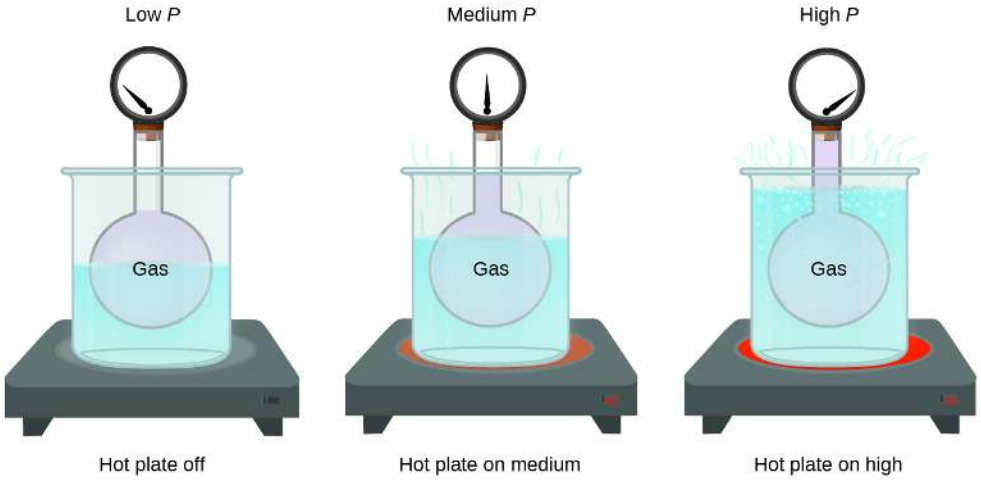


Figure 8.3.2: The effect of temperature on gas pressure: When the hot plate is off, the pressure of the gas in the sphere is relatively low. As the gas is heated, the pressure of the gas in the sphere increases.

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 8.3.3. 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

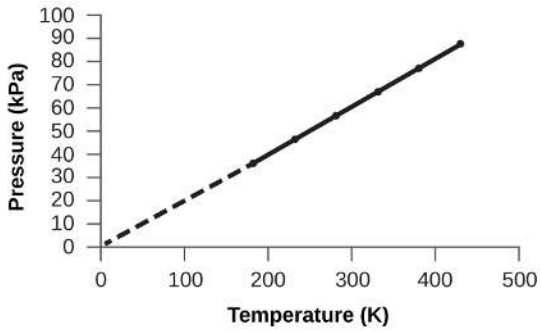


Figure 8.3.3: For a constant volume and amount of air, the pressure and temperature are directly proportional, provided the temperature is in kelvin. (Measurements cannot be made at lower temperatures because of the condensation of the gas.) When this line is extrapolated to lower pressures, it reaches 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 8.3.1: 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 8.3.1

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

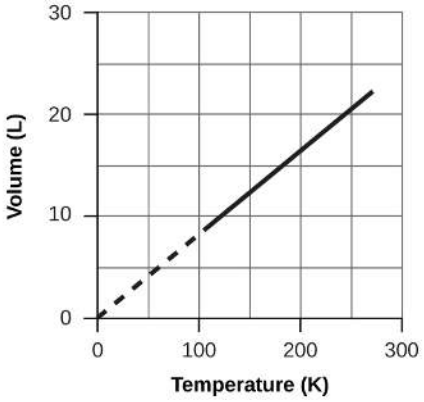


Figure 9.12 The volume and temperature are linearly related for 1 mole of methane gas at a constant pressure of 1 atm. If the temperature is in kelvin, volume and temperature are directly proportional. The line stops at 111 K because methane liquefies at this temperature; when extrapolated, it intersects 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:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}. \tag{8.3.1}$$

Example 8.3.2: 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 8.3.2

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 8.3.3: 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 8.3.3

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 8.3.4.

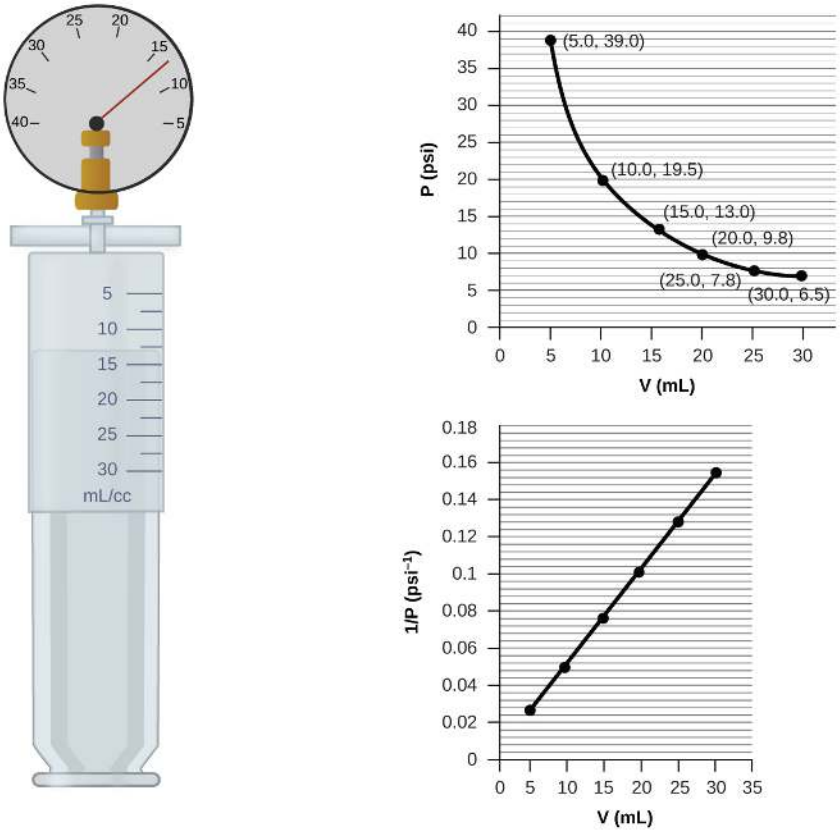


Figure 8.3.4: When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since P and V are inversely proportional, a graph of $\frac{1}{P}$ vs. V is linear.

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:

$$P \propto \frac{1}{V} \tag{8.3.2}$$

or

$$P = k \cdot \frac{1}{V} \tag{8.3.3}$$

or

$$P \cdot V = k \tag{8.3.4}$$

or

$$P_1V_1 = P_2V_2 \tag{8.3.5}$$

with *k* being a constant. Graphically, this relationship is shown by the straight line that results when plotting the inverse of the pressure

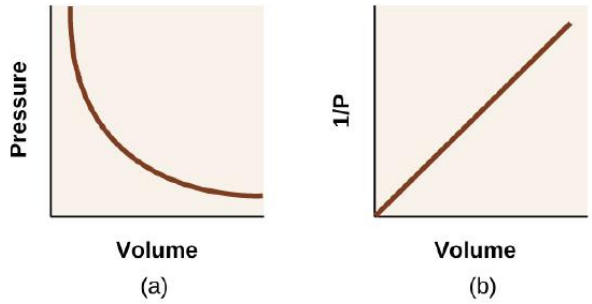


Figure 8.3.5: The relationship between pressure and volume is inversely proportional. (a) The graph of *P* vs. *V* is a hyperbola, whereas (b) the graph of $\frac{1}{P}$ vs. *V* is linear.

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 8.3.4: Volume of a Gas Sample

The sample of gas in Figure 8.3.4: 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 8.3.5
b. the $\frac{1}{P}$ vs. V graph in Figure 8.3.5
c. the Boyle's law equations (Equations 8.3.2 - 8.3.5)

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 8.3.4

The sample of gas in Figure 8.3.4 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 8.3.5
b. the $\frac{1}{P}$ vs. \(*V*\) graph in Figure 8.3.5
c. the Boyle's law equations (8.3.2 - 8.3.5)

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 8.3.6).

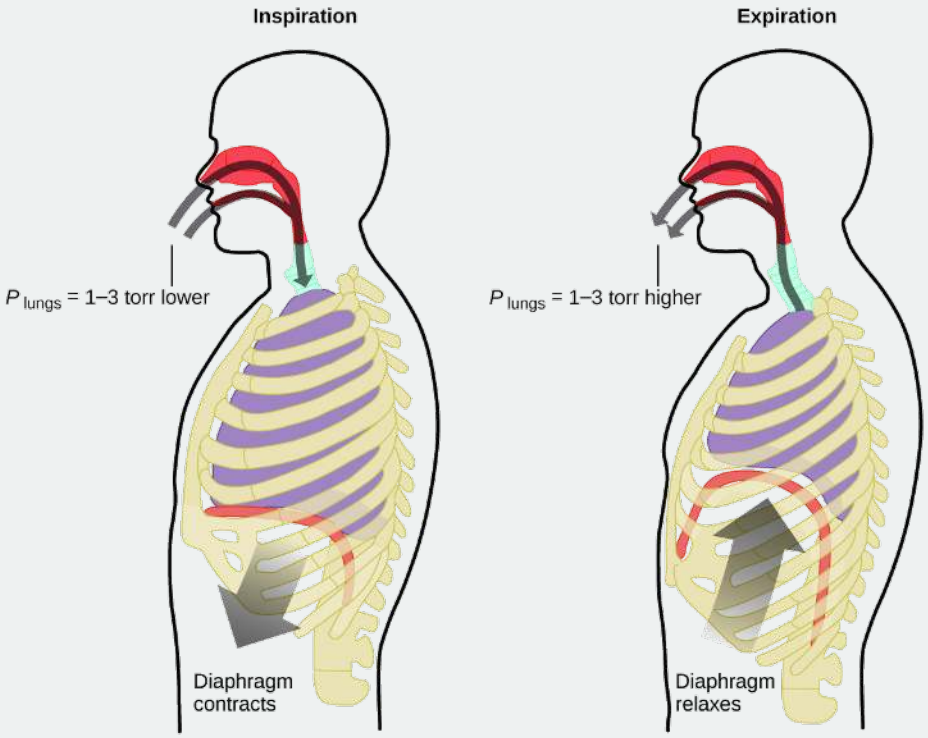


Figure 8.3.6: Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.

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:

$$V \propto n \tag{8.3.6}$$

or

$$V = k \times n \tag{8.3.7}$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \tag{8.3.8}$$

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 8.3.5: 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:

$$n = 655 \text{ g } \cancel{CH_4} \times \frac{1 \text{ mol}}{16.043 \text{ g } \cancel{CH_4}} = 40.8 \text{ mol}$$

$$T = 25^{\circ}C + 273 = 298 \text{ K}$$

$$P = 745 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.980 \text{ atm}$$

$$V = \frac{nRT}{P} = \frac{(40.8 \cancel{\text{ mol}}) \left(0.08206 \ L \text{ atm } \cancel{\text{mol}^{-1}\text{K}^{-1}}\right) (298 \cancel{\text{ K}})}{0.980 \cancel{\text{ atm}}} = 1.02 \times 10^3 \text{ L}$$

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 8.3.5

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 8.3.6: 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 8.3.7). 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 8.3.7: 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 8.3.6

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 8.3.8) or in the Caribbean, divers must understand how pressure affects a number of issues related to their comfort and safety.



Figure 8.3.8: 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 8.3.9).

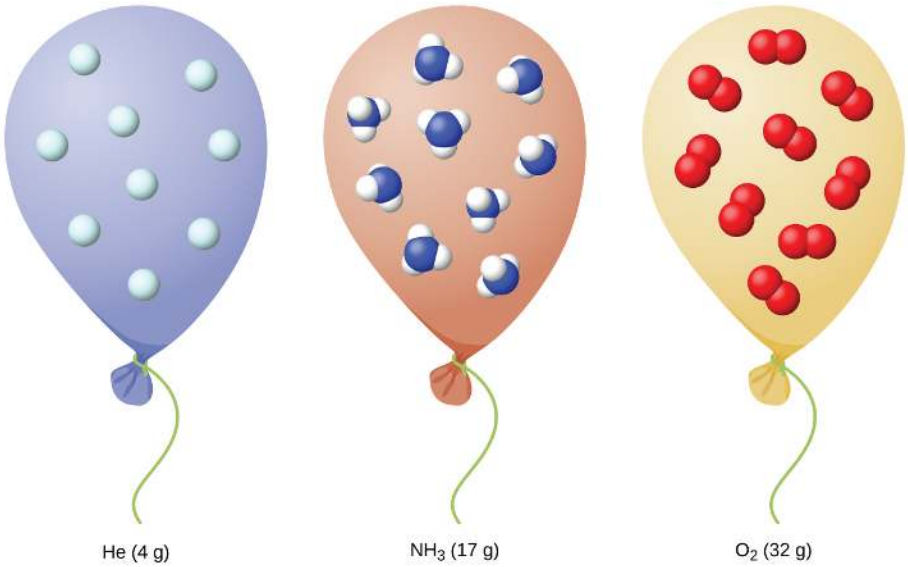


Figure 8.3.9: Regardless of its chemical identity, one mole of gas behaving ideally occupies a volume of ~22.4 L at STP.

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.

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8.4: 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 8.4.1.

Example 8.4.1: 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 8.4.1

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 8.4.2.

Example 8.4.2: 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 8.4.2

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 8.4.3: 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 8.4.1)

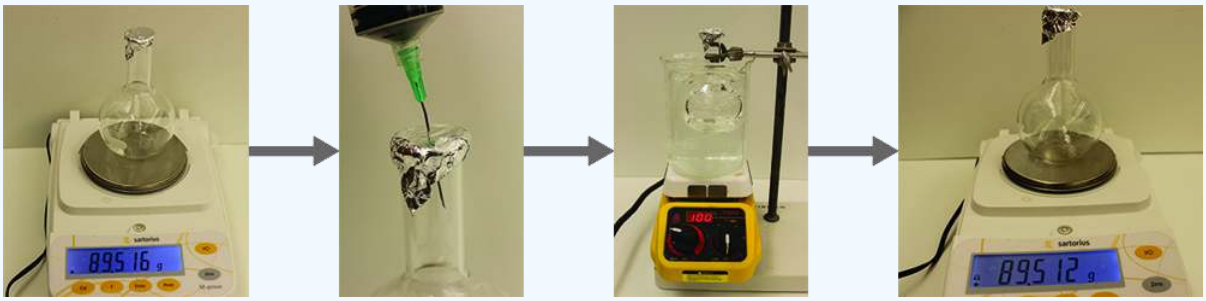


Figure 8.4.1: When the volatile liquid in the flask is heated past its boiling point, it becomes gas and drives air out of the flask. At the flask is filled with volatile liquid gas at the same pressure as the atmosphere. If the flask is then cooled to room temperature, the gas condenses and the mass 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 8.4.3

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 8.4.2). 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.

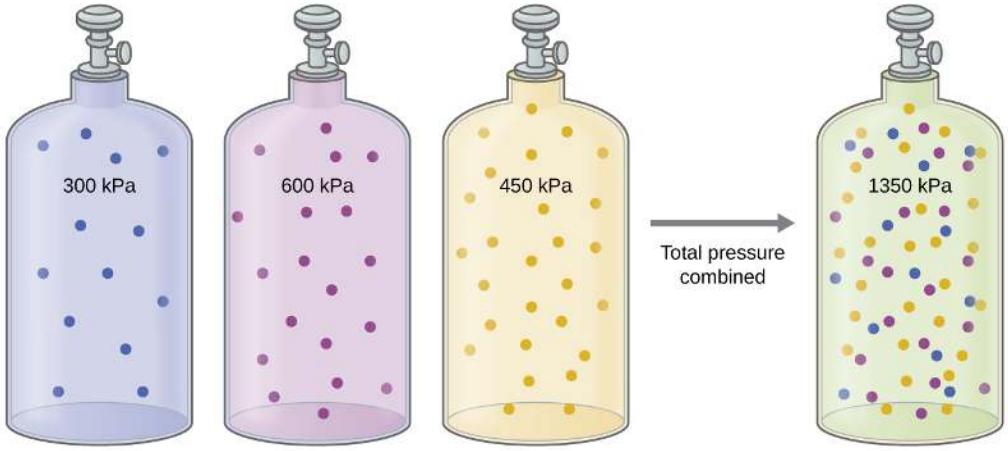


Figure 8.4.2: If equal-volume cylinders containing gasses at pressures of 300 kPa, 450 kPa, and 600 kPa are all combined in the same-size cylinder, the total pressure of the gas mixture is 1350 kPa.

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 8.4.4: 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 8.4.4

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 8.4.5: 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 8.4.5

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 8.4.3), the pressure of the gas is equal to the atmospheric pressure, which can be measured with a barometer.

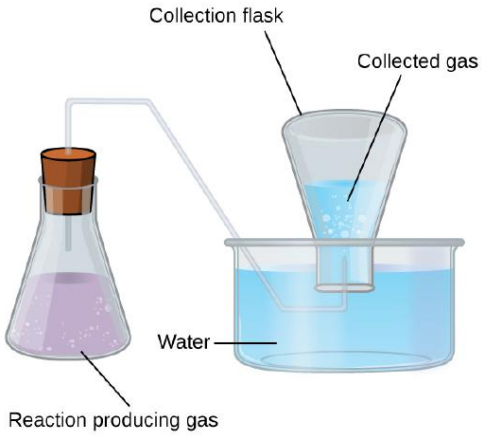


Figure 8.4.3: When a reaction produces a gas that is collected above water, the trapped gas is a mixture of the gas produced by the reaction and water vapor. If the collection flask is appropriately positioned to equalize the water levels both within and outside the flask, the pressure of the trapped 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 8.4.4); more detailed information on the temperature dependence of water vapor can be found in Table 8.4.1, and vapor pressure will be discussed in more detail in the next chapter on liquids.

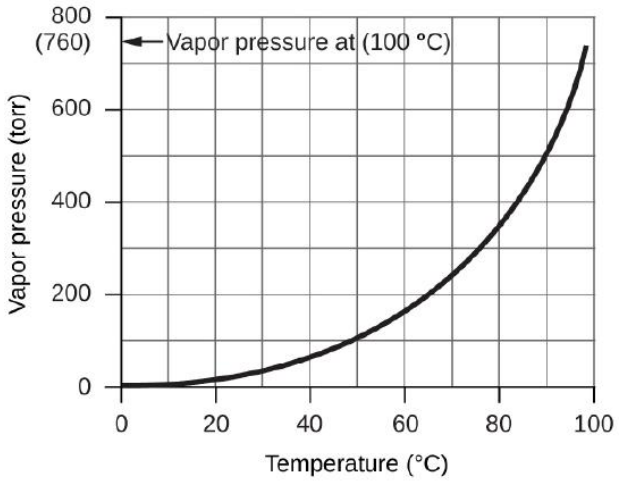


Figure 8.4.4: This graph shows the vapor pressure of water at sea level as a function of temperature.

Table 8.4.1: 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 8.4.6: 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 8.4.3, 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 8.4.6

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 8.4.5. 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.

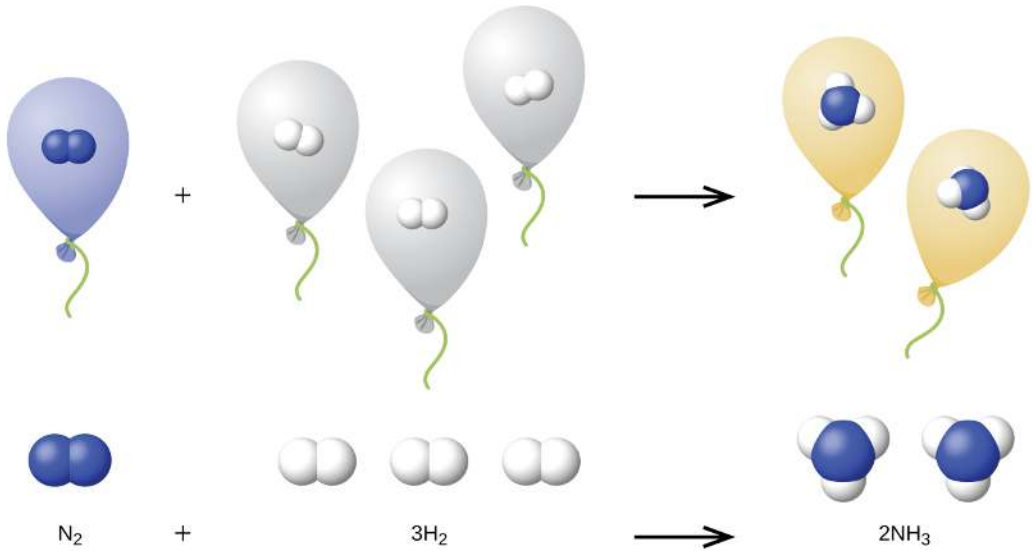


Figure 8.4.5: One volume of N_2 combines with three volumes of H_2 to form two volumes of NH_3.

Example 8.4.7: 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 8.4.7

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 8.4.8: 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 8.4.8

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 8.4.9: 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 8.4.9

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 8.4.6).

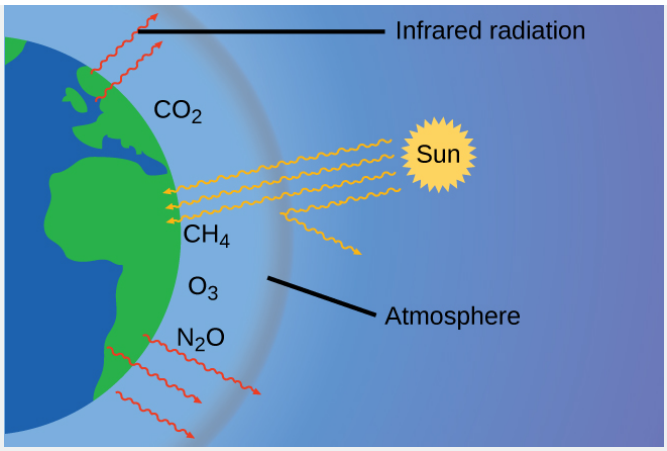


Figure 8.4.6: Greenhouse gases trap enough of the sun's energy to make the planet habitable—this is known as the greenhouse effect. Human activities are increasing greenhouse gas levels, warming the planet and causing more extreme weather events.

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 8.4.7).

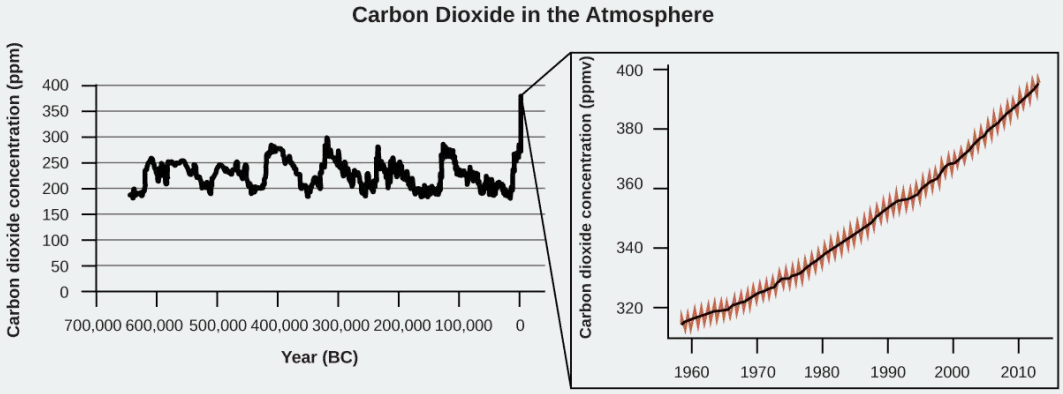


Figure 8.4.7: CO_2 levels over the past 700,000 years were typically from 200–300 ppm, with a steep, unprecedented increase over the past 50 years.

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 8.4.8) 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 8.4.8: 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

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8.5: 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 8.5.1). 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 8.5.1. 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).

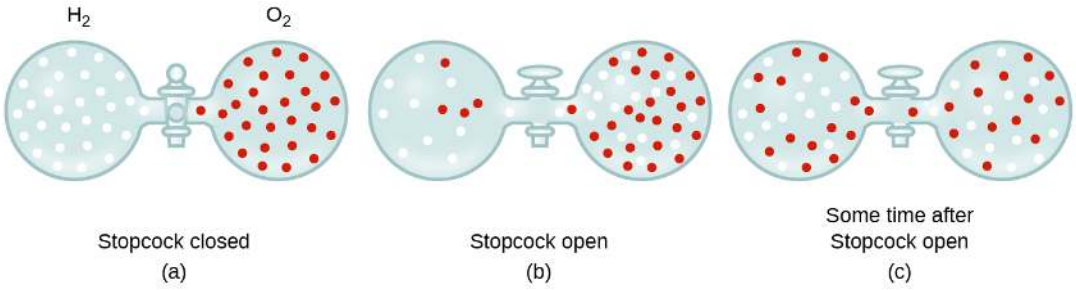


Figure 8.5.1: (a) Two gases, H_2 and O_2, are initially separated. (b) When the stopcock is opened, they mix together. The lighter gas, H_2, passes through the opening faster than O_2, so just after the stopcock is opened, more H_2 molecules move to the O_2 side than O_2 molecules move to the H_2 side. (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 8.5.2). 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.

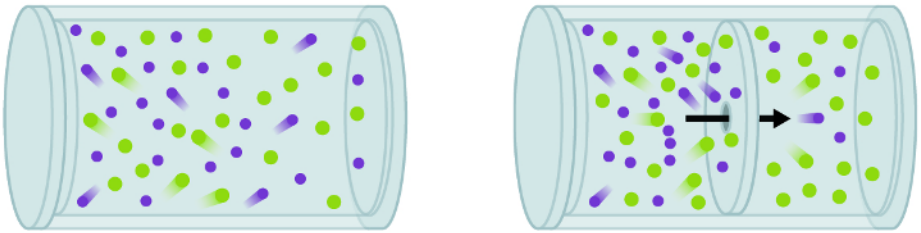


Figure 8.5.2: Diffusion involves the unrestricted dispersal of molecules throughout space due to their random motion. When this process is restricted to passage of molecules through very small openings in a physical barrier, the process is called effusion.

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 8.5.3). 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:

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

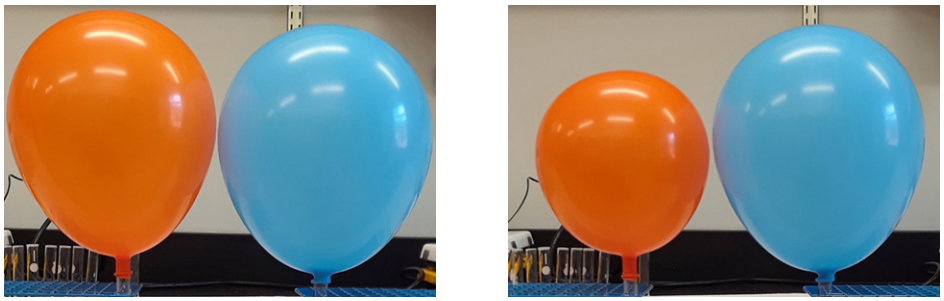


Figure 8.5.3: The left photograph shows two balloons inflated with different gases, helium (orange) and argon (blue).The right-side photograph shows the balloons approximately 12 hours after being filled, at which time the helium balloon has become noticeably more deflated than the argon balloon, due to the greater effusion rate of the lighter helium gas. (credit: modification of work by Paul Flowers)

Example 8.5.1: 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:

$$\frac{\text{rate of effusion of hydrogen}}{\text{rate of effusion of oxygen}} = \frac{\sqrt{32 g mol^{-1}}}{\sqrt{2 g mol^{-1}}} = \frac{\sqrt{16}}{\sqrt{1}} = \frac{4}{1}$$

Hydrogen effuses four times as rapidly as oxygen.

Exercise 8.5.1

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 8.5.2: 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 8.5.2

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 8.5.3: 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 8.5.3

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 8.5.4). But by connecting many diffusers in a sequence of stages (called a cascade), the desired level of enrichment can be attained.

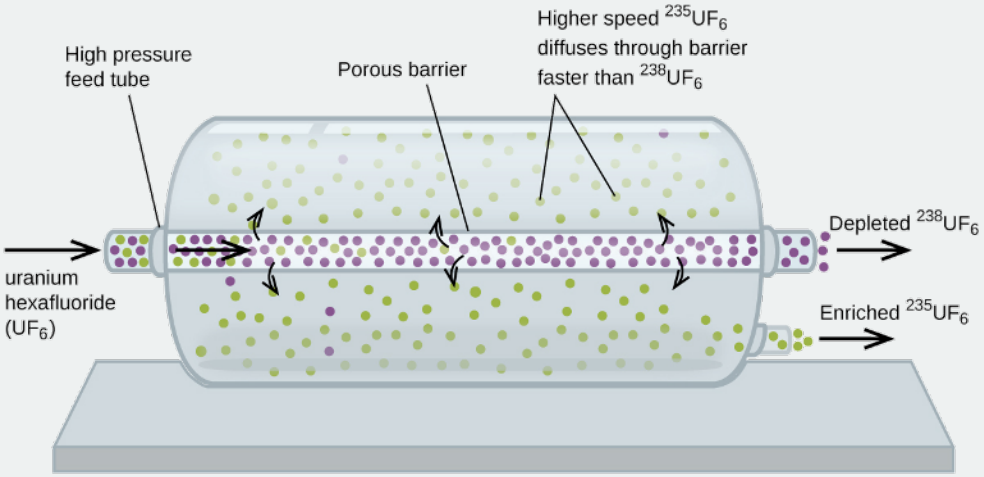


Figure 8.5.4: In a diffuser, gaseous UF_6 is pumped through a porous barrier, which partially separates $^{235}UF_6$ from $^{238}UF_6$ The UF_6 must pass through many large diffuser units to achieve sufficient enrichment in ^{235}U.

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.

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8.6: 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 8.6.1).
- *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 8.6.1).
- *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 8.6.1).
- *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.

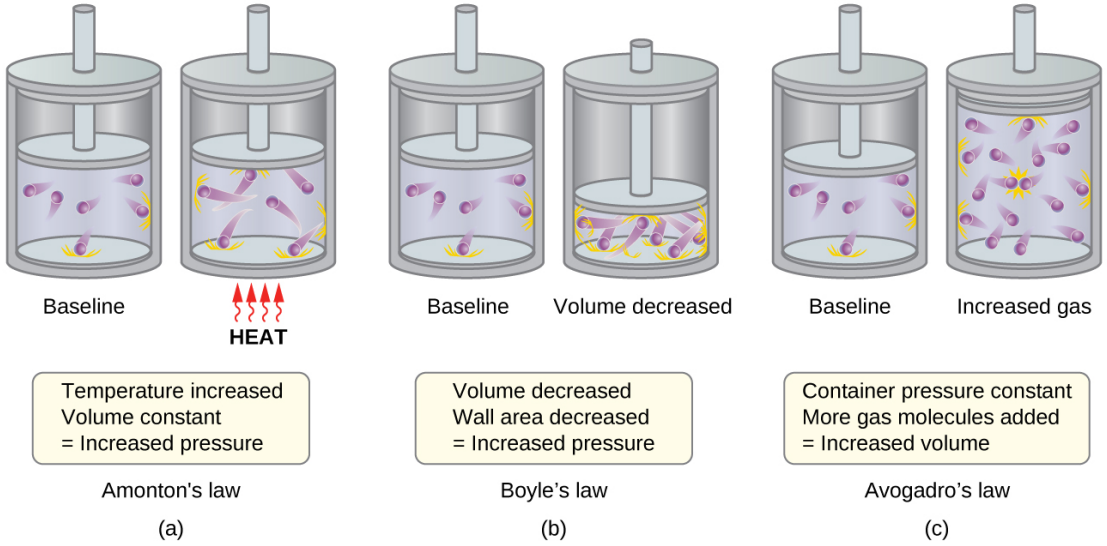


Figure 8.6.1: (a) When gas temperature increases, gas pressure increases due to increased force and frequency of molecular collisions. (b) When volume decreases, gas pressure increases due to increased frequency of molecular collisions. (c) When the amount of gas increases at a constant pressure, volume 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 8.6.2).

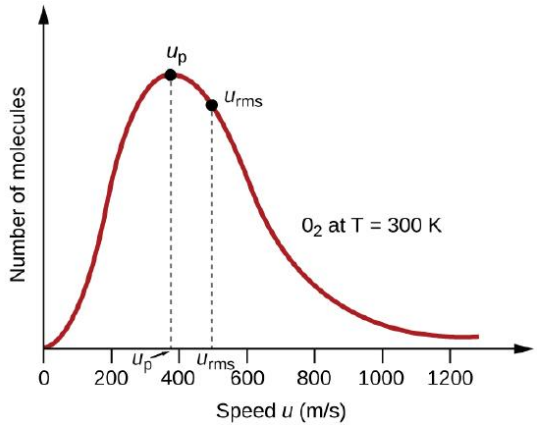


Figure 8.6.2: The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that 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:

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \dots}{n}}$$

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 8.6.1: 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 $m^2 s^{-2}$:

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

$$u_{rms} = \sqrt{\frac{3(8.314 J/molK)(303K)}{(0.028 kg/mol)}} = \sqrt{2.70 \times 10^5 m^2 s^{-2}} = 519 m/s$$

Exercise 8.6.1

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 8.6.3.

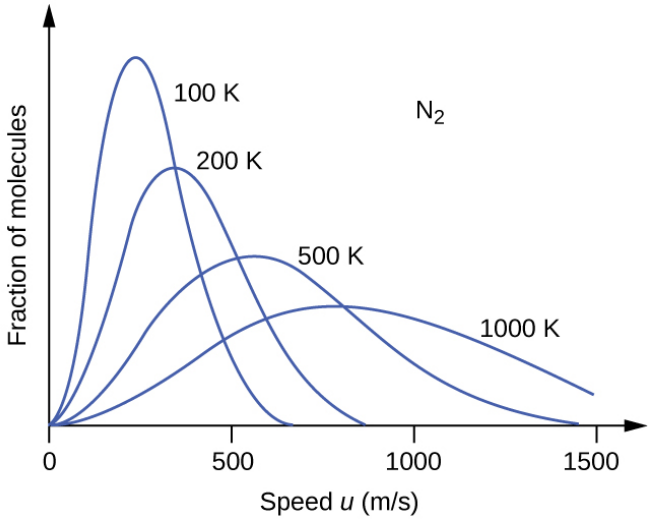


Figure 8.6.3: The molecular speed distribution for nitrogen gas (N_2) shifts to the right and flattens as the temperature increases; it shifts to the left and heightens as the temperature decreases.

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 8.6.4.

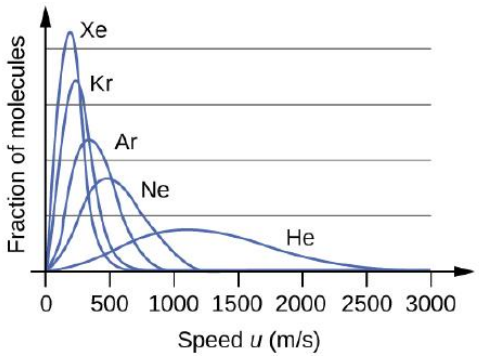


Figure 8.6.4: molecular speed is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules.

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.

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8.7: 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 8.7.1 shows plots of Z over a large pressure range for several common gases.

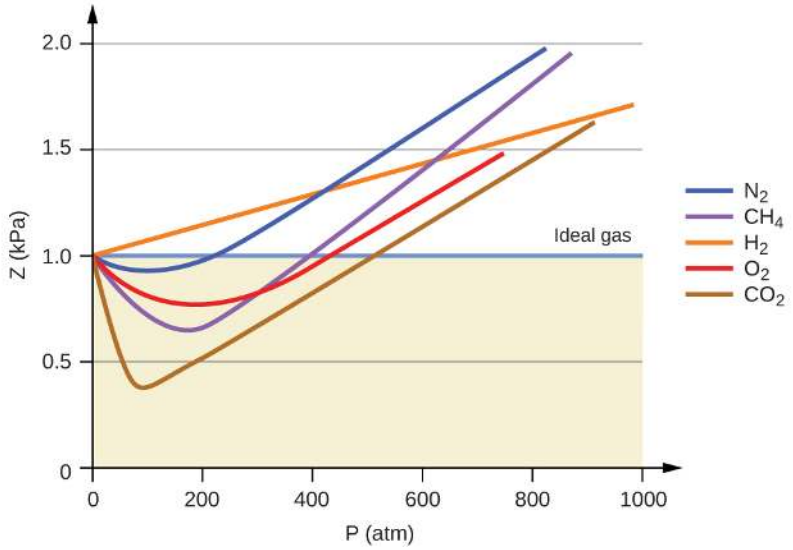


Figure 8.7.1: A graph of the compressibility factor (Z) vs. pressure shows that gases can exhibit significant deviations from the behavior predicted by the ideal gas law.

As is apparent from Figure 8.7.1, 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 8.7.2). 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.

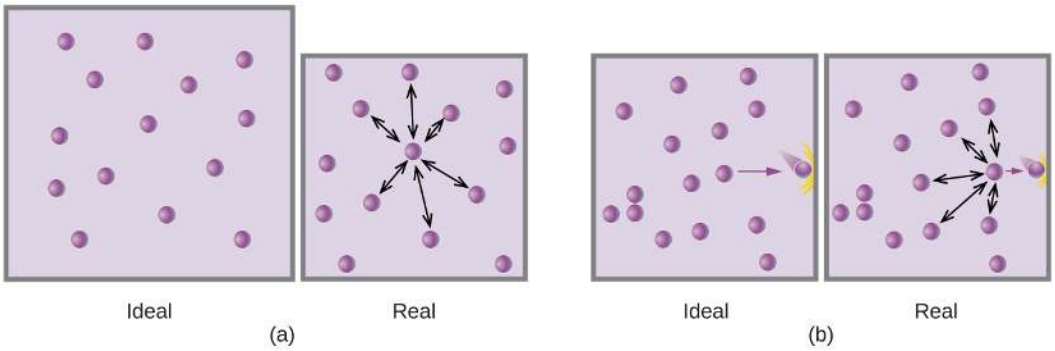


Figure 8.7.2: (a) Attractions between gas molecules serve to decrease the gas volume at constant pressure compared to an ideal gas whose molecules experience no attractive forces. (b) These attractive forces will decrease the force of collisions between the molecules and container walls, therefore reducing 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 8.7.1: 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 8.7.1. 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 8.7.1: 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:

$$P = \frac{nRT}{V} = \frac{3.46 \text{ mol} \times 0.08206 L atm mol^{-1} K^{-1} \times 502 K}{4.25 I} = 33.5 atm$$

(b) From the van der Waals equation:

$$\left(P + \frac{n^2 a}{V^2}\right) \times (V - nb) = nRT \longrightarrow P = \frac{nRT}{(V - nb)} - \frac{n^2 a}{V^2}$$

$$P = \frac{3.46 \text{ mol} \times 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 502 \text{ K}}{\left(4.25 \text{ L} - 3.46 \text{ mol} \times 0.0427 \text{ L mol}^{-1}\right)} - \frac{(3.46 \text{ mol})^2 \times 3.59 \text{ L}^2 \text{ atm mol}^2}{(4.24 \text{ L})^2}$$

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 8.7.1

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.

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8.8: 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

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8.9: Key Equations

$p = h\rho g$

$PV = nRT$

$P_{Total} = P_A + P_B + P_C + \ldots = \Sigma_i P_i$

$P_A = X_A P_{Total}$

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8.10: 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.

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CHAPTER OVERVIEW

9: 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.

9.1: Introduction
9.2: Energy Basics
9.3: Calorimetry
9.4: Enthalpy
9.5: Strengths of Ionic and Covalent Bonds
9.6: Key Terms
9.7: Key Equations
9.8: Summary

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9.1: 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 9.1.1: 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).

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9.2: 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 9.2.1). 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 9.2.1: 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 9.2.2). A battery has potential energy because the chemicals within it can produce electricity that can do work.



(a)



(b)

Figure 9.2.2: (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 9.2.3). 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.

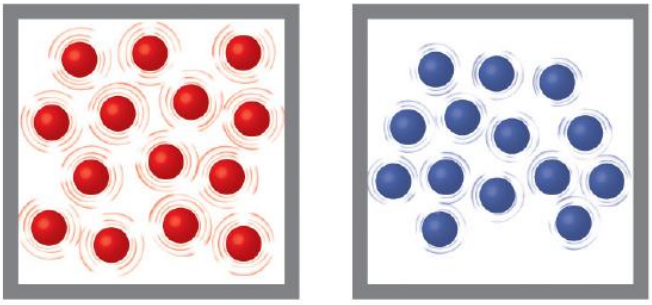


Figure 9.2.3: (a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.

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 9.2.4. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

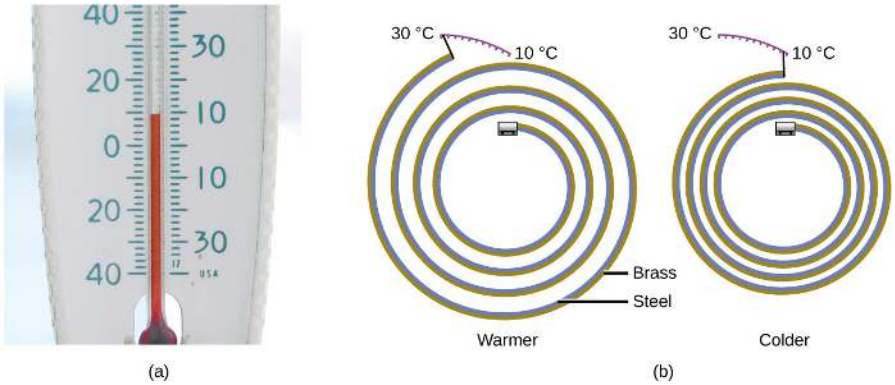


Figure 9.2.4: (a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. 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 9.2.5).

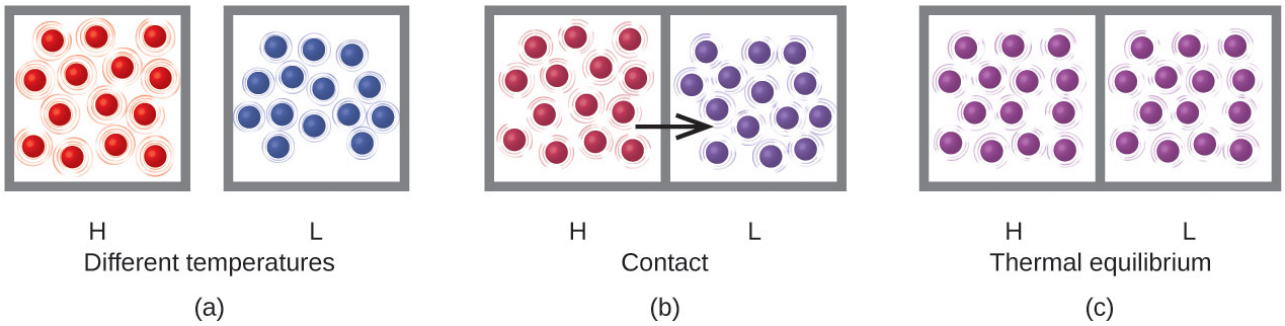


Figure 9.2.5: (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they contact each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two 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 9.2.6). 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 9.2.6: (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 9.2.7).



Figure 9.2.7: 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 9.2.1.

Table 9.2.1: 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:

$$\begin{aligned} q &= (\text{ specific heat }) \times (\text{ mass of substance }) \times (\text{ temperature change }) \\ &= c \times m \times \Delta T \\ &= c \times m \times (T_{\text{final}} - T_{\text{initial}}) \end{aligned} \tag{9.2.1}$$

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 9.2.1: 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 9.2.1:

$$\begin{aligned} q &= c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}}) \\ &= \left(4.184 J/\frac{g}{\circ}{}^{\circ}C\right) \times \left(8.0 \times 10^2 g\right) \times (85 - 21)^{\circ}C \\ &= (4.184 J/g^{\circ}C) \times \left(8.0 \times 10^2 g\right) \times (64)^{\circ}C \\ &= 210,000 J \left(= 2.1 \times 10^2 kJ\right) \end{aligned}$$

Because the temperature increased, the water absorbed heat and q is positive.

Exercise 9.2.1

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 9.2.2: 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 9.2.1:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

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 9.2.1, this value matches the specific heat of aluminum, which suggests that the unknown metal may be aluminum.

Exercise 9.2.2

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

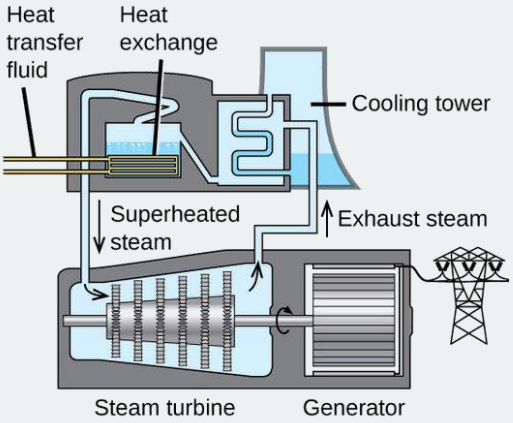
Chemistry in Everyday Life: Solar Thermal Energy Power Plants

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 9.2.8). 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 9.2.8: 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 9.2.9). 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 9.2.9: (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)

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9.3: 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 9.3.1). 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.

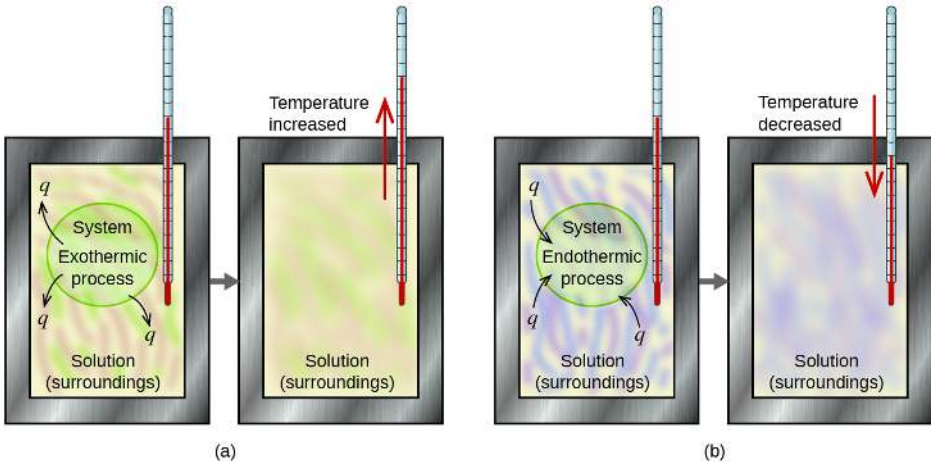


Figure 9.3.1: In a calorimetric determination, either (a) an exothermic process occurs and heat, q, is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q, is positive, indicating that thermal energy is transferred 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 9.3.2). These easy-to-use “coffee cup” calorimeters allow more heat exchange with the outside environment, and therefore produce less accurate energy values.

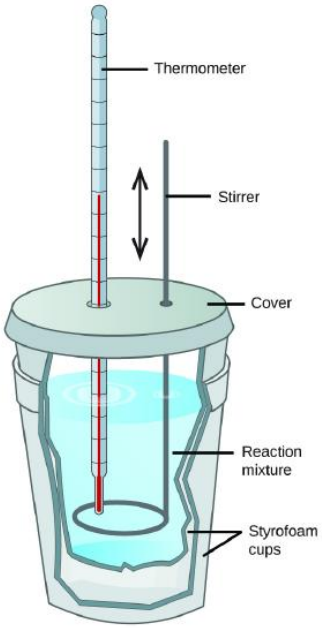


Figure 9.3.2: A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.

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 9.3.3).

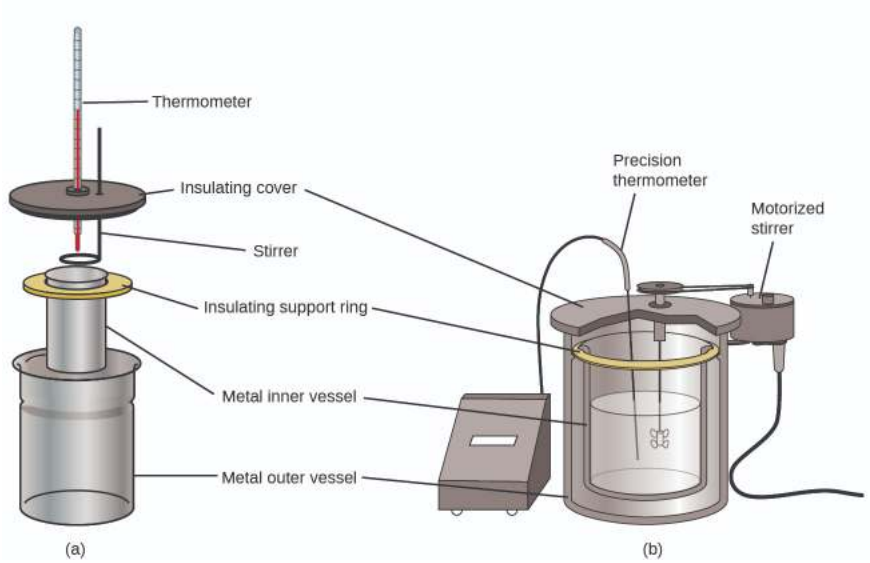


Figure 9.3.3: Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.

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 9.3.4). 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.

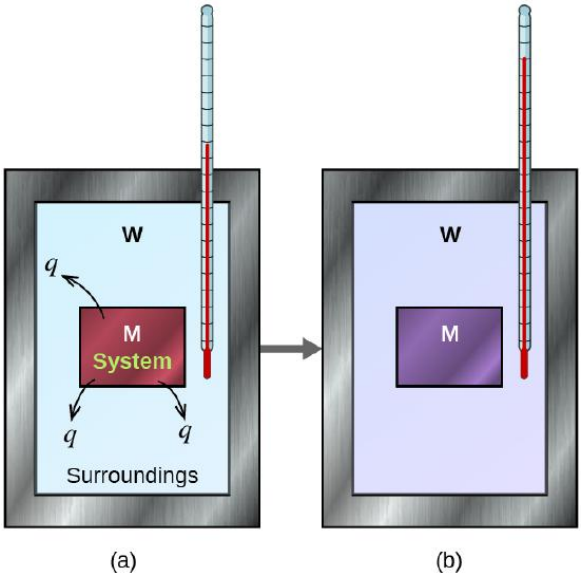


Figure 9.3.4: In a simple calorimetry process, (a) heat, q, is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.

Example 9.3.1: 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 9.3.1A

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 9.3.1B

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 9.3.2: 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 9.3.2

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 9.3.3: 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:

$$q_{solution} = \left(4.184 J/g^{\circ}C\right)\left(1.0 \times 10^2 g\right)\left(28.9^{\circ}C - 22.0^{\circ}C\right) = 2.9 \times 10^3 J$$

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 9.3.3

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 9.3.5). 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 9.3.5: 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 9.3.4: 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 9.3.6). 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 9.3.6: 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 9.3.4

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 9.3.7). 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.

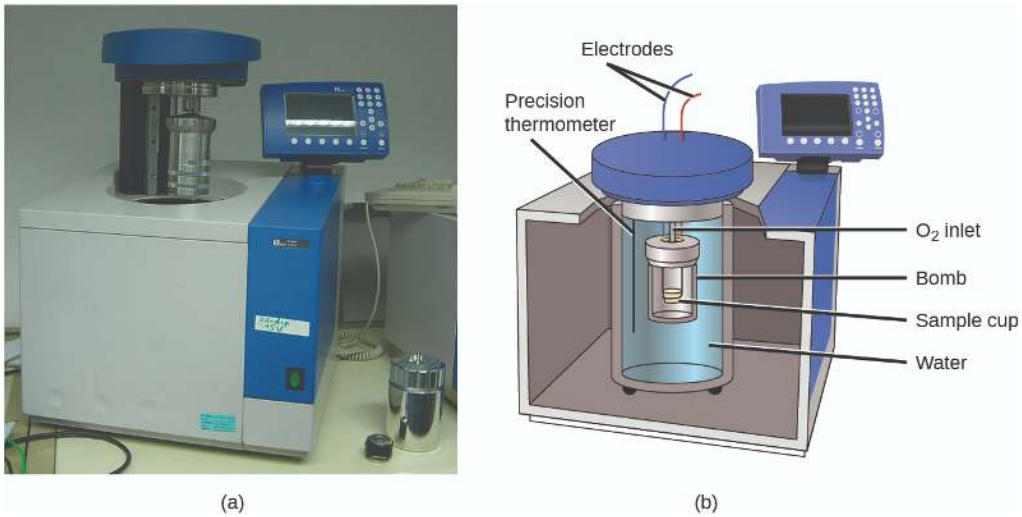


Figure 9.3.7: (a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight "bomb," which is submerged in water and surrounded by insulating materials. (credit a: modification of work 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 9.3.5: 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 9.3.5

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 9.3.8).

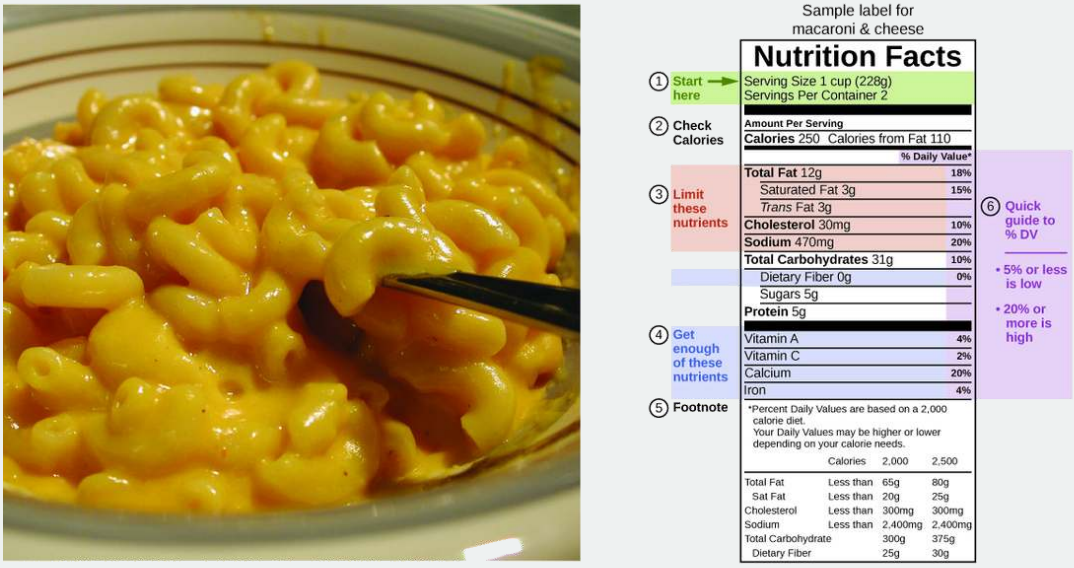


Figure 9.3.8: (a) Macaroni and cheese contain energy in the form of the macronutrients in the food. (b) The food's nutritional information is shown on the package label. In the US, the energy content is given in Calories (per serving); the rest of the world usually uses kilojoules. (credit a: modification 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.

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9.4: 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 9.4.1. 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.

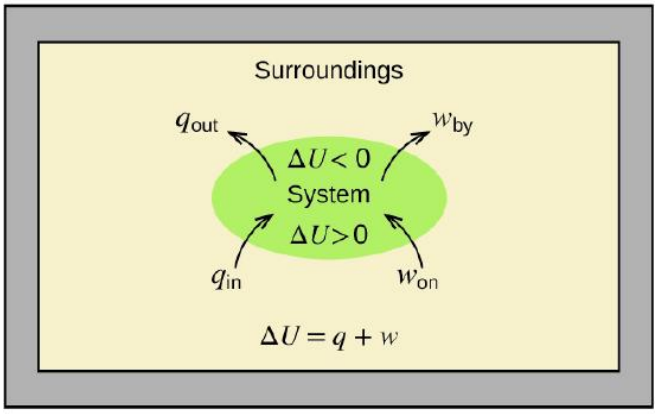


Figure 9.4.1: The internal energy, U, of a system can be changed by heat flow and work. If heat flows into the system, q_{in}, or work is done on the system, w_{on}, its internal energy increases, $\Delta U > 0$. If heat flows out of the system, q_{out}, or work is done by the system, w_{by}, its internal 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 9.4.2). The distances traveled would differ (distance is not a state function) but the elevation reached would be the same (altitude is a state function).

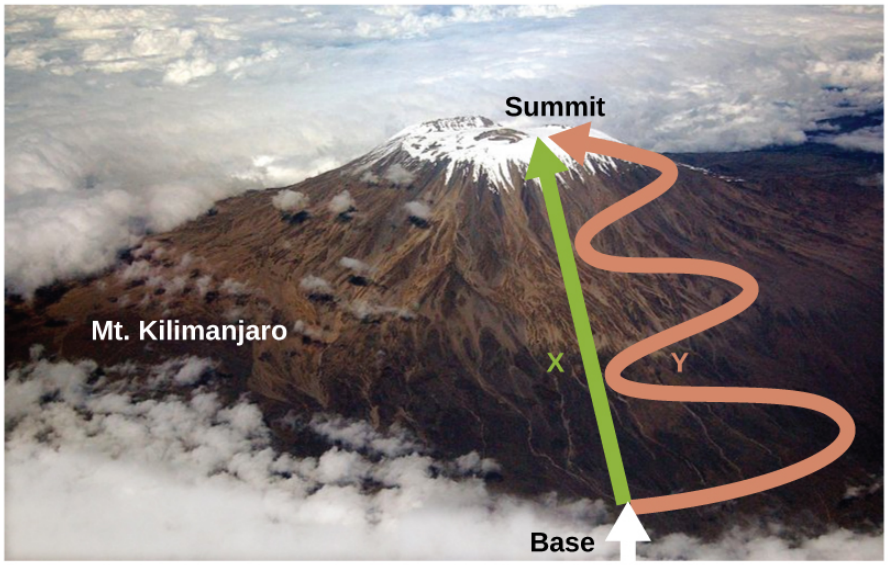


Figure 9.4.2: Paths X and Y represent two different routes to the summit of Mt. Kilimanjaro. Both have the same change in elevation (altitude or elevation on a mountain is a state function; it does not depend on path), but they have very different distances traveled (distance walked is not a state function; 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 9.4.1: 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 9.4.1

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 9.4.2: 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 9.4.1

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 9.4.1. 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 9.4.1: 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 9.4.3: Using Enthalpy of Combustion

As Figure 9.4.3 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 9.4.3: 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 9.4.1 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 9.4.3

How much heat is produced by the combustion of 125 g of acetylene?

Answer

6.25 10^3 kJ

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 9.4.4).



Figure 9.4.4: 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 9.4.4: 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 9.4.4

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 9.4.5: 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 9.4.5

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 9.4.5. 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.

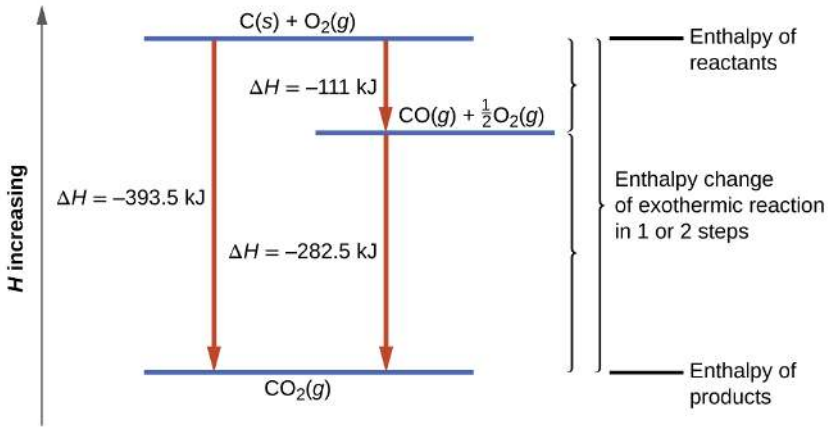


Figure 9.4.5: The formation of $\mathrm{CO_2(g)}$ from its elements can be thought of as occurring in two steps, which sum to the overall reaction, as described by Hess's law. The horizontal blue lines represent enthalpies. For an exothermic process, the products are at lower enthalpy than are the reactants.

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 9.4.6: 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 9.4.6

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 9.4.7: 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 9.4.7

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 9.4.8: 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 9.4.8

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.

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9.5: 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.

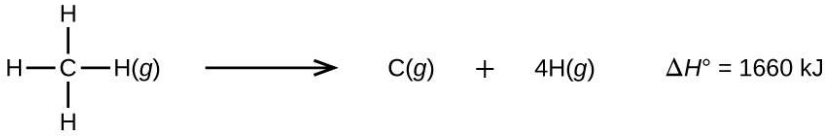
The energy required to break a specific covalent bond in one mole of gaseous molecules is called the **bond energy** or the **bond dissociation energy**. The bond energy for a diatomic molecule, $\mathrm{D_{X–Y}}$, is defined as the **standard enthalpy change** for the endothermic reaction:

$$\mathrm{XY(g) \longrightarrow X(g) + Y(g)} \qquad \mathrm{D_{\text{X}–\text{Y}}} = \Delta H^\circ$$

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 9.5.1, and a comparison of bond lengths and bond strengths for some common bonds appears in Table 9.5.2. 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 9.5.1: 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 9.5.2: 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 9.5.2) 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.

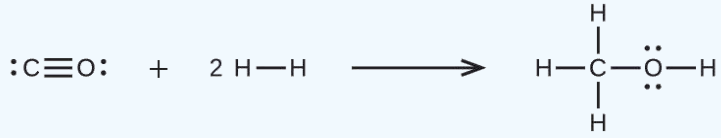
Example 9.5.1: Using Bond Energies to Calculate Approximate Enthalpy Changes

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 9.5.2, 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 9.5.2, we obtain:

$$\begin{aligned}\Delta H &= [1080 + 2(436)] - [3(415) + 350 + 464] \\ &= -107 kJ\end{aligned}$$

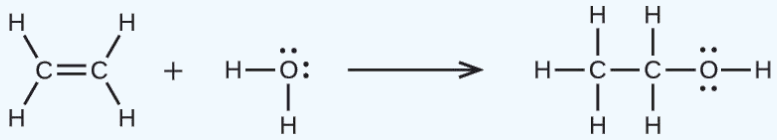
We can compare this value to the value calculated based on data from Appendix G:

$$\begin{aligned}\Delta H &= \left[\Delta H_f^\circ CH_3OH(g)\right] - \left[\Delta H_f^\circ CO(g) + 2 \times \Delta H_f^\circ H_2\right] \\ &= [-201.0] - [-110.52 + 2 \times 0] \\ &= -90.5 kJ\end{aligned}$$

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 9.5.1

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 9.5.2, calculate an approximate enthalpy change, Δ*H*, 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 9.5.2: 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 9.5.1

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 9.5.1: diagrams the Born-Haber cycle for the formation of solid cesium fluoride.

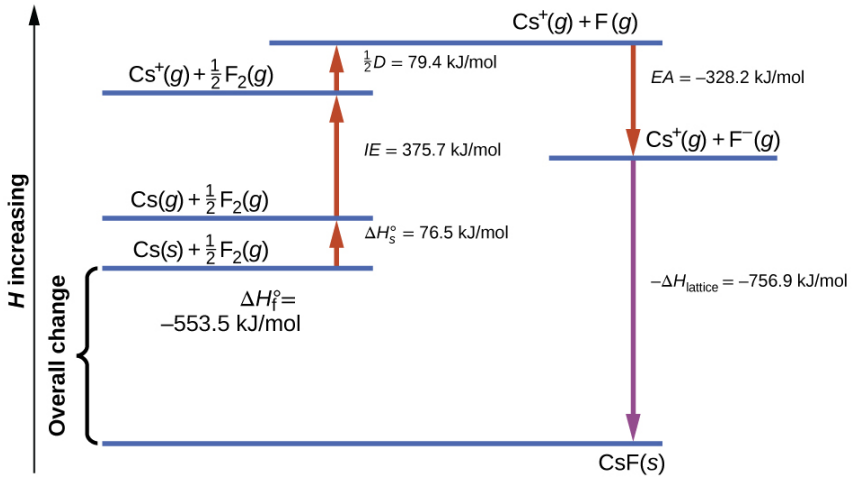


Figure 9.5.1: The Born-Haber cycle shows the relative energies of each step involved in the formation of an ionic solid from the necessary elements in their reference states.

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 $9.5.3$ shows this for fluoride, CsF.

Table 9.5.3

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.

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9.6: 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

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9.7: 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: $\Delta H_{\text{lattice}} = \frac{\text{C}(\text{Z}^{+})(\text{Z}^{-})}{}$	

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9.8: 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.

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CHAPTER OVERVIEW

10: Liquids and Solids

10.1: Introduction
10.2: Intermolecular Forces
10.3: Properties of Liquids
10.4: Phase Transitions
10.5: Phase Diagrams
10.6: The Solid State of Matter
10.7: Lattice Structures in Crystalline Solids
10.8: Key Terms
10.9: Key Equations
10.10: Summary

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10.1: 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 10.1.1: 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.

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10.2: 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 10.2.1 illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

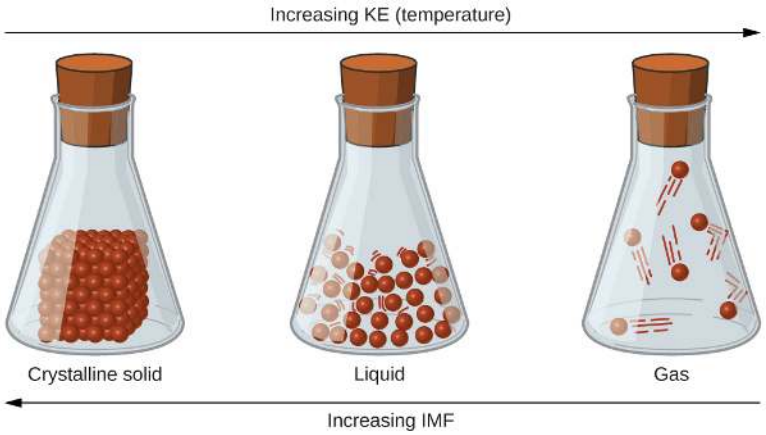


Figure 10.2.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

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 10.2.2



Figure 10.2.2: 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 10.2.3



Figure 10.2.3: 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 10.2.4 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.

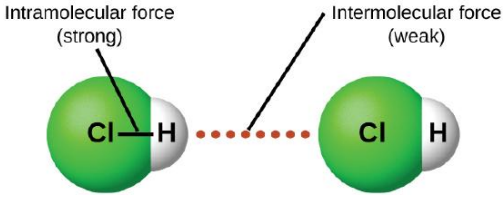


Figure 10.2.4: *Intra*molecular forces keep a molecule intact. *Inter*molecular forces hold multiple molecules together and determine many of a substance's properties.

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 10.2.5.

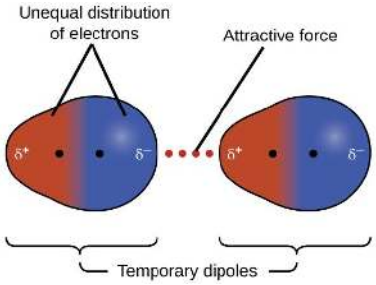


Figure 10.2.5: Dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules.

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 10.2.1.

Table 10.2.1: 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.

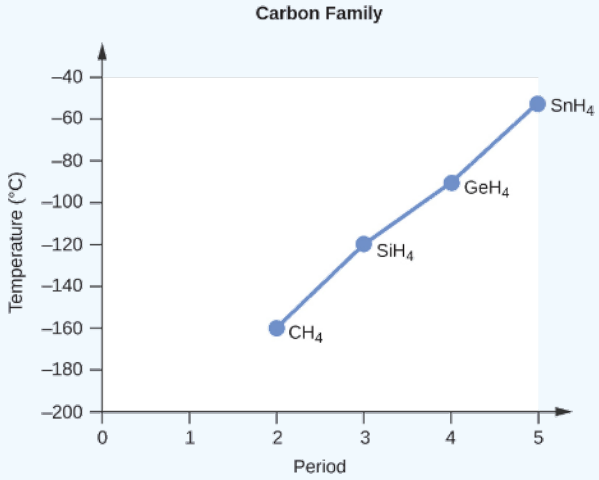
Example 10.2.1: London Forces and Their Effects

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 10.2.1

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 10.2.6) 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.

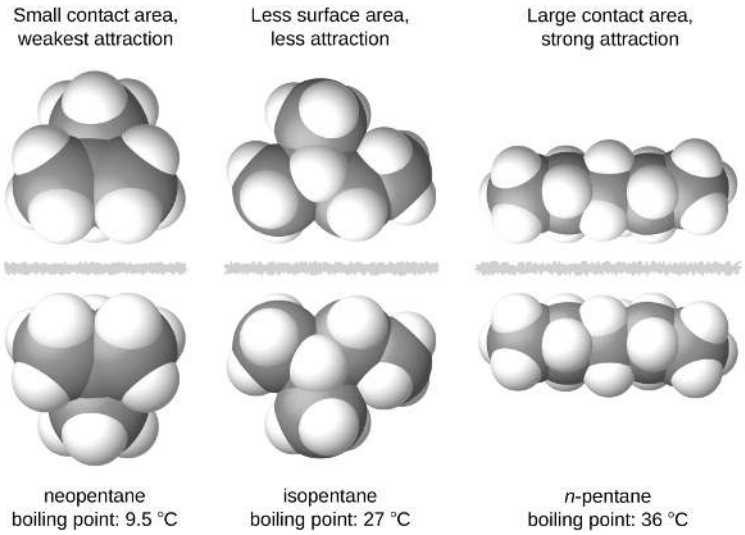


Figure 10.2.6: The strength of the dispersion forces increases with the contact area between molecules, as demonstrated by the boiling points of these pentane isomers.

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 10.2.7, 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.

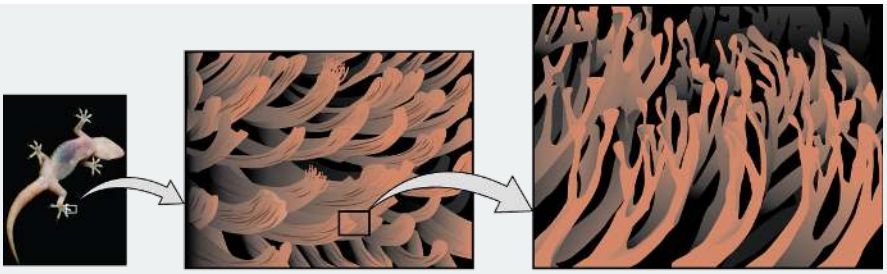


Figure 10.2.7: Geckos' toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko's millions of spatulae. By changing how the spatulae contact the surface, geckos can 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 10.2.8.

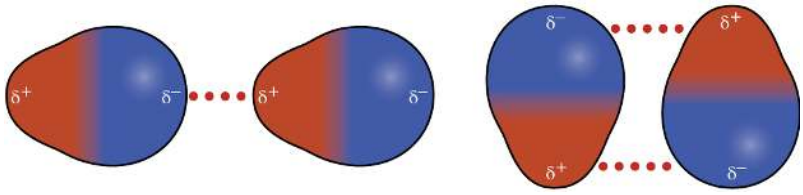


Figure 10.2.8: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

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 10.2.2: 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 10.2.2

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 10.2.9 illustrates hydrogen bonding between water molecules.

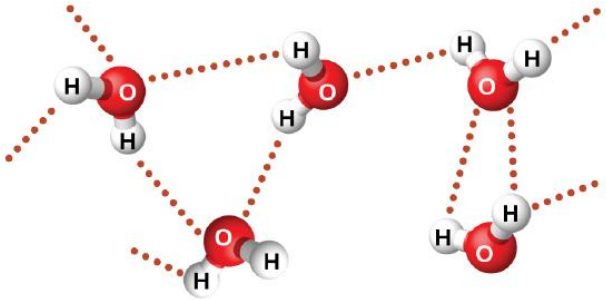


Figure 10.2.9: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

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 10.2.10 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.

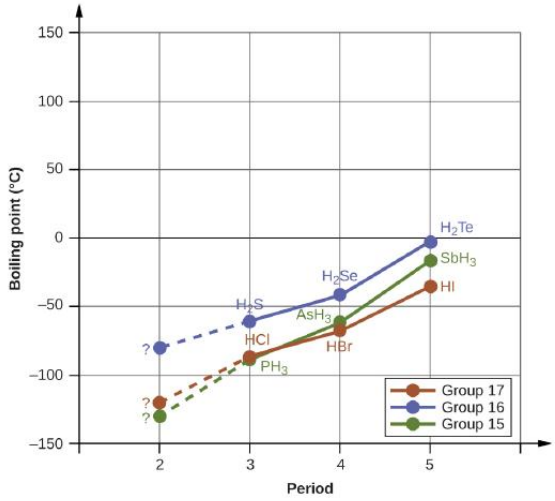


Figure 10.2.10: For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5.

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 10.2.11. The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

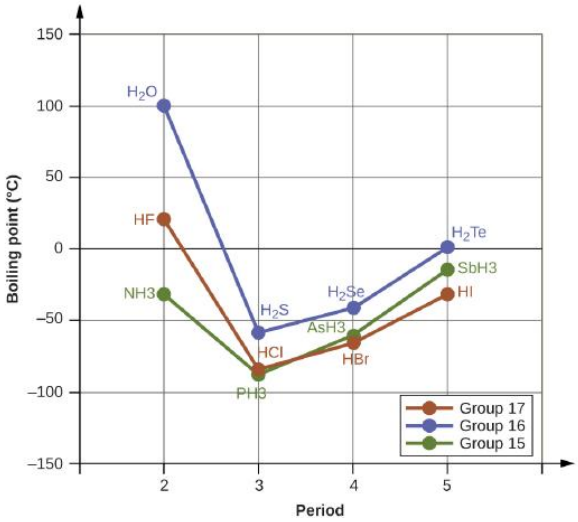


Figure 10.2.11: In comparison to periods 3−5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

Example 10.2.3: 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 10.2.3

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.

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 10.2.12

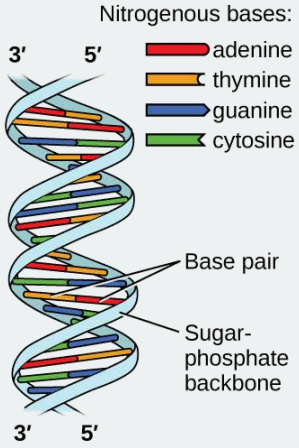


Figure 10.2.12: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

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 10.2.13

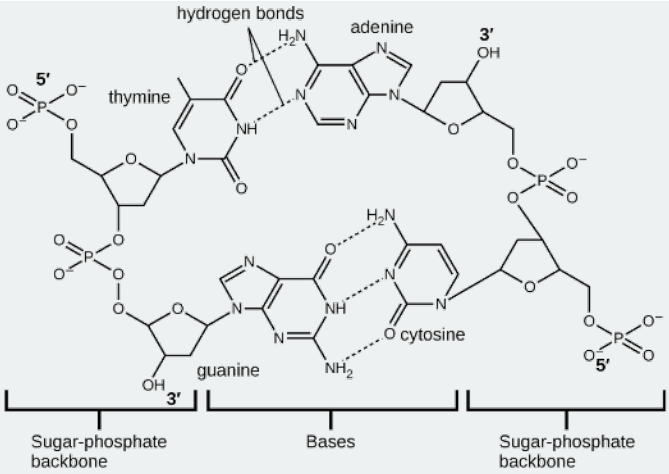


Figure 10.2.13: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called “complementary base pairs.”

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.

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10.3: 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 10.3.1, 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 10.3.1: (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 10.3.1 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 10.3.1: 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 10.3.2, 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.

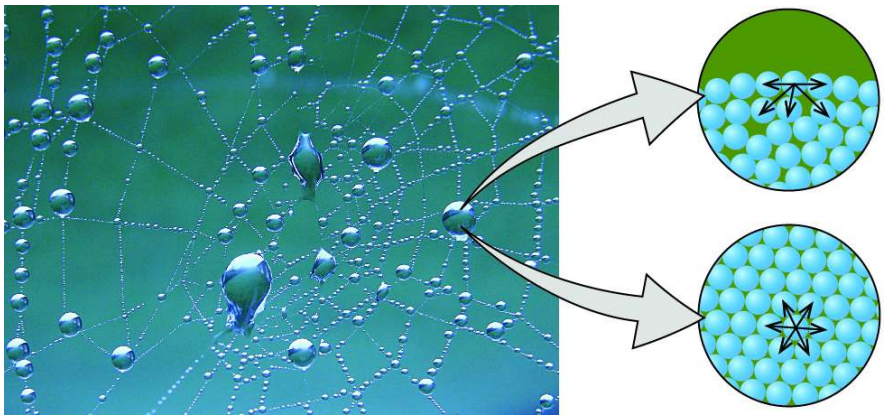


Figure 10.3.2: Attractive forces result in a spherical water drop that minimizes surface area; cohesive forces hold the sphere together; adhesive forces keep the drop attached to the web. (credit photo: modification of work by “OliBac”/Flickr)

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 10.3.2. 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 10.3.3, even though they are denser than water, move on its surface because they are supported by the surface tension.

Table 10.3.2: 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.

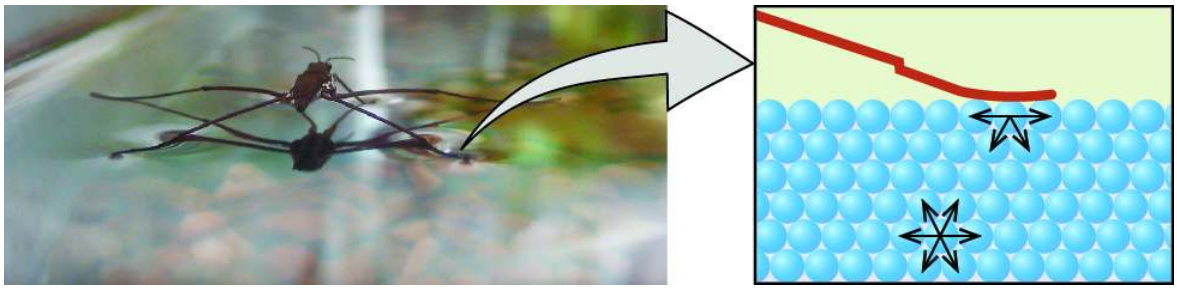


Figure 10.3.3: Surface tension (right) prevents this insect, a “water strider,” from sinking into the water.

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 10.3.4).

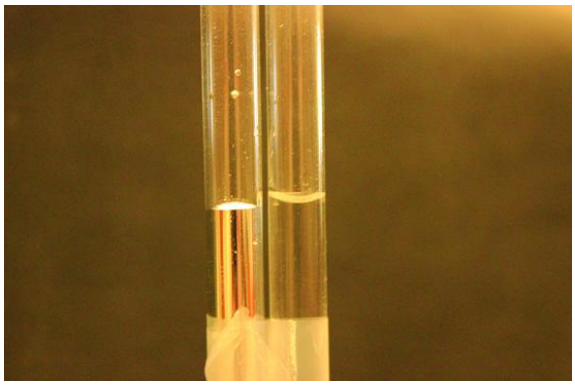


Figure 10.3.4: Differences in the relative strengths of cohesive and adhesive forces result in different meniscus shapes for mercury (left) and water (right) in glass tubes. (credit: Mark Ott)

If you place one end of a paper towel in spilled wine, as shown in Figure 10.3.5, 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.

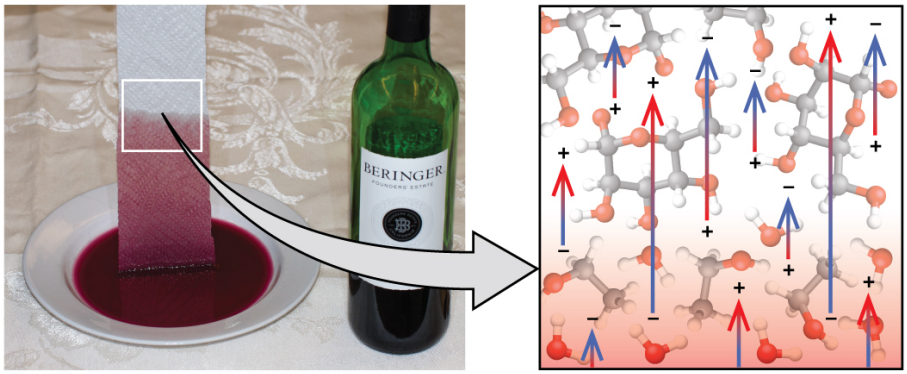


Figure 10.3.5: Wine wicks up a paper towel (left) because of the strong attractions of water (and ethanol) molecules to the −OH groups on the towel's cellulose fibers and the strong attractions of water molecules to other water (and ethanol) molecules (right). (credit photo: modification of work by Mark 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 10.3.6. 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.

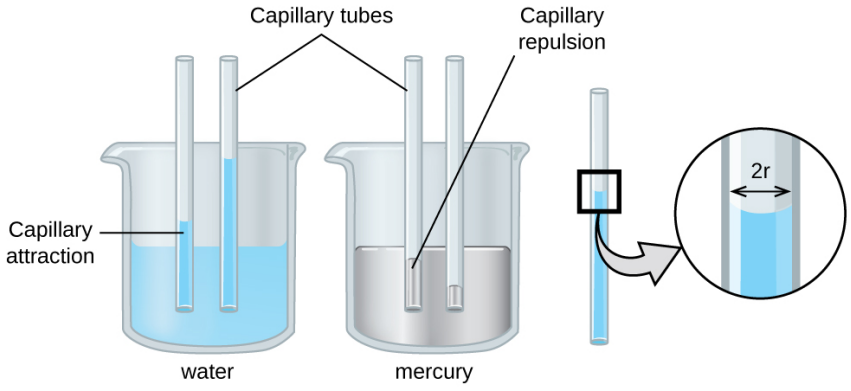


Figure 10.3.6: Depending upon the relative strengths of adhesive and cohesive forces, a liquid may rise (such as water) or fall (such as mercury) in a glass capillary tube. The extent of the rise (or fall) is directly proportional to the surface tension of the liquid and inversely proportional to the 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:

$$h = \frac{2T\cos\theta}{r\rho g} \tag{10.3.1}$$

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 10.3.1: 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 10.3.1:

$$h = \frac{2T\cos\theta}{r\rho g}$$

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 10.3.1

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 10.3.7. 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 10.3.7: 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)

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10.4: 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 10.4.1, and using a manometer to measure the increase in pressure that is due to the vapor in equilibrium with the condensed phase.

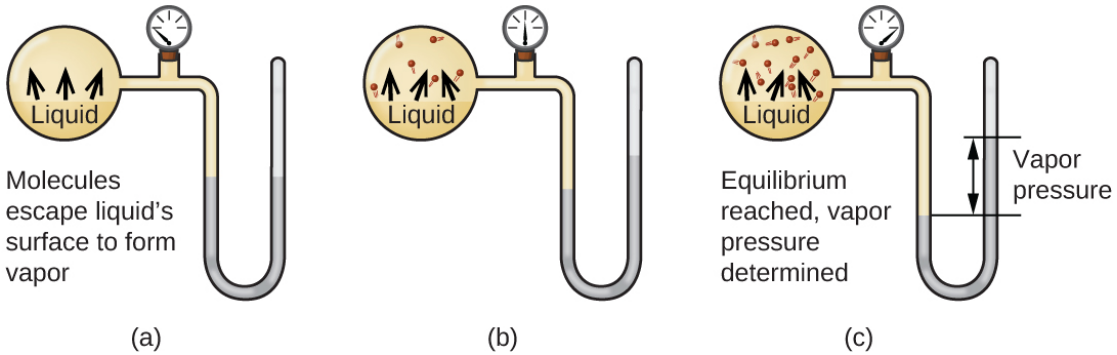
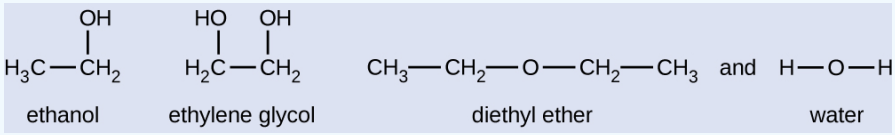


Figure 10.4.1: In a closed container, dynamic equilibrium is reached when (a) the rate of molecules escaping from the liquid to become the gas (b) increases and eventually (c) equals the rate of gas molecules entering the liquid. When this equilibrium is reached, the vapor pressure of the gas is constant, 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 10.4.1: 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 10.4.1

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 10.4.2 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.

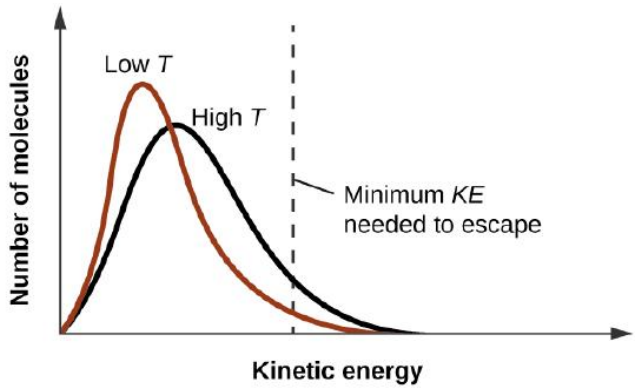


Figure 10.4.2: Temperature affects the distribution of kinetic energies for the molecules in a liquid. At the higher temperature, more molecules have the necessary kinetic energy, KE, to escape from the liquid into the gas phase.

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 10.4.3 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.

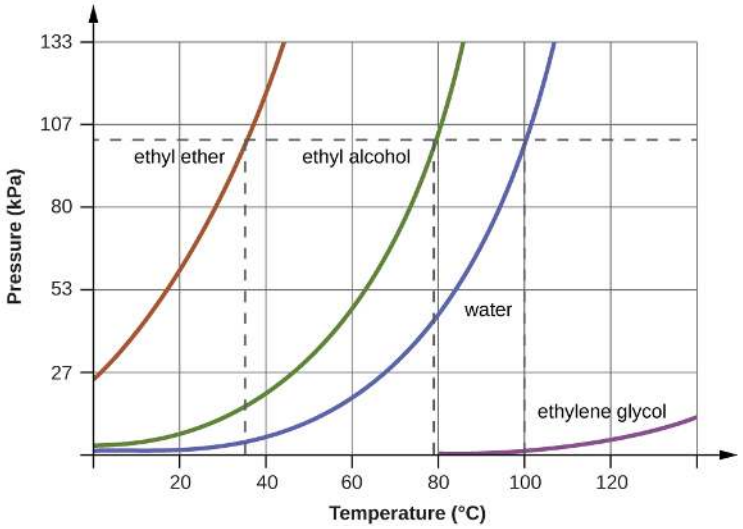


Figure 10.4.3: The boiling points of liquids are the temperatures at which their equilibrium vapor pressures equal the pressure of the surrounding atmosphere. Normal boiling points are those corresponding to a pressure of 1 atm (101.3 kPa.)

Example 10.4.2: 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 10.4.3 to determine the boiling point of water at this elevation.

Solution

The graph of the vapor pressure of water versus temperature in Figure 10.4.3 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 10.4.2

The boiling point of ethyl ether was measured to be 10 °C at a base camp on the slopes of Mount Everest. Use Figure 10.4.3to 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 10.4.3: 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 10.4.3

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 10.4.4: 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 10.4.4

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 10.4.5: Using Enthalpy of Vaporization

One way our body is cooled is by evaporation of the water in sweat (Figure 10.4.4). 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 10.4.4: 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:

$$1.5 \text{ L} \times \frac{1000 \text{ g}}{1 \text{ L}} \times \frac{1 \text{ mol}}{18 \text{ g}} \times \frac{43.46 \text{ kJ}}{1 \text{ mol}} = 3.6 \times 10^3 \text{ kJ}$$

Thus, 3600 kJ of heat are removed by the evaporation of 1.5 L of water.

Exercise 10.4.1

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 10.4.5).



Figure 10.4.5: (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 10.4.6). 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 10.4.6: 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:

$$\mathrm{CO_2(s) \longrightarrow CO_2(g)} \qquad \Delta H_{\mathrm{sub}} = 26.1 kJ/mol$$

Likewise, the enthalpy change for the reverse process of deposition is equal in magnitude but opposite in sign to that for sublimation:

$$\mathrm{CO_2(g) \longrightarrow CO_2(s)} \qquad \Delta H_{\mathrm{dep}} = -\Delta H_{\mathrm{sub}} = -26.1 kJ/mol$$

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 10.4.7. For example:

$$\begin{array}{ll} \mathrm{solid \longrightarrow\ liquid} & \Delta H_{\mathrm{fus}} \\ \mathrm{liquid \longrightarrow\ gas} & \Delta H_{\mathrm{vap}} \\ \hline \mathrm{solid \longrightarrow\ gas} & \Delta H_{\mathrm{sub}} = \Delta H_{\mathrm{fus}} + \Delta H_{\mathrm{vap}} \end{array}$$

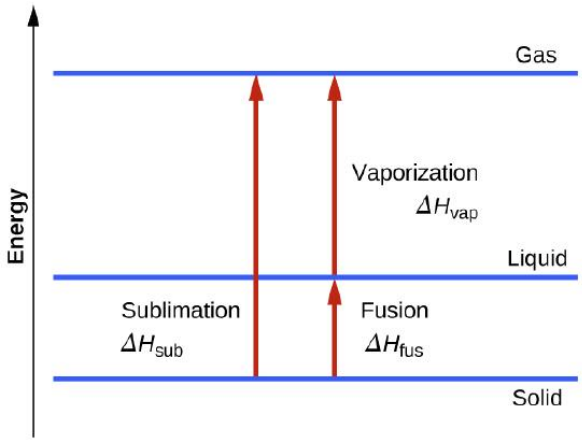


Figure 10.4.7: For a given substance, the sum of its enthalpy of fusion and enthalpy of vaporization is approximately equal to its enthalpy of sublimation.

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 10.4.8 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.

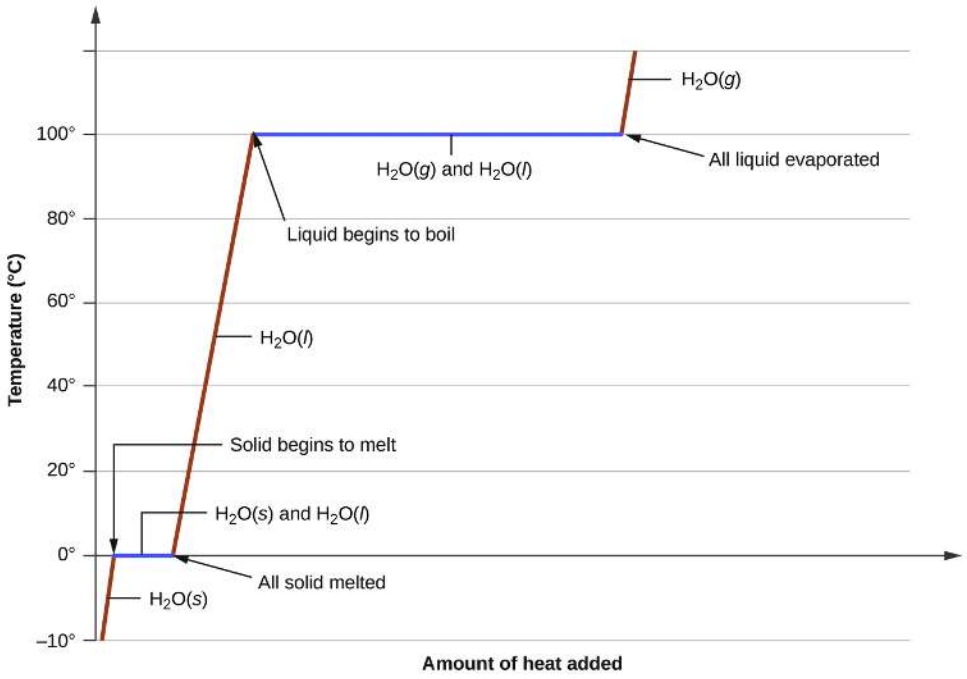


Figure 10.4.8: A typical heating curve for a substance depicts changes in temperature that result as the substance absorbs increasing amounts of heat. Plateaus in the curve (regions of constant temperature) are exhibited when the substance undergoes phase transitions.

Example 10.4.6: 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 10.4.6

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

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10.5: 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 10.5.1.

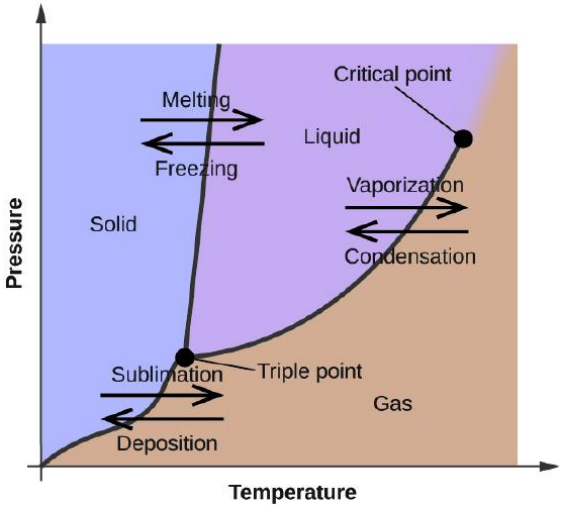


Figure 10.5.1: The physical state of a substance and its phase-transition temperatures are represented graphically in a phase diagram.

To illustrate the utility of these plots, consider the phase diagram for water shown in Figure 10.5.2.

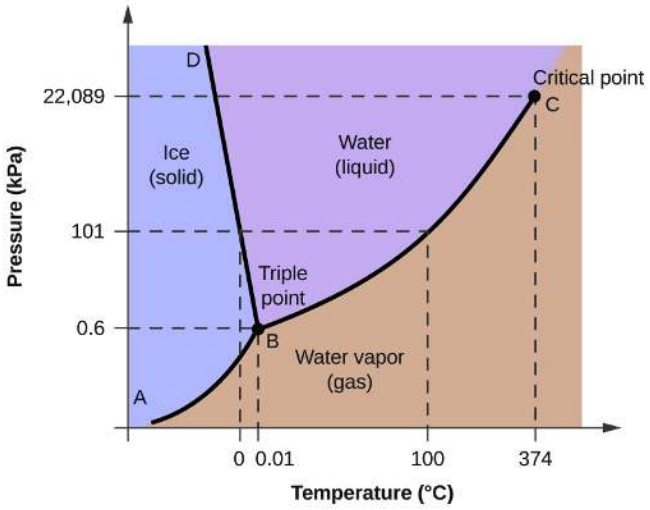


Figure 10.5.2: The pressure and temperature axes on this phase diagram of water are not drawn to constant scale in order to illustrate several important properties.

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 10.5.2is 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 10.5.2, 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 10.5.2, 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 10.5.3



Figure 10.5.3: 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 10.5.4 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.

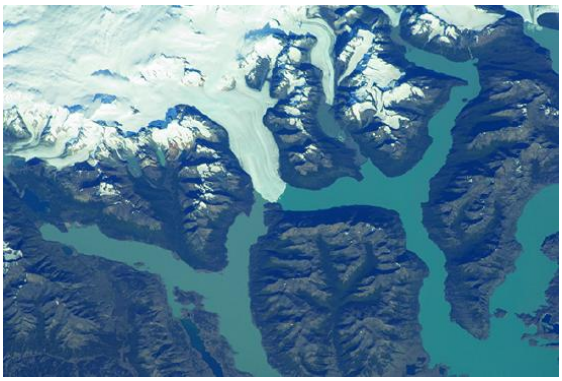


Figure 10.5.4: The immense pressures beneath glaciers result in partial melting to produce a layer of water that provides lubrication to assist glacial movement. This satellite photograph shows the advancing edge of the Perito Moreno glacier in Argentina. (credit: NASA)

The point of intersection of all three curves is labeled B in Figure 10.5.2 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 10.5.1: Determining the State of Water

Using the phase diagram for water given in Figure 10.5.2, 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 10.5.1

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 10.5.5 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.

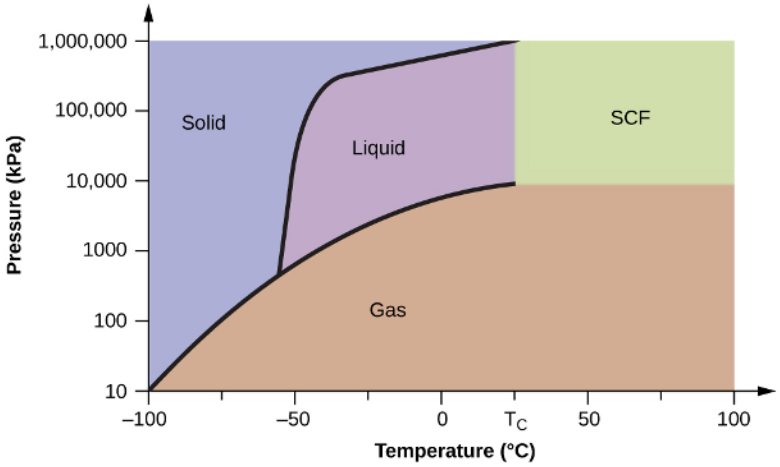


Figure 10.5.5: A phase diagram for carbon dioxide is shown. The pressure axis is plotted on a logarithmic scale to accommodate the large range of values.

Example 10.5.2: Determining the State of Carbon Dioxide

Using the phase diagram for carbon dioxide shown in Figure 10.5.5, 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 10.5.2

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 10.5.2), 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 10.5.6). 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

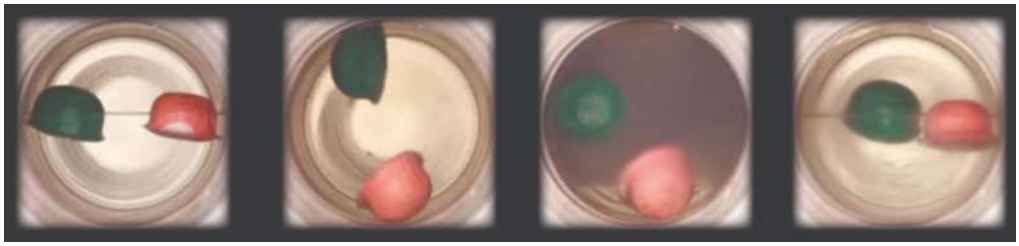


Figure 10.5.6: (a) A sealed container of liquid carbon dioxide slightly below its critical point is heated, resulting in (b) the formation of the supercritical fluid phase. Cooling the supercritical fluid lowers its temperature and pressure below the critical point, resulting in the reestablishment of 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 10.5.3: 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 10.5.3

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 10.5.7). 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.

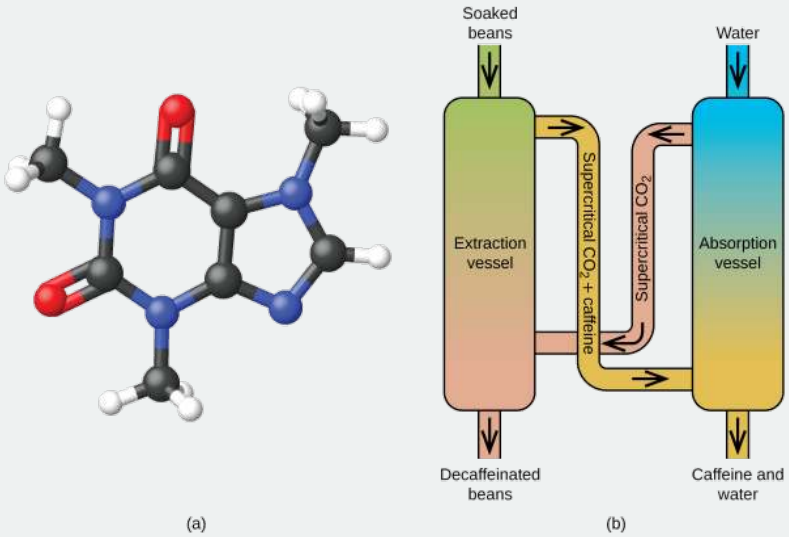


Figure 10.5.7: (a) Caffeine molecules have both polar and nonpolar regions, making it soluble in solvents of varying polarities. (b) The schematic shows a typical decaffeination process involving supercritical carbon dioxide.

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10.6: 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 10.6.1).

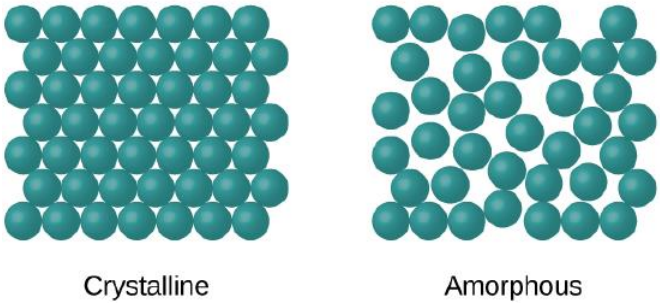


Figure 10.6.1: The entities of a solid phase may be arranged in a regular, repeating pattern (crystalline solids) or randomly (amorphous).

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 10.6.2), 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.

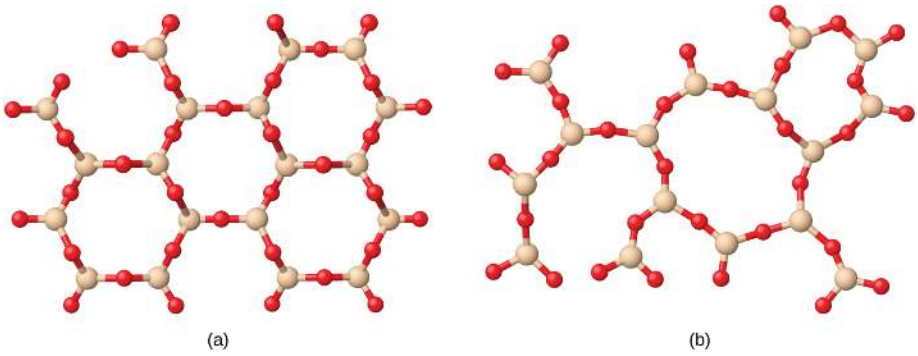


Figure 10.6.2: (a) Silicon dioxide, SiO_2, is abundant in nature as one of several crystalline forms of the mineral quartz. (b) Rapid cooling of molten SiO_2 yields an amorphous solid known as “fused silica”.

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 10.6.3). 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.

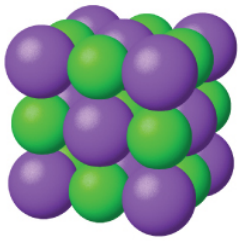


Figure 10.6.3: Sodium chloride is an ionic solid.

Metallic Solids

Metallic solids such as crystals of copper, aluminum, and iron are formed by metal atoms Figure 10.6.4 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.

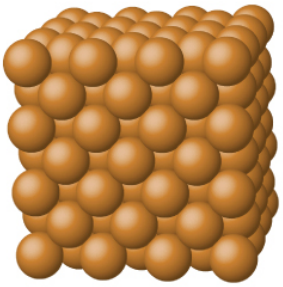


Figure 10.6.4: Copper is a metallic solid.

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 10.6.5 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.

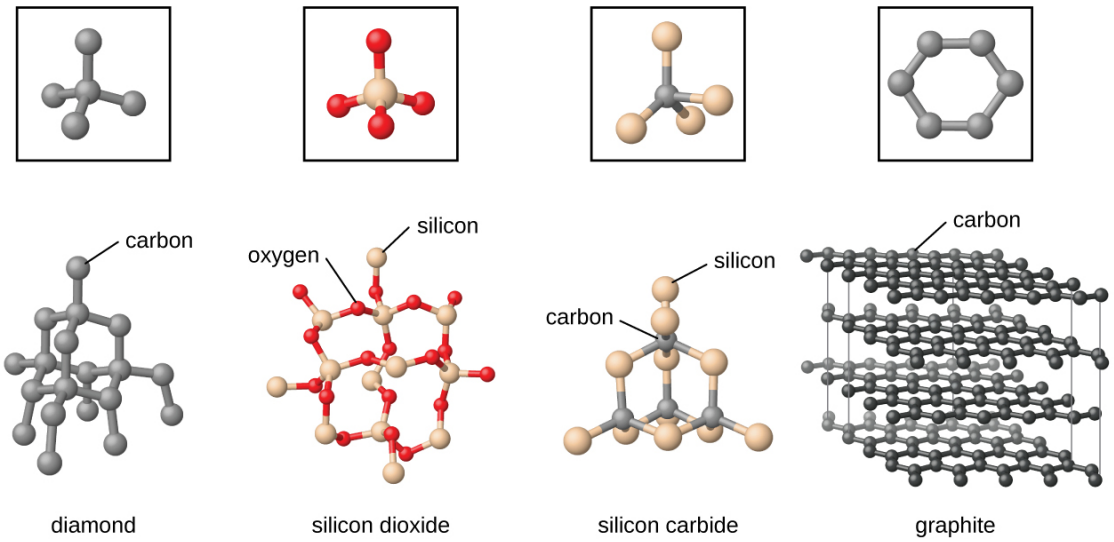


Figure 10.6.5: A covalent crystal contains a three-dimensional network of covalent bonds, as illustrated by the structures of diamond, silicon dioxide, silicon carbide, and graphite. Graphite is an exceptional example, composed of planar sheets of covalent crystals that are held together in layers by 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 10.6.6, 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).

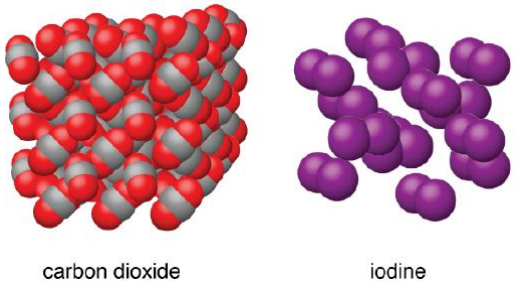


Figure 10.6.6: Carbon dioxide (CO_2) consists of small, nonpolar molecules and forms a molecular solid with a melting point of −78 °C. Iodine (I_2) consists of larger, nonpolar molecules and forms a molecular solid that melts at 114 °C.

Properties of Solids

A crystalline solid, like those listed in Table 10.6.1, 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 10.6.1: 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 10.6.7. 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.

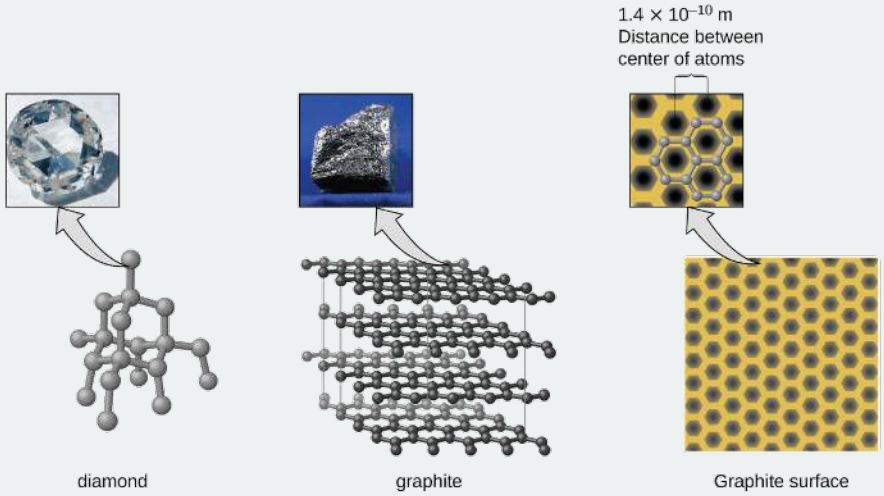


Figure 10.6.7: Diamond is extremely hard because of the strong bonding between carbon atoms in all directions. Graphite (in pencil lead) rubs off onto paper due to the weak attractions between the carbon layers. An image of a graphite surface shows the distance between the centers of adjacent carbon 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 10.6.8 , 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.

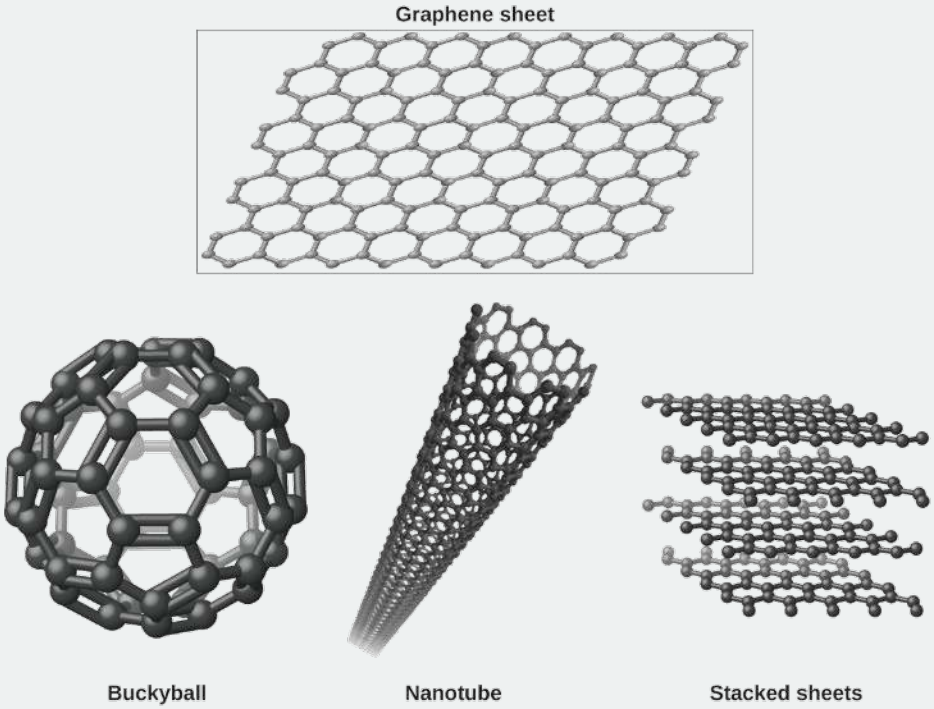


Figure 10.6.8: Graphene sheets can be formed into buckyballs, nanotubes, and stacked layers.

Crystal Defects

In a crystalline solid, the atoms, ions, or molecules are arranged in a definite repeating pattern, but occasional defects may occur in the pattern. Several types of defects are known, as illustrated in Figure 10.6.9. **Vacancies** are defects that occur when positions that should contain atoms or ions are vacant. Less commonly, some atoms or ions in a crystal may occupy positions, called **interstitial sites**, located between the regular positions for atoms. Other distortions are found in impure crystals, as, for example, when the cations, anions, or molecules of the impurity are too large to fit into the regular positions without distorting the structure. Trace amounts of impurities are sometimes added to a crystal (a process known as *doping)* in order to create defects in the structure that yield desirable changes in its properties. For example, silicon crystals are doped with varying amounts of different elements to yield suitable electrical properties for their use in the manufacture of semiconductors and computer chips.

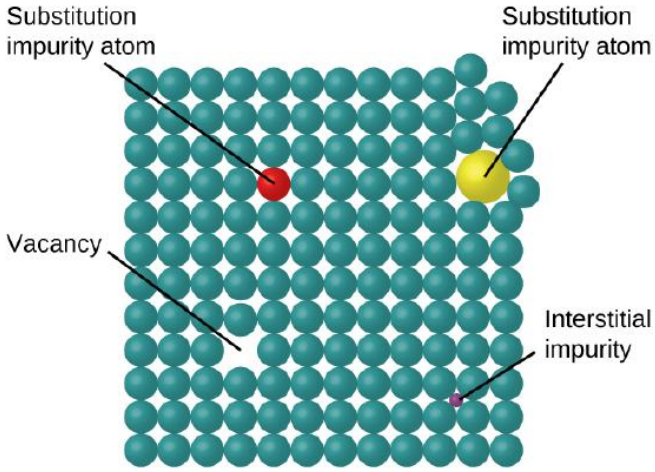


Figure 10.6.9: Types of crystal defects include vacancies, interstitial atoms, and substitutions impurities.

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10.7: Lattice Structures in Crystalline Solids

Learning Objectives

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

- Describe the arrangement of atoms and ions in crystalline structures
- Compute ionic radii using unit cell dimensions
- Explain the use of X-ray diffraction measurements in determining crystalline structures

Over 90% of naturally occurring and man-made solids are crystalline. Most solids form with a regular arrangement of their particles because the overall attractive interactions between particles are maximized, and the total intermolecular energy is minimized, when the particles pack in the most efficient manner. The regular arrangement at an atomic level is often reflected at a macroscopic level. In this module, we will explore some of the details about the structures of metallic and ionic crystalline solids, and learn how these structures are determined experimentally.

The Structures of Metals

We will begin our discussion of crystalline solids by considering elemental metals, which are relatively simple because each contains only one type of atom. A pure metal is a crystalline solid with metal atoms packed closely together in a repeating pattern. Some of the properties of metals in general, such as their malleability and ductility, are largely due to having identical atoms arranged in a regular pattern. The different properties of one metal compared to another partially depend on the sizes of their atoms and the specifics of their spatial arrangements. We will explore the similarities and differences of four of the most common metal crystal geometries in the sections that follow.

Unit Cells of Metals

The structure of a crystalline solid, whether a metal or not, is best described by considering its simplest repeating unit, which is referred to as its **unit cell**. The unit cell consists of lattice points that represent the locations of atoms or ions. The entire structure then consists of this unit cell repeating in three dimensions, as illustrated in Figure 10.7.1.

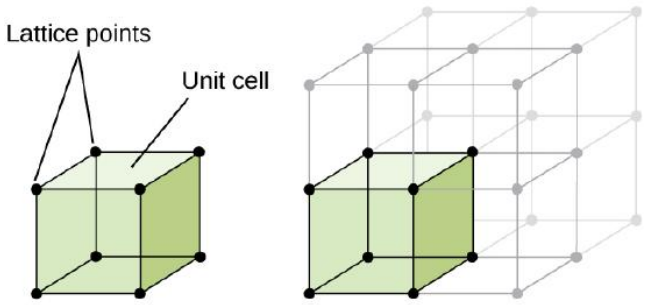


Figure 10.7.1: A unit cell shows the locations of lattice points repeating in all directions.

Let us begin our investigation of crystal lattice structure and unit cells with the most straightforward structure and the most basic unit cell. To visualize this, imagine taking a large number of identical spheres, such as tennis balls, and arranging them uniformly in a container. The simplest way to do this would be to make layers in which the spheres in one layer are directly above those in the layer below, as illustrated in Figure 10.7.2. This arrangement is called **simple cubic structure**, and the unit cell is called the **simple cubic unit cell** or primitive cubic unit cell.

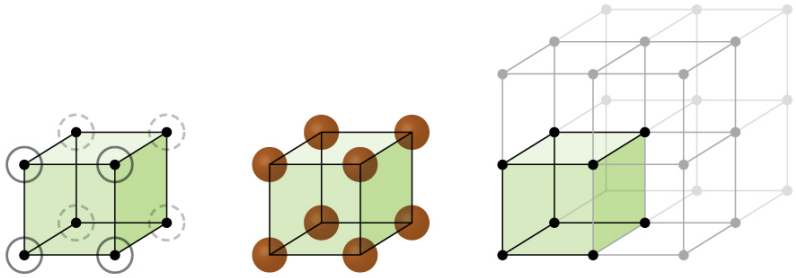


Figure 10.7.2: When metal atoms are arranged with spheres in one layer directly above or below spheres in another layer, the lattice structure is called simple cubic. Note that the spheres are in contact.

In a simple cubic structure, the spheres are not packed as closely as they could be, and they only "fill" about 52% of the volume of the container. This is a relatively inefficient arrangement, and only one metal (polonium, Po) crystallizes in a simple cubic structure. As shown in Figure 10.7.3, a solid with this type of arrangement consists of planes (or layers) in which each atom contacts only the four nearest neighbors in its layer; one atom directly above it in the layer above; and one atom directly below it in the layer below. The number of other particles that each particle in a crystalline solid contacts is known as its **coordination number**. For a polonium atom in a simple cubic array, the coordination number is, therefore, six.

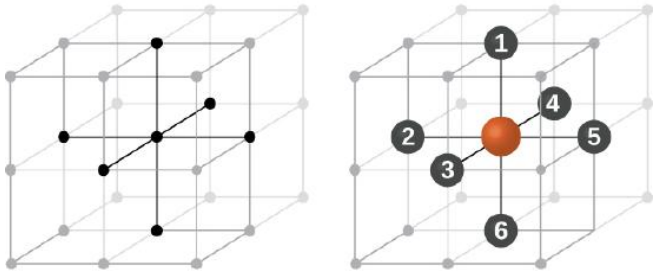


Figure 10.7.3: An atom in a simple cubic lattice structure contacts six other atoms, so it has a coordination number of six.

In a simple cubic lattice, the unit cell that repeats in all directions is a cube defined by the centers of eight atoms, as shown in Figure 10.7.4. Atoms at adjacent corners of this unit cell contact each other, so the edge length of this cell is equal to two atomic radii, or one atomic diameter. A cubic unit cell contains only the parts of these atoms that are within it. Since an atom at a corner of a simple cubic unit cell is contained by a total of eight unit cells, only one-eighth of that atom is within a specific unit cell. And since each simple cubic unit cell has one atom at each of its eight "corners," there is $8 \times \frac{1}{8} = 1$ atom within one simple cubic unit cell.

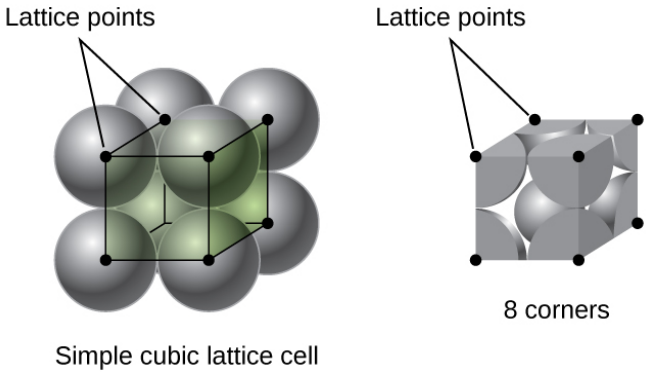


Figure 10.7.4: A simple cubic lattice unit cell contains one-eighth of an atom at each of its eight corners, so it contains one atom total.

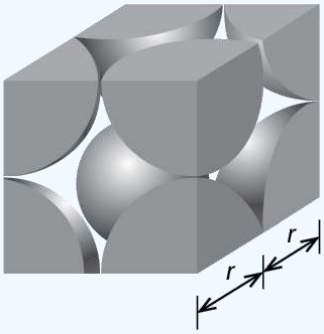
Example 10.7.1: Calculation of Atomic Radius and Density for Metals, Part 1

The edge length of the unit cell of alpha polonium is 336 pm.

a. Determine the radius of a polonium atom.
b. Determine the density of alpha polonium.

Solution

Alpha polonium crystallizes in a simple cubic unit cell:



a. Two adjacent Po atoms contact each other, so the edge length of this cell is equal to two Po atomic radii: $l = 2r$. Therefore, the radius of Po is:

$$r = \frac{l}{2} = \frac{336pm}{2} = 168pm.$$

b. Density is given by $\text{density} = \frac{\text{mass}}{\text{volume}}$. The density of polonium can be found by determining the density of its unit cell (the mass contained within a unit cell divided by the volume of the unit cell). Since a Po unit cell contains one-eighth of a Po atom at each of its eight corners, a unit cell contains one Po atom.

The mass of a Po unit cell can be found by

$$1 \text{ Po unit cell} \times \frac{1 \text{ Po atom}}{1 \text{ Po unit cell}} \times \frac{1molPo}{6.022 \times 10^{23} \text{ Po atoms}} \times \frac{208.998g}{1molPo} = 3.47 \times 10^{-22}g$$

The volume of a Po unit cell can be found by:

$$V = l^3 = \left(336 \times 10^{-10}cm\right)^3 = 3.79 \times 10^{-23}cm^3$$

(Note that the edge length was converted from pm to cm to get the usual volume units for density.) Therefore, the density is

$$\text{Po } = \frac{3.471 \times 10^{-22}g}{3.79 \times 10^{-23}cm^3} = 9.16g/cm^3$$

Exercise 10.7.1

The edge length of the unit cell for nickel is 0.3524 nm. The density of Ni is 8.90 g/cm^3. Does nickel crystallize in a simple cubic structure? Explain.

Answer

No. If Ni was simple cubic, its density would be given by:

$$1Ni \text{ atom} \times \frac{1molNi}{6.022 \times 10^{23}Ni \text{ atoms}} \times \frac{58.693g}{1molNi} = 9.746 \times 10^{-23}g$$

$$V = l^3 = \left(3.524 \times 10^{-8}cm\right)^3 = 4.376 \times 10^{-23}cm^3$$

Then the density of Ni would be

$$= \frac{9.746 \times 10^{-23} g}{4.376 \times 10^{-23} cm^3} = 2.23 g/cm^3$$

Since the actual density of Ni is not close to this, Ni does not form a simple cubic structure.

Most metal crystals are one of the four major types of unit cells. For now, we will focus on the three cubic unit cells: simple cubic (which we have already seen), **body-centered cubic unit cell**, and **face-centered cubic unit cell**—all of which are illustrated in Figure 10.7.5 (Note that there are actually seven different lattice systems, some of which have more than one type of lattice, for a total of 14 different types of unit cells. We leave the more complicated geometries for later in this module.)

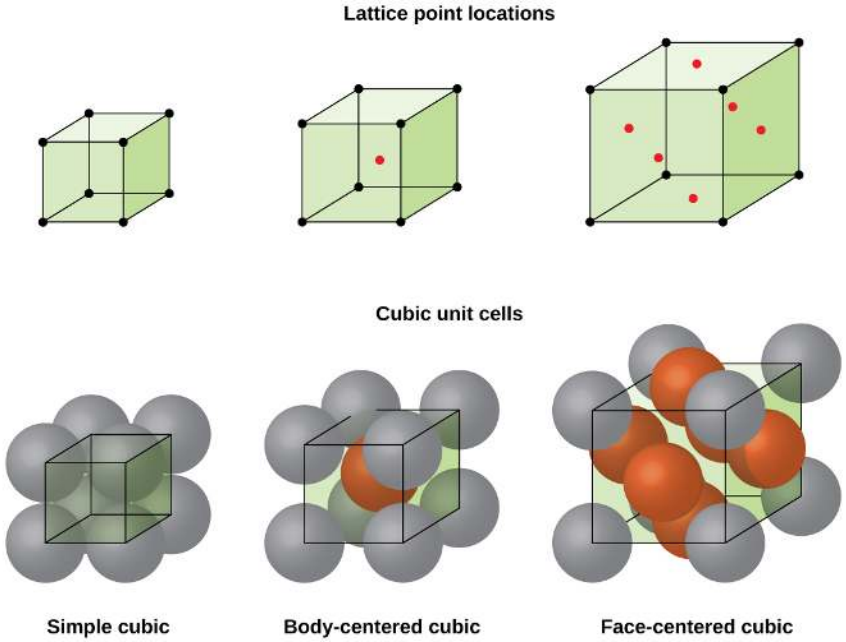


Figure 10.7.5: Cubic unit cells of metals show (in the upper figures) the locations of lattice points and (in the lower figures) metal atoms located in the unit cell.

Some metals crystallize in an arrangement that has a cubic unit cell with atoms at all of the corners and an atom in the center, as shown in Figure 10.7.6 This is called a **body-centered cubic** (BCC) solid. Atoms in the corners of a BCC unit cell do not contact each other but contact the atom in the center. A BCC unit cell contains two atoms: one-eighth of an atom at each of the eight corners atom from the corners ($8 \times \frac{1}{8} = 1$ atom from the corners) plus one atom from the center. Any atom in this structure touches four atoms in the layer above it and four atoms in the layer below it. Thus, an atom in a BCC structure has a coordination number of eight.

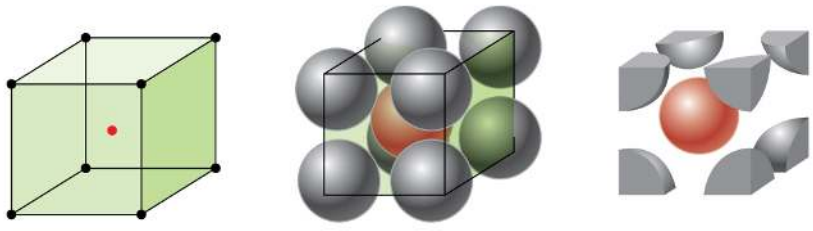


Figure 10.7.6: In a body-centered cubic structure, atoms in a specific layer do not touch each other. Each atom touches four atoms in the layer above it and four atoms in the layer below it.

Atoms in BCC arrangements are much more efficiently packed than in a simple cubic structure, occupying about 68% of the total volume. Isomorphous metals with a BCC structure include K, Ba, Cr, Mo, W, and Fe at room temperature. (Elements or compounds that crystallize with the same structure are said to be **isomorphous**.)

Many other metals, such as aluminum, copper, and lead, crystallize in an arrangement that has a cubic unit cell with atoms at all of the corners and at the centers of each face, as illustrated in Figure 10.7.7. This arrangement is called a **face-centered cubic** (FCC) solid. A FCC unit cell contains four atoms: one-eighth of an atom at each of the eight corners atom from the corners ($8 \times \frac{1}{8} = 1$ atom from the corners) and one-half of an atom on each of the six faces atoms from the faces ($6 \times \frac{1}{2} = 3$ atom from the faces). The atoms at the corners touch the atoms in the centers of the adjacent faces along the face diagonals of the cube. Because the atoms are on identical lattice points, they have identical environments.

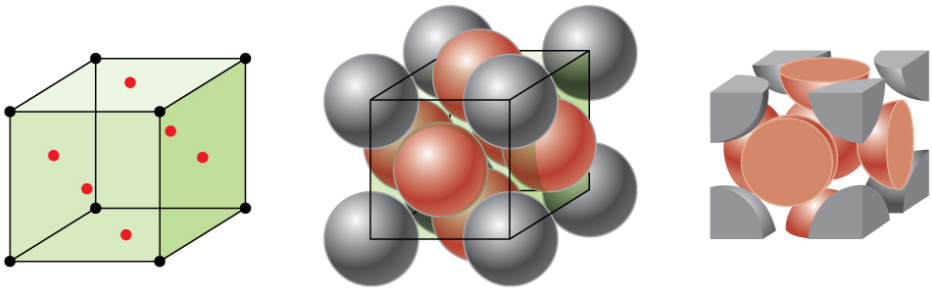


Figure 10.7.7: A face-centered cubic solid has atoms at the corners and, as the name implies, at the centers of the faces of its unit cells.

Atoms in an FCC arrangement are packed as closely together as possible, with atoms occupying 74% of the volume. This structure is also called **cubic closest packing** (CCP). In CCP, there are three repeating layers of hexagonally arranged atoms. Each atom contacts six atoms in its own layer, three in the layer above, and three in the layer below. In this arrangement, each atom touches 12 near neighbors, and therefore has a coordination number of 12. The fact that FCC and CCP arrangements are equivalent may not be immediately obvious, but why they are actually the same structure is illustrated in Figure 10.7.8.

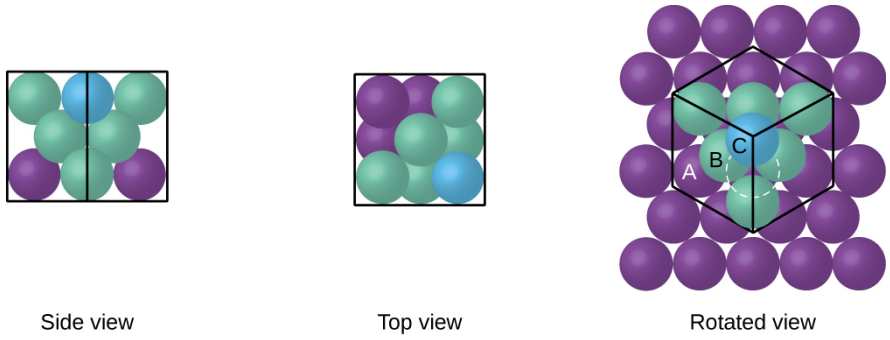


Figure 10.7.8: A CCP arrangement consists of three repeating layers (ABCABC…) of hexagonally arranged atoms. Atoms in a CCP structure have a coordination number of 12 because they contact six atoms in their layer, plus three atoms in the layer above and three atoms in the layer below. By rotating our perspective, we can see that a CCP structure has a unit cell with a face containing an atom from layer A at one corner, atoms from layer B across a diagonal (at two corners and in the middle of the face), and an atom from layer C at the remaining corner. This is the same as a face-centered cubic arrangement.

Because closer packing maximizes the overall attractions between atoms and minimizes the total intermolecular energy, the atoms in most metals pack in this manner. We find two types of closest packing in simple metallic crystalline structures: CCP, which we have already encountered, and hexagonal closest packing (HCP) shown in Figure 10.7.9. Both consist of repeating layers of **hexagonally arranged atoms**. In both types, a second layer (B) is placed on the first layer (A) so that each atom in the second layer is in contact with three atoms in the first layer. The third layer is positioned in one of two ways. In HCP, atoms in the third layer are directly above atoms in the first layer (i.e., the third layer is also type A), and the stacking consists of alternating type A and type B close-packed layers (i.e., ABABAB⋯). In CCP, atoms in the third layer are not above atoms in either of the first two layers (i.e., the third layer is type C), and the stacking consists of alternating type A, type B, and type C close-packed layers (i.e., ABCABCABC⋯). About two–thirds of all metals crystallize in closest-packed arrays with coordination numbers of 12. Metals that crystallize in an HCP structure include Cd, Co, Li, Mg, Na, and Zn, and metals that crystallize in a CCP structure include Ag, Al, Ca, Cu, Ni, Pb, and Pt.

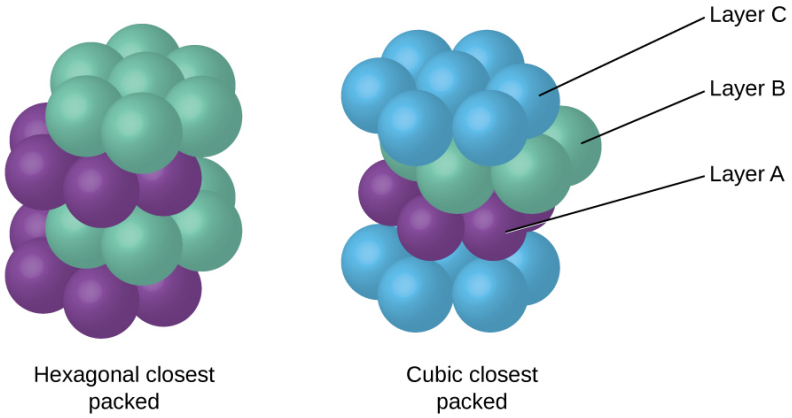


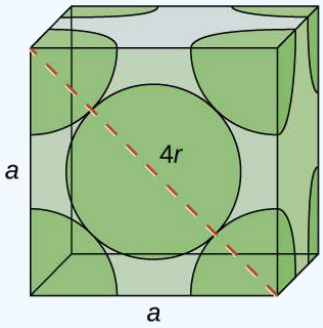
Figure 10.7.9: In both types of closest packing, atoms are packed as compactly as possible. Hexagonal closest packing consists of two alternating layers (ABABAB…). Cubic closest packing consists of three alternating layers (ABCABCABC…).

Example 10.7.2: Calculation of Atomic Radius and Density for Metals, Part 2

Calcium crystallizes in a face-centered cubic structure. The edge length of its unit cell is 558.8 pm.

a. What is the atomic radius of Ca in this structure?
b. Calculate the density of Ca.

Solution



(a) In an FCC structure, Ca atoms contact each other across the diagonal of the face, so the length of the diagonal is equal to four Ca atomic radii (d = 4*r*). Two adjacent edges and the diagonal of the face form a right triangle, with the length of each side equal to 558.8 pm and the length of the hypotenuse equal to four Ca atomic radii:

$$a^2 + a^2 = d^2 \longrightarrow (558.8pm)^2 + (558.5pm)^2 = (4r)^2$$

Solving this gives

$$r = \sqrt{\frac{(558.8pm)^2 + (558.5pm)^2}{16}} = 197.6pm \text{ for a Ca radius.}$$

(b) Density is given by $\text{density} = \frac{\text{mass}}{\text{volume}}$. The density of calcium can be found by determining the density of its unit cell: for example, the mass contained within a unit cell divided by the volume of the unit cell. A face-centered Ca unit cell has one-eighth of an atom at each of the eight corners ($8 \times \frac{1}{8} = 1$ atom from the corners) and one-half of an atom on each of the six faces ($6 \times \frac{1}{2} = 3$ atom), for a total of four atoms in the unit cell.

The mass of the unit cell can be found by:

$$1Ca \text{ unit cell} \times \frac{4Ca \text{ atoms}}{1Ca \text{ unit cell}} \times \frac{1molCa}{6.022 \times 10^{23} Ca \text{ atoms}} \times \frac{40.078g}{1molCa} = 2.662 \times 10^{-22} g$$

The volume of a Ca unit cell can be found by:

$$V = a^3 = \left(558.8 \times 10^{-10} cm\right)^3 = 1.745 \times 10^{-22} cm^3$$

(Note that the edge length was converted from pm to cm to get the usual volume units for density.)

Then, the density of Ca is

$$\frac{2.662 \times 10^{-22} g}{1.745 \times 10^{-22} cm^3} = 1.53 g/cm^3$$

? Exercise 10.7.2

Silver crystallizes in an FCC structure. The edge length of its unit cell is 409 pm.

a. What is the atomic radius of Ag in this structure?
b. Calculate the density of Ag.

Answer

a. 144 pm
b. 10.5 g/cm^3

In general, a unit cell is defined by the lengths of three axes (*a*, *b*, and *c*) and the angles (*α*, *β*, and *γ*) between them, as illustrated in Figure 10.7.10 The axes are defined as being the lengths between points in the space lattice. Consequently, unit cell axes join points with identical environments.

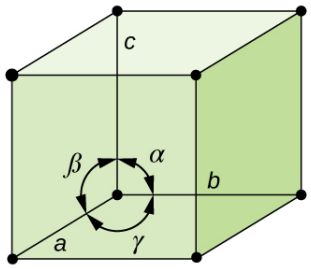


Figure 10.7.10: A unit cell is defined by the lengths of its three axes (*a*, *b*, and *c*) and the angles (*α*, *β*, and *γ*) between the axes.

There are seven different lattice systems, some of which have more than one type of lattice, for a total of fourteen different unit cells, which have the shapes shown in Figure 10.7.11

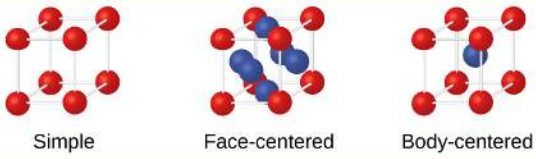
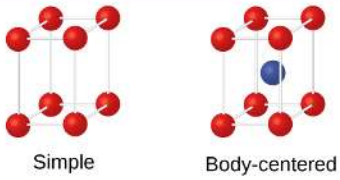
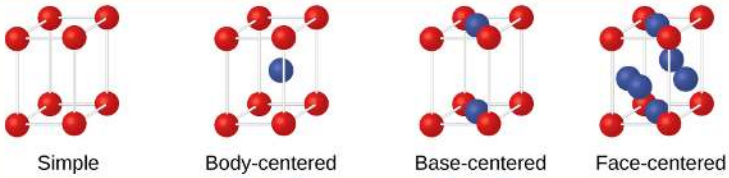
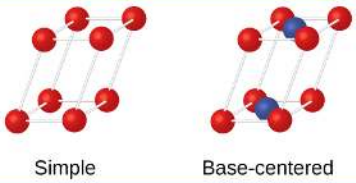
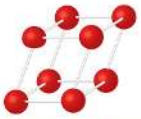
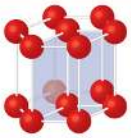
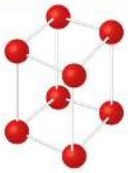
System/Axes/Angles	Unit Cells
Cubic $a = b = c$ $\alpha = \beta = \gamma = 90^\circ$	
Tetragonal $a = b \neq c$ $\alpha = \beta = \gamma = 90^\circ$	
Orthorhombic $a \neq b \neq c$ $\alpha = \beta = \gamma = 90^\circ$	
Monoclinic $a \neq b \neq c$ $\alpha = \gamma = 90^\circ; \beta \neq 90^\circ$	
Triclinic $a \neq b \neq c$ $\alpha \neq \beta \neq \gamma \neq 90^\circ$	
Hexagonal $a = b \neq c$ $\alpha = \beta = 90^\circ; \gamma = 120^\circ$	
Rhombohedral $a = b = c$ $\alpha = \beta = \gamma \neq 90^\circ$	

Figure 10.7.11: There are seven different lattice systems and 14 different unit cells.

The Structures of Ionic Crystals

Ionic crystals consist of two or more different kinds of ions that usually have different sizes. The packing of these ions into a crystal structure is more complex than the packing of metal atoms that are the same size.

Most monatomic ions behave as charged spheres, and their attraction for ions of opposite charge is the same in every direction. Consequently, stable structures for ionic compounds result (1) when ions of one charge are surrounded by as many ions as possible of the opposite charge and (2) when the cations and anions are in contact with each other. Structures are determined by two principal factors: the relative sizes of the ions and the ratio of the numbers of positive and negative ions in the compound.

In simple ionic structures, we usually find the anions, which are normally larger than the cations, arranged in a closest-packed array. (As seen previously, additional electrons attracted to the same nucleus make anions larger and fewer electrons attracted to the same nucleus make cations smaller when compared to the atoms from which they are formed.) The smaller cations commonly occupy one of two types of **holes** (or interstices) remaining between the anions. The smaller of the holes is found between three

anions in one plane and one anion in an adjacent plane. The four anions surrounding this hole are arranged at the corners of a tetrahedron, so the hole is called a tetrahedral hole. The larger type of hole is found at the center of six anions (three in one layer and three in an adjacent layer) located at the corners of an octahedron; this is called an **octahedral hole**. Figure 10.7.12illustrates both of these types of holes.

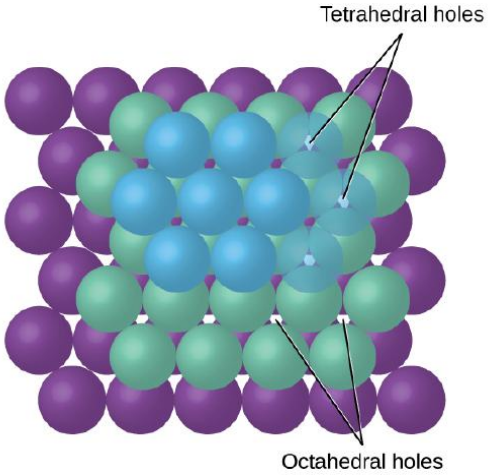


Figure 10.7.12: Cations may occupy two types of holes between anions: octahedral holes or tetrahedral holes.

Depending on the relative sizes of the cations and anions, the cations of an ionic compound may occupy tetrahedral or octahedral holes, as illustrated in Figure 10.7.13 Relatively small cations occupy tetrahedral holes, and larger cations occupy octahedral holes. If the cations are too large to fit into the octahedral holes, the anions may adopt a more open structure, such as a simple cubic array. The larger cations can then occupy the larger cubic holes made possible by the more open spacing.

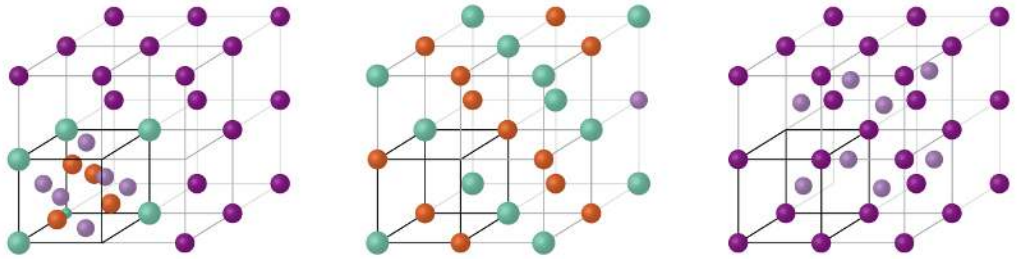


Figure 10.7.13: A cation's size and the shape of the hole occupied by the compound are directly related.

There are two tetrahedral holes for each anion in either an HCP or CCP array of anions. A compound that crystallizes in a closest-packed array of anions with cations in the tetrahedral holes can have a maximum cation:anion ratio of 2:1; all of the tetrahedral holes are filled at this ratio. Examples include Li_2O, Na_2O, Li_2S, and Na_2S. Compounds with a ratio of less than 2:1 may also crystallize in a closest-packed array of anions with cations in the tetrahedral holes, if the ionic sizes fit. In these compounds, however, some of the tetrahedral holes remain vacant.

Example 10.7.3: Occupancy of Tetrahedral Holes

Zinc sulfide is an important industrial source of zinc and is also used as a white pigment in paint. Zinc sulfide crystallizes with zinc ions occupying one-half of the tetrahedral holes in a closest-packed array of sulfide ions. What is the formula of zinc sulfide?

Solution

Because there are two tetrahedral holes per anion (sulfide ion) and one-half of these holes are occupied by zinc ions, there must be $\frac{1}{2} \times 2$ or 1, zinc ion per sulfide ion. Thus, the formula is ZnS.

Exercise 10.7.3

Lithium selenide can be described as a closest-packed array of selenide ions with lithium ions in all of the tetrahedral holes. What it the formula of lithium selenide?

Answer

Li_2Se

The ratio of octahedral holes to anions in either an HCP or CCP structure is 1:1. Thus, compounds with cations in octahedral holes in a closest-packed array of anions can have a maximum cation:anion ratio of 1:1. In NiO, MnS, NaCl, and KH, for example, all of the octahedral holes are filled. Ratios of less than 1:1 are observed when some of the octahedral holes remain empty.

Example 10.7.4: Stoichiometry of Ionic Compounds

Sapphire is aluminum oxide. Aluminum oxide crystallizes with aluminum ions in two-thirds of the octahedral holes in a closest-packed array of oxide ions. What is the formula of aluminum oxide?

Solution

Because there is one octahedral hole per anion (oxide ion) and only two-thirds of these holes are occupied, the ratio of aluminum to oxygen must be $\frac{2}{3}:1$, which would give $Al_{\frac{2}{3}}O$. The simplest whole number ratio is 2:3, so the formula is Al_2O_3.

Exercise 10.7.4

The white pigment titanium oxide crystallizes with titanium ions in one-half of the octahedral holes in a closest-packed array of oxide ions. What is the formula of titanium oxide?

Answer

TiO_2

In a simple cubic array of anions, there is one cubic hole that can be occupied by a cation for each anion in the array. In CsCl, and in other compounds with the same structure, all of the cubic holes are occupied. Half of the cubic holes are occupied in SrH_2, UO_2, $SrCl_2$, and CaF_2.

Different types of ionic compounds often crystallize in the same structure when the relative sizes of their ions and their stoichiometries (the two principal features that determine structure) are similar.

Unit Cells of Ionic Compounds

Many ionic compounds crystallize with cubic unit cells, and we will use these compounds to describe the general features of ionic structures.

When an ionic compound is composed of cations and anions of similar size in a 1:1 ratio, it typically forms a simple cubic structure. Cesium chloride, CsCl, (illustrated in Figure 10.7.14) is an example of this, with Cs^+ and Cl^- having radii of 174 pm and 181 pm, respectively. We can think of this as chloride ions forming a simple cubic unit cell, with a cesium ion in the center; or as cesium ions forming a unit cell with a chloride ion in the center; or as simple cubic unit cells formed by Cs^+ ions overlapping unit cells formed by Cl^- ions. Cesium ions and chloride ions touch along the body diagonals of the unit cells. One cesium ion and one chloride ion are present per unit cell, giving the l:l stoichiometry required by the formula for cesium chloride. Note that there is no lattice point in the center of the cell, and CsCl is not a BCC structure because a cesium ion is not identical to a chloride ion.

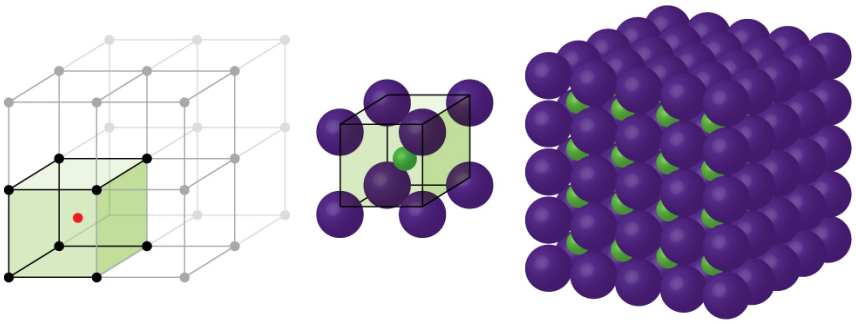


Figure 10.7.14: Ionic compounds with similar-sized cations and anions, such as CsCl, usually form a simple cubic structure. They can be described by unit cells with either cations at the corners or anions at the corners.

We have said that the location of lattice points is arbitrary. This is illustrated by an alternate description of the CsCl structure in which the lattice points are located in the centers of the cesium ions. In this description, the cesium ions are located on the lattice points at the corners of the cell, and the chloride ion is located at the center of the cell. The two unit cells are different, but they describe identical structures.

When an ionic compound is composed of a 1:1 ratio of cations and anions that differ significantly in size, it typically crystallizes with an FCC unit cell, like that shown in Figure 10.7.15 Sodium chloride, NaCl, is an example of this, with Na^+ and Cl^- having radii of 102 pm and 181 pm, respectively. We can think of this as chloride ions forming an FCC cell, with sodium ions located in the octahedral holes in the middle of the cell edges and in the center of the cell. The sodium and chloride ions touch each other along the cell edges. The unit cell contains four sodium ions and four chloride ions, giving the 1:1 stoichiometry required by the formula, NaCl.

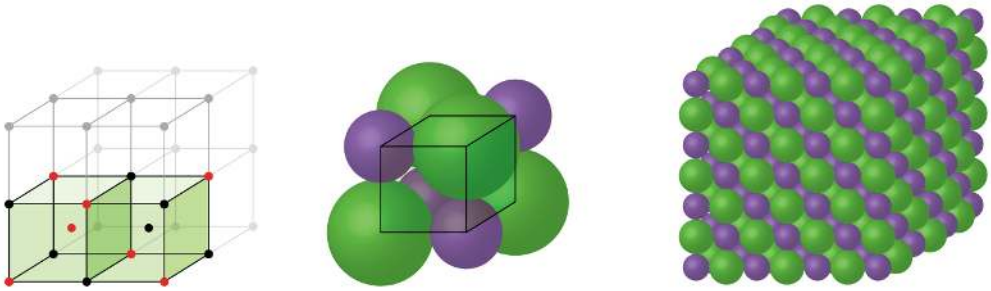


Figure 10.7.15: Ionic compounds with anions that are much larger than cations, such as NaCl, usually form an FCC structure. They can be described by FCC unit cells with cations in the octahedral holes.

The cubic form of zinc sulfide, zinc blende, also crystallizes in an FCC unit cell, as illustrated in Figure 10.7.16 This structure contains sulfide ions on the lattice points of an FCC lattice. (The arrangement of sulfide ions is identical to the arrangement of chloride ions in sodium chloride.) The radius of a zinc ion is only about 40% of the radius of a sulfide ion, so these small Zn^{2+} ions are located in alternating tetrahedral holes, that is, in one half of the tetrahedral holes. There are four zinc ions and four sulfide ions in the unit cell, giving the empirical formula ZnS.

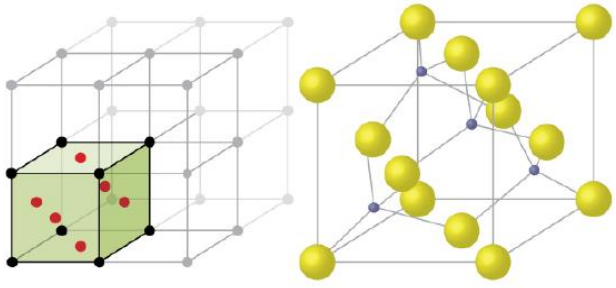
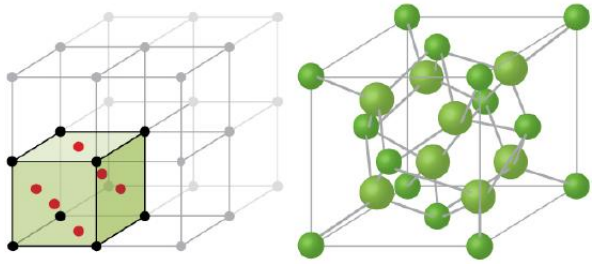


Figure 10.7.16: ZnS, zinc sulfide (or zinc blende) forms an FCC unit cell with sulfide ions at the lattice points and much smaller zinc ions occupying half of the tetrahedral holes in the structure.

A calcium fluoride unit cell, like that shown in Figure 10.7.17, is also an FCC unit cell, but in this case, the cations are located on the lattice points; equivalent calcium ions are located on the lattice points of an FCC lattice. All of the tetrahedral sites in the FCC array of calcium ions are occupied by fluoride ions. There are four calcium ions and eight fluoride ions in a unit cell, giving a calcium:fluorine ratio of 1:2, as required by the chemical formula, CaF_2. Close examination of Figure 10.7.17will reveal a simple cubic array of fluoride ions with calcium ions in one half of the cubic holes. The structure cannot be described in terms of a **space lattice** of points on the fluoride ions because the fluoride ions do not all have identical environments. The orientation of the four calcium ions about the fluoride ions differs.



CaF_2 face-centered unit cell

Figure 10.7.17: Calcium fluoride, CaF_2, forms an FCC unit cell with calcium ions (green) at the lattice points and fluoride ions (red) occupying all of the tetrahedral sites between them.

Calculation of Ionic Radii

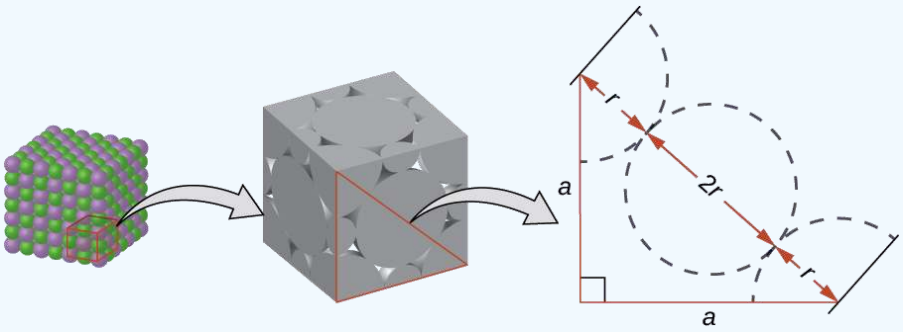
If we know the edge length of a unit cell of an ionic compound and the position of the ions in the cell, we can calculate ionic radii for the ions in the compound if we make assumptions about individual ionic shapes and contacts.

Example 10.7.5: Calculation of Ionic Radii

The edge length of the unit cell of LiCl (NaCl-like structure, FCC) is 0.514 nm or 5.14 Å. Assuming that the lithium ion is small enough so that the chloride ions are in contact, as in Figure 10.7.15, calculate the ionic radius for the chloride ion. Note: The length unit angstrom, Å, is often used to represent atomic-scale dimensions and is equivalent to 10^{-10} m.

Solution

On the face of a LiCl unit cell, chloride ions contact each other across the diagonal of the face:



Drawing a right triangle on the face of the unit cell, we see that the length of the diagonal is equal to four chloride radii (one radius from each corner chloride and one diameter—which equals two radii—from the chloride ion in the center of the face), so $d = 4r$. From the Pythagorean theorem, we have:

a^2+a^2=d^2 \nonumber \]

which yields:

$$(0.514nm)^2 + (0.514nm)^2 = (4r)^2 = 16r^2$$

Solving this gives:

$$r = \sqrt{\frac{(0.514nm)^2 + (0.514nm)^2}{16}} = 0.182nm(1.82\text{Å}) \text{ for a } Cl^- \text{ radius.}$$

Exercise 10.7.1

The edge length of the unit cell of KCl (NaCl-like structure, FCC) is 6.28 Å. Assuming anion-cation contact along the cell edge, calculate the radius of the potassium ion. The radius of the chloride ion is 1.82 Å.

Answer

The radius of the potassium ion is 1.33 Å.

It is important to realize that values for ionic radii calculated from the edge lengths of unit cells depend on numerous assumptions, such as a perfect spherical shape for ions, which are approximations at best. Hence, such calculated values are themselves approximate and comparisons cannot be pushed too far. Nevertheless, this method has proved useful for calculating ionic radii from experimental measurements such as X-ray crystallographic determinations.

X-Ray Crystallography

The size of the unit cell and the arrangement of atoms in a crystal may be determined from measurements of the *diffraction* of X-rays by the crystal, termed X-ray crystallography. Diffraction is the change in the direction of travel experienced by an electromagnetic wave when it encounters a physical barrier whose dimensions are comparable to those of the wavelength of the light. X-rays are electromagnetic radiation with wavelengths about as long as the distance between neighboring atoms in crystals (on the order of a few Å).

When a beam of monochromatic X-rays strikes a crystal, its rays are scattered in all directions by the atoms within the crystal. When scattered waves traveling in the same direction encounter one another, they undergo *interference*, a process by which the waves combine to yield either an increase or a decrease in amplitude (intensity) depending upon the extent to which the combining waves' maxima are separated (see Figure 10.7.18).

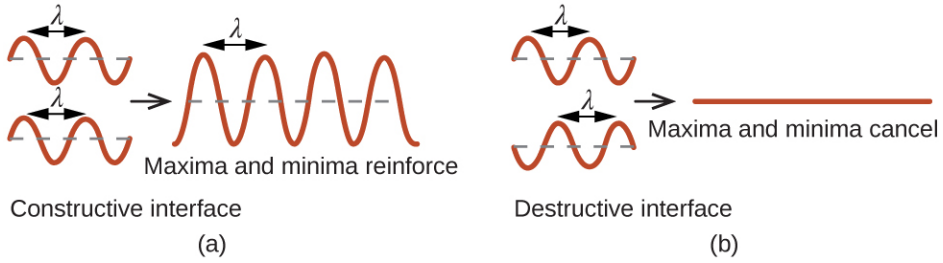


Figure 10.7.18: Light waves occupying the same space experience interference, combining to yield waves of greater (a) or lesser (b) intensity, depending upon the separation of their maxima and minima.

When X-rays of a certain wavelength, λ, are scattered by atoms in adjacent crystal planes separated by a distance, d, they may undergo constructive interference when the difference between the distances traveled by the two waves prior to their combination is an integer factor, n, of the wavelength. This condition is satisfied when the angle of the diffracted beam, θ, is related to the wavelength and interatomic distance by the equation:

$$n\lambda = 2d\sin\theta \tag{10.7.1}$$

This relation is known as the **Bragg equation** in honor of W. H. Bragg, the English physicist who first explained this phenomenon. Figure 10.7.19 illustrates two examples of diffracted waves from the same two crystal planes. The figure on the left depicts waves diffracted at the Bragg angle, resulting in constructive interference, while that on the right shows diffraction and a different angle that does not satisfy the Bragg condition, resulting in destructive interference.

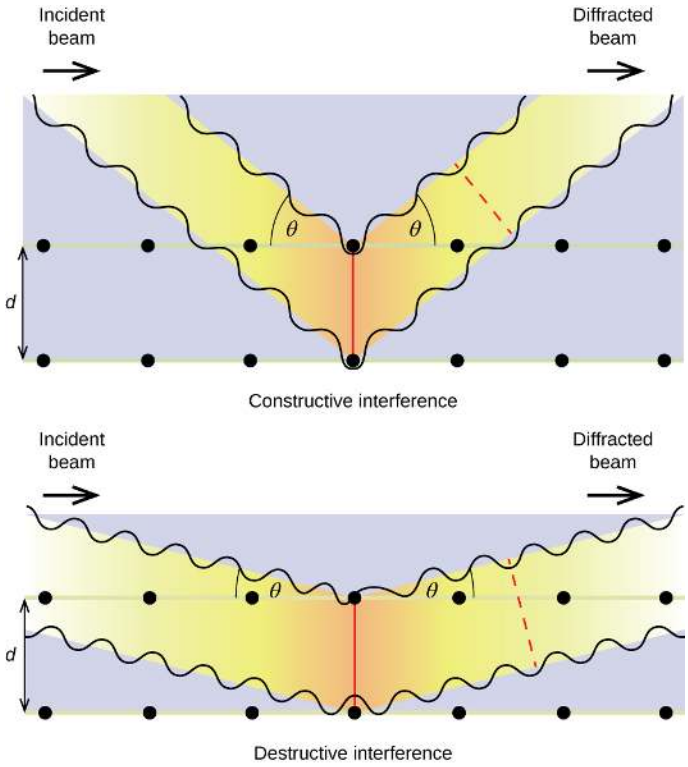


Figure 10.7.19: The diffraction of X-rays scattered by the atoms within a crystal permits the determination of the distance between the atoms. The top image depicts constructive interference between two scattered waves and a resultant diffracted wave of high intensity. The bottom image depicts destructive interference and a low intensity diffracted wave.

Link to Learning

Visit this site for more details on the Bragg equation and a simulator that allows you to explore the effect of each variable on the intensity of the diffracted wave.

An X-ray diffractometer, such as the one illustrated in Figure 10.7.20, may be used to measure the angles at which X-rays are diffracted when interacting with a crystal as described earlier. From such measurements, the Bragg equation may be used to compute distances between atoms as demonstrated in the following example exercise.

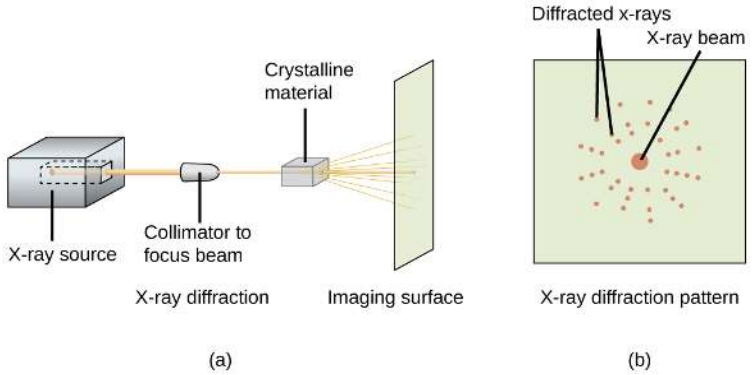


Figure 10.7.20: (a) In a diffractometer, a beam of X-rays strikes a crystalline material, producing (b) an X-ray diffraction pattern that can be analyzed to determine the crystal structure.

Example 10.7.6: Using the Bragg Equation

In a diffractometer, X-rays with a wavelength of 0.1315 nm were used to produce a diffraction pattern for copper. The first order diffraction ($n = 1$) occurred at an angle $\theta = 25.25°$. Determine the spacing between the diffracting planes in copper.

Solution

The distance between the planes is found by solving the Bragg equation (Equation 10.7.1), $n\lambda = 2d\sin\theta$, for d.

This gives:

$$d = \frac{n\lambda}{2\sin\theta} = \frac{1(0.1315nm)}{2\sin(25.25^\circ)} = 0.154nm$$

Exercise 10.7.6

A crystal with spacing between planes equal to 0.394 nm diffracts X-rays with a wavelength of 0.147 nm. What is the angle for the first order diffraction?

Answer

10.8°

Portrait of a Chemist: X-ray Crystallographer Rosalind Franklin

The discovery of the structure of DNA in 1953 by Francis Crick and James Watson is one of the great achievements in the history of science. They were awarded the 1962 Nobel Prize in Physiology or Medicine, along with Maurice Wilkins, who provided experimental proof of DNA's structure. British chemist Rosalind Franklin made invaluable contributions to this monumental achievement through her work in measuring X-ray diffraction images of DNA. Early in her career, Franklin's research on the structure of coals proved helpful to the British war effort. After shifting her focus to biological systems in the early 1950s, Franklin and doctoral student Raymond Gosling discovered that DNA consists of two forms: a long, thin fiber formed when wet (type "B") and a short, wide fiber formed when dried (type "A"). Her X-ray diffraction images of DNA (Figure 10.7.21) provided the crucial information that allowed Watson and Crick to confirm that DNA forms a double helix, and to determine details of its size and structure. Franklin also conducted pioneering research on viruses and the RNA that contains their genetic information, uncovering new information that radically changed the body of knowledge in the field. After developing ovarian cancer, Franklin continued to work until her death in 1958 at age 37. Among many posthumous recognitions of her work, the Chicago Medical School of Finch University of Health Sciences changed its name to the Rosalind Franklin University of Medicine and Science in 2004, and adopted an image of her famous X-ray diffraction image of DNA as its official university logo.

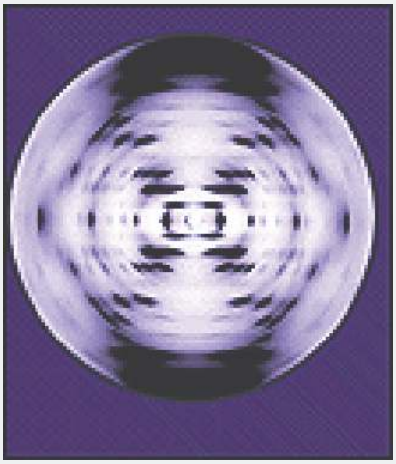


Figure 10.7.21: This illustration shows an X-ray diffraction image similar to the one Franklin found in her research. (credit: National Institutes of Health)

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10.8: 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

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10.9: Key Equations

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10.10: 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.

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CHAPTER OVERVIEW

11: Solutions and Colloids

In this chapter, we will consider the nature of solutions, and examine factors that determine whether a solution will form and what properties it may have. In addition, we will discuss colloids—systems that resemble solutions but consist of dispersions of particles somewhat larger than ordinary molecules or ions.

11.1: Introduction
11.2: The Dissolution Process
11.3: Electrolytes
11.4: Solubility
11.5: Colligative Properties
11.6: Colloids
11.7: Key Terms
11.8: Key Equations
11.9: Summary

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11.1: Introduction

Coral reefs are home to about 25% of all marine species. They are being threatened by climate change, oceanic acidification, and water pollution, all of which change the composition of the solution known as seawater. Dissolved oxygen in seawater is critical for sea creatures, but as the oceans warm, oxygen becomes less soluble. As the concentration of carbon dioxide in the atmosphere increases, the concentration of carbon dioxide in the oceans increases, contributing to oceanic acidification. Coral reefs are particularly sensitive to the acidification of the ocean, since the exoskeletons of the coral polyps are soluble in acidic solutions. Humans contribute to the changing of seawater composition by allowing agricultural runoff and other forms of pollution to affect our oceans.



Figure 11.1.1: Coral reefs, such as this one at the Palmyra Atoll National Wildlife Refuge, are vital to the ecosystem of earth's oceans. The health of coral reefs and all marine life depends on the specific chemical composition of the complex mixture known as seawater. (credit: modification of work by "USFWS – Pacific Region"/Wikimedia Commons)

Solutions are crucial to the processes that sustain life and to many other processes involving chemical reactions. This chapter considers the nature of solutions and examines factors that determine whether a solution will form and what properties it may have. The properties of colloids—mixtures containing dispersed particles larger than the molecules and ions of typical solutions—are also discussed.

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11.2: The Dissolution Process

Learning Objectives

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

- Describe the basic properties of solutions and how they form
- Predict whether a given mixture will yield a solution based on molecular properties of its components
- Explain why some solutions either produce or absorb heat when they form

An earlier chapter of this text introduced *solutions*, defined as homogeneous mixtures of two or more substances. Often, one component of a solution is present at a significantly greater concentration, in which case it is called the *solvent*. The other components of the solution present in relatively lesser concentrations are called *solutes*. Sugar is a covalent solid composed of sucrose molecules, $C_{12}H_{22}O_{11}$. When this compound dissolves in water, its molecules become uniformly distributed among the molecules of water:

$$C_{12}H_{22}O_{11}(s) \longrightarrow C_{12}H_{22}O_{11}(aq)$$

The subscript "*aq*" in the equation signifies that the sucrose molecules are solutes and are therefore *individually dispersed* throughout the *aqueous solution* (water is the solvent). Although sucrose molecules are heavier than water molecules, they remain dispersed throughout the solution; gravity does not cause them to "settle out" over time.

Potassium dichromate, $K_2Cr_2O_7$, is an ionic compound composed of colorless potassium ions, K^+, and orange dichromate ions, $Cr_2O_7^{2-}$. When a small amount of solid potassium dichromate is added to water, the compound dissolves and dissociates to yield potassium ions and dichromate ions uniformly distributed throughout the mixture (Figure 11.2.1), as indicated in this equation:

$$K_2Cr_2O_7(s) \longrightarrow 2\,K^+(aq) + Cr_2O_7^{2-}(aq)$$

As with the mixture of sugar and water, this mixture is also an aqueous solution. Its solutes, potassium and dichromate ions, remain individually dispersed among the solvent (water) molecules.



Figure 11.2.1: When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott)

Link to Learning

Visit this virtual lab to view simulations of the dissolution of common covalent and ionic substances (sugar and salt) in water.

Water is used so often as a solvent that the word solution has come to imply an aqueous solution to many people. However, almost any gas, liquid, or solid can act as a solvent. Many alloys are solid solutions of one metal dissolved in another; for example, US five-cent coins contain nickel dissolved in copper. Air is a gaseous solution, a homogeneous mixture of nitrogen, oxygen, and several other gases. Oxygen (a gas), alcohol (a liquid), and sugar (a solid) all dissolve in water (a liquid) to form liquid solutions. Table 11.2.1 gives examples of several different solutions and the phases of the solutes and solvents.

Table 11.2.1: Different Types of Solutions

Solution	Solute	Solvent
air	$O_2(g)$	$N_2(g)$
soft drinks[1]	$CO_2(g)$	$H_2O(l)$

Solution	Solute	Solvent
hydrogen in palladium	$H_2(g)$	Pd(*s*)
rubbing alcohol	$H_2O(l)$	$C_3H_8O(l)$ (2-propanol)
saltwater	NaCl(*s*)	$H_2O(l)$
brass	Zn(*s*)	Cu(*s*)

Solutions exhibit these defining traits:

- They are homogeneous; after a solution is mixed, it has the same composition at all points throughout (its composition is uniform).
- The physical state of a solution—solid, liquid, or gas—is typically the same as that of the solvent, as demonstrated by the examples in Table 11.2.1.
- The components of a solution are dispersed on a molecular scale; they consist of a mixture of separated solute particles (molecules, atoms, and/or ions) each closely surrounded by solvent species.
- The dissolved solute in a solution will not settle out or separate from the solvent.
- The composition of a solution, or the concentrations of its components, can be varied continuously (within limits determined by the *solubility* of the components, discussed in detail later in this chapter).

The Formation of Solutions

The formation of a solution is an example of a spontaneous process, a process that occurs under specified conditions without the requirement of energy from some external source. Sometimes a mixture is stirred to speed up the dissolution process, but this is not necessary; a homogeneous solution will form eventually. The topic of spontaneity is critically important to the study of chemical thermodynamics and is treated more thoroughly in a later chapter of this text. For purposes of this chapter's discussion, it will suffice to consider two criteria that *favor*, but do not guarantee, the spontaneous formation of a solution:

1. a decrease in the internal energy of the system (an exothermic change, as discussed in the previous chapter on thermochemistry)
2. an increased dispersal of matter in the system (which indicates an increase in the *entropy* of the system, as you will learn about in the later chapter on thermodynamics)

In the process of dissolution, an internal energy change often, but not always, occurs as heat is absorbed or evolved. An increase in matter dispersal always results when a solution forms from the uniform distribution of solute molecules throughout a solvent.

When the strengths of the intermolecular forces of attraction between solute and solvent species in a solution are no different than those present in the separated components, the solution is formed with no accompanying energy change. Such a solution is called an ideal solution. A mixture of ideal gases (or gases such as helium and argon, which closely approach ideal behavior) is an example of an ideal solution, since the entities comprising these gases experience no significant intermolecular attractions.

When containers of helium and argon are connected, the gases spontaneously mix due to diffusion and form a solution (Figure 11.2.2). The formation of this solution clearly involves an increase in matter dispersal, since the helium and argon atoms occupy a volume twice as large as that which each occupied before mixing.

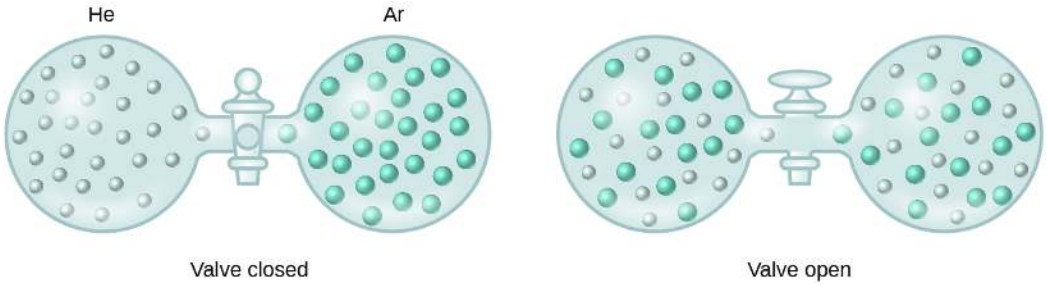


Figure 11.2.2: Samples of helium and argon spontaneously mix to give a solution.

Ideal solutions may also form when structurally similar liquids are mixed. For example, mixtures of the alcohols methanol (CH_3OH) and ethanol (C_2H_5OH) form ideal solutions, as do mixtures of the hydrocarbons pentane, C_5H_{12}, and hexane, C_6H_{14}. Placing methanol and ethanol, or pentane and hexane, in the bulbs shown in Figure 11.2.2 will result in the same diffusion and

subsequent mixing of these liquids as is observed for the He and Ar gases (although at a much slower rate), yielding solutions with no significant change in energy. Unlike a mixture of gases, however, the components of these liquid-liquid solutions do, indeed, experience intermolecular attractive forces. But since the molecules of the two substances being mixed are structurally very similar, the intermolecular attractive forces between like and unlike molecules are essentially the same, and the dissolution process, therefore, does not entail any appreciable increase or decrease in energy. These examples illustrate how increased matter dispersal alone can provide the driving force required to cause the spontaneous formation of a solution. In some cases, however, the relative magnitudes of intermolecular forces of attraction between solute and solvent species may prevent dissolution.

Three types of intermolecular attractive forces are relevant to the dissolution process: solute-solute, solvent-solvent, and solute-solvent. As illustrated in Figure 11.2.3 , the formation of a solution may be viewed as a stepwise process in which energy is consumed to overcome solute-solute and solvent-solvent attractions (endothermic processes) and released when solute-solvent attractions are established (an exothermic process referred to as solvation). The relative magnitudes of the energy changes associated with these stepwise processes determine whether the dissolution process overall will release or absorb energy. In some cases, solutions do not form because the energy required to separate solute and solvent species is so much greater than the energy released by solvation.

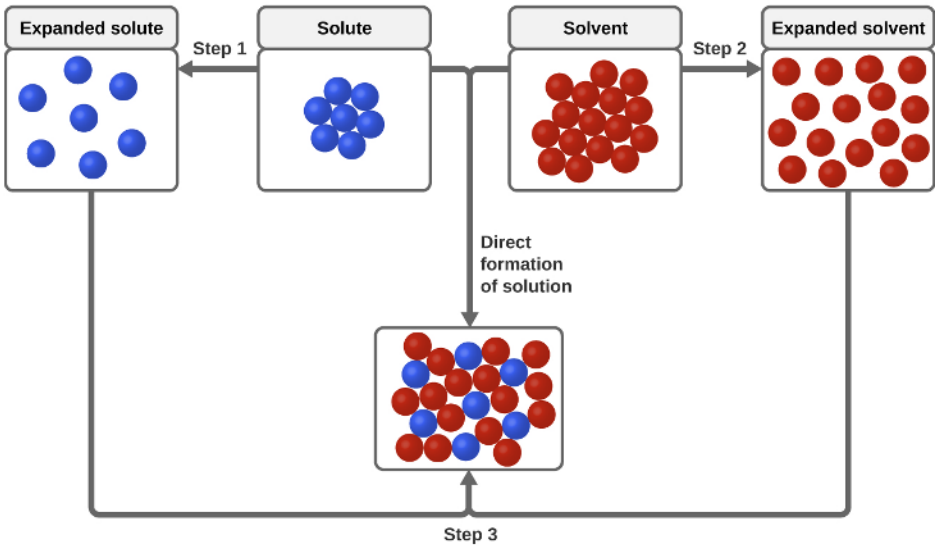


Figure 11.2.3: This schematic representation of dissolution shows a stepwise process involving the endothermic separation of solute and solvent species (Steps 1 and 2) and exothermic solvation (Step 3).

Consider the example of an ionic compound dissolving in water. Formation of the solution requires the electrostatic forces between the cations and anions of the compound (solute–solute) be overcome completely as attractive forces are established between these ions and water molecules (solute–solvent). Hydrogen bonding between a relatively small fraction of the water molecules must also be overcome to accommodate any dissolved solute. If the solute's electrostatic forces are significantly greater than the solvation forces, the dissolution process is significantly endothermic and the compound may not dissolve to an appreciable extent. Calcium carbonate, the major component of coral reefs, is one example of such an "insoluble" ionic compound. On the other hand, if the solvation forces are much stronger than the compound's electrostatic forces, the dissolution is significantly exothermic and the compound may be highly soluble. common example of this type of ionic compound is sodium hydroxide, commonly known as lye.

As noted at the beginning of this module, spontaneous solution formation is favored, but not guaranteed, by exothermic dissolution processes. While many soluble compounds do, indeed, dissolve with the release of heat, some dissolve endothermically. Ammonium nitrate (NH_4NO_3) is one such example and is used to make instant cold packs, like the one pictured in Figure 11.2.4, which are used for treating injuries. A thin-walled plastic bag of water is sealed inside a larger bag with solid NH_4NO_3. When the smaller bag is broken, a solution of NH_4NO_3 forms, absorbing heat from the surroundings (the injured area to which the pack is applied) and providing a cold compress that decreases swelling. Endothermic dissolutions such as this one require a greater energy input to separate the solute species than is recovered when the solutes are solvated, but they are spontaneous nonetheless due to the increase in disorder that accompanies formation of the solution.



Figure 11.2.4: An instant cold pack gets cold when certain salts, such as ammonium nitrate, dissolve in water—an endothermic process.

Link to Learning

Watch this brief video illustrating endothermic and exothermic dissolution processes.

Footnotes

- 1If bubbles of gas are observed within the liquid, the mixture is not homogeneous and, thus, not a solution.

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11.3: Electrolytes

Learning Objectives

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

- Define and give examples of electrolytes
- Distinguish between the physical and chemical changes that accompany dissolution of ionic and covalent electrolytes
- Relate electrolyte strength to solute-solvent attractive forces

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called electrolytes. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 11.3.1).

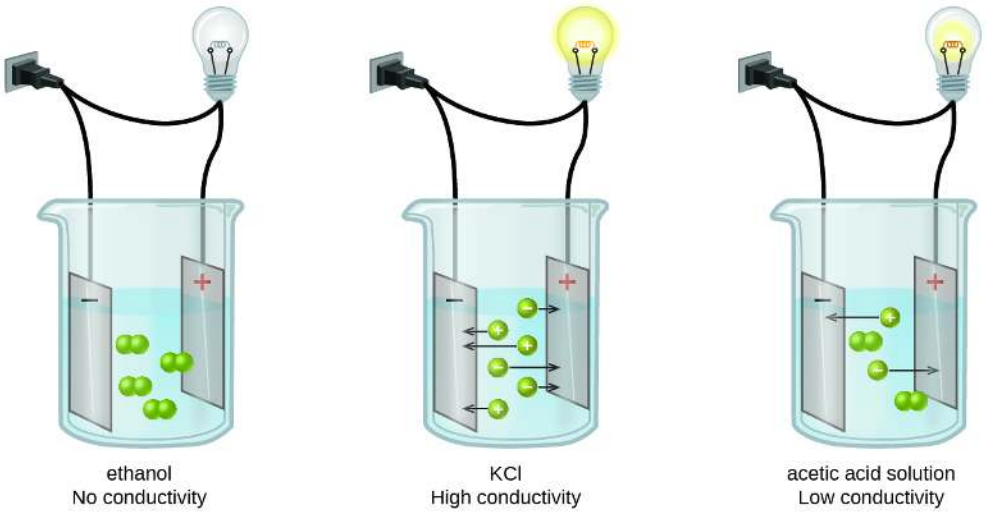


Figure 11.3.1: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 11.3.2 The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

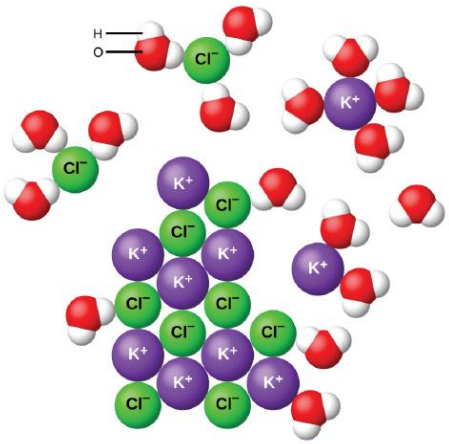


Figure 11.3.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes. Even sparingly, soluble ionic compounds are strong electrolytes, since the small amount that does dissolve will dissociate completely.

Consider what happens at the microscopic level when solid KCl is added to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules surround individual K^+ and Cl^- ions, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 11.3.2 shows. Overcoming the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution as the ions transition from fixed positions in the undissolved compound to widely dispersed, solvated ions in solution.

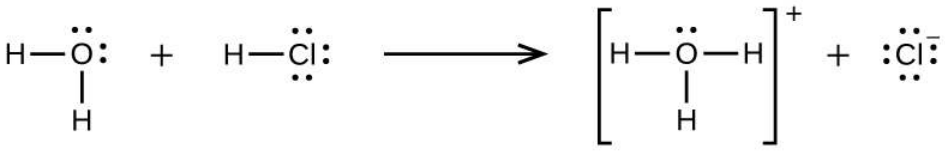
Covalent Electrolytes

Pure water is an extremely poor conductor of electricity because it is only very slightly ionized—only about two out of every 1 billion molecules ionize at 25 °C. Water ionizes when one molecule of water gives up a proton (H^+ ion) to another molecule of water, yielding hydronium and hydroxide ions.

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

In some cases, solutions prepared from covalent compounds conduct electricity because the solute molecules react chemically with the solvent to produce ions. For example, pure hydrogen chloride is a gas consisting of covalent HCl molecules. This gas contains no ions. However, an aqueous solution of HCl is a very good conductor, indicating that an appreciable concentration of ions exists within the solution.

Because HCl is an *acid*, its molecules react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-):



This reaction is essentially 100% complete for HCl (i.e., it is a *strong acid* and, consequently, a strong electrolyte). Likewise, weak acids and bases that only react partially generate relatively low concentrations of ions when dissolved in water and are classified as weak electrolytes. The reader may wish to review the discussion of strong and weak acids provided in the earlier chapter of this text on reaction classes and stoichiometry.

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11.4: Solubility

Learning Objectives

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

- Describe the effects of temperature and pressure on solubility
- State Henry's law and use it in calculations involving the solubility of a gas in a liquid
- Explain the degrees of solubility possible for liquid-liquid solutions

Imagine adding a small amount of sugar to a glass of water, stirring until all the sugar has dissolved, and then adding a bit more. You can repeat this process until the sugar concentration of the solution reaches its natural limit, a limit determined primarily by the relative strengths of the solute-solute, solute-solvent, and solvent-solvent attractive forces discussed in the previous two modules of this chapter. You can be certain that you have reached this limit because, no matter how long you stir the solution, undissolved sugar remains. The concentration of sugar in the solution at this point is known as its solubility.

The **solubility** of a solute in a particular solvent is the maximum concentration that may be achieved under given conditions when the dissolution process is *at equilibrium*.

When a solute's concentration is equal to its solubility, the solution is said to be **saturated** with that solute. If the solute's concentration is less than its solubility, the solution is said to be **unsaturated**. A solution that contains a relatively low concentration of solute is called dilute, and one with a relatively high concentration is called concentrated.

Link to Learning

Use this interactive simulation to prepare various saturated solutions.

Solutions may be prepared in which a solute concentration *exceeds* its solubility. Such solutions are said to be supersaturated, and they are interesting examples of *nonequilibrium* states (a detailed treatment of this important concept is provided in the text chapters on equilibrium). For example, the carbonated beverage in an open container that has not yet "gone flat" is supersaturated with carbon dioxide gas; given time, the CO_2 concentration will decrease until it reaches its solubility.

Link to Learning

Watch this impressive video showing the precipitation of sodium acetate from a supersaturated solution.

Solutions of Gases in Liquids

As for any solution, the solubility of a gas in a liquid is affected by the intermolecular attractive forces between solute and solvent species. Unlike solid and liquid solutes, however, there is no solute-solute intermolecular attraction to overcome when a gaseous solute dissolves in a liquid solvent (see Figure 11.4) since the atoms or molecules comprising a gas are far separated and experience negligible interactions. Consequently, solute-solvent interactions are the sole energetic factor affecting solubility. For example, the water solubility of oxygen is approximately three times greater than that of helium (there are greater dispersion forces between water and the larger oxygen molecules) but 100 times less than the solubility of chloromethane, $CHCl_3$ (polar chloromethane molecules experience dipole–dipole attraction to polar water molecules). Likewise note the solubility of oxygen in hexane, C_6H_{14}, is approximately 20 times greater than it is in water because greater dispersion forces exist between oxygen and the larger hexane molecules.

Temperature is another factor affecting solubility, with gas solubility typically decreasing as temperature increases (Figure 11.4.1). This inverse relation between temperature and dissolved gas concentration is responsible for one of the major impacts of thermal pollution in natural waters.

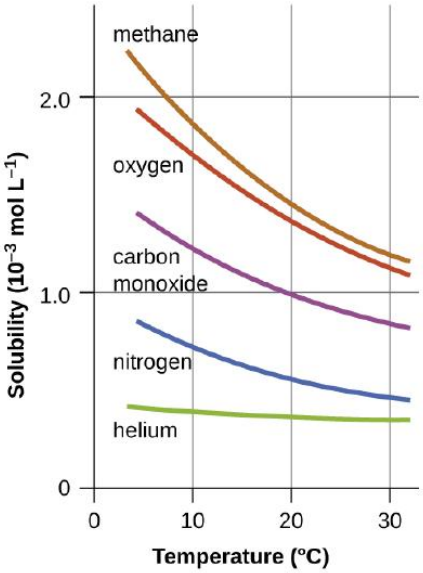


Figure 11.4.1: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised, 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 11.4.2).



Figure 11.4.2: (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 11.4.3). 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."

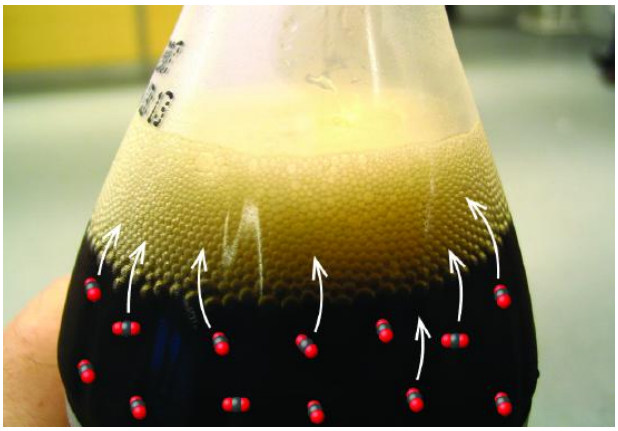


Figure 11.4.3: Opening the bottle of carbonated beverage reduces the pressure of the *gaseous* carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some *dissolved* carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

For many gaseous solutes, the relation between solubility, C_g, and partial pressure, P_g, is a proportional one:

$$C_g = kP_g$$

where *k* is a proportionality constant that depends on the identity of the gaseous solute, the identity of the solvent, and 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 11.4.1: 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 is 1.38 \times 10^{−3} mol L^{-1}. Use Henry's law to determine the solubility of oxygen when its partial pressure is 20.7 kPa, 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 \times 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 in this case). Because both C_g and P_g are known, this relation can be rearranged and used to solve for *k*.

$$\begin{aligned} C_g &= kP_g \\ k &= \frac{C_g}{P_g} \\ &= \frac{1.38 \times 10^{-3} molL^{-1}}{101.3kPa} \\ &= 1.36 \times 10^{-5} molL^{-1}kPa^{-1} \end{aligned}$$

Now, use k to find the solubility at the lower pressure.

$$\begin{aligned} C_g =& kP_g \\ & 1.36 \times 10^{-5} molL^{-1}kPa^{-1} \times 20.7kPa \\ =& 2.82 \times 10^{-4} molL^{-1} \end{aligned}$$

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 11.4.1

Exposing a 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 152 torr resulted in the dissolution of 1.45 \times 10^{−3} g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 760 torr.

Answer

7.25 \times 10^{−3} in 100.0 mL or 0.0725 g/L

Example 11.4.2: Thermal Pollution and Oxygen Solubility

A certain species of freshwater trout requires a dissolved oxygen concentration of 7.5 mg/L. Could these fish thrive in a thermally polluted mountain stream (water temperature is 30.0 °C, partial pressure of atmospheric oxygen is 0.17 atm)? Use the data in Figure 11.4.1to estimate a value for the Henry's law constant at this temperature.

Solution

First, estimate the Henry's law constant for oxygen in water at the specified temperature of 30.0 °C (Figure 11.4.1indicates the solubility at this temperature is approximately ~1.2 mol/L).

\[k=\frac{C_{ g }}{P_{ g }}=1.2 \times 10^{-3} mol / L / 1.00 atm =1.2 \times 10^{-3} mol / L atm]

Then, use this \(*k*\) value to compute the oxygen solubility at the specified oxygen partial pressure, 0.17 atm.

$$C_g = kP_g = \left(1.2 \times 10^{-3} mol/Latm\right)\left(0.17 atm\right) = 2.0 \times 10^{-4} mol/L$$

Finally, convert this dissolved oxygen concentration from mol/L to mg/L.

$$\left(2.0 \times 10^{-4} mol/L\right)\left(32.0 g/1 mol\right)\left(1000 mg/g\right) = 6.4 mg/L$$

This concentration is lesser than the required minimum value of 7.5 mg/L, and so these trout would likely not thrive in the polluted stream.

Exercise 11.4.2

What dissolved oxygen concentration is expected for the stream above when it returns to a normal summer time temperature of 15 °C?

Answer

8.2 mg/L

Chemistry in Everyday Life: Decompression Sickness or "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.

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 11.4.4). Researchers are also investigating related body reactions and defenses in order to develop better testing and treatment for decompression sickness. For example, Ingrid Eftedal, a barophysiologist specializing in bodily reactions to diving, has shown that white blood cells undergo chemical

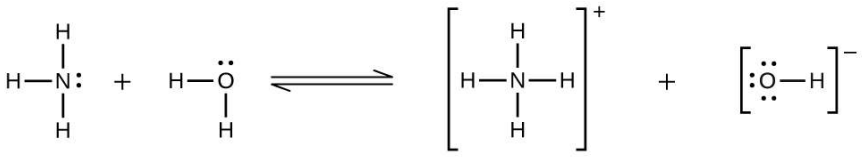
and genetic changes as a result of the condition; these can potentially be used to create biomarker tests and other methods to manage decompression sickness.



(a) (b)

Figure 11.4.4: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

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 increases more rapidly with increasing pressure than 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 11.4.5), 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 11.4.5: (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)

Solutions of Liquids in Liquids

Some liquids may be mixed in any proportions to yield solutions; in other words, they have infinite mutual solubility and are said to be **miscible**. Ethanol, sulfuric acid, and ethylene glycol (popular for use as antifreeze, pictured in Figure 11.4.6) are examples of

liquids that are completely miscible with water. Two-cycle motor oil is miscible with gasoline, mixtures of which are used as lubricating fuels for various types of outdoor power equipment (chainsaws, leaf blowers, and so on).



Figure 11.4.6: Water and antifreeze are miscible; mixtures of the two are homogeneous in all proportions. (credit: “dno1967”/Wikimedia commons)

Miscible liquids are typically those with very similar polarities. Consider, for example, liquids that are polar or capable of hydrogen bonding. For such liquids, the dipole-dipole attractions (or hydrogen bonding) of the solute molecules with the solvent molecules are at least as strong as those between molecules in the pure solute or in the pure solvent. Hence, the two kinds of molecules mix easily. Likewise, nonpolar liquids are miscible with each other because there is no appreciable difference in the strengths of solute-solute, solvent-solvent, and solute-solvent intermolecular attractions. The solubility of polar molecules in polar solvents and of nonpolar molecules in nonpolar solvents is, again, an illustration of the chemical axiom “like dissolves like.”

Two liquids that do not mix to an appreciable extent are called **immiscible**. Separate layers are formed when immiscible liquids are poured into the same container. Gasoline, oil (Figure 11.4.7), benzene, carbon tetrachloride, some paints, and many other nonpolar liquids are immiscible with water. Relatively weak attractive forces between the polar water molecules and the nonpolar liquid molecules are not adequate to overcome much stronger hydrogen bonding between water molecules. The distinction between immiscibility and miscibility is really one of extent, so that miscible liquids are of infinite mutual solubility, while liquids said to be immiscible are of very low (though not zero) mutual solubility.

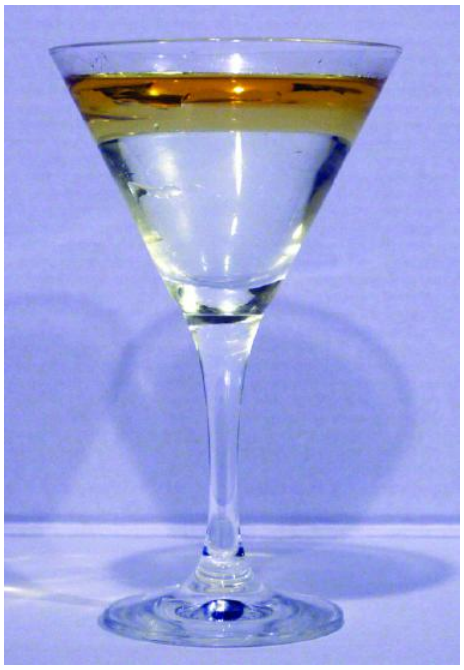


Figure 11.4.7: Water and oil are immiscible. Mixtures of these two substances will form two separate layers with the less dense oil floating on top of the water. (credit: “Yortw”/Flickr)

Two liquids, such as bromine and water, that are of *moderate* mutual solubility are said to be **partially miscible**. Two partially miscible liquids usually form two layers when mixed. In the case of the bromine and water mixture, the upper layer is water, saturated with bromine, and the lower layer is bromine saturated with water. Since bromine is nonpolar, and, thus, not very soluble in water, the water layer is only slightly discolored by the bright orange bromine dissolved in it. Since the solubility of water in bromine is very low, there is no noticeable effect on the dark color of the bromine layer (Figure 11.4.8).



Figure 11.4.8: Bromine (the deep orange liquid on the left) and water (the clear liquid in the middle) are partially miscible. The top layer in the mixture on the right is a saturated solution of bromine in water; the bottom layer is a saturated solution of water in bromine. (credit: Paul Flowers)

Solutions of Solids in Liquids

The dependence of solubility on temperature for a number of solids in water is shown by the solubility curves in Figure 11.4.9. Reviewing these data indicates a general trend of increasing solubility with temperature, although there are exceptions, as illustrated by the ionic compound cerium sulfate.

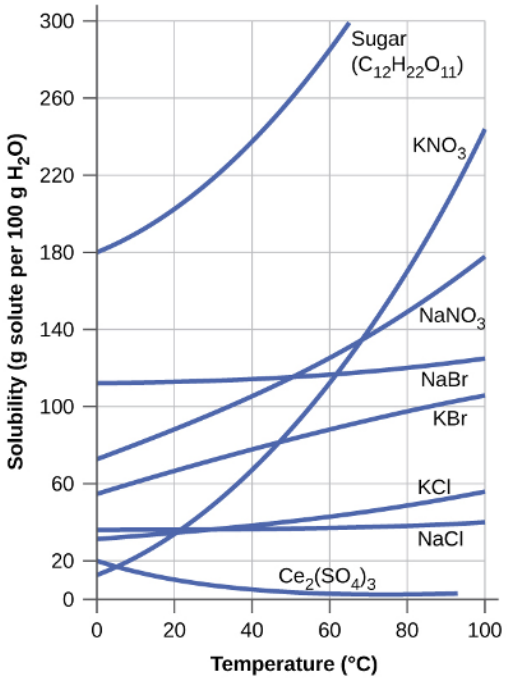


Figure 11.4.9: This graph shows how the solubility of several solids changes with temperature.

The temperature dependence of solubility can be exploited to prepare supersaturated solutions of certain compounds. A solution may be saturated with the compound at an elevated temperature (where the solute is more soluble) and subsequently cooled to a lower temperature without precipitating the solute. The resultant solution contains solute at a concentration greater than its equilibrium solubility at the lower temperature (i.e., it is supersaturated) and is relatively stable. Precipitation of the excess solute can be initiated by adding a seed crystal (see the video in the Link to Learning earlier in this module) or by mechanically agitating the solution. Some hand warmers, such as the one pictured in Figure 11.4.10, take advantage of this behavior.



Figure 11.4.10: This hand warmer produces heat when the sodium acetate in a supersaturated solution precipitates. Precipitation of the solute is initiated by a mechanical shockwave generated when the flexible metal disk within the solution is “clicked.” (credit: modification of work by “Velela”/Wikimedia Commons)

Link to Learning

This video shows the crystallization process occurring in a hand warmer.

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11.5: Colligative Properties

Learning Objectives

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

- Express concentrations of solution components using mole fraction and molality
- Describe the effect of solute concentration on various solution properties (vapor pressure, boiling point, freezing point, and osmotic pressure)
- Perform calculations using the mathematical equations that describe these various colligative effects
- Describe the process of distillation and its practical applications
- Explain the process of osmosis and describe how it is applied industrially and in nature

The properties of a solution are different from those of either the pure solute(s) or solvent. Many solution properties are dependent upon the chemical identity of the solute. Compared to pure water, a solution of hydrogen chloride is more acidic, a solution of ammonia is more basic, a solution of sodium chloride is more dense, and a solution of sucrose is more viscous. There are a few solution properties, however, that depend *only* upon the total concentration of solute species, regardless of their identities. These **colligative properties** include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure. This small set of properties is of central importance to many natural phenomena and technological applications, as will be described in this module.

Mole Fraction and Molality

Several units commonly used to express the concentrations of solution components were introduced in an earlier chapter of this text, each providing certain benefits for use in different applications. For example, molarity (*M*) is a convenient unit for use in stoichiometric calculations, since it is defined in terms of the molar amounts of solute species:

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

Because solution volumes vary with temperature, molar concentrations will likewise vary. When expressed as molarity, the concentration of a solution with identical numbers of solute and solvent species will be different at different temperatures, due to the contraction/expansion of the solution. More appropriate for calculations involving many colligative properties are mole-based concentration units whose values are not dependent on temperature. Two such units are *mole fraction* (introduced in the previous chapter on gases) and *molality*.

The mole fraction, X, of a component is the ratio of its molar amount to the total number of moles of all solution components:

$$X_A = \frac{molA}{\text{total mol of all components}}$$

By this definition, the sum of mole fractions for all solution components (the solvent and all solutes) is equal to one.

Molality is a concentration unit defined as the ratio of the numbers of moles of solute to the mass of the solvent in kilograms:

$$m = \frac{\text{mol solute}}{kg\,\text{solvent}}$$

Since these units are computed using only masses and molar amounts, they do not vary with temperature and, thus, are better suited for applications requiring temperature-independent concentrations, including several colligative properties, as will be described in this chapter module.

Example 11.5.1: Calculating Mole Fraction and Molality

The antifreeze in most automobile radiators is a mixture of equal volumes of ethylene glycol and water, with minor amounts of other additives that prevent corrosion. What are the (a) mole fraction and (b) molality of ethylene glycol, $C_2H_4(OH)_2$, in a solution prepared from 2.22 \times 10^{3} g of ethylene glycol and 2.00 \times 10^{3} g of water (approximately 2 L of glycol and 2 L of water)?

Solution

(a) The mole fraction of ethylene glycol may be computed by first deriving molar amounts of both solution components and then substituting these amounts into the definition of mole fraction.

$$molC_2H_4(OH)_2 = 2.22 \times 10^3 g \times \frac{1 molC_2H_4(OH)_2}{62.07 gC_2H_4(OH)_2} = 35.8 molC_2H_4(OH)_2$$

$$molH_2O = 2.00 \times 10^3 g \times \frac{1 molH_2O}{18.02 gH_2O} = 111 molH_2O$$

$$X_{\text{ethylene glycol}} = \frac{35.8 molC_2H_4(OH)_2}{(35.8 + 111)\text{ mol total}} = 0.244$$

Notice that mole fraction is a dimensionless property, being the ratio of properties with identical units (moles).

(b) Derive moles of solute and mass of solvent (in kg).

First, use the given mass of ethylene glycol and its molar mass to find the moles of solute:

$$2.22 \times 10^3 gC_2H_4(OH)_2 \left(\frac{molC_2H_2(OH)_2}{62.07 g}\right) = 35.8 molC_2H_4(OH)_2$$

Then, convert the mass of the water from grams to kilograms:

$$2.00 \times 10^3 gH_2O \left(\frac{1 kg}{1000 g}\right) = 2.00 kgH_2O$$

Finally, calculate molality per its definition:

$$\begin{aligned}\text{molality} &= \frac{\text{mol solute}}{\text{kg solvent}} \\ &= \frac{35.8 molC_2H_4(OH)_2}{2 kgH_2O} \\ &= 17.9 m\end{aligned}$$

Exercise 11.5.1

What are the mole fraction and molality of a solution that contains 0.850 g of ammonia, NH_3, dissolved in 125 g of water?

Answer

7.14 \times 10^{−3}; 0.399 *m*

Example 11.5.2: Converting Mole Fraction and Molal Concentrations

Calculate the mole fraction of solute and solvent in a 3.0 *m* solution of sodium chloride.

Solution

Converting from one concentration unit to another is accomplished by first comparing the two unit definitions. In this case, both units have the same numerator (moles of solute) but different denominators. The provided molal concentration may be written as:

$$\frac{3.0 molNaCl}{1 kgH_2O}$$

The numerator for this solution's mole fraction is, therefore, 3.0 mol NaCl. The denominator may be computed by deriving the molar amount of water corresponding to 1.0 kg

$$1.0 kgH_2O \left(\frac{1000 g}{1 kg}\right) \left(\frac{molH_2O}{18.02 g}\right) = 55 molH_2O$$

and then substituting these molar amounts into the definition for mole fraction.

$$\begin{aligned} X_{H_2O} &= \frac{molH_2O}{molNaCl + molH_2O} \\ &= \frac{55 molH_2O}{3.0 molNaCl + 55 molH_2O} \\ &= 0.95 \\ &= \frac{molNaCl}{molNaCl + molHH_2O} \\ &= \frac{3.0 molNaCl}{3.0 molNaCl + 55 molH_2O} \\ &= 0.052 \end{aligned}$$

Exercise 11.5.2

The mole fraction of iodine, I_2, dissolved in dichloromethane, CH_2Cl_2, is 0.115. What is the molal concentration, *m*, of iodine in this solution?

Answer

1.50 *m*

Example 11.5.3: Molality and Molarity Conversions

Intravenous infusion of a 0.556 *M* aqueous solution of glucose (density of 1.04 g/mL) is part of some post-operative recovery therapies. What is the molal concentration of glucose in this solution?

Solution

The provided molal concentration may be explicitly written as:

$$M = 0.556 mol \text{ glucose } / 1 L \text{ solution}$$

Consider the definition of molality:

$$m = \text{ mol solute } / kg \text{ solvent}$$

The amount of glucose in 1-L of this solution is 0.556 mol, so the mass of water in this volume of solution is needed.

First, compute the mass of 1.00 L of the solution:

$$(1.0 L \text{ soln })(1.04 g/mL)(1000 mL/1 L)(1 kg/1000 g) = 1.04 kg \text{ soln}$$

This is the mass of both the water and its solute, glucose, and so the mass of glucose must be subtracted. Compute the mass of glucose from its molar amount:

$$(0.556 mol \text{ glucose })(180.2 g/1 mol) = 100.2 g \text{ or } 0.1002 kg$$

Subtracting the mass of glucose yields the mass of water in the solution:

$$1.04 kg \text{ solution } - 0.1002 kg \text{ glucose } = 0.94 kg \text{ water}$$

Finally, the molality of glucose in this solution is computed as:

$$m = \frac{0.556 \text{ mol glucose}}{0.94 \text{ kg water}} = 0.59 \ m$$

Exercise 11.5.3

Nitric acid, $HNO_3(aq)$, is commercially available as a 33.7 m aqueous solution (density = 1.35 g/mL). What is the molarity of this solution?

Answer

14.6 M

Vapor Pressure Lowering

As described in the chapter on liquids and solids, the equilibrium vapor pressure of a liquid is the pressure exerted by its gaseous phase when vaporization and condensation are occurring at equal rates:

$$\text{liquid} \rightleftharpoons \text{gas}$$

Dissolving a nonvolatile substance in a volatile liquid results in a lowering of the liquid's vapor pressure. This phenomenon can be rationalized by considering the effect of added solute molecules on the liquid's vaporization and condensation processes. To vaporize, solvent molecules must be present at the surface of the solution. The presence of solute decreases the surface area available to solvent molecules and thereby reduces the rate of solvent vaporization. Since the rate of condensation is unaffected by the presence of solute, the net result is that the vaporization-condensation equilibrium is achieved with fewer solvent molecules in the vapor phase (i.e., at a lower vapor pressure) (Figure 11.5.1). While this interpretation is useful, it does not account for several important aspects of the colligative nature of vapor pressure lowering. A more rigorous explanation involves the property of entropy, a topic of discussion in a later text chapter on thermodynamics. For purposes of understanding the lowering of a liquid's vapor pressure, it is adequate to note that the more dispersed nature of matter in a solution, compared to separate solvent and solute phases, serves to effectively stabilize the solvent molecules and hinder their vaporization. A lower vapor pressure results, and a correspondingly higher boiling point as described in the next section of this module.

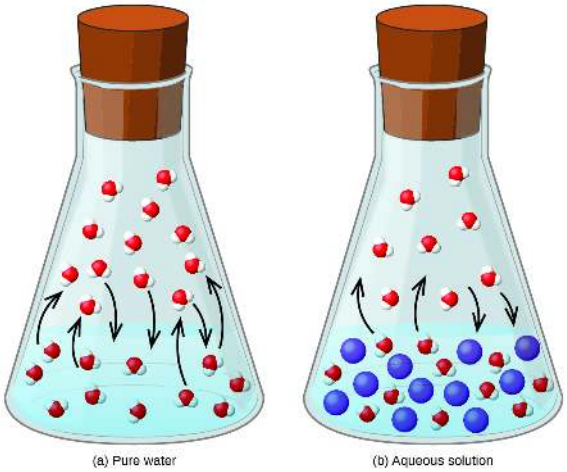


Figure 11.5.1: The presence of nonvolatile solutes lowers the vapor pressure of a solution by impeding the evaporation of solvent molecules.

The relationship between the vapor pressures of solution components and the concentrations of those components is described by **Raoult's law**: *The partial pressure exerted by any component of an ideal solution is equal to the vapor pressure of the pure component multiplied by its mole fraction in the solution.*

$$P_A = X_A P_A^* \tag{11.5.1}$$

where P_A is the partial pressure exerted by component A in the solution, P_A^* is the vapor pressure of pure A, and X_A is the mole fraction of A in the solution.

Recalling that the total pressure of a gaseous mixture is equal to the sum of partial pressures for all its components (Dalton's law of partial pressures), the total vapor pressure exerted by a solution containing *i* components is

$$P_{\text{solution}} = \sum_i P_i = \sum_i X_i P_i^*$$

A nonvolatile substance is one whose vapor pressure is negligible ($P^* \approx 0$), and so the vapor pressure above a solution containing only nonvolatile solutes is due only to the solvent:

$$P_{\text{solution}} = X_{\text{solvent}}\, P^*_{\text{solvent}}$$

Example 11.5.4: Calculation of a Vapor Pressure

Compute the vapor pressure of an ideal solution containing 92.1 g of glycerin, $C_3H_5(OH)_3$, and 184.4 g of ethanol, C_2H_5OH, at 40 °C. The vapor pressure of pure ethanol is 0.178 atm at 40 °C. Glycerin is essentially nonvolatile at this temperature.

Solution

Since the solvent is the only volatile component of this solution, its vapor pressure may be computed per Raoult's law (Equation 11.5.1) as:

$$P_{\text{solution}} = X_{\text{solvent}}\, P^*_{\text{solvent}}$$

First, calculate the molar amounts of each solution component using the provided mass data.

$$92.1 g_3 H_5(OH)_3 \times \frac{1 mol C_3H_5(OH)_3}{92.094 g_5 C_3H_5(OH)_3} = 1.00 mol C_3H_5(OH)_3$$

$$184.4 g_2 C_2H_5OH \times \frac{1 mol C_2H_5OH}{46.069 g_2 CH_5OH} = 4.000 mol C_2H_5OH$$

Next, calculate the mole fraction of the solvent (ethanol) and use Raoult's law to compute the solution's vapor pressure.

$$X_{C_2H_5OH} = \frac{4.000 mol}{(1.00 mol + 4.000 mol)} = 0.800$$

$$P_{\text{solv}} = X_{\text{solv}}\, P^*_{\text{solv}} = 0.800 \times 0.178 atm = 0.142 atm$$

Exercise 11.5.4

A solution contains 5.00 g of urea, $CO(NH_2)_2$ (a nonvolatile solute) and 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 assuming ideal behavior?

Answer

23.4 torr

Distillation of Solutions

Solutions whose components have significantly different vapor pressures may be separated by a selective vaporization process known as distillation. Consider the simple case of a mixture of two volatile liquids, A and B, with A being the more volatile liquid. Raoult's law can be used to show that the vapor above the solution is enriched in component A, that is, the mole fraction of A in the vapor is greater than the mole fraction of A in the liquid (see end-of-chapter Exercise 65). By appropriately heating the mixture, component A may be vaporized, condensed, and collected—effectively separating it from component B.

Distillation is widely applied in both laboratory and industrial settings, being used to refine petroleum, to isolate fermentation products, and to purify water. A typical apparatus for laboratory-scale distillations is shown in Figure 11.5.2

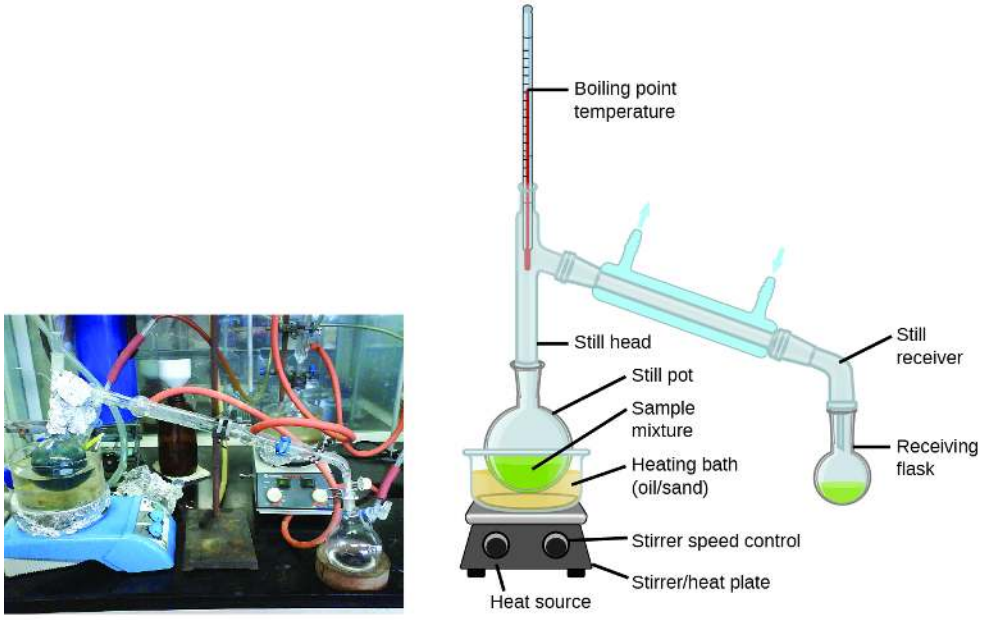


Figure 11.5.2: A typical laboratory distillation unit is shown in (a) a photograph and (b) a schematic diagram of the components. (credit a: modification of work by “Rifleman82”/Wikimedia commons; credit b: modification of work by “Slashme”/Wikimedia Commons)

Oil refineries use large-scale *fractional distillation* to separate the components of crude oil. The crude oil is heated to high temperatures at the base of a tall *fractionating column*, vaporizing many of the components that rise within the column. As vaporized components reach adequately cool zones during their ascent, they condense and are collected. The collected liquids are simpler mixtures of hydrocarbons and other petroleum compounds that are of appropriate composition for various applications (e.g., diesel fuel, kerosene, gasoline), as depicted in Figure 11.5.3

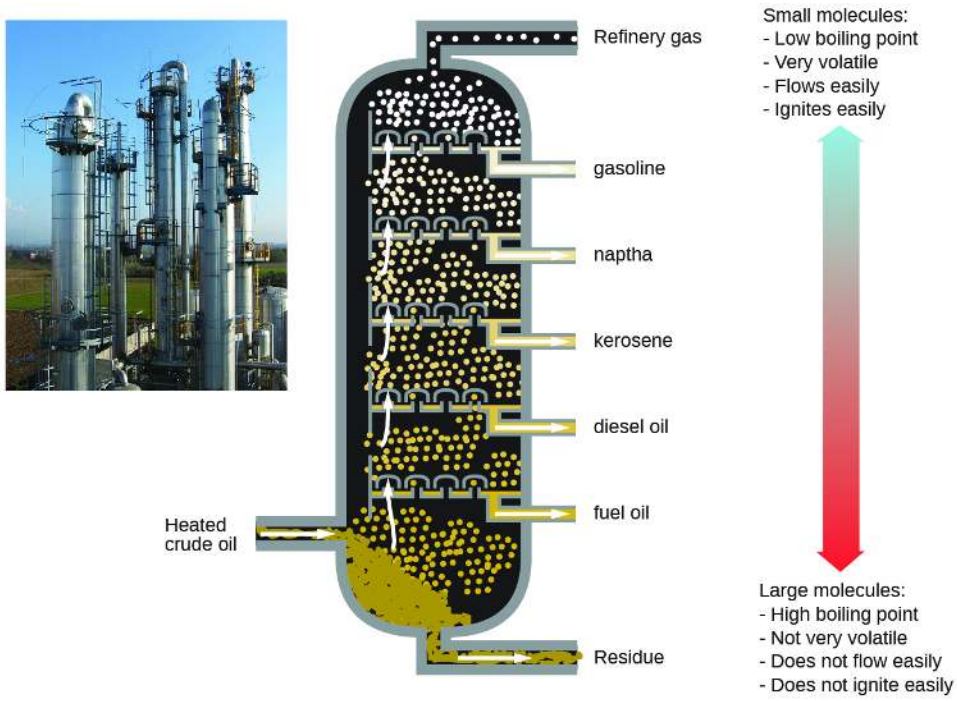


Figure 11.5.3: Crude oil is a complex mixture that is separated by large-scale fractional distillation to isolate various simpler mixtures.

Boiling Point Elevation

As described in the chapter on liquids and solids, the *boiling point* of a liquid is the temperature at which its vapor pressure is equal to ambient atmospheric pressure. Since the vapor pressure of a solution is lowered due to the presence of nonvolatile solutes, it stands to reason that the solution's boiling point will subsequently be increased. Vapor pressure increases with temperature, and so a solution will require a higher temperature than will pure solvent to achieve any given vapor pressure, including one equivalent to that of the surrounding atmosphere. The increase in boiling point observed when nonvolatile solute is dissolved in a solvent, ΔT_b, is called **boiling point elevation** and is directly proportional to the molal concentration of solute species:

$$\Delta T_b = K_b m$$

where K_b is the boiling point elevation constant, or the *ebullioscopic constant* and m is the molal concentration (molality) of all solute species.

Boiling point elevation constants are characteristic properties that depend on the identity of the solvent. Values of K_b for several solvents are listed in Table 11.5.1.

Table 11.5.1: Boiling Point Elevation and Freezing Point Depression Constants for Several Solvents

Solvent	Boiling Point (°C at 1 atm)	K_b (°Cm^{-1})	Freezing Point (°C at 1 atm)	K_f (°Cm^{-1})
water	100.0	0.512	0.0	1.86
hydrogen acetate	118.1	3.07	16.6	3.9
benzene	80.1	2.53	5.5	5.12
chloroform	61.26	3.63	−63.5	4.68
nitrobenzene	210.9	5.24	5.67	8.1

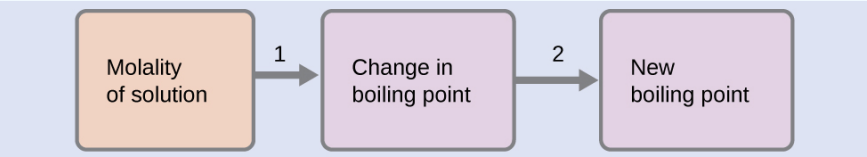
The extent to which the vapor pressure of a solvent is lowered and the boiling point is elevated depends on the total number of solute particles present in a given amount of solvent, not on the mass or size or chemical identities of the particles. A 1 *m* aqueous solution of sucrose (342 g/mol) and a 1 *m* aqueous solution of ethylene glycol (62 g/mol) will exhibit the same boiling point because each solution has one mole of solute particles (molecules) per kilogram of solvent.

Example 11.5.5: Calculating the Boiling Point of a Solution

Assuming ideal solution behavior, what is the boiling point of a 0.33 *m* solution of a nonvolatile solute in benzene?

Solution

Use the equation relating boiling point elevation to solute molality to solve this problem in two steps.



Step 1: Calculate the change in boiling point.

$$\Delta T_b = K_b m = 2.53^\circ C m^{-1} \times 0.33 m = 0.83^\circ C$$

Step 2: Add the boiling point elevation to the pure solvent's boiling point.

$$\text{Boiling temperature } = 80.1^\circ C + 0.83^\circ C = 80.9^\circ C$$

Exercise 11.5.1

Assuming ideal solution behavior, what is the boiling point of the antifreeze described in Example 11.5.1?

Answer

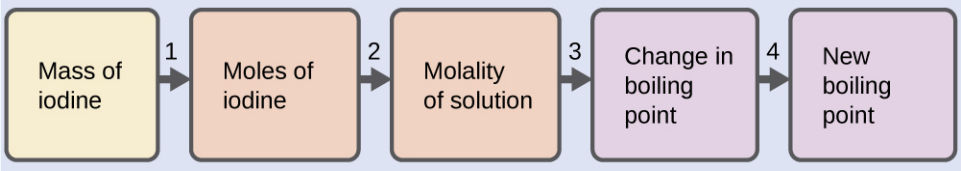
109.2 °C

Example 11.5.6: The Boiling Point of an Iodine Solution

Find the boiling point of a solution of 92.1 g of iodine, I_2, in 800.0 g of chloroform, $CHCl_3$, assuming that the iodine is nonvolatile and that the solution is ideal.

Solution

A four-step approach to solving this problem is outlined below.



Step 1. *Convert from grams to moles of* I_2 *using the molar mass of* I_2 *in the unit conversion factor.*

Result: 0.363 mol

Step 2. *Determine the molality of the solution from the number of moles of solute and the mass of solvent, in kilograms.*

Result: 0.454 *m*

Step 3. *Use the direct proportionality between the change in boiling point and molal concentration to determine how much the boiling point changes.*

Result: 1.65 °C

Step 4. *Determine the new boiling point from the boiling point of the pure solvent and the change.*

Result: 62.91 °C

Check each result as a self-assessment.

Exercise 11.5.5

What is the boiling point of a solution of 1.0 g of glycerin, $C_3H_5(OH)_3$, in 47.8 g of water? Assume an ideal solution.

Answer

100.12 °C

Freezing Point Depression

Solutions freeze at lower temperatures than pure liquids. This phenomenon is exploited in “de-icing” schemes that use salt (Figure 11.5.4), 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 11.5.4: 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)

The decrease in freezing point of a dilute solution compared to that of the pure solvent, ΔT_f, is called the **freezing point depression** and is directly proportional to the molal concentration of the solute

$$\Delta T_f = K_f m$$

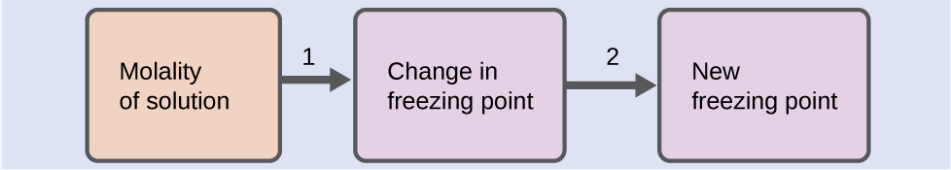
where *m* is the molal concentration of the solute and K_f is called the freezing point depression constant (or *cryoscopic constant*). Just as for boiling point elevation constants, these are characteristic properties whose values depend on the chemical identity of the solvent. Values of K_f for several solvents are listed in Table 11.5.1.

Example 11.5.7: Calculation of the Freezing Point of a Solution

Assuming ideal solution behavior, what is the freezing point of the 0.33 *m* solution of a nonvolatile nonelectrolyte solute in benzene described in Example 11.5.2?

Solution

Use the equation relating freezing point depression to solute molality to solve this problem in two steps.



Step 1: Calculate the change in freezing point.

$$\Delta T_f = K_f m = 5.12^\circ C m^{-1} \times 0.33 m = 1.7^\circ C$$

Step 2: Subtract the freezing point change observed from the pure solvent's freezing point.

$$\text{Freezing Temperature } = 5.5^\circ C - 1.7^\circ C = 3.8^\circ C$$

Exercise 11.5.1

Assuming ideal solution behavior, what is the freezing point of a 1.85 *m* solution of a nonvolatile nonelectrolyte solute in nitrobenzene?

Answer

−9.3 °C

Chemistry in Everyday Life: Colligative Properties and De-Icing

Sodium chloride and its group 2 analogs calcium and magnesium chloride are often used to de-ice roadways and sidewalks, due to the fact that a solution of any one of these salts will have a freezing point lower than 0 °C, the freezing point of pure water. The group 2 metal salts are frequently mixed with the cheaper and more readily available sodium chloride ("rock salt") for use on roads, since they tend to be somewhat less corrosive than the NaCl, and they provide a larger depression of the freezing point, since they dissociate to yield three particles per formula unit, rather than two particles like the sodium chloride.

Because these ionic compounds tend to hasten the corrosion of metal, they would not be a wise choice to use in antifreeze for the radiator in your car or to de-ice a plane prior to takeoff. For these applications, covalent compounds, such as ethylene or propylene glycol, are often used. The glycols used in radiator fluid not only lower the freezing point of the liquid, but they elevate the boiling point, making the fluid useful in both winter and summer. 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 (Figure 11.5.5).



Figure 11.5.5: Freezing point depression is exploited to remove ice from (a) roadways and (b) the control surfaces of aircraft.

Phase Diagram for a Solution

The colligative effects on vapor pressure, boiling point, and freezing point described in the previous section are conveniently summarized by comparing the phase diagrams for a pure liquid and a solution derived from that liquid (Figure 11.5.6).

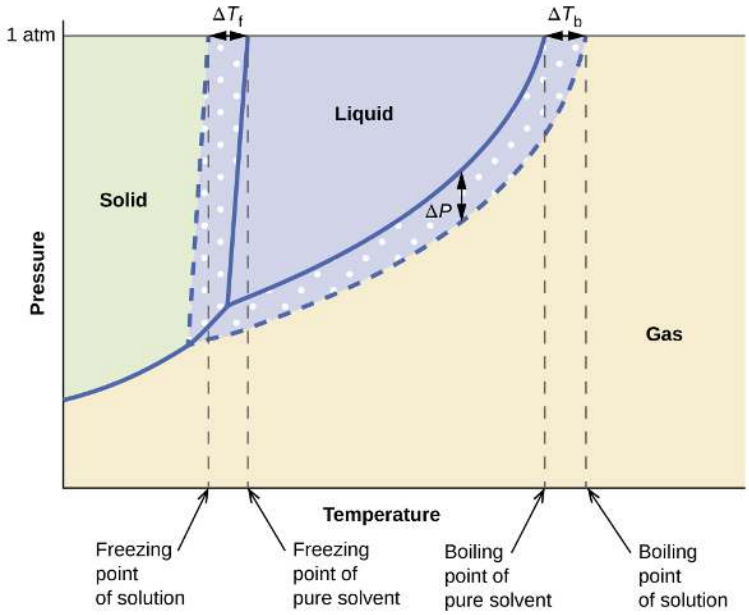


Figure 11.5.6: Phase diagrams for a pure solvent (solid curves) and a solution formed by dissolving nonvolatile solute in the solvent (dashed curves).

The liquid-vapor curve for the solution is located *beneath* the corresponding curve for the solvent, depicting the vapor pressure *lowering*, ΔP, that results from the dissolution of nonvolatile solute. Consequently, at any given pressure, the solution's boiling point is observed at a higher temperature than that for the pure solvent, reflecting the boiling point elevation, ΔT_b, associated with

the presence of nonvolatile solute. The solid-liquid curve for the solution is displaced left of that for the pure solvent, representing the freezing point depression, ΔT_f, that accompanies solution formation. Finally, notice that the solid-gas curves for the solvent and its solution are identical. This is the case for many solutions comprising liquid solvents and nonvolatile solutes. Just as for vaporization, when a solution of this sort is frozen, it is actually just the *solvent* molecules that undergo the liquid-to-solid transition, forming pure solid solvent that excludes solute species. The solid and gaseous phases, therefore, are composed of solvent only, and so transitions between these phases are not subject to colligative effects.

Osmosis and Osmotic Pressure of Solutions

A number of natural and synthetic materials exhibit *selective permeation*, meaning that only molecules or ions of a certain size, shape, polarity, charge, and so forth, are capable of passing through (permeating) the material. Biological cell membranes provide elegant examples of selective permeation in nature, while dialysis tubing used to remove metabolic wastes from blood is a more simplistic technological example. Regardless of how they may be fabricated, these materials are generally referred to as **semipermeable membranes**.

Consider the apparatus illustrated in Figure 11.5.7, in which samples of pure solvent and a solution are separated by a membrane that only solvent molecules may permeate. Solvent molecules will diffuse across the membrane in both directions. Since the concentration of *solvent* is greater in the pure solvent than the solution, these molecules will diffuse from the solvent side of the membrane to the solution side at a faster rate than they will in the reverse direction. The result is a net transfer of solvent molecules from the pure solvent to the solution. Diffusion-driven transfer of solvent molecules through a semipermeable membrane is a process known as **osmosis**.

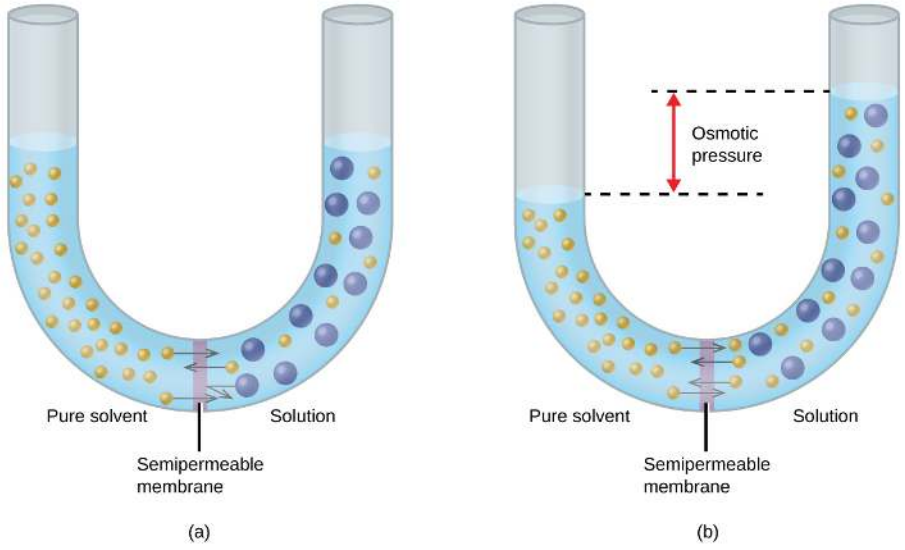


Figure 11.5.7: (a) A solution and pure solvent are initially separated by an osmotic membrane. (b) Net transfer of solvent molecules to the solution occurs until its osmotic pressure yields equal rates of transfer in both directions.

When osmosis is carried out in an apparatus like that shown in Figure 11.5.7, the volume of the solution increases as it becomes diluted by accumulation of solvent. This causes the level of the solution to rise, increasing its hydrostatic pressure (due to the weight of the column of solution in the tube) and resulting in a faster transfer of solvent molecules back to the pure solvent side. When the pressure reaches a value that yields a reverse solvent transfer rate equal to the osmosis rate, bulk transfer of solvent ceases. This pressure is called the **osmotic pressure** (Π) of the solution. The osmotic pressure of a dilute solution is related to its solute molarity, *M*, and absolute temperature, *T*, according to the equation

$$\Pi = MRT \tag{11.5.2}$$

where R is the universal gas constant.

Example 11.5.8: Calculation of Osmotic Pressure

Assuming ideal solution behavior, what is the osmotic pressure (atm) of a 0.30 M solution of glucose in water that is used for intravenous infusion at body temperature, 37 °C?

Solution

Find the osmotic pressure, Π, using Equation 11.5.2 (\ce{Π = MRT}), where T is on the Kelvin scale (310 K) and the value of R is expressed in appropriate units (0.08206 L atm/mol K).

$$\begin{aligned}\Pi &= MRT \\ &= 0.30 mol/L \times 0.08206 Latm/molK \times 310 K \\ &= 7.6 atm\end{aligned}$$

Exercise 11.5.8

Assuming ideal solution behavior, what is the osmotic pressure (atm) a solution with a volume of 0.750 L that contains 5.0 g of methanol, CH_3OH, in water at 37 °C?

Answer

5.3 atm

If a solution is placed in an apparatus like the one shown in Figure 11.5.8, applying pressure greater than the osmotic pressure of the solution reverses the osmosis and pushes solvent molecules from the solution into the pure solvent. This technique of reverse osmosis is used for large-scale desalination of seawater and on smaller scales to produce high-purity tap water for drinking.

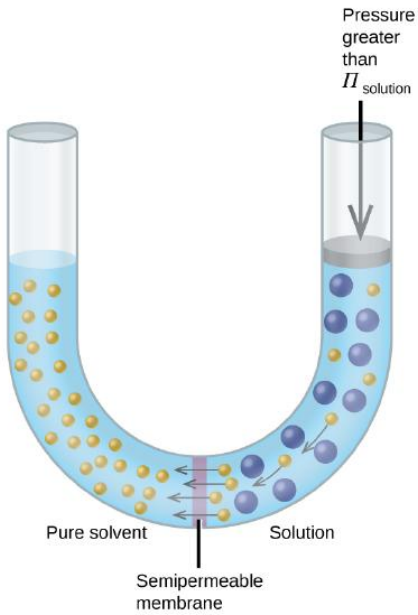


Figure 11.5.8: Applying a pressure greater than the osmotic pressure of a solution will reverse osmosis. Solvent molecules from the solution are pushed into the pure solvent.

Chemistry in Everyday Life: Reverse Osmosis Water Purification

In the process of osmosis, diffusion serves to move water through a semipermeable membrane from a less concentrated solution to a more concentrated solution. Osmotic pressure is the amount of pressure that must be applied to the more concentrated solution to cause osmosis to stop. If greater pressure is applied, the water will go from the more concentrated solution to a less concentrated (more pure) solution. This is called reverse osmosis. Reverse osmosis (RO) is used to purify water in many applications, from desalination plants in coastal cities, to water-purifying machines in grocery stores (Figure 11.5.9), and smaller reverse-osmosis household units. With a hand-operated pump, small RO units can be used in third-world

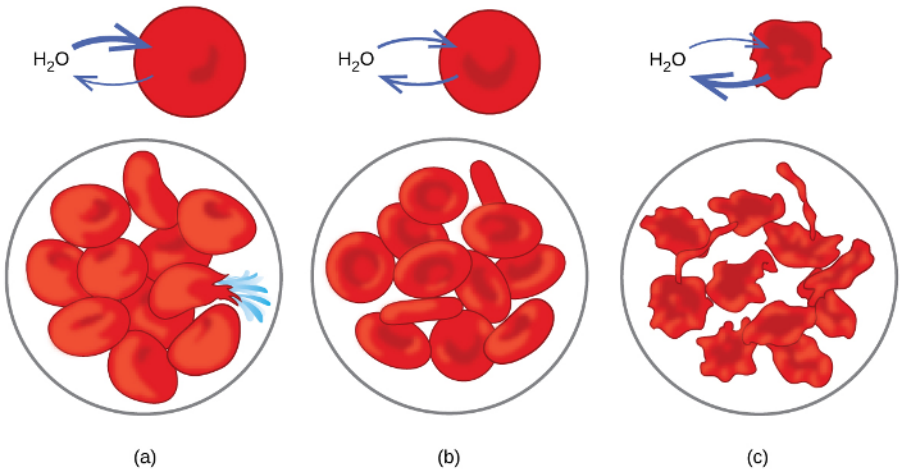
countries, disaster areas, and in lifeboats. Our military forces have a variety of generator-operated RO units that can be transported in vehicles to remote locations.



(a) (b)

Figure 11.5.9: Reverse osmosis systems for purifying drinking water are shown here on (a) small and (b) large scales. (credit a: modification of work by Jerry Kirkhart; credit b: modification of work by Willard J. Lathrop)

Examples of osmosis are evident in many biological systems because cells are surrounded by semipermeable membranes. Carrots and celery that have become limp because they have lost water can be made crisp again by placing them in water. Water moves into the carrot or celery cells by osmosis. A cucumber placed in a concentrated salt solution loses water by osmosis and absorbs some salt to become a pickle. Osmosis can also affect animal cells. Solute concentrations are particularly important when solutions are injected into the body. Solutes in body cell fluids and blood serum give these solutions an osmotic pressure of approximately 7.7 atm. Solutions injected into the body must have the same osmotic pressure as blood serum; that is, they should be **isotonic** with blood serum. If a less concentrated solution, a **hypotonic** solution, is injected in sufficient quantity to dilute the blood serum, water from the diluted serum passes into the blood cells by osmosis, causing the cells to expand and rupture. This process is called **hemolysis**. When a more concentrated solution, a hypertonic solution, is injected, the cells lose water to the more concentrated solution, shrivel, and possibly die in a process called **crenation**. These effects are illustrated in Figure 11.5.9.



(a) (b) (c)

Figure 11.5.9 Red blood cell membranes are water permeable and will (a) swell and possibly rupture in a hypotonic solution; (b) maintain normal volume and shape in an isotonic solution; and (c) shrivel and possibly die in a hypertonic solution. (credit a/b/c: modifications of work by “LadyofHats”/Wikimedia commons)

Determination of Molar Masses

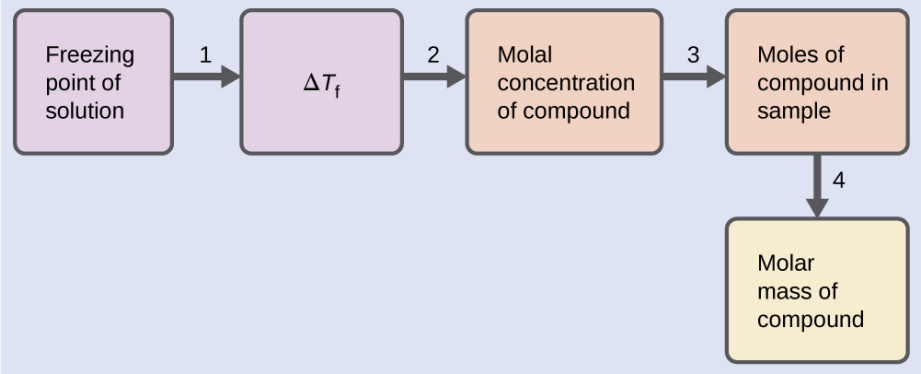
Osmotic pressure and changes in freezing point, boiling point, and vapor pressure are directly proportional to the number of solute species present in a given amount of solution. Consequently, measuring one of these properties for a solution prepared using a known mass of solute permits determination of the solute’s molar mass.

Example 11.5.9: Determination of a Molar Mass from a 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. Assuming ideal solution behavior, what is the molar mass of this compound?

Solution

Solve this problem using the following steps.



Step 1. Determine the change in freezing point from the observed freezing point and the freezing point of pure benzene (Table 11.5.1.

$$\Delta T_f = 5.5^\circ C - 2.32^\circ C = 3.2^\circ C$$

Step 2. Determine the molal concentration from Kf, the freezing point depression constant for benzene (Table 11.5.1), and ΔT_f.

$$\Delta T_f = K_f m$$

$$m = \frac{\Delta T_f}{K_f} = \frac{3.2^\circ C}{5.12^\circ C m^{-1}} = 0.63 m$$

Step 3. Determine the number of moles of compound in the solution from the molal concentration and the mass of solvent used to make the solution.

$$\text{Moles of solute } = \frac{0.63 \text{ mol solute}}{1.00 \text{ kg solvent}} \times 0.0550 \text{ kg solvent} = 0.035 mol$$

Step 4. Determine the molar mass from the mass of the solute and the number of moles in that mass.

$$\text{Molar mass } = \frac{4.00 \text{ g}}{0.035 \text{ mol}} = 1.1 \times 10^2 \text{ g / mol}$$

Exercise 11.5.9

A solution of 35.7 g of a nonelectrolyte in 220.0 g of chloroform has a boiling point of 64.5 °C. Assuming ideal solution behavior, what is the molar mass of this compound?

Answer

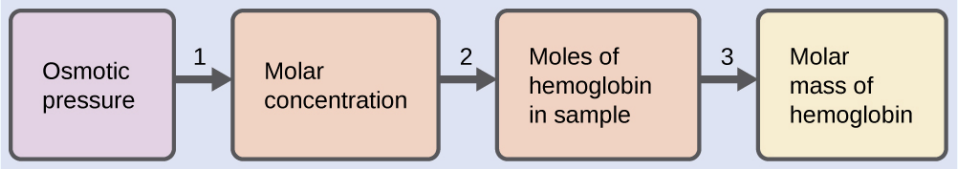
1.8 \times 10^{2} g/mol

Example 11.5.10: 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. Assuming ideal solution behavior, what is the molar mass of hemoglobin?

Solution

Here is one set of steps that can be used to solve the problem:



Step 1. Convert the osmotic pressure to atmospheres, then determine the molar concentration from the osmotic pressure.

$$\Pi = \frac{5.9 \text{ torr } \times 1atm}{760 \text{ torr}} = 7.8 \times 10^{-3} atm$$

$$\Pi = MRT$$

$$M = \frac{\Pi}{RT} = \frac{7.8 \times 10^{-3} atm}{(0.08206 Latm/molK)(295K)} = 3.2 \times 10^{-4} M$$

Step 2. Determine the number of moles of hemoglobin in the solution from the concentration and the volume of the solution.

$$\text{moles of hemoglobin } = \frac{3.2 \times 10^{-4} mol}{1 \text{ L solution}} \times 0.500 \text{ L solution} = 1.6 \times 10^{-4} \text{ mol}$$

Step 3. 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 11.5.10

Assuming ideal solution behavior, 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

3 \times 10^{4} g/mol

Colligative Properties of Electrolytes

As noted previously in this module, the colligative properties of a solution depend only on the number, not on the identity, of solute species dissolved. The concentration terms in the equations for various colligative properties (freezing point depression, boiling point elevation, osmotic pressure) pertain to *all solute species present in the solution*. For the solutions considered thus far in this chapter, the solutes have been nonelectrolytes that dissolve physically without dissociation or any other accompanying process. Each molecule that dissolves yields one dissolved solute molecule. The dissolution of an electroyte, however, is not this simple, as illustrated by the two common examples below:

$$\text{dissociation} \quad NaCl(s) \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

$$\text{ionization} \quad HCl(aq) + H_2O(l) \longrightarrow Cl^{-}(aq) + H3O^{+}(aq)$$

Considering the first of these examples, and assuming complete dissociation, a 1.0 *m* aqueous solution of NaCl contains 2.0 mole of ions (1.0 mol Na^{+} and 1.0 mol Cl^{-}) per each kilogram of water, and its freezing point depression is expected to be

$$\Delta T_f = 2.0 mol \text{ ions } / kg \text{ water } \times 1.86^{\circ} C kg \text{ water } / mol \text{ ion } = 3.7^{\circ} C.$$

When this solution is actually prepared and its freezing point depression measured, however, a value of 3.4 °C is obtained. Similar discrepancies are observed for other ionic compounds, and the differences between the measured and expected colligative property values typically become more significant as solute concentrations increase. These observations suggest that the ions of sodium chloride (and other strong electrolytes) are not completely dissociated in solution.

To account for this and avoid the errors accompanying the assumption of total dissociation, an experimentally measured parameter named in honor of Nobel Prize-winning German chemist Jacobus Henricus van 't Hoff is used. The van 't Hoff factor (i) is defined as the ratio of solute particles in solution to the number of formula units dissolved:

$$i = \frac{\text{moles of particles in solution}}{\text{moles of formula units dissolved}}$$

Values for measured van 't Hoff factors for several solutes, along with predicted values assuming complete dissociation, are shown in Table 11.5.2

Table 11.5.2: Predicated and Measured van't Hoff Factors for Several 0.050 *m* Aqueous Solutions

Formula unit	Classification	Dissolution products	*i* (predicted)	*i* (measured)
$C_{12}H_{22}O_{11}$ (glucose)	Nonelectrolyte	$C_{12}H_{22}O_{11}$	1	1.0
NaCl	Strong electrolyte	Na^+, Cl^-	2	1.9
HCl	Strong electrolyte (acid)	H_3O^+, Cl^-	2	1.9
$MgSO_4$	Strong electrolyte	Mg^{2+}, SO_4^{2-},	2	1.3
$MgCl_2$	Strong electrolyte	Mg^{2+}, $2Cl^-$	3	2.7
$FeCl_3$	Strong electrolyte	Fe^{3+}, $3Cl^-$	4	3.4

In 1923, the chemists Peter Debye and Erich Hückel proposed a theory to explain the apparent incomplete ionization of strong electrolytes. They suggested that although interionic attraction in an aqueous solution is very greatly reduced by solvation of the ions and the insulating action of the polar solvent, it is not completely nullified. The residual attractions prevent the ions from behaving as totally independent particles (Figure 11.5.11). In some cases, a positive and negative ion may actually touch, giving a solvated unit called an ion pair. Thus, the activity, or the effective concentration, of any particular kind of ion is less than that indicated by the actual concentration. Ions become more and more widely separated the more dilute the solution, and the residual interionic attractions become less and less. Thus, in extremely dilute solutions, the effective concentrations of the ions (their activities) are essentially equal to the actual concentrations. Note that the van't Hoff factors for the electrolytes in Table 11.5.2 are for 0.05 *m* solutions, at which concentration the value of *i* for NaCl is 1.9, as opposed to an ideal value of 2.

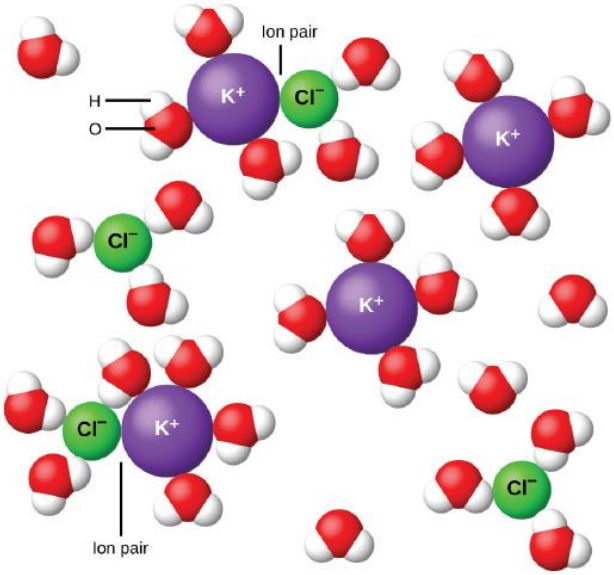


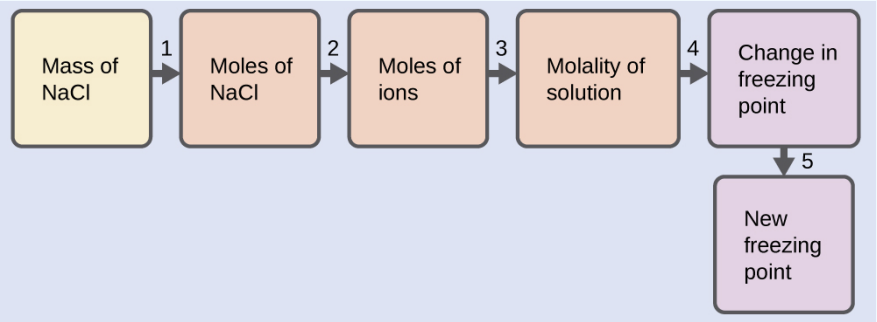
Figure 11.5.11: Dissociation of ionic compounds in water is not always complete due to the formation of ion pairs.

Example 11.5.11: The Freezing Point of a Solution of an Electrolyte

The concentration of ions in seawater is approximately the same as that in a solution containing 4.2 g of NaCl dissolved in 125 g of water. Use this information and a predicted value for the van't Hoff factor (Table 11.5.2) to determine the freezing temperature the solution (assume ideal solution behavior).

Solution

Solve this problem using the following series of steps.



1. Step 1. *Convert from grams to moles of NaCl using the molar mass of NaCl in the unit conversion factor.*
   Result: 0.072 mol NaCl
2. Step 2. *Determine the number of moles of ions present in the solution using the number of moles of ions in 1 mole of NaCl as the conversion factor (2 mol ions/1 mol NaCl).*
   Result: 0.14 mol ions
3. Step 3. *Determine the molality of the ions in the solution from the number of moles of ions and the mass of solvent, in kilograms.*
   Result: 1.2 *m*
4. Step 4. *Use the direct proportionality between the change in freezing point and molal concentration to determine how much the freezing point changes.*
   Result: 2.1 °C
5. Step 5. *Determine the new freezing point from the freezing point of the pure solvent and the change.*
   Result: −2.1 °C

Check each result as a self-assessment, taking care to avoid rounding errors by retaining guard digits in each step's result for computing the next step's result.

Exercise 11.5.11

Assuming complete dissociation and ideal solution behavior, calculate the freezing point of a solution of 0.724 g of $CaCl_2$ in 175 g of water.

Answer

−0.208 °C

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11.6: Colloids

Learning Objectives

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

- Describe the composition and properties of colloidal dispersions
- List and explain several technological applications of colloids

As a child, you may have made suspensions such as mixtures of mud and water, flour and water, or a suspension of solid pigments in water, known as tempera paint. These **suspensions** are heterogeneous mixtures composed of relatively large particles that are visible (or that can be seen with a magnifying glass). They are cloudy, and the suspended particles settle out after mixing. On the other hand, a solution is a homogeneous mixture in which no settling occurs and in which the dissolved species are molecules or ions. Solutions exhibit completely different behavior from suspensions. A solution may be colored, but it is transparent, the molecules or ions are invisible, and they do not settle out on standing. Another class of mixtures called **colloids** (or **colloidal dispersions**) exhibit properties intermediate between those of suspensions and solutions (Figure 11.6.1). The particles in a colloid are larger than most simple molecules; however, colloidal particles are small enough that they do not settle out upon standing.

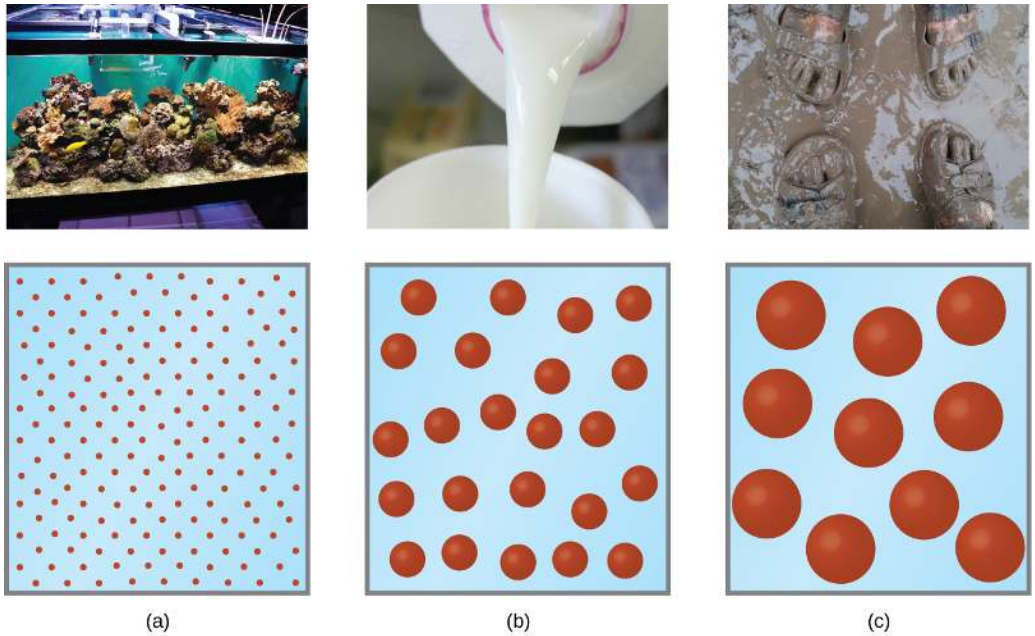


Figure 11.6.1: (a) A solution is a homogeneous mixture that appears clear, such as the saltwater in this aquarium. (b) In a colloid, such as milk, the particles are much larger but remain dispersed and do not settle. (c) A suspension, such as mud, is a heterogeneous mixture of suspended particles that appears cloudy and in which the particles can settle. (credit a photo: modification of work by Adam Wimsatt; credit b photo: modification of work by Melissa Wiese; credit c photo: modification of work by Peter Burgess)

The particles in a colloid are large enough to scatter light, a phenomenon called the Tyndall effect. This can make colloidal mixtures appear cloudy or opaque, such as the searchlight beams shown in Figure 11.6.2 Clouds are colloidal mixtures. They are composed of water droplets that are much larger than molecules, but that are small enough that they do not settle out.



Figure 11.6.2: The paths of searchlight beams are made visible when light is scattered by colloidal-size particles in the air (fog, smoke, etc.). (credit: “Bahman”/Wikimedia Commons)

The term “colloid”—from the Greek words *kolla*, meaning “glue,” and *eidos*, meaning “like”—was first used in 1861 by Thomas Graham to classify mixtures such as starch in water and gelatin. Many colloidal particles are aggregates of hundreds or thousands of molecules, but others (such as proteins and polymer molecules) consist of a single extremely large molecule. The protein and synthetic polymer molecules that form colloids may have molecular masses ranging from a few thousand to many million atomic mass units.

Analogous to the identification of solution components as “solute” and “solvent,” the components of a colloid are likewise classified according to their relative amounts. The particulate component typically present in a relatively minor amount is called the dispersed phase and the substance or solution throughout which the particulate is dispersed is called the dispersion medium. Colloids may involve virtually any combination of physical states (gas in liquid, liquid in solid, solid in gas, etc.), as illustrated by the examples of colloidal systems given in Table 11.6.1.

Table 11.6.1: Examples of Colloidal Systems

Dispersed Phase	Dispersion Medium	Common Examples	Name
solid	gas	smoke, dust	—
solid	liquid	starch in water, some inks, paints, milk of magnesia	sol
solid	solid	some colored gems, some alloys	—
liquid	gas	clouds, fogs, mists, sprays	aerosol
liquid	liquid	milk, mayonnaise, butter	emulsion
liquid	solid	jellies, gels, pearl, opal (H_2O in SiO_2)	gel
gas	liquid	foams, whipped cream, beaten egg whites	foam
gas	solid	pumice, floating soaps	—

Preparation of Colloidal Systems

Colloids are prepared by producing particles of colloidal dimensions and distributing these particles throughout a dispersion medium. Particles of colloidal size are formed by two methods:

1. Dispersion methods: breaking down larger particles. For example, paint pigments are produced by dispersing large particles by grinding in special mills.
2. Condensation methods: growth from smaller units, such as molecules or ions. For example, clouds form when water molecules condense and form very small droplets.

A few solid substances, when brought into contact with water, disperse spontaneously and form colloidal systems. Gelatin, glue, starch, and dehydrated milk powder behave in this manner. The particles are already of colloidal size; the water simply disperses them. Powdered milk particles of colloidal size are produced by dehydrating milk spray. Some atomizers produce colloidal dispersions of a liquid in air.

An emulsion may be prepared by shaking together or blending two immiscible liquids. This breaks one liquid into droplets of colloidal size, which then disperse throughout the other liquid. Oil spills in the ocean may be difficult to clean up, partly because wave action can cause the oil and water to form an emulsion. In many emulsions, however, the dispersed phase tends to coalesce, form large drops, and separate. Therefore, emulsions are usually stabilized by an emulsifying agent, a substance that inhibits the coalescence of the dispersed liquid. For example, a little soap will stabilize an emulsion of kerosene in water. Milk is an emulsion of butterfat in water, with the protein casein serving as the emulsifying agent. Mayonnaise is an emulsion of oil in vinegar, with egg yolk components as the emulsifying agents.

Condensation methods form colloidal particles by aggregation of molecules or ions. If the particles grow beyond the colloidal size range, drops or precipitates form, and no colloidal system results. Clouds form when water molecules aggregate and form colloid-sized particles. If these water particles coalesce to form adequately large water drops of liquid water or crystals of solid water, they settle from the sky as rain, sleet, or snow. Many condensation methods involve chemical reactions. A red colloidal suspension of iron(III) hydroxide may be prepared by mixing a concentrated solution of iron(III) chloride with hot water:

$$\mathrm{Fe^{3+}(aq) + 3\,Cl^-(aq) + 6\,H_2O(l) \longrightarrow Fe(OH)_3(s) + H_3O^+(aq) + 3\,Cl^-(aq)}$$

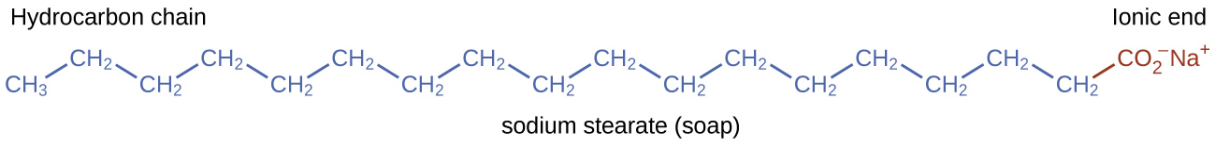
A colloidal gold sol results from the reduction of a very dilute solution of gold(III) chloride by a reducing agent such as formaldehyde, tin(II) chloride, or iron(II) sulfate:

$$\mathrm{Au^{3+} + 3\,e^- \longrightarrow Au}$$

Some gold sols prepared in 1857 are still intact (the particles have not coalesced and settled), illustrating the long-term stability of many colloids.

Soaps and Detergents

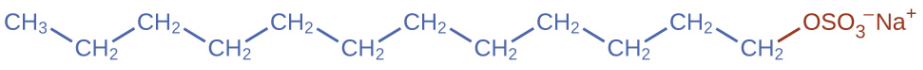
Pioneers made soap by boiling fats with a strongly basic solution made by leaching potassium carbonate, K_2CO_3, from wood ashes with hot water. Animal fats contain polyesters of fatty acids (long-chain carboxylic acids). When animal fats are treated with a base like potassium carbonate or sodium hydroxide, glycerol and salts of fatty acids such as palmitic, oleic, and stearic acid are formed. The salts of fatty acids are called *soaps*. The sodium salt of stearic acid, sodium stearate, has the formula $C_{17}H_{35}CO_2Na$ and contains an uncharged nonpolar hydrocarbon chain, the $C_{17}H_{35}$— unit, and an ionic carboxylate group, the —



sodium stearate (soap)

Figure 11.6.3: Soaps contain a nonpolar hydrocarbon end (blue) and an ionic end (red). The ionic end is a carboxylate group. The length of the hydrocarbon end can vary from soap to soap.

Detergents (soap substitutes) also contain nonpolar hydrocarbon chains, such as $C_{12}H_{25}$—, and an ionic group, such as a sulfate—



sodium lauryl sulfate (detergent)

Figure 11.6.4: Detergents contain a nonpolar hydrocarbon end (blue) and an ionic end (red). The ionic end can be either a sulfate or a sulfonate. The length of the hydrocarbon end can vary from detergent to detergent.

The cleaning action of soaps and detergents can be explained in terms of the structures of the molecules involved. The hydrocarbon (nonpolar) end of a soap or detergent molecule dissolves in, or is attracted to, nonpolar substances such as oil, grease, or dirt

particles. The ionic end is attracted by water (polar), illustrated in Figure 11.6.5. As a result, the soap or detergent molecules become oriented at the interface between the dirt particles and the water so they act as a kind of bridge between two different kinds of matter, nonpolar and polar. Molecules such as this are termed **amphiphilic** since they have both a hydrophobic ("water-fearing") part and a hydrophilic ("water-loving") part. As a consequence, dirt particles become suspended as colloidal particles and are readily washed away.

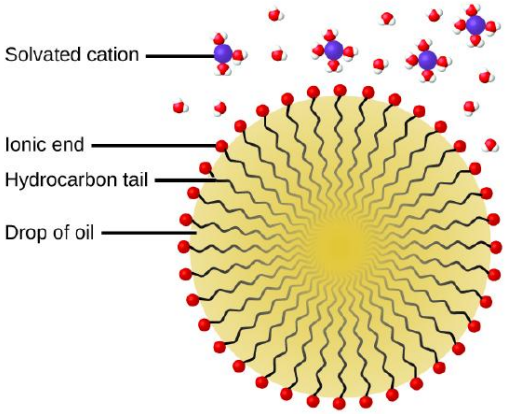


Figure 11.6.5: This diagrammatic cross section of an emulsified drop of oil in water shows how soap or detergent acts as an emulsifier.

Chemistry in Everyday Life: Deepwater Horizon Oil Spill

The blowout of the Deepwater Horizon oil drilling rig on April 20, 2010, in the Gulf of Mexico near Mississippi began the largest marine oil spill in the history of the petroleum industry. In the 87 days following the blowout, an estimated 4.9 million barrels (210 million gallons) of oil flowed from the ruptured well 5000 feet below the water's surface. The well was finally declared sealed on September 19, 2010.

Crude oil is immiscible with and less dense than water, so the spilled oil rose to the surface of the water. Floating booms, skimmer ships, and controlled burns were used to remove oil from the water's surface in an attempt to protect beaches and wetlands along the Gulf coast. In addition to removal of the oil, attempts were also made to lessen its environmental impact by rendering it "soluble" (in the loose sense of the term) and thus allowing it to be diluted to hopefully less harmful levels by the vast volume of ocean water. This approach used 1.84 million gallons of the oil dispersant Corexit 9527, most of which was injected underwater at the site of the leak, with small amounts being sprayed on top of the spill. Corexit 9527 contains 2-butoxyethanol ($C_6H_{14}O_2$), an amphiphilic molecule whose polar and nonpolar ends are useful for emulsifying oil into small droplets, increasing the surface area of the oil and making it more available to marine bacteria for digestion (Figure 11.6.6). While this approach avoids many of the immediate hazards that bulk oil poses to marine and coastal ecosystems, it introduces the possibility of long-term effects resulting from the introduction of the complex and potential toxic components of petroleum into the ocean's food chain. A number of organizations are involved in monitoring the extended impact of this oil spill, including the National Oceanic and Atmospheric Administration (visit this website for additional details).

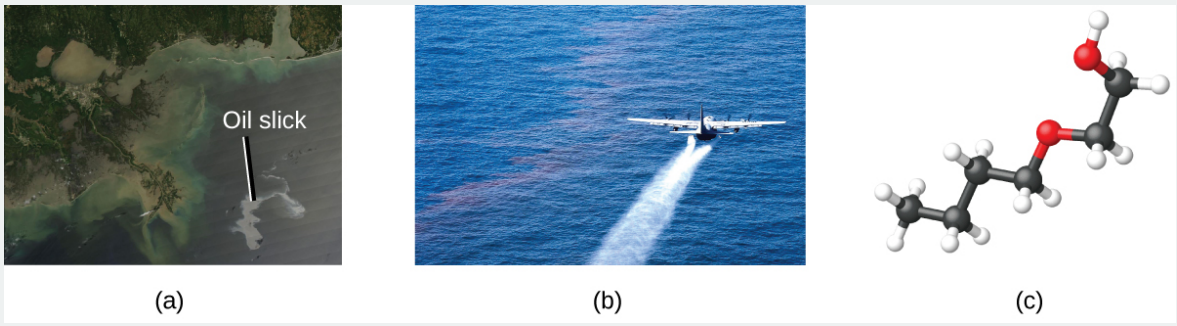


Figure 11.6.6: (a) This NASA satellite image shows the oil slick from the Deepwater Horizon spill. (b) A US Air Force plane sprays Corexit, a dispersant. (c) The molecular structure of 2-butoxyethanol is shown. (credit a: modification of work by "NASA, FT2, demis.nl"/Wikimedia Commons; credit b: modification of work by "NASA/MODIS Rapid Response Team"/Wikimedia Commons)

Electrical Properties of Colloidal Particles

Dispersed colloidal particles are often electrically charged. A colloidal particle of iron(III) hydroxide, for example, does not contain enough hydroxide ions to compensate exactly for the positive charges on the iron(III) ions. Thus, each individual colloidal particle bears a positive charge, and the colloidal dispersion consists of charged colloidal particles and some free hydroxide ions, which keep the dispersion electrically neutral. Most metal hydroxide colloids have positive charges, whereas most metals and metal sulfides form negatively charged dispersions. All colloidal particles in any one system have charges of the same sign. This helps keep them dispersed because particles containing like charges repel each other.

The charged nature of some colloidal particles may be exploited to remove them from a variety of mixtures. For example, the particles comprising smoke are often colloidally dispersed and electrically charged. Frederick Cottrell, an American chemist, developed a process to remove these particles. The charged particles are attracted to highly charged electrodes, where they are neutralized and deposited as dust (Figure 11.6.8). This is one of the important methods used to clean up the smoke from a variety of industrial processes. The process is also important in the recovery of valuable products from the smoke and flue dust of smelters, furnaces, and kilns. There are also similar electrostatic air filters designed for home use to improve indoor air quality.

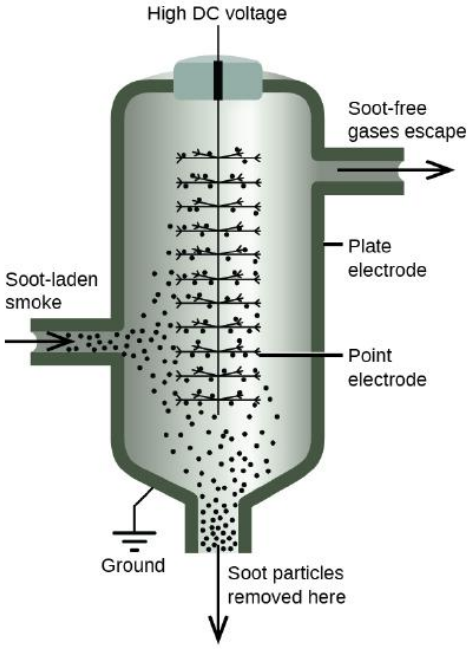


Figure 11.6.8: In a Cottrell precipitator, positively and negatively charged particles are attracted to highly charged electrodes, where they are neutralized and deposited as dust.

Portrait of a Chemist: Frederick Gardner Cottrell

Born in Oakland, CA, in 1877, Frederick Cottrell devoured textbooks as if they were novels and graduated from high school at the age of 16. He then entered the University of California (UC), Berkeley, completing a Bachelor's degree in three years. He saved money from his $1200 annual salary as a chemistry teacher at Oakland High School to fund his studies in chemistry in Berlin with Nobel prize winner Jacobus Henricus van't Hoff, and in Leipzig with Wilhelm Ostwald, another Nobel awardee. After earning his PhD in physical chemistry, he returned to the United States to teach at UC Berkeley. He also consulted for the DuPont Company, where he developed the electrostatic precipitator, a device designed to curb air pollution by removing colloidal particles from air. Cottrell used the proceeds from his invention to fund a nonprofit research corporation to finance scientific research.

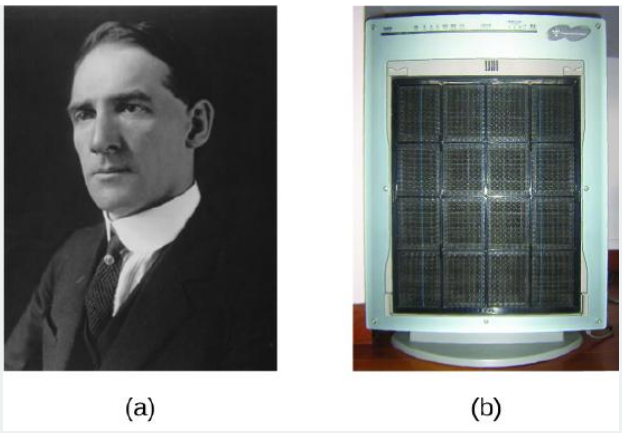


Figure 11.6.7: (a) Frederick Cottrell developed (b) the electrostatic precipitator, a device designed to curb air pollution by removing colloidal particles from air. (credit b: modification of work by “SpLot”/Wikimedia Commons)

Gels

Gelatin desserts, such as Jell-O, are a type of colloid (Figure 11.6.9). Gelatin sets on cooling because the hot aqueous mixture of gelatin coagulates as it cools, yielding an extremely viscous body known as a **gel**. A gel is a colloidal dispersion of a liquid phase throughout a solid phase. It appears that the fibers of the dispersing medium form a complex three-dimensional network, the interstices being filled with the liquid medium or a dilute solution of the dispersing medium.

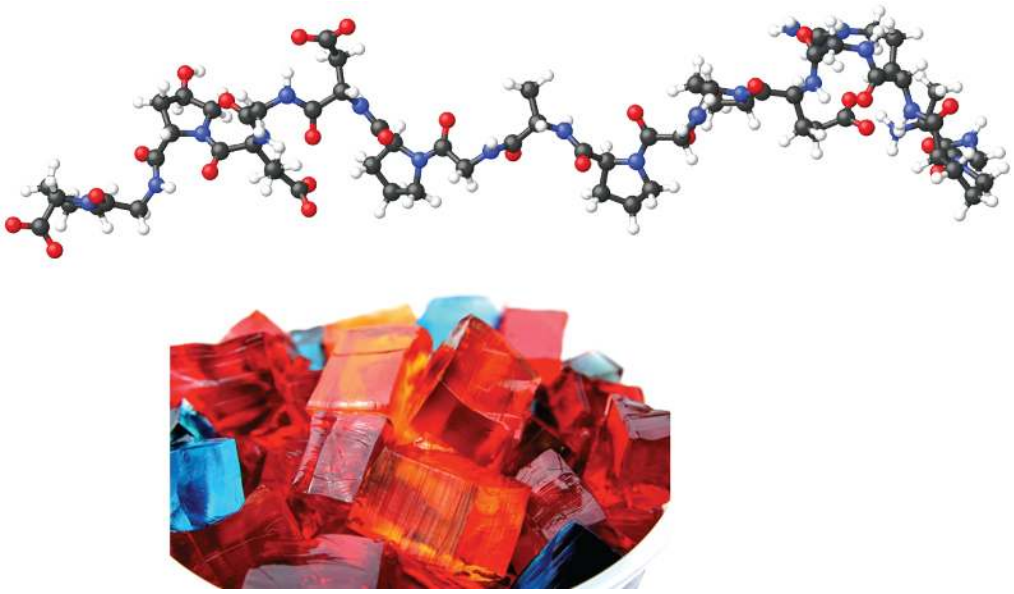


Figure 11.6.9: Gelatin desserts are colloids in which an aqueous solution of sweeteners and flavors is dispersed throughout a medium of solid proteins. (credit photo: modification of work by Steven Depolo)

Pectin, a carbohydrate from fruit juices, is a gel-forming substance important in jelly making. Silica gel, a colloidal dispersion of hydrated silicon dioxide, is formed when dilute hydrochloric acid is added to a dilute solution of sodium silicate. Canned Heat is a flammable gel made by mixing alcohol and a saturated aqueous solution of calcium acetate.

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11.7: Key Terms

alloy | solid mixture of a metallic element and one or more additional elements

amphiphilic | molecules possessing both hydrophobic (nonpolar) and a hydrophilic (polar) parts

boiling point elevation | elevation of the boiling point of a liquid by addition of a solute

boiling point elevation constant | the proportionality constant in the equation relating boiling point elevation to solute molality; also known as the ebullioscopic constant

colligative property | property of a solution that depends only on the concentration of a solute species

colloid | (also, colloidal dispersion) mixture in which relatively large solid or liquid particles are dispersed uniformly throughout a gas, liquid, or solid

crenation | process whereby biological cells become shriveled due to loss of water by osmosis

dispersed phase | substance present as relatively large solid or liquid particles in a colloid

dispersion medium | solid, liquid, or gas in which colloidal particles are dispersed

dissociation | physical process accompanying the dissolution of an ionic compound in which the compound's constituent ions are solvated and dispersed throughout the solution

electrolyte | substance that produces ions when dissolved in water

emulsifying agent | amphiphilic substance used to stabilize the particles of some emulsions

emulsion | colloid formed from immiscible liquids

freezing point depression | lowering of the freezing point of a liquid by addition of a solute

freezing point depression constant | (also, cryoscopic constant) proportionality constant in the equation relating freezing point depression to solute molality

gel | colloidal dispersion of a liquid in a solid

hemolysis | rupture of red blood cells due to the accumulation of excess water by osmosis

Henry's law | the proportional relationship between the concentration of dissolved gas in a solution and the partial pressure of the gas in contact with the solution

hypertonic | of greater osmotic pressure

hypotonic | of less osmotic pressure

ideal solution | solution that forms with no accompanying energy change

immiscible | of negligible mutual solubility; typically refers to liquid substances

ion pair | solvated anion/cation pair held together by moderate electrostatic attraction

ion-dipole attraction | electrostatic attraction between an ion and a polar molecule

isotonic | of equal osmotic pressure

miscible | mutually soluble in all proportions; typically refers to liquid substances

molality (*m*) | a concentration unit defined as the ratio of the numbers of moles of solute to the mass of the solvent in kilograms

nonelectrolyte | substance that does not produce ions when dissolved in water

osmosis | diffusion of solvent molecules through a semipermeable membrane

osmotic pressure (Π) | opposing pressure required to prevent bulk transfer of solvent molecules through a semipermeable membrane

partially miscible | of moderate mutual solubility; typically refers to liquid substances

Raoult's law | the relationship between a solution's vapor pressure and the vapor pressures and concentrations of its components

saturated | of concentration equal to solubility; containing the maximum concentration of solute possible for a given temperature and pressure

semipermeable membrane | a membrane that selectively permits passage of certain ions or molecules

solubility | extent to which a solute may be dissolved in water, or any solvent

solvation | exothermic process in which intermolecular attractive forces between the solute and solvent in a solution are established

spontaneous process | physical or chemical change that occurs without the addition of energy from an external source

strong electrolyte | substance that dissociates or ionizes completely when dissolved in water

supersaturated | of concentration that exceeds solubility; a nonequilibrium state

suspension | heterogeneous mixture in which relatively large component particles are temporarily dispersed but settle out over time

Tyndall effect | scattering of visible light by a colloidal dispersion

unsaturated | of concentration less than solubility

van't Hoff factor (*i*) | the ratio of the number of moles of particles in a solution to the number of moles of formula units dissolved in the solution

weak electrolyte | substance that ionizes only partially when dissolved in water

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11.8: Key Equations

$\Delta T_b = K_b m$
$\Delta T_f = K_f m$
$\Pi = MRT$

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11.9: Summary

A solution forms when two or more substances combine physically to yield a mixture that is homogeneous at the molecular level. The solvent is the most concentrated component and determines the physical state of the solution. The solutes are the other components typically present at concentrations less than that of the solvent. Solutions may form endothermically or exothermically, depending upon the relative magnitudes of solute and solvent intermolecular attractive forces. Ideal solutions form with no appreciable change in energy.

Substances that dissolve in water to yield ions are called electrolytes. Electrolytes may be covalent compounds that chemically react with water to produce ions (for example, acids and bases), or they may be ionic compounds that dissociate to yield their constituent cations and anions, when dissolved. Dissolution of an ionic compound is facilitated by ion-dipole attractions between the ions of the compound and the polar water molecules. Soluble ionic substances and strong acids ionize completely and are strong electrolytes, while weak acids and bases ionize to only a small extent and are weak electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water.

The extent to which one substance will dissolve in another is determined by several factors, including the types and relative strengths of intermolecular attractive forces that may exist between the substances' atoms, ions, or molecules. This tendency to dissolve is quantified as a substance's solubility, its maximum concentration in a solution at equilibrium under specified conditions. A saturated solution contains solute at a concentration equal to its solubility. A supersaturated solution is one in which a solute's concentration exceeds its solubility—a nonequilibrium (unstable) condition that will result in solute precipitation when the solution is appropriately perturbed. Miscible liquids are soluble in all proportions, and immiscible liquids exhibit very low mutual solubility. Solubilities for gaseous solutes decrease with increasing temperature, while those for most, but not all, solid solutes increase with temperature. The concentration of a gaseous solute in a solution is proportional to the partial pressure of the gas to which the solution is exposed, a relation known as Henry's law.

Properties of a solution that depend only on the concentration of solute particles are called colligative properties. They include changes in the vapor pressure, boiling point, and freezing point of the solvent in the solution. The magnitudes of these properties depend only on the total concentration of solute particles in solution, not on the type of particles. The total concentration of solute particles in a solution also determines its osmotic pressure. This is the pressure that must be applied to the solution to prevent diffusion of molecules of pure solvent through a semipermeable membrane into the solution. Ionic compounds may not completely dissociate in solution due to activity effects, in which case observed colligative effects may be less than predicted.

Colloids are mixtures in which one or more substances are dispersed as relatively large solid particles or liquid droplets throughout a solid, liquid, or gaseous medium. The particles of a colloid remain dispersed and do not settle due to gravity, and they are often electrically charged. Colloids are widespread in nature and are involved in many technological applications.

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CHAPTER OVERVIEW

12: Lab Manual

12.1: Lab 1 - Using Excel for Graphical Analysis of Data
12.2: Lab 2 - Measurements in the Laboratory
12.3: Lab 3 - The Densities of Solutions and Solids
12.4: Lab 4 - Paper Chromatography- Separation and Identification of Five Metal Cations
12.5: Lab 5 - Flame Tests of Metal Cations
12.6: Lab 6 - Atomic Emission
12.7: Lab 7 - Chemical Nomenclature Dry Lab
12.8: Lab 8 - Measuring Manganese Concentration Using Spectrophotometry
12.9: Lab 9 - VSEPR Theory and Shapes of Molecules Dry Lab
12.10: Lab 10 - Titration of Vinegar
12.11: Lab 11 - Types of Chemical Reactions
12.12: Lab 12 - Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions
12.13: Lab 13 - Mole Ratios and Reaction Stoichiometry
12.14: Lab 14 - Properties of Hydrates
12.15: Lab 15 - The Properties of Oxygen Gas
12.16: Lab 16 - The Dumas Method
12.17: Lab 17 - The Molecular Weight of Carbon Dioxide
12.18: Lab 18 - Calorimetry and Hess's Law

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12.1: Lab 1 - Using Excel for Graphical Analysis of Data

Using Excel for Graphical Analysis of Data

Last Name	First Name	Partner Name(s)	Date

1.0 INTRODUCTION

In several upcoming labs, a primary goal will be to determine the mathematical relationship between two variable physical parameters. Graphs are useful tools that can elucidate such relationships. First, plotting a graph provides a visual image of data and any trends therein. Second, using appropriate analysis, graphs provide us with the ability to predict the results of any changes to the system.

An important technique in graphical analysis is the transformation of experimental data to produce a straight line. If there is a direct, linear relationship between two variables, the data may be fitted to the equation of a line with the familiar form $y = mx + b$ through a technique known as linear regression. Here m represents the slope of the line, and b represents the y-intercept, as shown in the figure below. This equation expresses the mathematical relationship between the two variables plotted and allows for the prediction of unknown values within the parameters.

Computer spreadsheets are powerful tools for manipulating and graphing quantitative data. In this exercise, the spreadsheet program **Microsoft Excel**© will be used for this purpose. Students will learn to use Excel in order to explore a number of linear graphical relationships. Please note that although Excel can fit curves to nonlinear data sets, this form of analysis is usually not as accurate as linear regression.

NOTE: The navigation prompts (commands) below are based on using MS Excel for PC (downloaded version).

2.0 PROCEDURE

Part A. Simple Linear Plot

Experimental Data: Different phosphate concentrations are analyzed using colorimetric analysis and the following data are generated.

mg Phosphate	Absorbance
4.984	0.097
9.968	0.174
19.935	0.303
29.903	0.500
49.838	0.774

1. Launch the program Microsoft Excel© in your computer. Go to the ‘Start’ button (at the bottom left or top right on the screen), then click Programs, followed by Microsoft Excel or look for this icon:

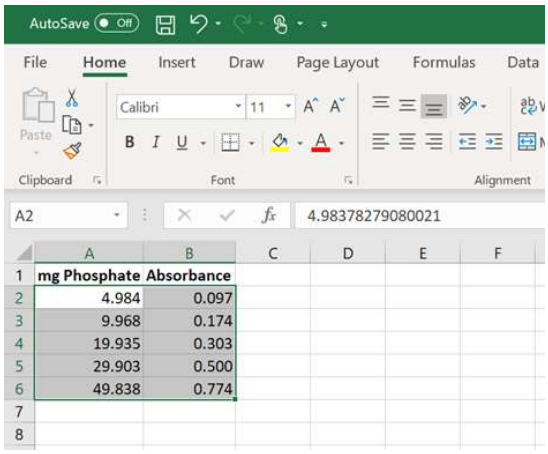


2. Select ‘Blank Workbook’ and enter the above data into the first two columns in the spreadsheet.

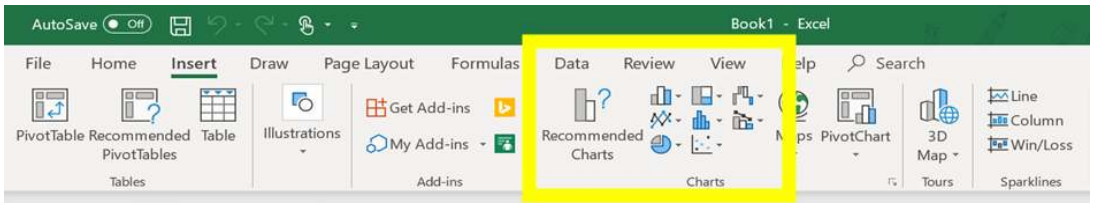
· Reserve the first row for column labels.

· The x values must be entered to the left of the y values in the spreadsheet. Remember that the independent variable (the one that you, as the experimenter, have control of) goes on the x-axis while the dependent variable (the measured data) goes on the y-axis.

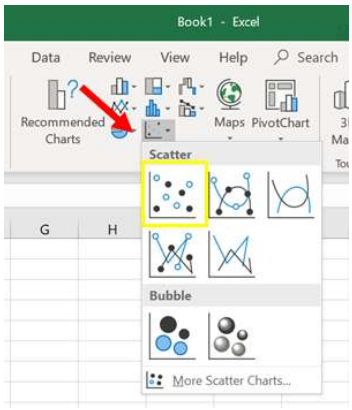
3. Highlight the set of data (not the column labels) that you wish to plot (Figure 1). Click and drag your mouse from the top-left corner of the data set (e.g. cell A2) to the bottom right corner.



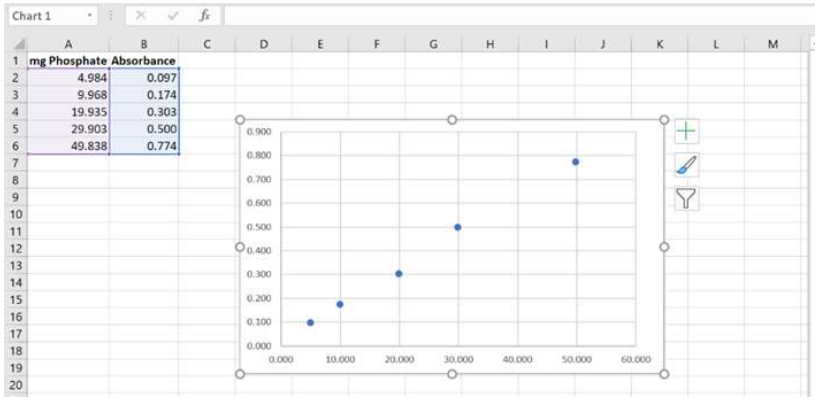
4. Click on ‘Insert’. It’s one of the tabs near the top of the Excel window. A toolbar below the ‘Insert’ tab should open (Figure 2).



In the Recommended Charts, select the ‘Scatter’ graph that shows data points only, with no connecting lines (Figure 3).



You should now see a scatter plot on your Excel workbook, which provides a preview of your graph (Figure 4).



5. If all looks well, you can now add title and label the axes of your graph.

a) To add title, click inside the chart; switch to the Design tab, and click Add Chart Element > Chart Title > Above Chart (Figure 5)

Note: The graph should be given a meaningful, explanatory title that starts out “Y versus X” followed by a description of your system.

b) To label the axes of your graph, click on Axis Titles > select Primary Horizontal Axis Title and > Primary Vertical Axis Title (Figure 6).

Note: It is important to label axes with both the measurement and the units used.

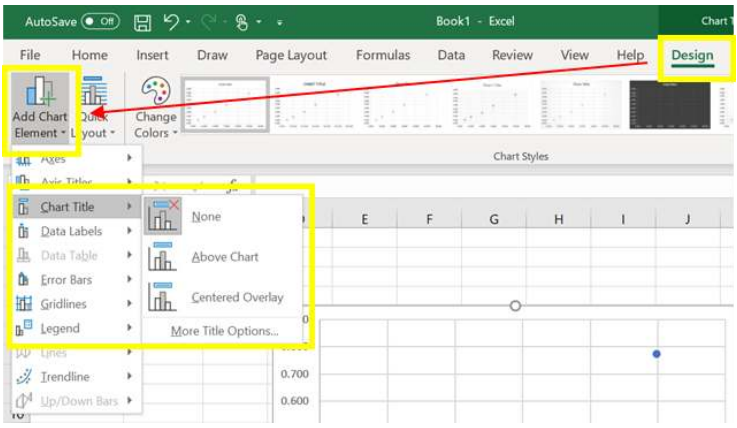


Figure 5

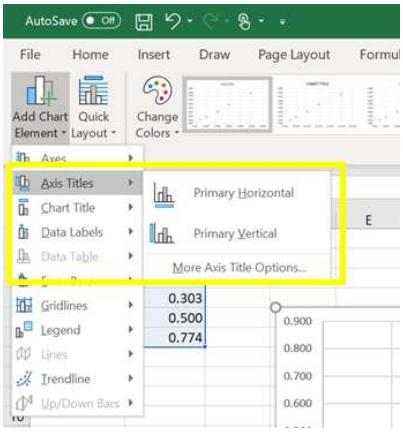


Figure 6

· To change the titles, click the text box for each title, highlight the text and type in your new title (Figure 7).

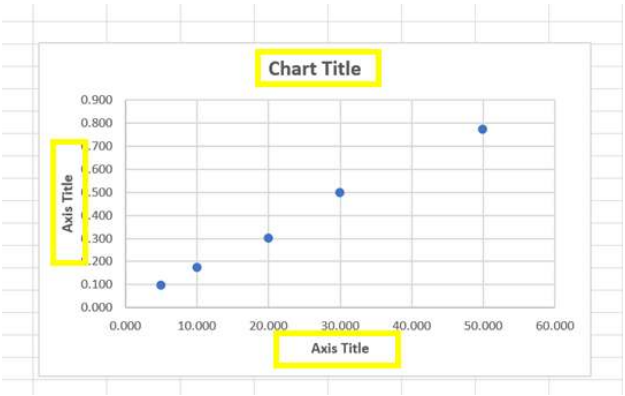
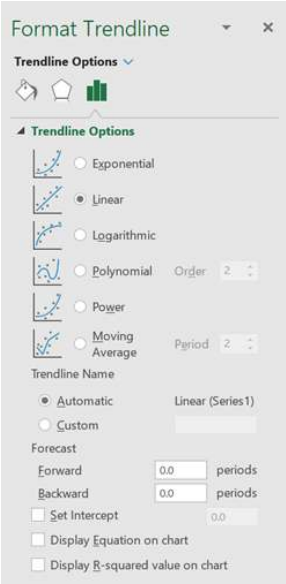


Figure 7

6. Your next step is to add a trendline to the plotted data points. A trendline represents the best possible linear fit to your data. To do this, "activate" the graph by clicking on any one of the data points. When you do this, all the data points will appear highlighted.



· Click the Chart Elements button next to the upper-right corner of the chart.

· Check the 'Trendline' box.

· Click 'More Options'. This will display the options shown in Figure 8. Notice that the 'Linear' button is already selected.

· If you like to see the algebraic equation of the trendline, select the 'Display Equation' option. The equation (in the form y = *m*x + *b*) should appear on your graph.

· If you like to know how well the data is fit by the equation, select the 'Display R-squared value' option. The closer the R^2 value is to 1, the better the fit.

· Note that the program will always fit a trendline to the data no matter how good or awful the data is. You must judge the quality of the fit and the suitability of this type of fit to your data set.

7. Print a full-sized copy of your prepared graph and attach to your data recording sheet. Then, record the following information on the data recording sheet.

· The equation of the best-fit trendline to your data

· The slope of the trendline

· The y-intercept of the trendline

· Whether the fit of the line to the data is good or bad, and why.

8. The graph contains a visual representation of the relationship (the graph) of the dependent and independent variables in your data and a mathematical expression of the relationship (the equation). It can now be used to make certain predictions.

For example, using the data above, you are asked to determine the phosphate concentration, in mg, in a different water sample. You used the same spectrometer and recorded an absorbance of 0.487. How can you find the concentration of phosphate in the water sample?

You can plug this value into the equation of the trendline and solve for the unknown concentration. Calculate and record your answer into your report. Note that this calculation will generally give a more precise value than inserting the data point (also called interpolating) and 'eyeballing' from the graph.

Part B. Two Data Sets with Overlay

Experimental Data: Two sets metal slugs of different sizes are measured. The masses and volumes are recorded, and the two sets of data collected are presented in the table below:

Metal A		**Metal B**	
Volume (cm^3)	**Mass (g)**	**Volume (cm^3)**	**Mass (g)**
1.00	11.87	1.20	7.64

2.50	21.66	4.00	8.71
4.25	32.27	5.00	15.72
5.00	41.69	6.00	19.93
7.00	54.47	7.00	15.75
8.00	59.46	8.50	23.88

You would like to see how these two sets of data relate to each other. To do this you will have to place both sets of data, as independent relationships, on the same graph. Note that this process only works when you have the same axis values and magnitudes.

1. Enter this new data on a new worksheet (Sheet 2) in Excel. Be sure to label your data columns A and B. Again, remember to enter the x values to the left of the y values.

2. First, plot data for Metal A only as an XY Scatter plot (the same way you did with the data in Part A). Fit a trendline to this data using linear regression and obtain the equation of this line.

3. Now you need to add data for Metal B to this graph.

· Activate the graph by clicking on one of the data points, or anywhere on the chart.

· Right-click the chart, and then choose ‘Select Data’ from the drop-down menu. The Select Data Source box appears on the worksheet with the source data of the chart.

· Click the ‘Add’ tab and type “Metal B” for the Series Name.

· Click on the box under Series X values, then highlight the x-axis values of Metal B. Press enter, then repeat this procedure for the Series Y Values, highlighting the y-axis values of Metal B. For each of these steps, you should see a display similar to what is shown in Figure 9. Note that slight differences may appear due to the version of Microsoft Excel installed on your computer.

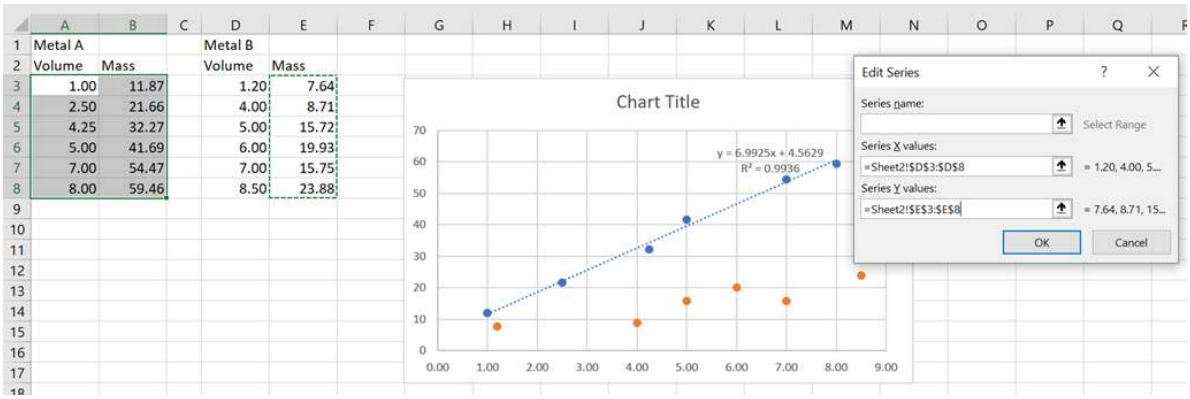


Figure 9

· Click ‘OK’ twice to return to the main Excel window.

· At this point, you should see the new data points as shown in Figure 10. You can now independently analyze the second dataset by inserting a trendline as before.

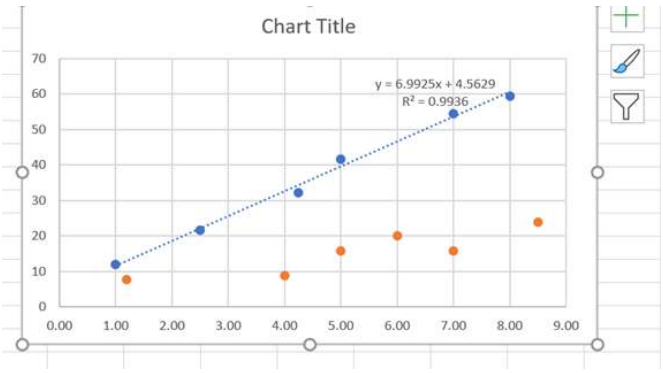


Figure 10

4. Print a full-sized copy of your prepared graph and attach to your data recording sheet. Then, record the following information on the data recording sheet.

· The equation of the best-fit trendline for Metal A

· The equation of the best-fit trendline for Metal B

· If these trendlines were extrapolated, they would intersect. Determine the values of x and y for the point of intersection using simultaneous equations.

Part C: Statistical Analysis

When many independent measurements are made for one variable, there is inevitably some scatter (noise) in the data. This is usually the result of random errors over which the experimenter has little control.

Experimental Data: Ten different students at two different colleges each measure the sulfate ion concentration in a sample of tap water:

College #1	35.9 ppm	43.2 ppm	33.5 ppm	35.1 ppm	32.8 ppm	37.6 ppm	31.9 ppm	36.6 ppm	35.0 ppm	32.0 ppm
College #2	45.1 ppm	34.2 ppm	36.8 ppm	31.0 ppm	40.7 ppm	29.6 ppm	35.4 ppm	32.5 ppm	43.5 ppm	38.8 ppm

Simple statistical analyses of these datasets might include calculations of the mean and median concentration, and the standard deviation. The mean () is simply the average value, defined as the sum (å) of each of the measurements ( ) in a data set divided by the number of measurements ( ).

The median ( ) is the midpoint value of a numerically ordered dataset, where half of the measurements are above the median and half are below. The median location of N measurements can be found using:



Standard deviation (s) is a measure of the variation in a dataset and is defined as the square root of the sum of squares divided by the number of measurements minus one.



While the mean, median and standard deviation can be calculated by hand using the above equations, it is often more convenient to use a calculator or computer to determine these values. Microsoft Excel© is particularly well suited for such statistical analyses, especially on large datasets.

1. Enter the data acquired by the students from College #1 (only) into a single column of cells (e.g. column A) on a new worksheet (Sheet 3) in Excel. Then, click on any empty cell (usually one close to the data cells) to enter the mean, median, or standard deviation for the data set.

2. Select the ‘Formulas’ tab on the top menu and click on ‘Insert Function’ button to perform the required functions on the data.

3. To compute the mean or average of the data entered in cells A1 through A10, on the box below ‘Search for a function’, type ‘average’ and then click ‘Go’. Select ‘average’ under ‘Select a function’ and then click ‘OK’.

Similarly, to compute the median, type ‘median’ and to obtain the standard deviation, type ‘stdev’.

4. Record on your data recording sheet the mean, median, and standard deviation for the College #1 data set.

Rejecting Outliers

Do all the measurements in the College #1 data set look equally good to you, or are there any values that do not seem to fit with the others? If so, are you allowed to reject these measurements?

Outliers are data points which lie far outside the range defined by the rest of the measurements and may skew your results to a great extent. If you determine that an outlier resulted from an obvious experimental error (e.g., you incorrectly read an instrument or prepared a solution), you may reject the point without hesitation. If, however, none of these errors is evident, you must use caution in making your decision to keep or reject a point. One **rough criterion** for rejecting a data point is if it lies *beyond two standard deviations from the mean or average*.

5. Using the above criteria, determine if there are any outliers in the College #1 dataset.

· Record these outlier measurements (if any) on your data recording sheet.

· Then, excluding the outliers, re-calculate the mean, median and standard deviation of this data set.

6. Repeat steps 1 – 5 to analyze the data from College #2.

Important Note: Rejecting data points cannot be done just because you want your results to look better. If you choose to reject an outlier for any reason, you must always include documentation in your lab report which clearly states:

· that you did reject a point

· which point you rejected

· why you rejected it

Failure to disclose this could constitute scientific fraud.

3.0 DATA RECORDING SHEET

Attach the graphs you made for Parts A and B in this assignment. For each graph, make sure the following components are in the printout:

· Graph Title

· Labels for x and y axes (along with appropriate units when applicable)

· Best-fit line equation and R^2 when appropriate

Part A. Simple Linear Plot

1. Which set of data is plotted on the y-axis? _____

the x-axis? _____

2. From the graph:

Equation of the fitted trendline _____

Value of the slope of this line _____

Value of the y-intercept of this line _____

3. Is the fit of the trendline to your data good (circle one)? Yes / No

Explain why you think the line is a good fit to the data.

4. Determine the concentration of phosphate in the water sample with a recorded absorbance of 0.487 using

a) Interpolation and “eyeballing” _____

b) The equation of the trendline _____

Show your calculations for b) below.

Part B. Two Data Sets and Overlay

1. From the graph:

Equation of the trendline for Metal A data set _____

Equation of the trendline for Metal B data set _____

2. Perform a simultaneous equations calculation to determine the x and y values for the point of intersection between these lines. Show your work below.

Part C. Statistical Analysis

1. For the College #1 data set, record the following values:

the mean SO_4^{2-} concentration _____

the median SO_4^{2-} concentration _____

the standard deviation in the data set _____

2. Are there any outliers in the College #1 data set (circle one)? Yes / No

If yes, which measurements are the outliers? _____

Show the calculations you used to identify the outliers (or, if none, how you determined that there were none).

3. Re-calculate the following values (using Excel) excluding the outliers:

the mean SO_4^{2-} concentration _____

the median SO_4^{2-} concentration _____

the standard deviation in the data set _____

4. For the College #2 data set, record the following values:

the mean SO_4^{2-} concentration _____

the median SO_4^{2-} concentration _____

the standard deviation in the data set _____

5. Are there any outliers in the College #2 data set (circle one)? Yes / No

If yes, which measurements are the outliers? _____

Show the calculations you used to identify the outliers (or, if none, how you determined that there were none).

6. For College #2 data set, re-calculate the following values (using Excel) excluding the outliers:

the mean SO_4^{2-} concentration _____

the median SO_4^{2-} concentration _____

the standard deviation in the data set _____

7. Which data set give the more precise measurements? _____

How do you know?

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12.2: Lab 2 - Measurements in the Laboratory

Objectives

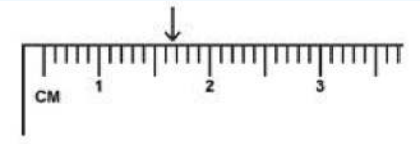
- To use standard laboratory measurement devices to measure length, volume and mass amounts.
- To use these measurements to determine the areas of shapes and volumes
- To determine the density of water.
- To determine the density of a solid and use this to determine further quantities.
- To determine the density of aluminum (applying the technique of water displacement) and use that value to determine the thickness of a piece of aluminum foil.

Chemistry is the study of matter. Our understanding of chemical processes thus depends on our ability to acquire accurate information about matter. Often, this information is quantitative, in the form of *measurements*. In this lab, you will be introduced to some common measuring devices, and learn how to use them to obtain correct measurements, each with correct precision. A metric ruler will be used to measure length in centimeters (cm).

All measuring devices are subject to error, making it impossible to obtain exact measurements. Students will record all the digits of the measurement using the markings that we know exactly and one further digit that we estimate and call uncertain. The *uncertain digit* is our best estimate using the smallest unit of measurement given and estimating between two of these values. These digits are collectively referred to as significant figures. Note, the electronic balance is designed to register these values and the student should only record the value displayed.

When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true, accepted value. Precision refers to how close repeated measurements (using the same device) are to each other.

Example 1.1 : Measuring length



Here the “ruler” markings are every 0.1-centimeter. The correct reading is 1.67 cm. The first 2 digits **1.6**7 are known exactly. The last digit 1.6**7** is uncertain. You may have instead estimated it as 1.68 cm.

The measuring devices used in this lab may have different scale graduations than the ones shown Precision is basically how many significant figures you have in your measurement. To find the precision, you basically take the smallest unit on your measuring device, and add a decimal place (the uncertain digit).

Note

In general, the more decimal places provided by a device, the more precise the measurement will be.

Measurements obtained in lab will often be used in subsequent calculations to obtain other values of interest. Thus, it is important to consider the number of significant figures that should be recorded for such calculated values. If multiplying or dividing measured values, the result should be reported with the lowest number of *significant figures* used in the calculation. If adding or subtracting measured values, the result should be reported with the lowest number of *decimal places* used in the calculation.

Example 1.2 : Significant Figures in Calculated Values

(a) A student runs 18.752 meters in 54.2 seconds. Calculate his velocity (or speed).

$$velocity = \frac{distance}{time} \tag{12.2.1}$$

$$= \frac{18.752m}{54.2s} \tag{12.2.2}$$

$$= 0.345978m/s \text{ from calculator} \tag{12.2.3}$$

$$= 0.346m/s \text{ to 3 significant figures} \tag{12.2.4}$$

(b) The mass of a glass is measured to be 12.456 grams. If 10.33 grams of water are added to this glass, what is the total combined mass?

$$\text{total mass} = 12.456g + 10.33g \tag{12.2.5}$$

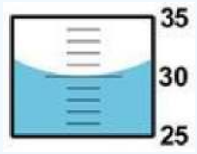
$$= 22.786g \text{ from calculator} \tag{12.2.6}$$

$$= 22.79g \text{ to 2 decimal places} \tag{12.2.7}$$

In this lab, students will also determine the density of water as well as aluminum. Volume is the amount of space occupied by matter. An extensive property is one that is dependent on the amount of matter present. Volume is an extensive property.

The *volume* of a liquid can be directly measured with specialized glassware, typically in units of milliliters (mL) or liters (L). In this lab, a beaker, two graduated cylinders and a burette will be used to measure liquid volumes, and their precision will be compared. Note that when measuring liquid volumes, it is important to read the graduated scale from the lowest point of the curved surface of the liquid, known as the liquid *meniscus*.

Example 1.3 : Measuring the Volume of a liquid



Here, the graduated cylinder markings are every 1-milliliter. When read from the lowest point of the meniscus, the correct volume reading is 30.0 mL. The first 2 digits **30**.0 are known exactly. The last digit 30.**0** is uncertain. Even though it is a zero, it is significant and must be recorded.

The volume of a solid must be measured indirectly based on its shape. For regularly shaped solids, such as a cube, sphere, cylinder, or cone, the volume can be calculated from its measured dimensions (length, width, height, diameter) by using an appropriate equation.

Example 12.2.4: Calculating the Volume of a Solid

$$\text{Volume of a cube} = l \times w \times h \tag{12.2.8}$$

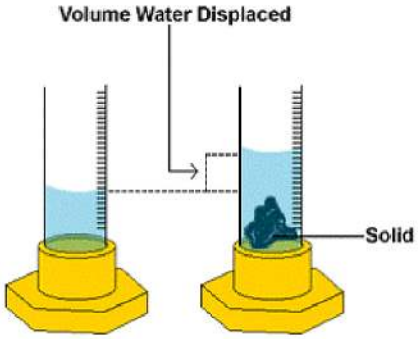
$$\text{Volume of a sphere} = \frac{4}{3}\pi r^3 \tag{12.2.9}$$

$$\text{Volume of a cylinder} = \pi r^2 h \tag{12.2.10}$$

(where r = radius = 1⁄2 the diameter)

For irregularly shaped solids, the volume can be indirectly determined via the volume of water (or any other liquid) that the solid displaces when it is immersed in the water (*Archimedes Principle*). The units for solid volumes are typically cubic centimeters (cm^3) or cubic meters (m^3). Note that 1 mL = 1 cm^3.

Measuring the Volume of an Irregularly Shaped Solid



The volume water displaced is equal to the difference between the final volume and the initial volume , or:

$$V = V_f - V_i \tag{12.2.11}$$

where the volume water displaced is equal to the volume of solid.

Density is defined as the mass per unit volume of a substance. Density is a physical property of matter. Physical properties can be measured without changing the chemical identity of the substance. Since pure substances have unique density values, measuring the density of a substance can help identify that substance. Density is also an intensive property. An intensive property is one that is independent of the amount of matter present. For example, the density of a gold coin and a gold statue are the same, even though the gold statue consists of the greater quantity of gold. Density is determined by dividing the mass of a substance by its volume:

$$density = \frac{mass}{volume} \tag{12.2.12}$$

Density is commonly expressed in units of g/cm^3 for solids, g/mL for liquids, and g/L for gases.

Procedure

Materials and Equipment

Metric ruler, shape sheet, electronic balance, triple-beam balance, 250 mL Erlenmeyer flask, 100-mL beaker, sugar, 400-mL beaker, scoopula, burette, 10 mL and 100 mL graduated cylinders, 100 mL beaker, wooden blocks, aluminum pellets, aluminum foil, electronic balance, distilled water.

Safety

Be careful when adding the aluminum to your graduated cylinder, as the glass could break.

Personal protective equipment (PPE) needed: lab coat, safety goggles, closed-toe shoes

Part A: Measuring the Dimensions of Regular Geometric Shapes

1. Borrow a ruler and “shape sheet” from the front bench. Record the ID code on your report form. Measure the dimensions of the two geometric shapes: length and width of the rectangle, and the diameter of the circle. Record these values on your lab report.
2. When finished, return the ruler and sheet to the front bench.
3. Use your measurements to calculate the area of each shape:

- Area of a rectangle: $A = l \times w$
- Area of a circle: $A = \pi r^2$ (r = radius = 1⁄2 the diameter)

Part B: Measuring the Mass of Solids

Comparing the Precision of two types of Balances

1. Obtain a 250-mL Erlenmeyer flask from your locker. Use the triple-beam balance to determine the mass of this flask.
2. Use the electronic balance in the weigh room to determine the mass of the same Erlenmeyer flask. Be sure to record your measured masses on your lab report.

Weighing by Difference

3. Obtain a 100-mL beaker from your locker. Use the electronic balance in the weigh room to determine the mass of this beaker.
4. Add two spoonfuls of sugar to this beaker, using your scoopula. **Do not do this over the balance!** Determine the new combined mass of both the beaker and the sugar. Be sure to use the same electronic balance as before. When finished, dispose the used sugar in the sink.
5. Use your two measurements to calculate the mass of sugar (only) by difference.

Part C: Volumes of Liquids and Solids

Volumes of Liquids

1. At the front bench, you will find a burette, 10-mL graduated cylinder, 100-mL graduated cylinder and 100-mL beaker, each filled with a certain quantity of water. Measure the volume of water in each. Remember to read the volume **at the bottom of the meniscus.** It is useful to hold a piece of white paper behind the burette/cylinder/beaker to make it clearer.

Volume of a Regularly Shaped Solid

2. Borrow a wooden block or cylinder and ruler from the front bench.
3. Measure the dimensions of the block. If it is a cube or a rectangular box, measure its length, width and height. If it is a cylinder or cone, measure its height and the diameter of its circular base.
4. Return the block or cylinder and ruler to the front bench when finished.

Part D: The Density of Water

1. Using the electronic balance in the weigh room, determine the mass of a clean, dry, 100-mL graduated cylinder.
2. Pour 40-50 mL of **distilled water** into the graduated cylinder and weigh. Make sure that the outside of the graduated cylinder is dry before placing it on the electronic balance.
3. Measure the liquid volume in the cylinder
4. Use the mass and volume to calculate the density of water.

Part E: The Density of Aluminum and the Thickness of Foil

Density of Aluminum

1. Using the electronic balance in the weigh room to determine the mass of a clean, dry, small beaker.
2. Obtain 20-25 aluminum pellets from the front bench. Transfer pellets to the beaker weighed in the previous step, and measure the mass of the beaker and pellets together.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
4. Carefully add all the aluminum pellets to the water, making sure not to lose any water to splashing. Also make sure that the pellets are all completely immersed in the water. Measure the new volume of the water plus the pellets.
5. When finished, retrieve and dry the aluminum pellets and return them to the front bench.
6. *Analysis:* Use your measured mass and volume (obtained via water displacement) of the aluminum pellets to calculate the density of aluminum.

The Thickness of Aluminum Foil

7. Borrow a rectangular piece of aluminum foil and ruler from the front bench. Use the ruler to measure the length and width of the piece of foil.
8. Fold the foil up into a small square and measure its mass using the electronic balance in the weigh room.
9. When finished, return the foil and ruler to the front bench.
10. *Analysis:* Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Lab Report: Measurements in the Laboratory

Part A: Measuring the Dimensions of Regular Geometric Shapes

Experimental Data

Shape Sheet ID Code:

Shape	Dimensions	Precision	Measurement	# Significant Figures

Shape	Dimensions	Precision	Measurement	# Significant Figures
Rectangle	Length			
	Width			
Circle	Diameter			

Data Analysis

1. Perform the conversions indicated. Show your work, and report your answers in scientific notation.

- Convert the measured rectangle length to hm.
- Convert the measured circle diameter to nm.

2. Calculate the areas of your rectangle and circle in cm^2. Show your work, and report your answers to the correct number of significant figures.

- Area of rectangle
- Area of circle

Part B: Measuring the Mass of Solids

Experimental Data

Table 1: Mass of an Erlenmeyer Flask

Measuring Device	Mass Measurement	# Significant Figures
Triple-Beam Balance		
Electronic Balance		

Table 2: Weighing by Difference

	Mass Measurement	# Significant Figures
Mass of Empty Beaker		
Mass of Beaker + Sugar		

Data Analysis

1. Compare your mass values obtained for the Erlenmeyer flask (in Table 1). Which balance, triple- beam or electronic, provides the more precise measurement? Explain.
2. Consider the data you obtained in Table 2.

- Calculate the mass of sugar weighed out. Show your work.
- Circle one: When performing the above calculation, significant figures / decimal places are the primary consideration.

3. Perform the conversions indicated below. Show your work, and report your answers in scientific notation.

- Convert the mass of the sugar weighed out to dg.
- Convert the mass of sugar weighed out to lbs.

Part C: The Volumes of Liquids and Solids

Table 1: The Volume of Liquid Water

Measuring Device	Precision	Volume Measurement	# Significant Figures
Burette			

Measuring Device	Precision	Volume Measurement	# Significant Figures
Beaker			
100-mL Graduated Cylinder			
10-mL Graduated Cylinder			

Table 2: The Volume of a Regular Solid, shaped as a

Dimensions Measured	Measurement	# Significant Figures

Data Analysis

Use your measured block dimensions (in Table 2) to calculate the block volume, in cm^3. Show your work, and report your answer to the correct number of significant figures.

Part D: The Density of Water

Table 1: The Density of Water

Mass of Empty, Dry Graduated Cylinder	
Mass of Graduated Cylinder + Water	
Mass of Distilled Water	
Volume of the Distilled Water in Graduated Cylinder	

Calculate the density of water, in g/mL. Show your work, and report your answer to the correct number of significant figures.

Part E: The Density of Aluminum and the Thickness of Foil

Experimental Data

Table 1: The Density of Aluminum

Mass of empty beaker	
Mass of beaker and pellets	
Mass of pellets	
Initial volume of water in cylinder	
Final volume of water and pellets	
Volume of pellets	

Table 2: The Thickness of Aluminum Foil

Mass of foil	
Length of foil	
Width of foil	

Data Analysis

1. Use your measured mass and volume of the aluminum pellets (in Table 1) to calculate the density of aluminum, in g/cm^3. Show your work, and report your answer to the correct number of significant figures.
2. Use your measurements for the aluminum foil (in Table 2) along with the true density of aluminum ($\boxed{\_D\_\{Al\}}$ = 2.70 g/cm^3) to calculate the foil thickness, in cm. Consider the foil to be a very flat rectangular box, where Volume of foil = V = length $\times$ width$\times$ height (thickness). Show your work, and report your answer in scientific notation.

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12.3: Lab 3 - The Densities of Solutions and Solids

Objectives

- To compare the precision of various types of volumetric glassware.
- To determine, and then compare, the densities of regular and diet de-gassed sodas.
- To determine the density of an unknown metal and use this value to identify the metal.
- To use a common spreadsheet program, Microsoft Excel©, to perform graphical analysis of experimental data.

Density is a fundamental physical property of matter. Physical properties are those characteristics of a substance that can be determined without changing the chemical identity of the substance. Other physical properties include melting point and solubility. In general, since different substances have unique densities, determining the density of an unknown substance can help identify it. Density is specifically defined as the ratio of a substance's mass to its volume:

$$Density = \frac{Mass}{Volume} \tag{12.3.1}$$

The S.I. unit of density is kg/m^3, but in chemistry it is more often expressed in units of g/cm^3 for solids, and g/mL for liquids and solutions.

In Part A of this lab, a study of density values will be conducted in order to introduce the concept of precision and the statistical nature of experimental data. Specifically, the density of water will be determined using measurements obtained from three different types of volumetric glassware: a buret, a 100-mL graduated cylinder and a volumetric pipet. Multiple measurements will be performed so that a large ensemble of density values is acquired for each type of glassware used. Simple statistical analyses of the three data sets will then be carried out, including calculations of the average density and standard deviation. Note that an average value ($\bar{x}$) is defined as the sum (Σ) of each of the measurements (x_i) divided by the number of measurements (N):

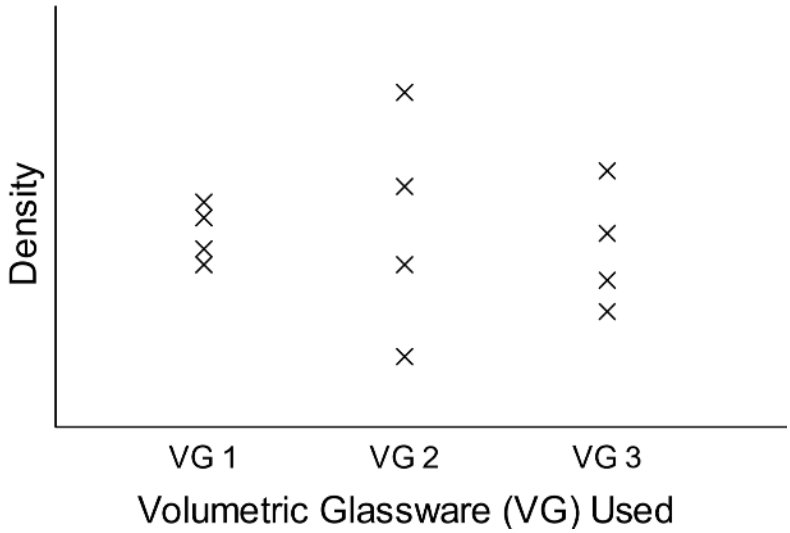
$$\bar{x} = \frac{\sum x_i}{N} \tag{12.3.2}$$

Standard deviation (s) is defined as:

$$s = \sqrt{\frac{\sum (x_i - \bar{x})^2}{N-1}} \tag{12.3.3}$$

Standard deviation essentially indicates the degree to which a set of measured values deviate from the average value. Data sets with a wide range of values (for example, see VG 2 data in the following figure) will have a larger standard deviation and are associated with less precise measurements than data sets with a narrow range of values. With this information, the precision of the measurements obtained using the three types of volumetric glassware can be compared.

The three density data sets will also be displayed in a Scatter Plot, as shown below:

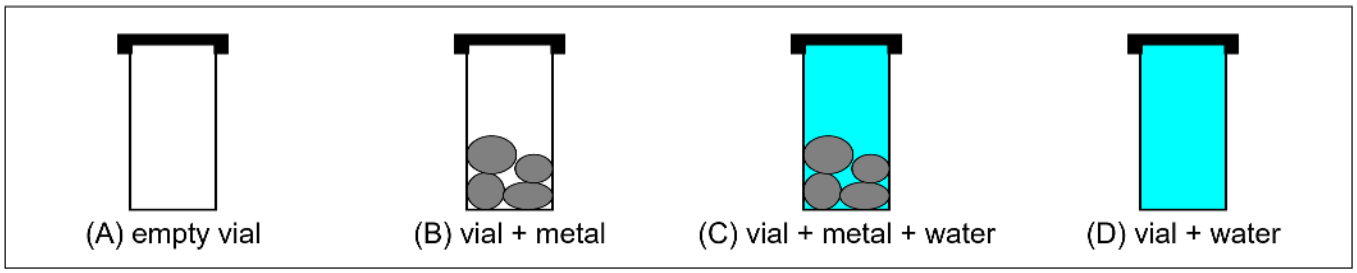


Occasionally one or more measurements in a data set will appear not to “fit in” with the others. This will often be evident upon close examination of a scatter plot. These points are called outliers – values that occur far outside the range defined by the rest of the measurements. One rough criterion for identifying an outlier is that it lies beyond two standard deviations from the average value. Such values may be legitimately excluded from a data set, as they can skew results to a great extent.

In Part B, the densities of two de-gassed sodas (regular and diet) will be determined and compared. Here, a buret will be used to dispense precise soda volumes. The masses of several different volumes of each soda will be measured, and this data will be used to determine the density of each soda.

This section of the lab will also clearly illustrate the intensive nature of density. An intensive property is one that is independent of the amount of matter present. This means that a pure substance will have the same density whether it is a larger sample or a small sample of that substance. In contrast, an extensive property is one that is dependent on the amount of matter present. Note that both mass and volume are extensive, however, since density is a ratio of these properties, density is an intensive property.

In Part C, the density of an unknown metal will be determined, and the metal identified using this value. Since mass can be more precisely measured than volume (due to the limitations of the instrumentation available), the technique employed here will involve mass measurements only. Using a capped glass vial, the following four mass measurements will be obtained:



The difference in masses A and B is the mass of the metal sample. The volume of the metal is obtained by taking the difference between the water volumes in C and D. These water volumes will be calculated using the water masses and the known density of water (see table below). Although simple, this method can yield density results accurate to 0.1%.

Once the density is determined, this experimental value will be compared to the true densities of several known metals (obtained from suggested reference sources) in order to identify the assigned unknown metal. The percent error between this experimental value (EV) and the true density value (TV) of the metal will also be calculated.

$$\text{Percent Error} = \frac{|EV - TV|}{TV} \times 100 \tag{12.3.4}$$

Note that a more accurate experimental value will yield a lower percent error (< 5% is desirable) than a less accurate value.

*Density of Liquid Water Measured for a Range of Temperatures**

https://chem.libretexts.org/@go/page/533413

Temperature (°C)	Density (g/mL)
16	0.9989
17	0.9988
18	0.9986
19	0.9984
20	0.9982
21	0.9980
22	0.9978
23	0.9975
24	0.9973
25	0.9970
26	0.9968
27	0.9965
28	0.9962
29	 0.9959

*obtained from the CRC Handbook of Chemistry and Physics, 53rd Edition

Procedure

Materials: De-gassed sodas (regular and diet), unknown metal sample.

Equipment: 10-mL volumetric pipet and pipet bulb*, 50-mL buret*, two 50-mL Erlenmeyer flasks*, 100-mL graduated cylinder, two 100-mL beakers, 50-mL beaker, electronic balance, capped glass vial, thermometer, wash bottle filled with distilled water.

*Items with an asterisk must be checked out from the stockroom.

Safety

There are no major safety considerations in this lab.

Part A: The Precision of Volumetric Glassware

All glassware should be cleaned and rinsed with distilled water before use. Pay special attention to significant figures in your recorded measurements.

1. Using a medium 100-mL beaker, obtain approximately 50 mL of distilled water from the supply in the lab.
2. Weigh a small **dry** 50-mL beaker on an electronic balance, and record this mass.
3. Graduated Cylinder: Fill the 100-mL graduated cylinder with approximately 10 mL of distilled water, and record the actual volume used. To how many significant figures should the volume be recorded? Transfer this distilled water into the small pre-weighed beaker, and measure the combined mass. When finished, empty the distilled water out of the small beaker, then carefully dry it.
4. Volumetric Pipet: Your instructor will demonstrate the correct use of the volumetric pipet at the beginning of the lab session. Use the pipet to transfer the *exactly* 10 mL of distilled water from the medium beaker into the small pre-weighed beaker. Record the volume used and the combined mass of the beaker and water. To how many significant figures should the volume be recorded? Again, empty the water out of the small beaker when finished, then carefully dry it.
5. Buret: Your instructor will demonstrate the correct use of the buret at the beginning of the lab session. Fill the buret with the remaining distilled water in your medium beaker and note the initial buret reading. Then drain approximately 10 mL of this

water into the small pre- weighed beaker. Record the actual volume used (= final – initial buret reading) and the combined mass of the beaker and water. To how many significant figures should the volume be recorded?

6. For each of the three sets of data collected (using the graduated cylinder, volumetric pipet and buret), calculate the density of distilled water to the correct number of significant figures. Then share your three density values with all the students in your lab section, and record the results of the entire class on your report form. You should have at least ten density values for each type of volumetric glassware used in this part of the lab.

Part B: The Densities of Degassed Sodas

The following procedure should be performed for both a regular soda and a diet soda.

1. A buret will be used to dispense precise volumes of each of the two sodas used. The buret should first be rinsed with distilled water, and then rinsed with a small quantity of the relevant soda before filling it with that soda.
2. Using a medium 100-mL beaker, obtain approximately 40 mL of the *regular* soda from the stock bottle. Fill the buret with this soda, and record the initial buret reading.
3. Weigh a small **dry** 50-mL Erlenmeyer flask using an electronic balance, and record this mass. Then drain approximately 5 mL of the soda from the buret into this flask, and record the new buret reading. Finally, measure and record the combined mass of the flask and soda.
4. Next, add an *additional* 5 mL of soda from the buret to the flask. Again, note the new buret reading and the new combined mass of the flask and soda. Then repeat this step four more times. You will have obtained a total of six measurements and used approximately 30 mL of the soda when you are finished.
5. Now perform the same series of steps 1-4 using the diet soda. Use your second small (dry) Erlenmeyer flask for this soda.

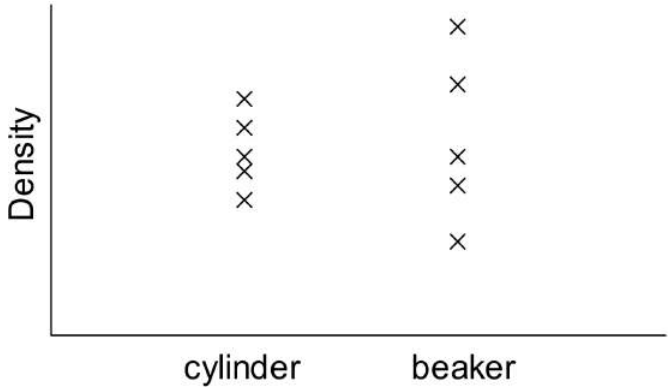
Part C: The Density of an Unknown Metal

1. Obtain a capped glass vial and an unknown metal sample from your instructor. The cap on the vial should have a small hole pierced through it. This hole will allow air and excess water to be expelled from the vial. Record the ID Code of the metal on your report form.
2. All mass measurements are to be performed on an electronic balance, and should be recorded on your report form. First, weigh the empty, dry capped vial. Then add the entire sample of your unknown metal to the vial, and weigh it again (with cap).
3. Now fill the vial (with the metal still in it) to the brim with distilled water. Gently tap the vial to ensure that no air remains between the metal pieces. Place the cap on firmly, pressing out any air or excess water. Wipe off any drops of water on the outside of the vial, and then weigh it.
4. Next, remove the metal from the vial and then fill it to the brim with distilled water only. Place the cap on firmly, wipe off excess water, and weigh.
5. Finally, using your thermometer, measure and record the temperature of the water in the vial. When finished, dry the metal sample and vial, and return them to your instructor.

Pre-laboratory Assignment: The Densities of Solutions and Solids

1. Five students decide to determine the density of a sample of liquid chloroform. Each student uses a graduated cylinder to obtain the chloroform volume and a balance to obtain the chloroform mass. Their density values are shown in the table below.

Student	Density (g/mL)
1	1.45
2	1.55
3	1.6
4	1.38
5	1.44



- Determine the average density and the standard deviation in this density dataset.

  Avg:

  Std Dev:

- The students then decide to determine the chloroform density using a beaker to measure volume. Their density results are displayed in the graph above, along with those obtained using the cylinder. Which glassware yielded the more precise measurements? Explain.
- In Part A of this lab, you will perform a similar analysis of density data obtained for distilled water. What three types of volumetric glassware will you be using?

2. A student decides to use the technique employed in Part B of this lab to determine the density of vodka. He first fills a buret with vodka to an initial volume of 1.52 mL. He then drains out some of the vodka into an empty beaker with a mass of 13.687 g. The final buret volume is 9.62 mL. If the combined mass of the beaker and vodka is 21.123 g, what is the the density of the vodka? Give your answer to the correct number of significant figures.
3. Another student decides to use the technique employed in Part C of this lab to determine the density of gold. She obtains the following data:

(A) Mass of Empty Vial	10.495 g
(B) Mass of Vial + Gold Sample	70.773 g
(C) Mass of Vial + Gold + Water	76.286 g
(D) Mass of Vial + Water	19.119 g
(E) Temperature of Water	25.0 °C
(F) Density of Water at above Temp.	0.9970 g/mL

Using her data, perform the calculations below. Show your work clearly, and pay attention to significant figures and units.

- What is the mass of the gold sample?
- What is the mass of water in the vial with the gold (see measurement C)? What is volume of water in this vial (hint: use the density of water)?
- What is the mass of water in the vial filled with water only (see measurement D)? What is the volume of water in this filled vial (hint: use the density of water)?
- The volume of the gold is equal to the difference in the two water volumes. What is the volume of the gold?
- Determine the density of the gold, in g/cm^3.

Lab Report: The Densities of Solutions and Solids

Part A: The Precision of Volumetric Glassware

Experimental Data

	Graduated Cylinder	Volumetric Pipet	Buret
Mass of Empty Beaker			
Mass of Beaker + Water			
Mass of Water only			
Actual Volume of Water used			
Density of Water			

Share your three density values with all the students in your lab section, and record the results of the entire class in the table below (pay attention to significant figures). You should acquire at least ten density values for each type of volumetric glassware used. Use an asterisks symbol (*) to denote which set of measurements you and your partner obtained.

Measurement	Class Density Values (in g/mL) obtained using:		
	Graduated Cylinder	Volumetric Pipet	Buret
1			
2			
3			
4			
5			
6			
7			
8			
9			
10			
11			
12			
13			
14			

Data Analysis

1. For each of the three density data sets, calculate the average density and the standard deviation in the data set.

	Average	Standard Deviation
Graduated Cylinder		
Volumetric Pipet		
Buret		

2. Using Microsoft Excel© plot all three density data sets on a single Scatter Graph as “Density of Water versus Type of Volumetric Glassware used”. Attach this graph to your report.

3. Now closely examine all three data sets collected by your class.

- Circle any identified outliers in the three data sets on both your *plot* and in the *data table*. Then explain, using supporting calculations, how you identified these outliers.

Now re-calculate the average and standard deviation for each data set, *omitting the outliers*.

	Average	Standard Deviation
Graduated Cylinder		
Volumetric Pipet		
Buret		

4. Based on your analyses of the three data sets without the outliers, rank the volumetric glassware used from least to most precise. Briefly explain how you determined this order.

Part B: The Densities of De-gassed Sodas

Experimental Data

Regular Soda

- Name of Soda:
- Mass of Empty Flask:
- Initial Buret Reading:

Data	Combined Mass of Flask + Soda	New Buret Reading	Total Mass of Beverage	Total Volume of Beverage	Density of Soda
1					
2					
3					
4					
5					
6					

- Average density of regular soda:

Diet Soda

- Name of Soda:
- Mass of Empty Flask:
- Initial Buret Reading:

Data	Combined Mass of Flask + Soda	New Buret Reading	Total Mass of Beverage	Total Volume of Beverage	Density of Soda
1					
2					
3					
4					
5					
6					

- Average density of diet soda:

Data Analysis

1. Which of the two sodas has the higher density? Suggest a possible reason for this result.

Part C: The Density of an Unknown Metal

Experimental Data

ID Code of Assigned Metal:

A. Mass of Empty Vial	
B. Mass of Vial + Metal	
C. Mass of Vial + Metal + Water	
D. Mass of Vial + Water	
E. Temperature of Water	
F. Density of Water at above Temperature	

Data Analysis

1. Using the data collected, calculate the density (in g/cm^3) of your unknown metal. Show your work clearly, and circle your final result.
2. Your metal could be any of the following metals: lead, aluminum, nickel or zinc. Look up the true densities of these metals and identify your unknown metal. Suggested resources include your text, websites (e.g., www.webelements.com, www.chemicool.com) and the CRC Handbook of Chemistry and Physics.

- Identity of unknown metal:
- True density of metal:
- Source of density data:

3. Calculate the percent error between your experimental density and the true metal density.

Questions

1. In Part A of this lab, you used your experimental results t rank the precision of three types of glassware from least to most precise. Did you expect these results? If yes, explain why. If no, explain why not, and give the order that you did expect.
2. Consider your density results for the diet soda in Part B of this lab. Clearly explain how they demonstrate that density is an intensive property.
3. Consider your results for Part C. Suppose your metal sample is hammered into the shape of a thin coin with a diameter of 8.62 cm. Determine the thickness (in mm) of this metal sheet. Use the measured metal mass and the true density of your metal for this calculation.

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12.4: Lab 4 - Paper Chromatography- Separation and Identification of Five Metal Cations

Objectives

- Known and unknown solutions of the metal ions Ag^{+}, Fe^{3+}, Co^{2+}, Cu^{2+} and Hg^{2+} will be analyzed using paper chromatography.
- An unknown solution containing some of these cations will be identified by comparison to the R_f values and colors of the stained spots of known solutions.

Most chemists and many other scientists must routinely separate mixtures and identify their components. The ability to qualitatively identify the substances found in a sample can be critical. For example, an environmental chemist investigating samples of polluted ground water will want to know which toxic ions might be present in a sample.

Chromatography is one of the first tools used in such situations. In this technique, many types of mixtures can be separated into the component pure substances; by comparison to a standard sample, each component substance can also be tentatively identified.

Many varieties of chromatography exist, each one designed to separate specific types of mixtures. The common feature of each type of chromatography is that a mobile phase (a liquid or gas) is pushed through a stationary phase (a solid). Table 1 lists several varieties of chromatography and typical identities of the phases. Paper chromatography will be used in this experiment.

Type of Chromatography	Mobile Phase	Stationary Phase
Gas (GC)	inert gas (helium)	waxy liquid or silicone inside narrow tubing
Liquid (LC ,HPLC, column)	solvent/solvent Mixture (organic or aqueous)	solid packing (silica, alumina)
Paper	solvent/solvent Mixture (organic or aqueous)	paper
Thin-Layer (TLC)	solvent/solvent Mixture (organic or aqueous)	silica/alumina coated glass, plastic or metal

The example of column chromatography (Figure 1) demonstrates the typical features found in this analytical technique. The diagram shows an experiment where a two-component mixture is subjected to column chromatography. The column is packed with a solid material called the stationary phase. A liquid solvent or *eluting solution* is poured into the column and completely wets the solid packing material. Then the mixture is loaded onto the top of the wet column and more eluent is added. Gravity pulls the mobile phase down through the stationary phase and the components in the mixture start to move through the column at different rates. In the diagram, component A moves faster than component B; thus component B is retained on the column for a longer time than component A. Usually this is due to a difference in solubility of the two compounds in the solvent and/or to a difference in attraction to the solid packing material. As more eluent is added to the top of the column, the components will eventually exit the column separately. The time taken to exit the column, called *retention time*, will be reproducible for each component under the given set conditions—mobile and stationary phase identities, temperature and column width. Once the components exit the column, the solvent can be removed by evaporation and the pure components can be further analyzed or identified.

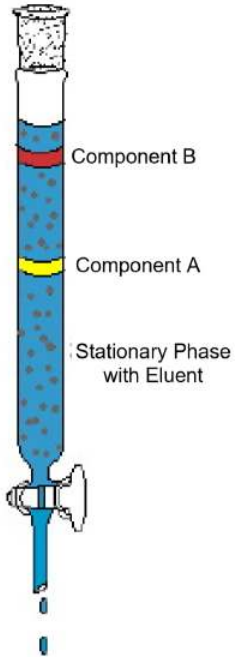


Figure 1: A typical column chromatography experiment demonstrates the separation of a two-component mixture.

Tentative identification of the components can be achieved by comparing the unknown mixture a carefully prepared known mixture: if a known component has the same retention time as an unknown component under the same conditions, it is probable—but not conclusive—that the two components are the same. Further analysis may be needed to confirm this hypothesis. If the known and the unknown have different retention times, then it is not likely that the two components are identical.

Other variations of chromatography use capillary action—the attraction of a liquid to a solid surface—to pull a solvent through solid material. An informal version of paper chromatography can be observed when an ink-written page comes in contact with water or other liquids. The ink runs and several colors are separated in the ink streak.

The diagram below (Figure 2) shows the result of a thin-layer chromatography experiment. Two black ink spots on the solid surface have had a solvent passed through them. The solvent is water or another liquid that is pulled through the stationary phase by capillary action. In this example, a piece of plastic coated with a powdered solid is used as the stationary phase. Alternatively a piece of filter paper can be used as the stationary phase. The experiment shows that the black ink is a mixture containing several different colored substances. Each component has a slightly different solubility in the mobile phase, so when the liquid is pulled through the stationary phase, each component moves at a different rate, separating the ink into spots of different colors.



Figure 2: Thin-layer chromatography of black ink after development. This picture demonstrates a common problem where the spots widen as they move up the plate, eventually merging at the top of the plate.

In this experiment, similar principles are used to separate several metal cations by a paper chromatography procedure. The metal ions—Ag^{+}, Fe^{3+},Co^{2+}, Cu^{2+}, and Hg^{2+}—have differing solubility in the mobile phase—aqueous HCl with ethyl and butyl

alcohol—and will move at different rates up the paper. The different metal-ion solubilities are probably due to the formation of various compounds with the chloride ion and their varying ability to dissolve in the organic solvent.

A diagram showing how to prepare the paper is shown below. Standard solutions containing each of these ions will be spotted onto the paper using a capillary tube, along with a standard solution containing all five ions. An unknown will also be spotted onto the paper. Once the paper is prepared, it will be developed by placing the paper into the eluent. After 75-90 minutes, the paper is visualized by wetting it with an aqueous solution containing potassium iodide, KI, and potassium ferrocyanide, $K_4[Fe(CN)_6]$. The unique color observed for each ion is produced by a chemical reaction with the visualization solution. This is one useful way to identify which ions are present in an unknown mixture.

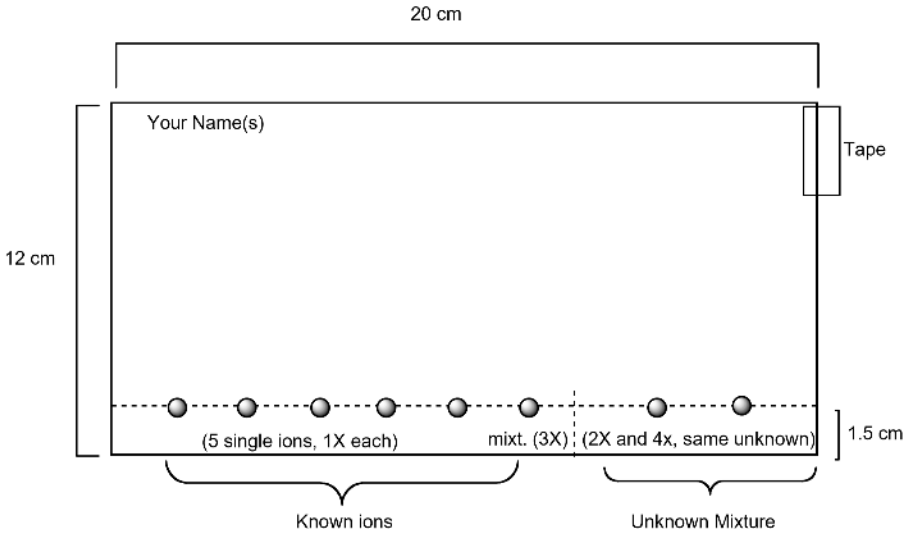


Figure 3: Diagram showing how to prepare the paper for the chromatography experiment

The distance the ion moves up the paper can also be used to identify the ion. However, since students will develop their chromatography experiments for different amounts of time and under slightly different conditions, each student will have somewhat different measured distance for a given ion. The ratio of the distance moved by an ion (D) to the distance moved by the solvent (F, solvent front) is characteristic and should be nearly the same for all students. This ratio is called R_f, or **"retention factor."**

$$R_f = \frac{D}{F} \tag{12.4.1}$$

Procedure

Materials and Equipment

Chemicals: 0.1 M aqueous solutions of $AgNO_3$, $Hg(NO_3)_2$, $Fe(NO_3)_3$, $Co(NO_3)_2$, and $Cu(NO_3)_2$, each with dedicated capillary tubes; eluting solution (aqueous HCl with ethyl and butyl alcohol); visualizing solution (aqueous solution of KI and $K_4[Fe(CN)_6]$).

Equipment: Clean piece of chromatography paper; disposable Latex gloves (nitrile gloves are vailable in the stockroom for people with allergies to Latex); 600 mL beaker; plastic wrap; forceps or beaker tongs; ruler *

*Items obtained from stockroom

Safety

Avoid contact with the metal ion solutions, the eluting solvent, and the visualizing solution. Wear disposable gloves to touch your chromatogram after the elution occurs and for the remainder of the experiment. Do not breathe the vapors of the eluting solvent or the visualizing solution. Place the wet chromatogram on a paper towel, not directly on the laboratory bench. Use the visualizing solution only in the space provided by your instructor. Dispose of the gloves and chromatogram in the specified waste container after the experiment is finished. Wash your hands thoroughly after contact with all solutions in this lab.

Preparation of the paper for chromatography

1. Each pair of students should obtain a piece of filter paper with the dimensions shown in Figure 3. Make sure the paper is clean and without tears or folds. Use a *pencil*—not a pen—and a ruler to draw a line across the paper one cm from the long edge of the paper. You will spot the metal ion solutions on this line. Write your name in pencil in the upper left-hand corner of the paper.
2. Practice spotting water and/or ion solutions onto a strip of filter paper so that you know how to create spots of the correct size. Use glass capillary tubes to spot the ions onto the paper. Solution is applied by lightly and quickly touching a capillary tube containing the solution to the line you drew on the paper. The spots should be between 5–8 mm in diameter. Spots larger than this will excessively spread out during the experiment and make analysis difficult.
3. Known 0.1 M aqueous solutions of $AgNO_3$, $Hg(NO_3)_2$, $Fe(NO_3)_3$, $Co(NO_3)_2$, and $Cu(NO_3)_2$ are provided in test tubes, each containing two or three capillary tubes. Starting on the left, mark the identity of the ion underneath each spot with a pencil; then spot each known ion carefully onto the line. Be careful to avoid contaminating the capillary tube with other ions and replace the capillary tubes back into the correct test tube. A test tube containing a known mixture of all five ions is also provided with a set of capillary tubes. Spot this mixture onto the line as well. Because this solution is more dilute than the single-ion known solutions, apply the known mixture three times, letting the spot dry between each application. A heat lamp will help to dry the spot more quickly.
4. Several unknowns are also provided in test tubes, along with capillary tubes. Your instructor will tell you which unknown should be used. The unknowns will contain between one and four cations, and are more dilute than the single-ion known solutions. The unknown will also need to be applied two and four times for the two trials, letting the spot dry between each application. In case of error, you should spot the unknown in two places along the line so that two trials are available for analysis.

Developing the chromatography paper

1. Place a piece of tape along the upper right edge, as shown in Figure 3. Then form a cylinder by connecting the two short edges of the paper with the tape. Make sure the edges do not touch. The paper should look similar to Figure 4.

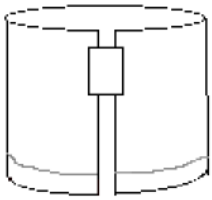


Figure 4: Folded paper should look like this prior to developing the experiment.

2. Obtain 15 mL of the eluting solution. Carefully pour some of this solvent into a 600 mL beaker and carefully swirl for a second or two. **Caution:** Do not breathe the vapors from this solution! Make sure that the level of the liquid will be *below* the spot line on the paper once the paper is placed in the developing chamber.
3. Place the paper cylinder into the beaker with the marked edge down. The spots should be *above* the level of the solvent. The paper should not be touching the sides of the beaker. Carefully cover the beaker with plastic wrap and place it in the hood for 75-90 minutes. The solvent should start to move up the paper. Once the beaker is covered, make sure it is level and do not disturb it during the development period. Your instructor may have an assignment for you to work on while you wait.

Visualization and analysis of the paper

1. Once the development period is over, wear disposable gloves and remove the paper from the beaker. Latex gloves are available in the lab and nitrile gloves are available in the stockroom for people with Latex allergies. Let any solvent drip back into the beaker, then remove the tape. Lay the chromatography paper on a paper towel and immediately mark the solvent front with a pencil. Pour the used eluting solvent into the waste container provided. Dry the paper under a heat lamp in the hood. **Caution:** Do not breathe the vapors! Be careful not to burn the paper under the lamp.
2. Once the paper is dry, bring it to the visualization station on the paper towel. *Briefly* dip the paper into the visualizing solution located in a shallow dish in the fume hood. Lift the paper out of the solution immediately and let any excess drip off at the station. Place the wet paper onto a dry paper towel and dry it under a heat lamp immediately, then carry it to your bench for analysis.

3. Find each known single-ion first and record the colors you observe. Some spots may fade over time, so record the colors while the paper is still wet. Measure the distance each spot moved, D, with a ruler. Measure to the center of each spot. Record your data in the data table.
4. Measure the distance to the solvent front, F. The value of F should be approximately the same across the entire paper. Use these values to calculate the Rf for each ion. Make your measurements as shown in Figure 5. Each observed spot has its own R_f value. Record your results in the data table.

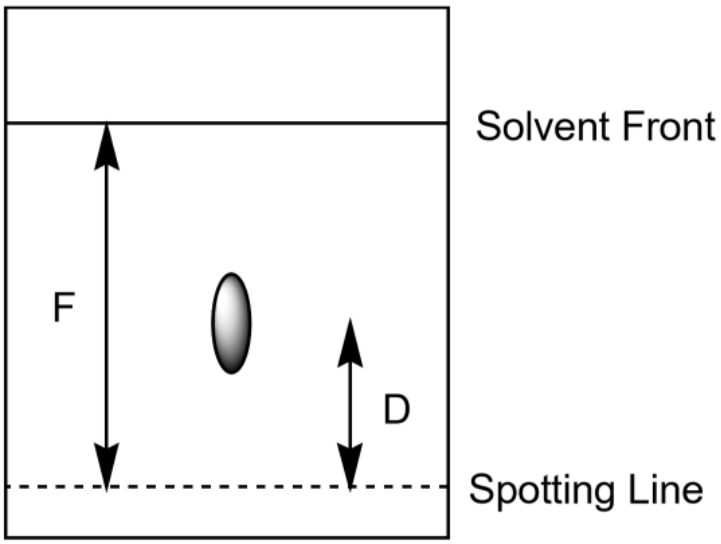


Figure 5: Measurement of distances used in the calculation of R_f for a spot.

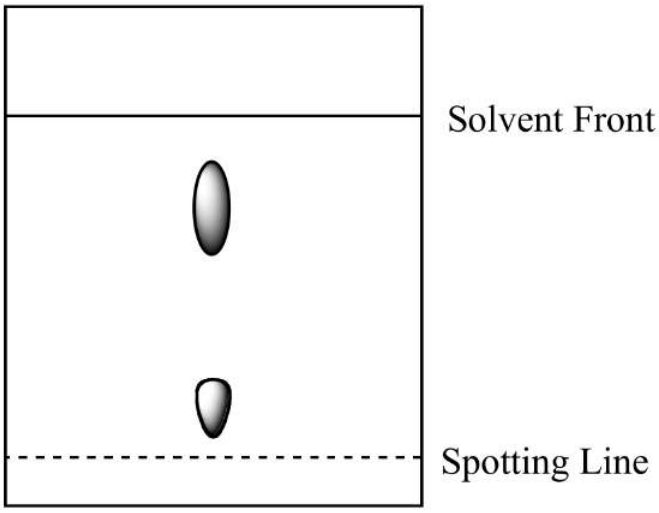
5. In the lane containing the mixture, find each ion and record the distance moved by each ion. Calculate the R_f for each ion in this lane. The values should closely match those observed in the single-ion knowns.
6. In the lane containing the unknowns, locate the center of each spot observed and record its distance and calculate the R_f values. Use the lane that has the clearest spots. The color and R_f values for the unknown spots should closely match some of the known ions. You should now be able to identify which ion or ions are found in your unknown. Record your data in the corresponding table.
7. Make a sketch of your chromatogram in the space provided on your lab report form, being sure to indicate the position and approximate size and shape of each spot on the paper. Dispose of the paper in the designated waste container.

Cleanup

Place the chromatography paper and the used gloves in the waste container provided. The used eluting solution should already have been placed into another waste container. Note that two different waste containers are provide for this experiment so be sure to read the labels so you will use the correct one! Be sure to wash your hands thoroughly before leaving the laboratory.

Pre-laboratory Assignment: Paper Chromatography

1. A two-component mixture is analyzed by paper chromatography. Component A is more soluble in the mobile phase than component B. The following result is obtained. Calculate the R_f for each component and label the identity of each spot.



2. The mixture from question 1 (above) is analyzed by column chromatography using the same mobile phase and a silica gel stationary phase. Which component has the shorter retention time? Explain your answer. (Assume the same type of binding forces between the components and the stationary phase are present in both paper chromatography and column chromatography.)
3. An unknown liquid sample is analyzed using paper chromatography using solvent X as the mobile phase. One spot is observed after the plate is developed and visualized. The same unknown substance is re-analyzed using solvent Y as the mobile phase. This time, three spots are observed after the plate is developed and visualized.

- Is the unknown sample a pure substance or a mixture? Explain your answer, including a possible reason for the different observations in the two experiments.

Lab Report for Paper Chromatography

Data, Observations, Calculations and Analysis Known Ions

Known Ions

Ion	Spot Color (Stained)	D (Single-Ion)	F (Single-Ion)	R_f	D (Ion Mixt.)	F (Ion Mixt.)	R_f
Ag^{+}							
Co^{2+}							
Cu^{2+}							
Fe^{3+}							
Hg^{2+}							

Unknown ID Code _____

Unknowns

Spot Number (from lowest R_r)	Spot Color (Stained)	D (Unknown)	F (Unknown)	Rf	Identity of Spot
1.					
2.					
3.					
4.					

Spot Number (from lowest R_r)	Spot Color (Stained)	D (Unknown)	F (Unknown)	Rf	Identity of Spot
5					

- Sketch of Chromatogram:

Summary:

Unknown ID Code	Ions Identified

Questions and Conclusions

1. What criteria were used to identify the ion(s) found in your unknown? Explain your answer in at least three sentences. Include any difficulties in identifying any ions.
2. If you let the experiment run for only 30 minutes, what would be the likely result? Would any problems arise in identification of the unknown?
3. If Co^{2+} and Cu^{2+} spots were the same color, would the identification of an unknown be any more difficult? Explain your answer.

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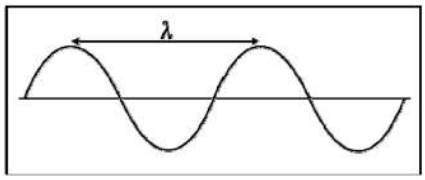
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12.5: Lab 5 - Flame Tests of Metal Cations

Objectives

- To perform flame tests of metal cations in order to observe their characteristic colors,
- To match the flame colors observed to an appropriate wavelength of visible light, and then perform calculations to determine the frequency and energy of the emitted photons,
- To relate these results to the types of electronic transitions occurring in these elements,
- To practice writing electron configurations for these (and other) elements.

Electromagnetic radiation is composed of perpendicular waves oscillating in the electric and magnetic fields (through space or matter). These waves are characterized by their wavelength (λ) and frequency (ν). Wavelength is defined as the distance between successive crests (or troughs) on a wave, and is measured in meters. Frequency is defined as the number of waves that pass a given point every second, and is measured in 1/seconds, or Hertz (Hz).



All electromagnetic waves travel at the speed of light (c), or $2.998 \times 10^8 m/s$. The relationship between the wavelength, frequency and speed of an electromagnetic wave is given by the equation:

$$c = \lambda \times \nu \tag{12.5.1}$$

Electromagnetic radiation also occurs as discreet packets of energy (or quanta) called photons. The energy per photon (in Joules) is given by the equation:

$$E_{photon} = h \times \nu \tag{12.5.2}$$

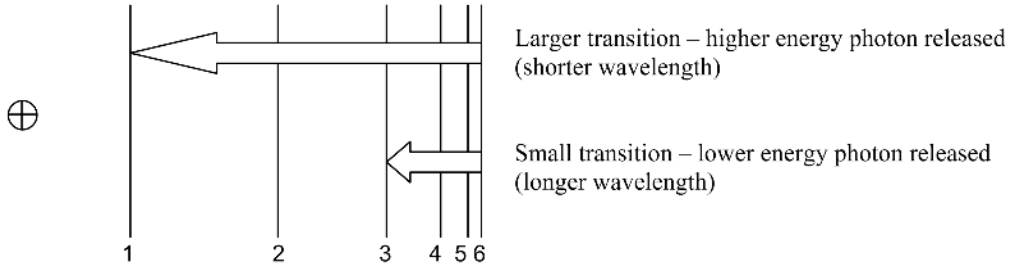
Here, h is Planck's constant, which has a value of $6.626 \times 10^{-34} J \cdot s$.

Visible light is the most familiar example of electromagnetic radiation. Differences in the wavelengths of visible light are manifested as different colors, shown in the Color Spectrum below (colors can be seen in the PDF document on-line). Other examples of electromagnetic radiation include X-rays, ultraviolet light, infrared light, microwaves and radio waves.

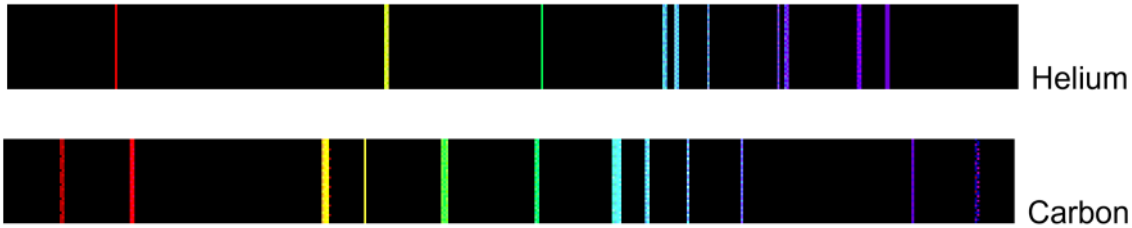


So, how does electromagnetic radiation relate to flame tests? Well, when an atom (or ion) absorbs energy, its electrons can make transitions from lower energy levels to higher energy levels. The energy absorbed could be in the form of heat (as in flame tests), or electrical energy, or electromagnetic radiation. However, when electrons subsequently return from higher energy levels to lower energy levels, energy is released predominantly in the form of *electromagnetic radiation*.

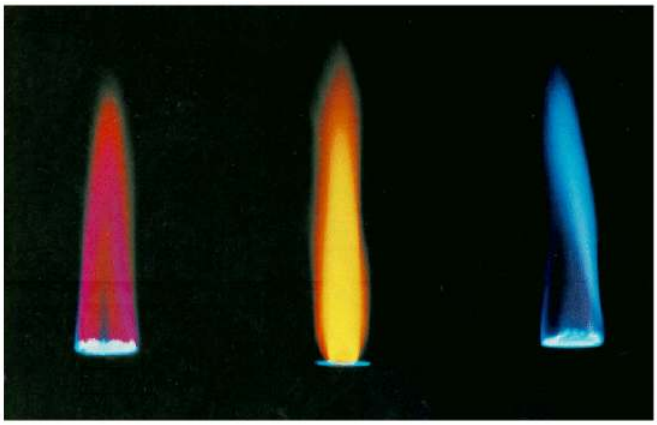
The spacing between energy levels in an atom determines the sizes of the transitions that occur, and thus the energy and wavelengths of the collection of photons emitted:



If emitted photons are in the visible region of the spectrum, they may be perceived as lines of different colors (note that photons outside the visible spectrum may also be emitted, but cannot be seen by eye). The result is called a line *emission spectrum*, and can serve as a 'fingerprint' of the element to which the atoms belong. For example, the line spectra shown below for the elements helium and carbon are clearly quite different.



Unfortunately, techniques more sophisticated than those used in this lab are required to obtain such line spectra. To the naked eye, when an element is vaporized in a flame (or an electrical discharge) the emission spectrum will appear to be just *one color*. For example, helium gas when excited by an electrical discharge emits light that appears an orange-peach color. This one color results from a *combination of all lines* of the emission spectrum, in proportion to their intensities. As many elements will still produce distinctive colors under such conditions, simple flame tests can be used to identify these elements. In fact, flame tests were used to identify elements long before the invention of modern techniques, such as emission spectroscopy.



Procedure

Materials and Equipment

Looped platinum or nichrome wires, wash bottle with distilled water, Bunsen burner, and the following solutions: $\mathrm{LiCl}\,(aq)$, $\mathrm{NaCl}\,(aq)$, $\mathrm{KCl}\,(aq)$, $\mathrm{CuCl_2}\,(aq)$, $\mathrm{BaCl_2}\,(aq)$, $\mathrm{CaCl_2}\,(aq)$.

Safety

Exercise appropriate caution when using the Bunsen burner.

Experimental Procedure

This experiment will be performed as an instructor demonstration only.

Your instructor will dip a looped wire into one of the solutions supplied, and then hold it in the Bunsen burner flame. Students will record the dominant flame color observed. The table below contains a list of appropriate colors to choose from. Your instructor will then repeat this for the remaining five solutions, using a fresh looped wire each time.

Analysis: For each metal cation tested, obtain the wavelength of light corresponding to the observed flame color from the table below. Note that the wavelengths supplied here are in nanometers. Using these wavelengths, calculate the frequency and energy of the photons emitted during the flame tests. Finally, answer the questions and perform the exercises as indicated on your Report form.

Dominant Color	Approximate Wavelength (in nm)*
Red	701
Red-Orange	622
Orange	609
Orange-Yellow	597
Yellow	587
Yellow-Green	577
Green	535
Green-Blue	492
Blue	474
Blue-Violet	455
Violet	423

*Wavelength values here are given for the mid- range of the color indicated.

Pre-laboratory Assignment: Flame Tests of Metal Cations

1. In this lab, you will perform flame tests of several different metal cations. The characteristic colors observed are due to emitted electromagnetic radiation from the excited metal cations.

- In this lab, how do the metal cations become "excited"?
- Circle the correct responses to complete the following statement:

EM radiation is emitted when electrons make transitions from low / high to low / high energy levels.

2. In a flame test, the element Boron emits EM radiation that is predominantly green in color.

- Use the Table in the Procedure to obtain the wavelength of this emitted radiation (in nm). Then convert this wavelength from nm to m.
- Calculate the frequency of this emitted green radiation, in s^{-1}. Show your work.
- Calculate the energy of this emitted green radiation, in J. Show your work.

3. Circle the type of EM radiation that has the property indicated:

- Higher energy: UV or Visible light?
- Longer wavelength: Orange or Blue light?
- Higher frequency: IR or Yellow light?

Lab Report: Flame Tests of Metal Cations

Experimental Data and Observations

Solution	Dominant Flame Color	Wavelength (nm)
$LiCl$		
$NaCl$		
KCl		
$CuCl_2$		
$BaCl_2$		
$CaCl_2$		

Data Analysis

Using the wavelengths recorded above, calculate the corresponding frequencies and photon energies for the emitted radiation observed for each compound tested. Record the results of your calculations in the table below.

Solution	Wavelength (m)	Frequency (s^{-1})	Photon Energy (J)
$LiCl$			
$NaCl$			
KCl			
$CuCl_2$			
$BaCl_2$			
$CaCl_2$			

Show a set of sample calculations for $LiCl$ only below. Clearly show any equations you have used.

- Wavelength (in m):
- Frequency (in s^{-1}):
- Photon Energy (in J):

Questions

1. Complete the following paragraph by circling the correct responses:

In this experiment, the metal cations in the solutions were initially in the (ground, excited) state. When placed in the flame, the metals then (absorbed, emitted) energy as (electricity, heat, EM radiation). When this occurred, electrons made transitions from (low, high) energy levels to (low, high) energy levels. The metals were then in the (ground, excited) state. The electrons in these metals then made transitions from (low, high) energy levels to (low, high) energy levels, resulting in the (absorption, emission) of energy as (electricity, heat, EM radiation).

2. What evidence is there that the colors observed in the flame tests are due to the metals, and not the nonmetals in the compounds tested?
3. Which metal cation was observed to emit radiation with the *shortest* wavelength?

Compared to the other metals studied, did the radiation emitted by this metal cation (identified above) have

- the highest or lowest frequency?
- the highest or lowest photon energy?

From this, would you conclude that the relationships between the following are direct or inverse?

- wavelength and frequency
- frequency and photon energy
- wavelength and photon energy

4. The energy, wavelength and frequency of an emitted photon are all related to the size of the electronic transition (high → low energy levels) occurring in the metal cation. Based on your observations, in which metal did the *smallest* electronic transitions occur? Briefly explain your response.
5. When heated in a flame, the element Indium emits electromagnetic radiation with a distinctive indigo blue color (the name indium is derived from the word indigo). The emitted photons that give rise to this color have energies of $4.405 \times 10^{-19} J$. Calculate the wavelength of this radiation in *nanometers*.

Exercise

In the Bohr Model of the atom, electrons occupy fixed orbits around the nucleus called energy levels. However in the Quantum Mechanical Model of the atom, electrons occupy orbitals. Orbitals are grouped by size and shape into shells and subshells (or, levels, and sublevels). Electron configurations and orbital diagrams are used to show the arrangement of electrons in shells (levels), subshells (sublevels) and orbitals for specific atoms.

Write complete electron configurations and abbreviated orbital diagrams for each of the elements given below. Circle the valence electrons in your complete electron configurations.

Chlorine

complete configuration:

abbreviated orbital diagram:

Tin

complete configuration:

abbreviated orbital diagram:

Selenium

complete configuration:

abbreviated orbital diagram:

Cobalt

complete configuration:

abbreviated orbital diagram:

Boron

complete configuration:

abbreviated orbital diagram:

Bismuth

complete configuration:

abbreviated orbital diagram:

Magnesium

complete configuration:

abbreviated orbital diagram:

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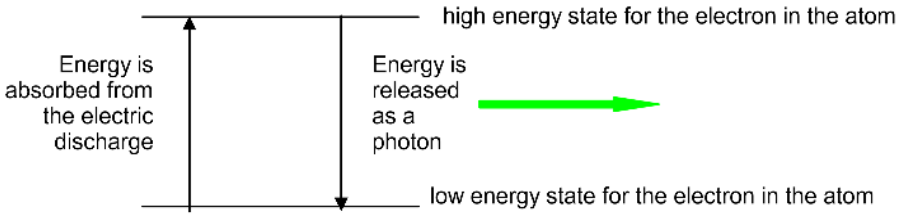
12.6: Lab 6 - Atomic Emission

Learning and Experimental Objectives

- define and evaluate the electromagnetic properties of light
- assess the relationship between electromagnetic radiation and color
- evaluate the color of light in terms of its interaction with matter
- measure the spectral position of atomic lines using gas lamps and spectroscopes
- compare experimentally observed spectral lines with known values
- calculate the change in principal quantum number of observed and theoretical lines

Atomic Emission Spectra Introduction

Electrons in atoms normally occupy the lowest energy states possible. Such an atom is said to be in the ground state. However, electrons can be excited to high energy states when they absorb excess energy. The excess energy can be provided by heat, light, or electrical discharge. The electrons then return to lower energy states, eventually returning all the way to the ground state. As the electrons return to lower energy states, they release their excess energy. Often, this excess energy is released in the form of light, with each atom or molecule releasing a single photon of light for each electron energy transition it makes.



In the hydrogen discharge tubes used in this experiment, the energy of the electric discharge first dissociates the H_2 molecules into H atoms, then excites the electrons in the H atoms into high energy states.

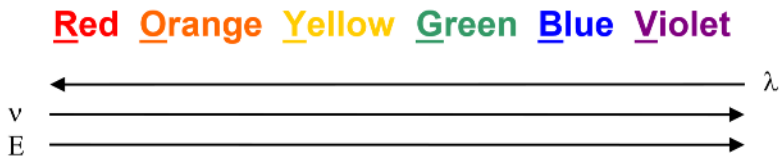
Due to conservation of energy, the amount of energy in an emitted photon will exactly match the amount of energy lost by the electron as it moves to the lower energy state. Different colors of light are associated with different photon energies. For example, a single photon of blue light has more energy than a single photon of red light. Thus, the color of the light emitted by a particular atom depends upon how much energy the electron releases as it moves down to a lower energy level. The energy levels that are allowed for each atom depend upon the number and arrangement of protons and electrons in the atom. Thus, each element has different energy states available to it, so each element releases photons of different color when its atoms return to their lower energy states. Since each atom has many excited states (high energy levels) available to it, several colors of light can be emitted by each element. The set of individual colors emitted by an element is called its spectrum. Since the spectrum of each element is unique, spectra can be used like fingerprints to identify unknown elements.

Wavelengths of Light

Light is one kind of electromagnetic radiation. The wavelength of radiation determines what kind of radiation it is. The human eye is able to detect only a narrow range of wavelengths of electromagnetic radiation, those from about 400nm to about 700nm. Radiation with wavelengths less than 400nm is classified as ultraviolet, x-ray, or γ-rays, while radiation with wavelengths longer than 700 nm is classified as infrared radiation, microwaves, and radio waves. In this experiment, we use our eyes to detect the radiation emitted by excited atoms and, therefore, we work only with visible light

The color of light is related to its wavelength (λ), which is related to its frequency (ν) and the energy of its photons (E). Shorter wavelengths of light (at the blue end of the visible spectrum) have higher frequencies and higher photon energies while longer wavelengths of light (at the red end of the spectrum) have lower frequencies and less energy per photon.

Visible Light



It is easy to convert between photon energy, wavelength, and frequency using the following relationships where c = the speed of light = $2.998 \times 10^{8} m/s$ and h = Planck's Constant = $6.626 \times 10^{-34} Js$.

$$\lambda \nu = c \tag{12.6.1}$$

and

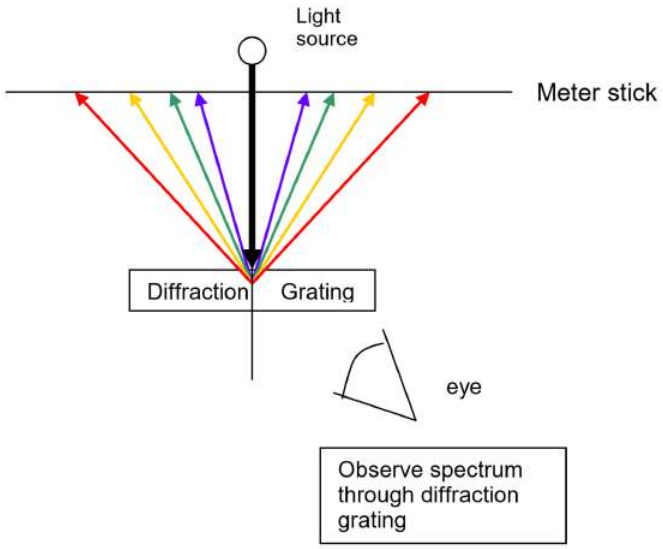
$$E = h\nu \tag{12.6.2}$$

These two relationships combine to give a third:

$$E = \frac{hc}{\lambda} \tag{12.6.3}$$

Thus, the spectrum of an element can be stated by listing the particular wavelengths of light that its atoms emit.

To measure these wavelengths in the laboratory, we must first separate them. To the naked eye, the various wavelengths (colors) of light emitted by an element are mixed together and appear as a single color that is a combination of the component colors. If we view the light through a prism or a diffraction grating, however, the individual wavelengths are separated. A diffraction grating is a piece of glass or clear plastic with many very narrow and closely spaced lines on it. As the light emerges after being reflected by the grating, these tiny lines cause the reflected light to interfere with itself in such a way that the different wavelengths of the light to appear in different positions to the left and right of the original direction in which the light was traveling. See the figure 1 below.



Using a light source that contains known wavelengths of light, we can measure exactly where each known wavelength appears along a meter stick. Since this position depends upon the wavelength in a linear way, a graph of wavelength vs. position of the spectral line will yield a straight line. Once the best fit straight line has been determined, the equation of this line can then be used to convert positions of other spectral lines to wavelength. For example, using the same apparatus and without moving the relative positions of the meter sticks, diffraction grating and lamp, it is possible to view the spectrum of a new element, measure where its spectral lines occur on the meter stick, and then read the graph or use the equation of the line to determine the wavelength to which each of those positions corresponds. The calibration graph is therefore an integral part of the spectroscope. Positions are measured using the meter sticks, then wavelengths are determined from the positions using the graph itself or the equation of the best fit line for that graph.

Bohr's Theory Introduction

For atoms that contain only one electron, the theory of atomic structure proposed by Niels Bohr can be used to calculate wavelengths for transitions between particular electronic energy levels of the atom. In this experiment, the only one-electron atom we will consider is hydrogen. (Note, there are other one-electron "atoms" if you consider ions such as He^{+}, Li^{2+}, etc.)

Using Bohr's theories for hydrogen, you should find a close match between the calculated wavelengths and those that you measure experimentally. To calculate the wavelengths of light emitted by hydrogen atoms, recall that the energy of an electron in the n-th energy level of a one-electron atom is given by:

$$E_n = -\frac{Z^2 R}{n^2} \tag{12.6.4}$$

where R is the Rydberg constant = $2.18 \times 10^{-18} J$, Z is the nuclear charge, and n = 1, 2, 3, ..., ∞. For hydrogen, the nuclear charge is 1 so this equation becomes:

$$E_n = -\frac{R}{n^2} \tag{12.6.5}$$

The change in energy for the electron when it makes a transition from one level to another is given by its subtracting its initial energy from its final energy:

$$\Delta E_{\text{electron}} = E_f - E_i \tag{12.6.6}$$

By conservation of energy, the energy of the photon emitted as this electron drops to a lower energy level must equal the change in energy for the electron. However, since photon energies must be a positive quantity, the absolute value of the change in energy for the electron must be used:

$$E_{\text{photon}} = |\Delta E_{\text{electron}}| \tag{12.6.7}$$

Once the energy of the photon is known, it is readily converted into a wavelength as discussed earlier:

$$E_{\text{photon}} = \frac{hc}{\lambda} \tag{12.6.8}$$

or

$$\lambda = \frac{hc}{E_{\text{photon}}} \tag{12.6.9}$$

Because there are many energy levels possible for the electron in a hydrogen atom, and because the electron could jump from any higher n to any lower n, there are many lines in the spectrum of hydrogen. However, most of these lines occur at wavelengths which our eyes cannot detect (either infrared or ultraviolet). The visible portion of the spectrum which you will observe in this experiment was the first to be studied by scientists since it is the only portion which can be seen with the naked eye. This series of spectral lines is named for one of the first scientists to study it and is called the Balmer series. Note that all of the spectral lines in the Balmer series involve transitions from a higher n level to the n=2 level. You will need that information to complete the calculations for your lab report.

Procedure

Materials and Equipment

Spectroscope; Power Supply; Gas Lamps: Mercury, Hydrogen, Polyatomic Element (Ar), Unknown (Kr, Ne, He)

Safety

- Gas lamps break easily so handle them with care.
- The lamps will also be very hot to the touch after they have been on for a while. Turn off the power supply when you are not using the lamp!
- Only change lamps with the power supply turned off.
- The power supplies are high voltage devices.

- Never put anything other than the lamp into the power supply socket.

Part A: Verification of Hg Line Position

1. Set up the power supply and the spectroscope as shown in Figure 2 Place the power supply on a stir plate so that the lamp will be at the right height for the spectroscope.


Figure 2. Experimental Apparatus.

2. Carefully place the Hg lamp in the power supply by inserting the bottom of the lamp into the lower socket and pressing down on the spring until the top of the lamp clears the housing of the top socket and can be inserted into the top socket. Turn on the power.
3. Open the slit on the end of the spectroscope by loosening the set screw. Place the spectroscope in front of the lamp
4. Looking through the eye piece down the left of the spectroscope, adjust the position of the base so that the maximum amount of light is coming through the slit. Make sure the spectroscope is screwed down tightly to the base.
5. Taking care not to shift the position of the base, close the slit until the lines on the scale off to the right are clearly separated. With the slit wide open they will be blurred, but as you close it down they will become crisper and more defined.
6. Record the positions of the Hg lines in your notebook. The Hg line spectrum should appear as shown in Figure 3. Each tick mark on the scale represents 20 nm. Verify that the positions observed are close to the known values. You can use a smart phone to take a photograph of the spectrum and then zoom in to try to get a more accurate estimate.
7. Turn off the power supply and allow the lamp to cool.
8. Within how many nanometers can you estimate your line positions?

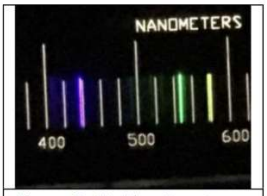


Figure 3. Hg Line Spectrum

Part B: Hydrogen

1. With the power supply unplugged, carefully exchange the lamp for hydrogen. The previous lamp used may be hot!!
2. Record the colors and positions of the lines for hydrogen.

Part C: Polyelectronic Elements

1. Exchange the lamp for a known polyelectronic lamp such as argon, or helium.
2. Record the positions of the lines for this element.

Part D: Unknown

1. Exchange the lamp for an unknown.
2. Record the positions of the lines for this element.

Pre-laboratory Assignment: Atomic Spectra

1. Calculate the energy of the n=1 level for an electron in a hydrogen atom.
2. Calculate the energy of the n=2 level for an electron in a hydrogen atom.
3. Calculate the energy change when an electron in a hydrogen atom moves from n=2 to n=1.
4. Calculate the wavelength of the light that an electron in a hydrogen atom would emit if it moved from n=2 to n=1.
5. We can't see the light emitted by hydrogen atoms when the electrons move from any upper level to the n=1 level. Why not?
6. What is the Balmer series equation?
7. Calculate the energies of the first four lines of the Balmer series
8. In some observations of atomic emission spectra, two lines of the same color appear very close to each other. Consider the difference in the energy value of a spin up and spin down electron upon transition. Explain the cause of the “doublet” of lines that occur in some spectra
9. Find images of the atomic spectra of Kr, Ne, and He on the internet. Sketch these images in your lab notebook either in color or with notes on their color. You may also want to save images of the spectra to your phone for reference during your lab.

Post-Lab:

Record the following tables and calculations in your notebook.

Part A: Mercury

Table 1: Emission Spectrum of Mercury Lines

Color	Position from Center of Spectroscope (cm)	Wavelength (nm)
violet		404.7
blue		435.8
green		546.1
yellow		579.0

1. What is the approximate margin of error: ± nm of your observed positions compared to the known values for mercury.
2. What does this margin of error tell you about how accurately you will be able to match measured values to theoretical values?

Part B: Spectrum of a Single Electron Element: Hydrogen

1. Record observed wavelengths for each color in hydrogen spectrum

Table 2: Position of Hydrogen Lines

Color	Observed Position (Measured Wavelength)
red	
green-blue	
violet 1	
violet 2	

2. Calculate the energy values and wavelengths for the following transitions in Table 3

3. Show step-by-step calculations and formulas used in one of the transitions from Table 3

Table 3: Calculated energies of the first four transitions of the hydrogen atom

$n_{initial}$	n_{final}	Change in Energy (J)	Theoretical Wavelength
3	2		
4	2		
5	2		
6	2		

4. How do your observed and theoretical values compare? What is an approximate percent error? (Fill out Table 4 below to answer these questions)

5. Show step-by-step calculations and formula used for one of the percentage error from Table 4

Table 4: Comparison of Observed and Theoretical Results for Hydrogen

Observed Wavelength (from Table 2)	Calculated Wavelength (from Table 3)	Percentage Error in Measured Wavelength

Observed Wavelength (from Table 2)	Calculated Wavelength (from Table 3)	Percentage Error in Measured Wavelength

Part C: Spectrum of a Polyelectronic Element

Table 5: Atomic Spectrum of _____

Color	Observed Position	Known Position (From the website)

Go to the NIST website http://physics.nist.gov/PhysRefData/...ement_name.htm and select your element. Click the yellow box that says Strong Lines. The values listed are in angstroms, Å. 10 Å = 1 nm. There are many more lines listed than the those you observed. Your observed lines will be the most intense ones from the visible region (4000 to 7000 Å). Record the wavelengths that best match your observed positions in the "Known Position" column of your Table 5. If there is more than one entry for a given wavelength; your value will correspond to the most intense one.

Part D: Unknown

Unknown ID :_____

1. Record colors and observed positions of them for the given unknown in Table 6

2. Using your sketches from the pre-lab, match your unknown to either Kr, Ne, and He.

3. Lookup the most intense lines on the NIST website the same as you did for Part C. Record the values as your "Known Positions".

4. What is the identity of your unknown? _____

Table 6: Positions of the lines of the unknown

Color	Observed Position	Known Position (From the website)

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- **14A: Atomic Emission Spectra (Experiment)** by Santa Monica College is licensed CC BY-NC 4.0.

12.7: Lab 7 - Chemical Nomenclature Dry Lab

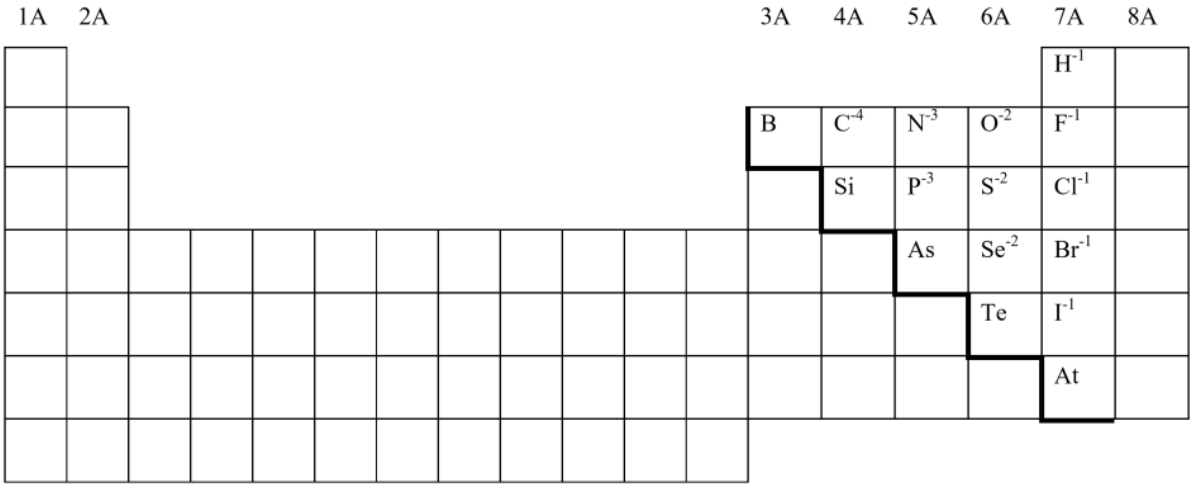
Nomenclature of Ionic Compounds Overview

Ionic compounds are composed of ions. An ion is an atom or molecule with an electrical charge. Monatomic ions are formed from single atoms that have gained or lost electrons. Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons.

Negative ions are called anions, and are formed when an atom or molecule gains electrons. All non- metals form negatively charged ions. Positive ions are called cations, and are formed when an atom or molecule loses electrons. All metals form positively charged cations. Ions with opposite charges (positive metal cations and negative non-metal anions) will experience a strong electrostatic attraction and form an ionic bond, which leads to the formation of the ionic compound.

Non-metal Anions

Non-metals will form anions with only one possible negative charge. The following Periodic Table shows the charges for non-metal anions commonly found in ionic compounds:

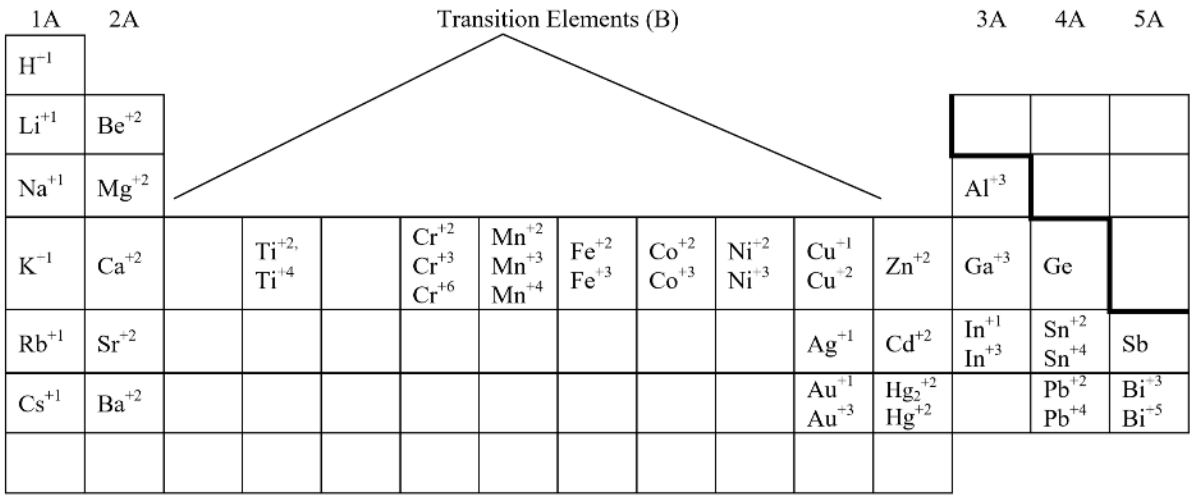


Note that

- The magnitude of the negative charges on these anions is equal to 8 minus their Group Number.
- The names of these anions are based on the element names, but the endings are all changed to –ide

Metal Cations

Most (but not all) main group metals will form cations with only one possible charge. Most (but not all) transition metals will form cations with more than one possible charge. The following Periodic Table shows the charges for metal cations commonly found in ionic compounds:



Note that

- The magnitude of the positive charge on the main group metal cations is generally equal to their Group Number.
- The names of metal cations with only one possible charge are the same as the names of the metals themselves.
- For metal cations with more than one possible charge, the ion charge must be indicated in the ion name. In the IUPAC system, the ion charge is indicated in the name as Roman numerals in brackets.

Polyatomic Ions

Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons. The table below includes a list of common polyatomic ions that must be memorized.

OH^{-1}	Hydroxide	O_2^{-2}	Peroxide
CN^{-1}	Cyanide	CO_3^{-2}	Carbonate
SCN^{-1}	Thiocyanate	SO_3^{-2}	Sulfite
HCO_3^{-1}	Bicarbonate (Hydrogen Carbonate)	SO_4^{-2}	Sulfate
HSO_3^{-1}	Bisulfite (Hydrogen Sulfite)	$S_2O_3^{-2}$	Thiosulfate
HSO_4^{-1}	Bisulfate (Hydrogen Sulfate)	$C_2O_4^{-2}$	Oxalate
$C_2H_3O_2^{-1}$	Acetate	CrO_4^{-2}	Chromate
NO_2^{-1}	Nitrite	$Cr_2O_7^{-2}$	Dichromate
NO_3^{-1}	Nitrate		
MnO_4^{-1}	Permanganate	PO_3^{-3}	Phosphite
ClO^{-1}	Hypochlorite	PO_4^{-3}	Phosphate
ClO_2^{-1}	Chlorite		
ClO_3^{-1}	Chlorate	NH_4^{+1}	Ammonium
ClO_4^{-1}	Perchlorate	Hg_2^{+2}	Mercury (I)

Note that

- Almost all the polyatomic ions are negatively charged anions.
- Most of the names of polyatomic anions end in either –ate or –ite. The –ate's always have one more oxygen than the -ite's.

Formulas and Names of Ionic Compounds

Ionic compounds are formed when positive cations and negative anions are attracted to each other via strong electrostatic forces. This attraction is called an ionic bond. The following are the basic rules for writing the formulas and names of ionic compounds:

Writing Formulas

1. Determine the formulas and charges on the cation and anion involved in the compound.
2. Combine the ions in a ratio that results in the formation of a neutral ionic compound. In other words, the total charge of all the positive cations must equal the total charge of all the negative anions in the compound. The numbers of each element present in the compound are shown as subscripts after the element symbol.

Writing Names

1. Both the cation and anion must be named.
2. Always name the cation first, then the anion.

Nomenclature of Simple Covalent Compounds Overview

Covalent compounds are compounds formed between non-metals only. Simple binary covalent compounds contain just two different types of non-metal elements. When non-metals combine they can form several different covalent compounds. These compounds must therefore be identified with unique names and formulas.

Formulas and Names of Simple Covalent Compounds

1. Always write/name the element with more metallic character first. Metallic character increases going from right to left, and top to bottom on the Periodic Table.
2. Then write/name the second (less metallic) element, changing the ending of its name to -ide.
3. Since nonmetals often combine in different proportions to form a number of different compounds, prefixes must be included in the names to indicate the numbers of each kind of atom present.Prefixes for 1-10 atoms are given in the following table.

Number	Prefix	Number	Prefix
1	Mono	6	Hexa
2	Di	7	Hepta
3	Tri	8	Octa
4	Tetra	9	Nona
5	Penta	10	Deca

There are two important exceptions to the naming rules outlined so far:

- Never use the prefix “mono” for the first element, even if just one atom is present.
- Never use any prefixes at all for simple covalent compounds containing Hydrogen.

Please note that many simple covalent compounds have common, rather than systematic names. Please memorize the common names of the following three compounds:

- H_2O water
- NH_3 ammonia
- CH_4 methane

Covalent compounds containing more than two non-metal elements become increasingly more difficult to name, and common names for these compounds are more extensively used. You will not have to learn these yet.

Nomenclature of Acids Overview

Acids are compounds that release hydrogen cations (H^{+1}) when dissolved in water. They are all found in the aqueous state (aq).

$$HCl(aq) \longrightarrow H^{+1}(aq) + Cl^{-1}(aq) \tag{12.7.1}$$

$$\text{hydrochloric acid} \longrightarrow \text{hydrogen ions} + \text{chloride ions} \tag{12.7.2}$$

$$HNO_3(aq) \longrightarrow H^{+1}(aq) + NO_3^{-1}(aq) \tag{12.7.3}$$

$$\text{nitric acid} \longrightarrow \text{hydrogen ions} + \text{nitrate ions} \tag{12.7.4}$$

$$H_2SO_3(aq) \longleftrightarrow 2\,H^{+1}(aq) + SO_3^{-2}(aq) \tag{12.7.5}$$

$$\text{sulfurous acid} \longleftrightarrow \text{hydrogen ions} + \text{sulfite ions} \tag{12.7.6}$$

In acids the element hydrogen actually behaves like a Group 1A metal cation. Since it behaves like a +1 cation, hydrogen is always written first in the formulas of all acids. The anion in the acid can be either monatomic or polyatomic, and affects how the acid is named.

Acids containing Non-Metal Anions

These acids contain the H^{+1} cation and a monatomic non-metal anion.

Acids containing monatomic anions are named using the prefix hydro + the name of the anion with the suffix -ic + the word acid.

The formulas of these acids are obtained in an identical fashion to regular ionic compounds. The H^{+1} cation and the monatomic anion are combined in a ratio to yield a neutral compound.

Acids containing Polyatomic Anions

These acids contain the H^{+1} cation and a polyatomic anion.

These acids have names that are based on the name of the polyatomic ion in the acid. If the polyatomic ion has the ending -ate, in the acid the ending is changed to -ic + acid. If the polyatomic ion has the ending -ite, in the acid the ending is changed to -ous + acid.

Again, the formulas of these acids are obtained in an identical fashion to regular ionic compounds. The H^{+1} cation and the polyatomic anion are combined in a ratio to yield a neutral compound.

Nomenclature of Hydrates Overview

A hydrate is typically an ionic compound with a certain number of water molecules loosely bound to it.

The general formula of a hydrate is $MX \cdot nH_2O$ (s), where M is the cation in the ionic compound, X is the anion in the ionic compound and nH_2O are the water molecules loosely bound to the ionic compound.

Hydrates are named by writing the name of the ionic compound first, followed by the word “hydrate”. To indicate the number of water molecules present, prefixes must be used.

The water molecules in a hydrate can be removed with relative ease by heating the hydrate. The ionic compound that remains after heating is called an anhydrous salt.

$$MX \cdot nH_2O(s) \longrightarrow MX(s) + nH_2O \tag{12.7.7}$$

$$\text{Hydrate} \longrightarrow \text{Anhydrous salt} + \text{Free Water} \tag{12.7.8}$$

Often the anhydrous salt has a completely different color and texture from the hydrate.

Lab Report: Chemical Nomenclature

Write the names and formulas for the following inorganic compounds in the spaces provided.

Part 1: Ions and Ionic Compounds

Write formulas/charges or names as appropriate for each of the following monatomic ions.

1. Calcium ion
2. Phosphide ion
3. Iodide ion
4. Gallium ion
5. Titanium(IV) ion
6. C^{-4}
7. Rb^{+1}
8. Pb^{+4}
9. S^{-2}
10. Cr^{+2}

Write formulas or names as appropriate for each of the following ionic compounds.

1. Magnesium nitride
2. Lithium oxide
3. Aluminum sulfite
4. Copper(II) bicarbonate
5. Sodium nitrate
6. SrI_2
7. $Ba_3(PO_4)_2$
8. $(NH_4)_2O$
9. $Fe(ClO)_3$
10. $ZnCrO_4$

Part 2: Covalent Compounds

Write formulas or names as appropriate for each of the following covalent compounds.

1. Dichlorine monoxide
2. Disulfur dichloride
3. Carbon tetrafluoride
4. Phosphorus pentachloride
5. Nitrogen tribromide
6. AsI_3
7. P_4O_{10}
8. Cl_2O_7
9. $SeCl_6$
10. NO

Part 3: Acids

Write formulas or names as appropriate for each of the following acids.

1. Hydroiodic acid
2. Carbonic acid
3. Chlorous acid
4. Sulfuric acid
5. Phosphorous acid
6. HCN (*aq*)
7. $H_2C_2O_4$ (*aq*)
8. HNO_2 (*aq*)
9. $H_2Cr_2O_7$ (*aq*)
10. $HMnO_4$ (*aq*)

Part 4: Hydrates (optional – check with your instructor to see if you are responsible for this section)

Write formulas or names as appropriate for each of the following hydrates.

1. Magnesium sulfate heptahydrate
2. Copper(I) sulfate pentahydrate
3. Potassium phosphate decahydrate
4. Calcium chloride hexahydrate
5. Iron(III) nitrate nonahydrate
6. $CoSO_4 \cdot H_2O$
7. $Na_2CrO_4 \cdot 4\,H_2O$
8. $CuF_2 \cdot 2\,H_2O$
9. $Sr(NO_3)_2 \cdot 6\,H_2O$
10. $ZnSO_4 \cdot 7\,H_2O$

Part 5: Nomenclature of Ionic Compounds, Covalent Compounds and Acids

	Classification	Name or Formula
C_3O_2		
IF_7		
Rb_2CO_3		
SnS_2		
$Au(CN)_3$		
H_2CrO_4 (*aq*)		
H_3P (*aq*)		
Li_3PO_4		
Mg_3N_2		
$Ti(C_2H_3O_2)_4$		
Fe_2O_3		
NaH		
Br_3O_8		
MnS_2O_3		
NH_4NO_2		
$Cd(ClO_2)_2$		
$Ba(HSO_3)_2$		
Cu_2O		
$NiBr_3$		
$Sr(OH)_2$		
Perchloric acid		
Potassium permanganate		
Calcium hydride		
Vanadium(II) bicarbonate		
Bismuth(V) nitrate		

	Classification	Name or Formula
Rubidium peroxide		

Strontium hydrogen phosphite		
Hydrofluoric acid		
Chromium(III) thiocyanate		
Acetic acid		
Molybdenum(IV) carbonate		
Tetraiodine nonaoxide		
Diphosphorus tetrafluoride		
Aluminum sulfate		
Ammonium hydroxide		
Sodium dichromate		
Carbon disulfide		
Nickel(II) oxalate		
Barium selenide		
Silver bisulfate		

Questions

1. How are the following types of compounds recognized from their *formulas*?

- Ionic
- Covalent
- Acid

2. When do *parentheses* appear in the formulas of ionic compounds?
3. Do *Roman Numerals* appear in the names of ionic or covalent compounds? Explain why they are used.
4. Do *Greek Prefixes* appear in the names of ionic or covalent compounds? Explain why they are used.
5. What is the relationship between the *number of hydrogens* in an acid and the *charge on the anion* that they are combined with?

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12.8: Lab 8 - Measuring Manganese Concentration Using Spectrophotometry

Objectives

- To use spectroscopy to determine the amount of Manganese is an unknown sample.

Scenario: you have just joined a "Green Team" at SMC and have been asked by your faculty mentor to determine the concentration of Manganese in a sample of groundwater taken from behind the "Old Steel Works" downtown. This is one site being considered by the city of LA for the construction of a new elementary school. Your faculty mentor is on a committee hired by the city council to perform a preliminary investigation of the site for potential environmental hazards. Other green teams on campus will be looking for lead, mercury, chromium, etc. The results of your analysis may help clear the way for the safe building of a new school; require the city council to recommend remediation (cleaning up) of the site before the school is built; or perhaps result in the selection of a new site for the construction.

It is your task to determine the exact quantity of manganese in the sample you have been given. People's lives, jobs, and safety will depend on the answers you provide!

Health Effects of Manganese:

Eating a small amount of manganese each day is important in maintaining your health. The amount of manganese in a normal diet (about 2,000-9,000 µg/day) seems to be enough to meet your daily need, and no cases of illness from eating too little manganese have been reported in humans. In animals, eating too little manganese can interfere with normal growth, bone formation, and reproduction.

Too much manganese, however, can cause serious illness. Although there are some differences between different kinds of manganese, most manganese compounds seem to cause the same effects. Manganese miners or steel workers exposed to high levels of manganese dust in air may have mental and emotional disturbances, and their body movements may become slow and clumsy. This combination of symptoms is a disease called manganism. Workers usually do not develop symptoms of manganism unless they have been exposed for many months or years. Manganism occurs because too much manganese injures a part of the brain that helps control body movements. Some of the symptoms of manganism can be reduced by medical treatment, but the brain injury is permanent.

It is not certain whether eating or drinking too much manganese can cause manganism or not. In one report, humans who drank water containing high levels of manganese developed symptoms similar to those seen in manganese miners or steel workers, but it is not certain if the effects were caused by manganese alone. In another report, people who drank water with above average levels of manganese seemed to have a slightly higher frequency of symptoms such as weakness, stiff muscles, and trembling of the hands. However, these symptoms are not specific for manganese, and might have been caused by other factors. Studies in animals have shown that very high levels of manganese in food or water can cause changes in the brain. This information suggests that high levels of manganese in food or water might cause brain injury, but it does not appear that this is of concern to people exposed to the normal amounts of manganese in food, water, or air. The chances of harm from exposure near a waste site can only be evaluated on a site-by-site basis.

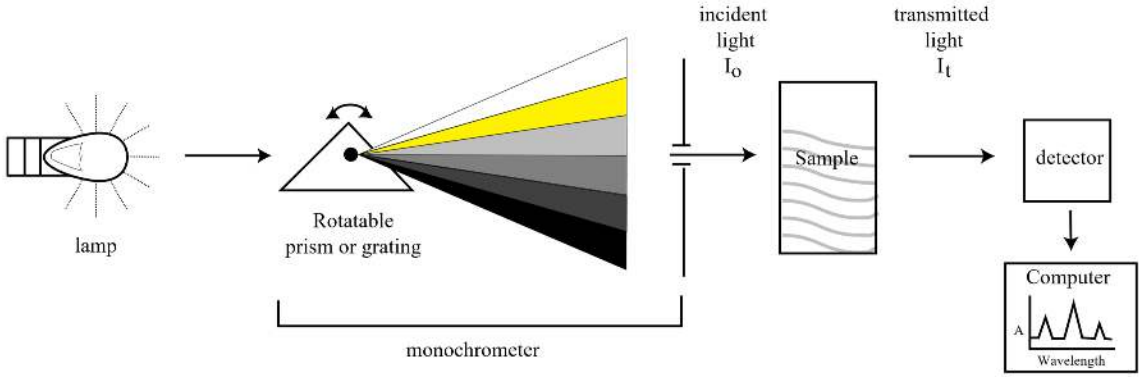
Information excerpted from: Toxicological Profile for Manganese July 1992, Agency for Toxic Substances and Disease Registry, United States Public Health Service.

Background

Spectroscopy is one of the most powerful analytical techniques in modern science. Before the advent of spectrophotometric techniques, a chemist interested in determining the amount of a particular substance present in a sample had to analyze the sample via a series of chemical reactions specific to that species and then carefully weigh the products (entire tomes exist detailing such analytical reactions). This process was extremely time consuming, prone to error, and generally impractical for measuring trace amounts. Today, most routine assaying is done quickly and efficiently by means of spectroscopy.

Spectroscopy works by correlating the concentration of a species in solution to the amount of light it absorbs. In this experiment we will determine the quantity of manganese using visible absorption spectroscopy. Because the wavelengths of light we will use are in the visible portion of the electromagnetic spectrum, our solutions will all be colored. However, this technique can also be used in other regions of the spectrum where the wavelength is not visible to our eyes, but can be measured using a photo-detector.

In a solution containing a colored compound, the intensity of the color can be used to measure the concentration of the compound; the more intense the color, the higher the concentration. The spectrophotometer measures how much light is absorbed at a given wavelength. The light absorption at that wavelength is expressed as a numerical value that can be related directly to the concentration of colored compound in the solution. In this experiment we will use spectroscopy to measure the amount of manganese in an unknown solution.



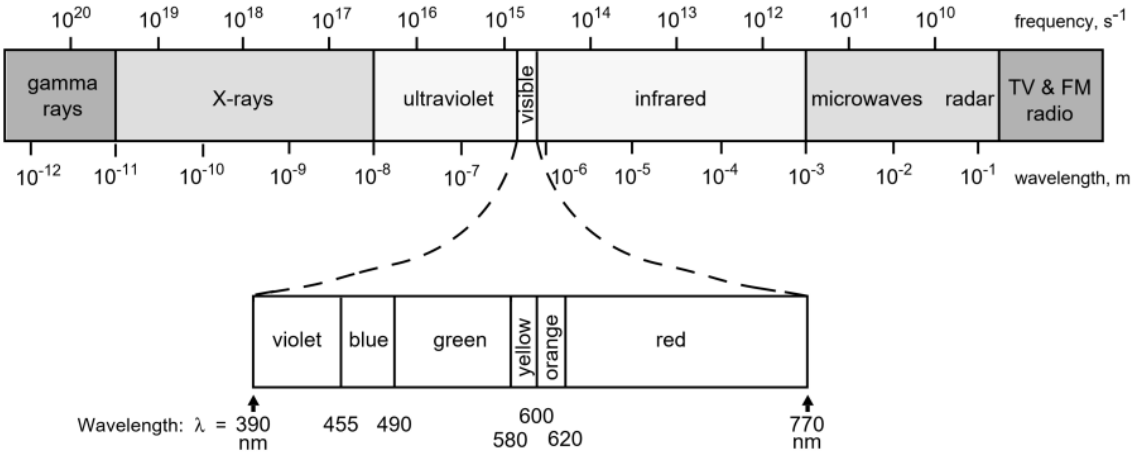
Schematic diagram of a typical absorption spectrometer

Experimental Considerations

Two important questions to answer before using a spectrophotometer are: "What wavelength will I use to measure the absorption of my compounds?" and, "How will I calibrate my instrument to get accurate readings?"

a) Choosing a wavelength

The exact color of a solution (red, blue, etc.) is an indication of how much light is absorbed at different wavelengths by its constituent molecules when white light is passed through the solution. For example, a solution that appears red in color to the eye looks red because it absorbs most colors *except* for red. This may sound backwards, but if white light (containing all colors) illuminates a "red" solution, then blue, green, yellow, etc. are absorbed by the solution; leaving only red light to reach our eyes. Thus, we perceive the solution as being red because all the other colors were absorbed. Similarly a blue solution absorbs most colors *except* blue. Some solutions have colors that are mixtures of more than one color transmitted, for example a "turquoise" colored solution may result from a mixture of blue and green light coming through the sample, etc.



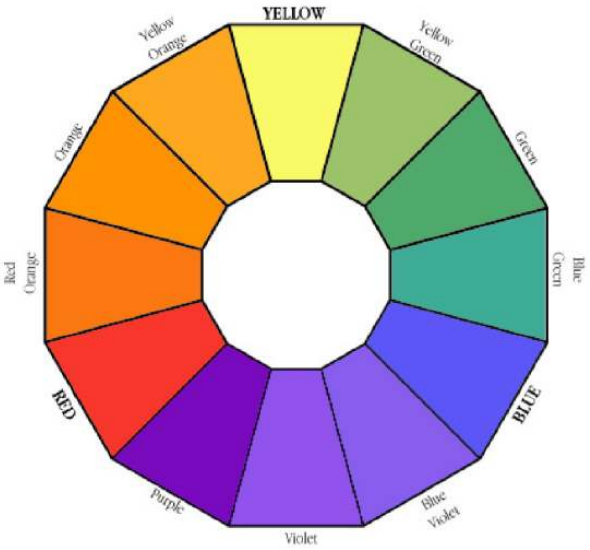
The Electromagnetic Spectrum

Thus to answer the question, *"What wavelength will I use to measure the absorption of my compounds?"* It is important to choose a wavelength where the solution strongly absorbs light. For example when studying a red solution it would be far better to use green light (red's complement), which is strongly absorbed by the solution; rather than orange or red light because these colors would be less strongly absorbed. The stronger the absorption at a particular wavelength the more sensitive the instrument will be at that

wavelength and the more accurate your results. Thus, for a red-colored sample a wavelength in the green would probably be best; conversely if a sample is blue, then a wavelength in the red would probably be best.

In practice one first determines the best range of wavelengths at which to test a solution and then selects one "best" wavelength for the experiment by measuring the absorption of the solution in question across this range and choosing the wavelength that gives the greatest absorption. For example, for the red solution described above, the experimenter might choose to measure the absorption of the sample at a range of wavelengths between 480nm and 560nm, and then narrow-in on the wavelength that gives the greatest numeric value of the absorbance.

Please note that while some spectrometers (so-called scanning spectrometers) can do this automatically; others require changing the wavelength manually and each time the wavelength is adjusted these instruments need to be re-zeroed or they will not work properly.



The Color Wheel is a guide to choosing complementary colors. A solution that is red is best analyzed using a wavelength in the green; a solution that is blue absorbs most strongly in the orange.

b) Calibrating the instrument

Once the specific wavelength that will be used for the experiment is chosen the experimenter needs to calibrate his or her spectrometer so that the absorbance readings can be converted into useful data such as molarity (M) of the sample.

The absorbance of light (A) is directly proportional to the concentration of the species in solution (C). This relationship is known as the **Beer-Lambert-Bouguer law** (or sometimes just "Beer's law") and may be expressed as:

$$A = E \times L \times C \tag{12.8.1}$$

where E is the molar absorptivity (or the molar extinction coefficient), L is the path length or distance the light travels through the sample, and C is the concentration of the solution in units of molarity (M). In most instruments L is a constant (1.0 cm in our experiment) and can therefore be factored into E.

Thus the Beer-Lambert equation can be simplified as:

$$A = E \times C \tag{12.8.2}$$

$$y = mx \tag{12.8.3}$$

Which is the equation of a line where the intercept is through the origin (i.e. zero).

The molar absorptivity, E, can therefore be determined by finding the slope of a plot of the absorbance as a function of concentration for a series of standard solutions of known concentrations. This is known as a calibration curve.

Once E has been determined from the calibration curve, the Beer-Lambert equation can be used to determine the concentration of an unknown solution by measuring its absorbance under the same conditions.

Note

A new calibration curve is required if a different instrument, wavelength, type of solution, or procedure is used. It is also a good idea to check an instrument's calibration on a periodic basis.

c) Preparing your known standards and unknown solution for analysis

The solutions you will be given contain manganese ions, Mn^{2+} (*aq*), and is an almost colorless solution. The manganese ions are easily oxidized in acidic solution to form permanganate ions, MnO_4^- (*aq*), an intensely purple species. The very intense color means that the analysis can be very sensitive because the light absorption will be relatively large, even with small amounts of manganese in the sample.

Potassium periodate, KIO_4 (*s*), will be used to oxidize Mn^{2+} (*aq*) to the purple MnO_4^- (*aq*) ion, according to the following balanced chemical equation (all species aqueous):

$$\underbrace{2\,Mn^{2+}}_{\text{Colorless}} + 5\,KIO_4 + 3\,H_2O \longrightarrow \underbrace{2\,MnO_4^-}_{\text{Purple}} + 5\,KIO_3 + 6\,H^+ \tag{12.8.4}$$

Procedure

Materials and Equipment

Obtain the following equipment from the stockroom:

- 1 100.0-mL volumetric flask
- 3 25.00-mL volumetric flasks (one or two per group will be OK if supplies are short)
- 2 cuvettes for the spectrophotometer
- 1 1.00 mL volumetric pipette
- 1 5.00 mL volumetric pipette
- 1 10.00 mL volumetric pipette
- 1 rubber pipette bulb
- 1 beaker tongs (*not* the same as the crucible tongs in your locker)

Your instructor will provide you with the spectrophotometer, stock solution, and unknowns. All other equipment is in your lab drawer.

Instructions

In order to complete this experiment in a timely manner it is suggested that one student from your team begin preparing the unknown for analysis starting with Step 16, while the rest of the team begins at Step 1 below.

1. *Be sure to rinse your pipette with water followed by the stock solution before using.* Pipette 5.00 mL of Mn^{2+} (*aq*) stock solution into a clean 400-mL beaker.
2. Using a graduated cylinder to measure volumes, add 30 mL of deionized water to the 5.00-mL stock solution. Then add 10 mL of 9 M phosphoric acid, H_3PO_4 (*aq*). (It is important to add the acid before adding the KIO_4 (*s*) in step 5 below).
3. Based on the concentration and volume of the known solution and the balanced reaction between Mn^{2+} and KIO_4, calculate the minimum mass of KIO_4 (*s*) required to oxidize all of the Mn^{2+} (*aq*) in solution to MnO_4^- (*aq*). Note: you will use double this amount to be certain that all of the Mn2+(aq) is actually oxidized. *Show the results of this calculation to your instructor before actually proceeding.*
4. Using a watch glass, weigh out double the amount of solid potassium periodate, KIO_4 (*s*), you just calculated. *Do not use paper for weighing as* KIO_4 *reacts with paper.*
5. Add the KIO_4 (*s*) to your solution in the hood and stir to dissolve using a glass stirring rod. Rinse any KIO_4 (*s*) that sticks to the watch glass into the beaker using deionized water and your squirt bottle. Leave the stirring rod in the beaker to help control bubble formation when boiling.
6. Boil the solution **gently**. A purple color should appear. If it does not add more KIO_4 (*s*) (add an additional 25% of your calculated value) and continue heating. Avoid splattering and boiling over! *If you lose any material due to splattering you must start over repeating steps 1-4.* Continue boiling gently for one to two minutes after the color changes to deep purple to be certain that the color change is complete.

7. Cool the solution and transfer the entire contents of the beaker to a 100.0-mL volumetric flask using a clean funnel. Rinse the inside of the beaker and the funnel with deionized water from your wash bottle and then transfer the rinse water to the volumetric flask as well to be sure you have transferred all of the manganese solution.
8. Dilute the solution in the volumetric flask with deionized water to exactly the index mark and swirl to mix thoroughly.
9. Calculate the concentration of MnO_4^- (*aq*) in this standard solution. Label this solution "standard." You may want to transfer some of this solution to a separate container in case of "accidents."
10. Label three clean, dry, 25.00-mL volumetric flasks: 1, 2, and 3. Pipette accurately 1.00, 5.00, and 10.00-mL portions of your standard solution into these flasks. *Be sure to rinse the pipettes with deionized water and standard solution first.* Fill each flask with deionized water exactly to the index mark and mix well. Calculate the concentration (molarity, M) of MnO_4^- (*aq*) ions in each of these flasks based on the molarity of your standard solution, the volume of this that you pipetted, and the total volume of your new diluted solution (25.00 mL).
11. Your group will only be given two cuvettes. Rinse these using deionized water and dry the outside of each cuvette using a Kim Wipe.
12. Use one cuvette for the blank and fill it 3⁄4 full with deionized water. You will use the other to measure the absorbance of each of the three solutions in the 25.00-mL volumetric flasks and the pure standard solution in the 100.0-mL volumetric flask. Be sure to rinse the second cuvette well between trials. You should first rinse this cuvette using a small amount of deionized water and then a second time using a small amount of the sample to be added.
13. Before measuring the absorbance of your solutions you will need to zero your spectrometer and choose a wavelength. Your instructor will show you how to zero the spectrometer. *It is up to your team to determine the best wavelength to use for this experiment – a poor choice of wavelengths will result in poor data. See the introduction to this experiment for help in choosing a wavelength.* If time is limited, your instructor may provide you with a specific wavelength to use.
14. Once you have zeroed the instrument and selected a wavelength, measure the absorbance of each of the three solutions in the 25.00-mL volumetric flasks and the pure standard solution in the 100.0-mL volumetric flask at this wavelength. You will have five measured data points including the deionized water blank (which should have a zero absorbance if you properly zeroed the spectrometer).
15. Make a calibration curve plotting the concentration (M) of the five solutions on the x-axis and the absorbance of each on the y-axis. Use a full sheet of graph paper to make your curve – *the bigger the plot the more accurate your estimate of the slope will be.* Draw a best-fit-line through your points and determine the molar absorptivity, E, the slope of your line. (Be sure to measure the slope of the best-fit-line; not specific data points – see your instructor if this is unclear). Your intercept should be at the origin, since your blank solution counts as one of the points. You may either plot this curve by hand using graph paper or you may use a computer program such as Excel.

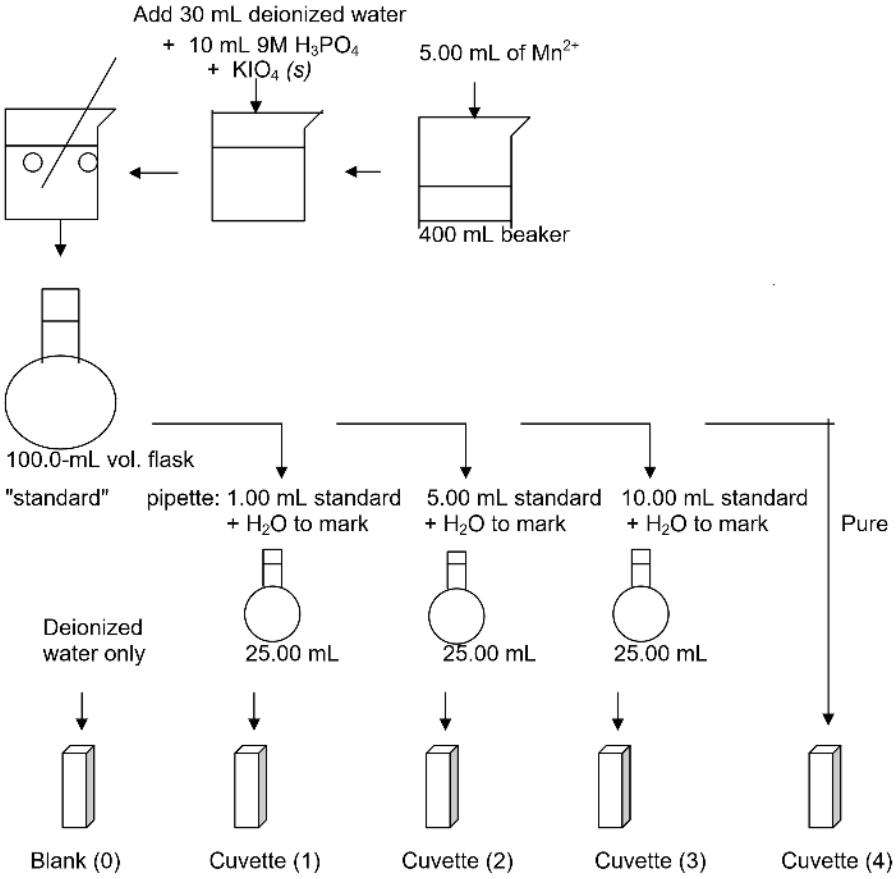
Analyzing your unknown sample:

16. Prepare your unknown sample in exactly the same way you prepared your standard solution following steps 1-8 above using 5.00 mL of your unknown solution instead of the known stock solution in step 1. You do not need to repeat steps 9 - 15 with the unknown.
17. Measure the absorbance of the 100.0-mL purple solution prepared from your unknown at the same wavelength that you used to construct your standard curve. You can now use the equation of your best-fit-line to determine the concentration of MnO_4^- (*aq*) ions in this solution.
18. Because this is a diluted solution, you will now need to calculate backwards from this concentration to determine the actual concentration of manganese in your original unknown sample.
19. ALL WASTE MUST GO INTO THE CONTAINER IN THE HOOD!

Procedure for making the known solutions and calibration curve:

1. Pipette 5.00 mL of Mn^{2+} (*aq*) Stock Solution (~0.3 mg Mn^{2+}/mL) into a clean 400 mL beaker. Record the exact concentration.
2. Add 30 mL deionized water + 10 mL 9M H_3PO_4 + KIO_4 (*s*) Add the acid IN THE FUME HOOD *before* you add the KIO_4 (*s*). (you will need to determine how many grams of KIO_4 (*s*) to add using the balanced equation given)
3. Use a glass stirring rod to mix. Place IN THE FUME HOOD and Boil **gently** for 1-2 minutes after a the solution turns purple. Do not splatter! After heating, Remove from hood and allow to cool. If the solution does not turn purple add some additional KIO_4 (*s*).
4. Transfer the cool solution + rinse to a 100.0-mL volumetric flask using a funnel. Add deionized water to the mark and mix so the solution volume is 100.0 mL.

5. Calculate the molarity (M) of MnO_4^- (*aq*) in this new solution. This will be your "standard" solution.
6. Finally: calculate the molarity of each diluted solution, choose λ, measure absorbances, and make a calibration curve.



Pre-laboratory Assignment: Measuring Manganese Using Spectrophotometry

1. List 3 health effects associated with the intake of excessive manganese:
2. In absorption spectroscopy the amount of light absorbed by a given sample at a specific wavelength is proportional to what two factors:
3. Suppose you need to do an absorption spectroscopy experiment using an orange colored solution. Approximately what wavelength(s) of light would be the worst to measure the absorbance at; what would be the best wavelength(s) of light to measure the absorbance at. Why? Explain your answers in one or two sentences.

- Worst wavelength(s):
- Best wavelength(s):
- Explain:

4. A solution contains 0.60 mg/mL Mn^{2+}. What minimum mass of KIO_4 must me added to 5.00 mL of the solution in order to completely oxidize the Mn^{2+} to MnO_4^-? Show all work.
5. A student is working to determine the concentration of a dissolved protein in a biological sample using absorption spectroscopy. She prepares a blank and two standard solutions containing known concentrations of dissolved protein. She chooses an appropriate wavelength, calibrates her instrument, and then obtains the following absorption data for her standard solutions:

Concentration (M)	Absorbance
0	0
0.013	0.046
0.039	0.138

Assuming that the path length L is 1.0 cm, what is the value of E? (show all work).

6. The same student now takes the biological sample containing an unknown quantity of dissolved protein and inserts it into the spectrometer. She obtains a reading that is off-scale (in other words her biological sample is too concentrated to measure). Undaunted by this result, she performs a dilution! A more dilute sample will have an absorption that can be measured by her spectrometer. She pipettes 20.00 mL of her sample into a 250.00 mL volumetric flask, adds deionized water to the index mark, and mixes the solution. She then pours a small sample of this diluted unknown solution into a cuvette and measures an absorbance of 0.228. What is the concentration of dissolved protein in her diluted unknown sample? Show all work.
7. Based on this result, what is the concentration of dissolved protein in her original (undiluted) unknown biological sample? Hint: calculate the dilution backwards. Show all work.

Lab Report: Measuring Manganese Using Spectrophotometry

- Concentration of Mn^{2+} (*aq*) stock solution from bottle: _____ mg/mL
- Concentration of Mn^{2+} (*aq*) stock solution in units of molarity:_____ M Mn^{2+}
  - show calculation:
- Double minimum mass of KIO_4 (*s*) required: show calculation:_____ g _____(Instructor's Initials)
  - show calculation:
- Size of volumetric flask used:_____ mL
- Volume of stock solution used (quantity pipetted into beaker): _____ mL
- Concentration of resulting standard solution:_____ M MnO_4^-
  - show calculation:
- Wavelength chosen for analysis: _____ nm
  - Justification for choice of wavelength:

Data Table:

Solution	Concentration (M)	Absorbance	Appearance
Cuvette (0) (blank)			
Cuvette (1) (1 mL / 25 mL)			
Cuvette (2) (5 mL / 25 mL)			
Cuvette (3) (10 mL / 25 mL)			
Cuvette (4) (pure standard)			

Show calculations for the concentrations of cuvettes (1-3) below:

Make a graph of Concentration in Molarity on the x-axis vs. the Measured Absorbance on the y- axis and attach it on the next page. The graph should fill one sheet of paper and may be done by hand or on a computer program such as Excel. Draw a best-fit-line on your graph and be sure it goes through the origin because zeroing your instrument ensured that at zero concentration the absorbance was zero.

Calculate the equation of your best-fit-line and show the calculation and final equation below (use points from the ends of your line and not two actual data points when doing this calculation or use linear regression analysis such as the "trend-line" function in Excel or on a graphing calculator):

- Slope of the line:
- E = _____ $M^{-1}cm^{-1}$

Unknown Analysis

- Unknown ID Code:_____
- Volume of unknown used:_____ mL

- Volume of volumetric flask:_____ mL
- Mass of KIO_4 (*s*) added:_____ g
- Observed color of unknown solution after boiling: _____
- Wavelength used to analyze unknown solution: _____ nm
- Absorbance of unknown solution:_____
- Concentration of MnO_4^- in diluted unknown solution:_____ M MnO_4^-
  - show calculations (hint: this is where you use the slope of your best-fit-line):
- Concentration of Mn^{2+} (*aq*) in original undiluted unknown sample: _____ M
  - show calculations:
- Concentration of Mn^{2+} (*aq*) in units of mg/mL: _____ mg / mL
  - show calculations:

On the back of this page (or on a separate sheet of paper), write a brief one to two paragraph conclusion based on your findings. Include the concentration of manganese present in the ground water sample and your recommendations to the city based on these findings. Support your final recommendation using published standards and not merely the information in this experiment. Be sure to properly cite all references consulted. Keep in mind that ground water is not the same as drinking water or well water when performing your research and writing your conclusions for this report.

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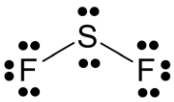
12.9: Lab 9 - VSEPR Theory and Shapes of Molecules Dry Lab

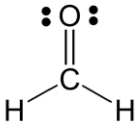
Objectives

- To build a variety of molecules and ions using molecular model kits.
- To draw Lewis structures (both projection and perspective drawings) for each of these molecules and ions.
- To determine the hybridization of the central atoms, the number and types of bonds, the geometries, and the polarities of the molecules and ions.

Lewis Structures

A Lewis Structure is a representation of covalent molecules (or polyatomic ions) where all the valence electrons are shown distributed about the bonded atoms as either shared electron pairs (bond pairs) or unshared electron pairs (lone pairs). A shared pair of electrons is represented as a short line (a single bond). Sometimes atoms can share two pairs of electrons, represented by two short lines (a double bond). Atoms can even share three pairs of electrons, represented by three short lines (a triple bond). Pairs of dots are used to represent lone pair electrons. Examples are shown for the molecules $ceSF2$ and $\mathrm{CH_2O}$ below.





Please review (in your text or notes) the rules for drawing Lewis structures before performing this exercise. This includes rules for structures which obey the octet rule, as well as those which involve expanded or reduced octets.

Resonance Structures

Resonance refers to bonding in molecules or ions that cannot be correctly represented by a single Lewis structure. These structures are often equivalent, meaning that they contain the same number of bonds at different locations. The molecule $\mathrm{SO_2}$ (shown above) has two such resonance forms. Resonance structures can also be non-equivalent, in which case they will have different numbers and/or locations of bonds. Note that any valid resonance structure of a molecule can be used to determine its shape and polarity.

VSEPR Theory

The VSEPR (**V**alence **S**hell **E**lectron **P**air **R**epulsion) model is used to predict the geometry of molecules based on the number of effective electron pairs around a central atom. The main postulate for the VSEPR theory is that the geometrical structure around a given atom is principally determined by minimizing the repulsion between effective electron pairs. Both the molecular geometry and the polarity of individual bonds then determine whether the molecule is polar or not.

Before determining the shape of a molecule, the Lewis structure must be properly drawn. The shape of a molecule is then determined by the number of areas of electron density (or, number of effective electron pairs) around a central atom. Areas of electrons density include:

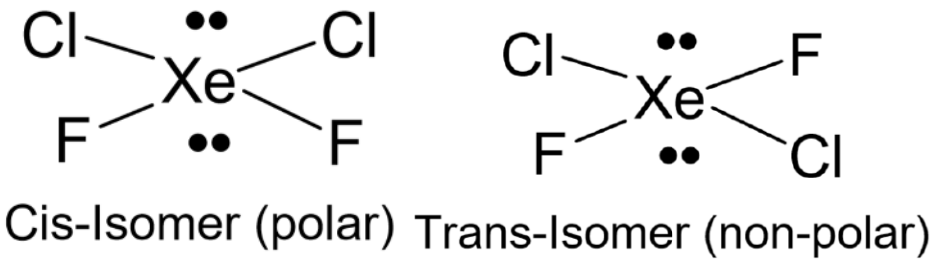
- Lone pairs of electrons: these electrons tend to take more space than the bonded pairs in space leading into somewhat distorted structures.
- Bonds (single, double and triple bonds count as one (1) area of electron density or one effective electron pair).

Before performing this exercise, please review (in your text or notes) the various geometries and bond angles that can be produced by different numbers of effective electron pairs around the central atom.

Molecular Polarity

A polar bond is one in which the electron cloud is closer to the nucleus of one atom (the more electronegative one) than the other (the less electronegative one). Knowledge of both the bond polarities and the shape are required in the determination of the molecule's overall polarity (dipole moment). A polar molecule is one that shows an imbalance in its electron distribution. When placed in an electric field, these molecules tend to align themselves with the electric field. Some molecules have polar bonds but no dipole moment. This happens when the bonds in a molecule are arranged in a way in which polarities cancel each other out. One way this occurs is when molecules have all identical bonds and there is no lone pair on the central atom (for example, $\mathrm{CO_2}$).

Molecules that do not fit both these criteria may be polar or not depending on how atoms are bonded and the electron pairs arranged around the central atom (for example, $XeCl_2F_2$ shown below).



Cis-Isomer (polar) Trans-Isomer (non-polar)

Procedure

Materials and Equipment

Molecular model kits from the stockroom

Instructions

Complete the following for each of the molecules and ions on your Report form:

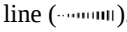
1. Draw Lewis structures, including all resonance structures if applicable (1).
2. Build models and then draw perspective structures (2) that accurately represent bond angles and molecular shapes.

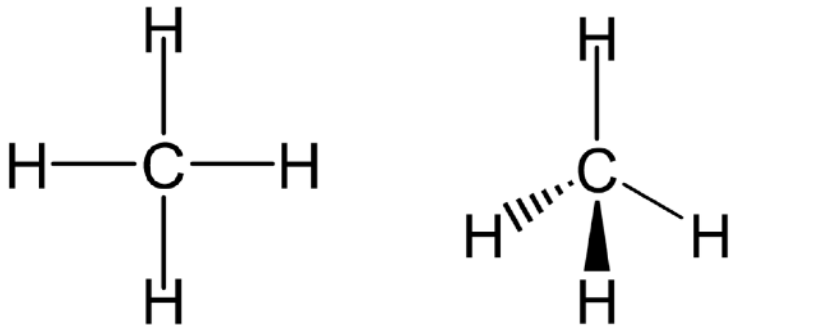
The molecular model kits contain different colored balls and different size stick connectors. Three-dimensional models will be constructed from these balls and sticks.

The stick connectors represent bonds. Use the short rigid sticks for single bonds. The long flexible sticks must be used to create double (2 sticks) and triple (3 sticks) bonds.

The different colored balls represent different atoms. There is a key on the inside of the lid of the model kit which indicates which colors correspond to which atoms. There are, however, some exceptions to this. For example, the kit indicates that the green balls with just one hole are to be used for the halogens. But if a halogen (such as Cl) appears in a molecule as a central atom with an expanded octet, you would need to use a ball that has 5 or 6 holes to build the model (brown or silver ball). Another example is oxygen. The kit indicates that the red balls with two holes should be be used for oxygen. However, if an oxygen atom in a compound requires more than two bonds, the red balls cannot be used. In this case you would substitute a blue ball for oxygen.

The 3-D models will serve as a visual guide to help you with your perspective structures. Use the following guidelines to draw them correctly:

- For bonds lying in the plane of the paper, use a regular solid line.
- For bonds that project down into the paper away from you, use a hatched wedge-shaped line ().
- For bonds that project up out of the paper towards you, use a solid wedge-shaped line ().
- An example showing both the Lewis structure and perspective representation of CH_4 is provided below.



Lewis Structure Perspective Diagram

3. Determine the number of atoms bonded to the central atom (or, number of σ-bonds) (3).

4. Determine the number of lone electron pairs on the central atom (4).
5. Predict the electronic geometry using all areas of electron density (or, effective electron pairs) and the ideal bond angles associated with this geometry (5).
6. Predict the actual geometry of the molecule or ion (6).
7. Determine the hybridization of the central atom (7).
8. Determine the polarity of the molecule (8). Use an arrow to show the direction of electron density for polar molecules on the perspective drawing.

Please be sure to return all balls and stick connectors to the model kit when finished.

Pre-laboratory Assignment: VSEPR Theory and Shapes of Molecules

Complete the following table

Chemical Species	KrF_2	PH_4^+	$TeCl_6$	$ClBr_3$	H_2S
Lewis Structure					
Perspective Drawing					
Number of atoms bonded to central atom					
Number of lone pairs on central atom					
Electronic geometry					
Molecular Geometry					
Polarity					

Lab Report: VSEPR Theory and Shapes of Molecules

HCN:

1. Lewis Structure		2. Perspective drawing
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

CH_3OH:

1. Lewis Structure		2. Perspective drawing
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

SeF_6:

1. Lewis Structure	2. Perspective drawing

3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

NO_2^-:

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

AsF_5:

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

XeF_2:

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

TeF_5^-:

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

H_2CO

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

SF_4

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

XeF_4

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

PO_4^{3-}

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

BrF_3

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

NH_3

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

CH_3NH_2

1. Lewis Structure	2. Perspective drawing	
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

Molecule:

1. Lewis Structure		2. Perspective drawing
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

Molecule:

1. Lewis Structure		2. Perspective drawing
3. Number of atoms bonded to central atom	4. Number of non-bonding electron pairs on the central atom	5. Electronic geometry:
6. Molecular geometry with ideal bond angles	7. Hybridization of central atom	8. Polarity:

Questions

1. For each one of the molecules with lone pairs of electrons on the central atom that is non polar, give an explanation why they have no dipole moment.
2. Find all molecules (or ions) with resonance structures and draw them in the box below:

1			
2			
3			

Summary of Types of Shapes

Areas of electron density			Molecular geometry	Example of species	Polarity of the example
Number atoms bonded to cenral atom (σ-bonds)	Number of lone pairs	Electronic geometry			
2	0				
3	0				
2	1				
4	0				
3	1				
2	2				
5	0				
4	1				
3	2				
2	3				
6	0				
5	1				
4	2				

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12.10: Lab 10 - Titration of Vinegar

Objectives

- To determine the molarity and percent by mass of acetic acid in vinegar.

Vinegar is essentially a solution of acetic acid ($HC_2H_3O_2$) in water. The concentration of acetic acid in vinegar may be expressed as a molarity (in mol/L):

$$\text{Molarity} = \frac{\text{Moles of Acetic Acid}}{\text{Volume of Vinegar (in L)}} \tag{12.10.1}$$

or as a mass percent

$$\text{Mass \%} = \left(\frac{\text{Mass of Acetic Acid}}{\text{Mass of Vinegar}}\right) \times 100\% \tag{12.10.2}$$

In this experiment, a technique known as a titration will be used to determine the concentration of acetic acid in vinegar. A titration involves performing a controlled reaction between a solution of known concentration (the titrant) and a solution of unknown concentration (the analyte). Here, the titrant is an aqueous solution of ~0.1 M sodium hydroxide (NaOH) and the analyte is vinegar. When mixed, a neutralization reaction occurs between sodium hydroxide and the acetic acid in vinegar:

$$NaOH(aq) + HC_2H_3O_2(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) \tag{12.10.3}$$

The sodium hydroxide will be gradually added to the vinegar in small amounts from a burette. A burette is a device that allows the precise delivery of a specific volume of a solution. The NaOH will be added to the vinegar sample until all the acetic acid in the vinegar has been exactly consumed (reacted away). At this point the reaction is completed, and no more NaOH is required. This is called the *equivalence point* of the titration.

In order to know when the equivalence point is reached, an *indicator solution* called phenolphthalein is added to the vinegar at the beginning of the titration. Phenolphthalein is a pH sensitive organic dye. Phenolphthalein is colorless in acidic solutions like vinegar, and deep pink in basic solutions like sodium hydroxide. At the equivalence point of the titration, just one drop of NaOH will cause the entire solution in the Erlenmeyer flask to change from colorless to a very pale pink.


In acidic solution


In basic solution


equivalence point

As the titration is performed, the following data will be collected:

- The molarity of NaOH (*aq*) used
- The volume of NaOH (*aq*) used to neutralize the vinegar
- The volume of vinegar used.

Using this data, the molarity and mass percent of acetic acid in vinegar can be determined by performing a series of solution stoichiometry calculations (see Calculations Section).

Procedure

Materials and Equipment

50-mL burette*, 5-mL volumetric pipette*, pipette bulb*, ~ 0.1 M NaOH (*aq*), vinegar, phenolphthalein, burette stand, two 250-mL (or 125 mL) Erlenmeyer flasks, wash bottle with distilled water, funnel

Safety

> Be especially careful when handling the sodium hydroxide base (NaOH), as it is corrosive and can cause chemical burns to the skin. If any NaOH spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor.

Titration Procedure

Your instructor will demonstrate the correct use of the volumetric pipette and burette at the beginning of the lab session. Detailed instructions on how to use a pipette are also found on the last page of this handout. *Note that three titrations must be performed.*

1. Obtain a 50-mL burette, 5-mL volumetric pipette and a pipette bulb from the stockroom.

Setting up the burette and preparing the NaOH

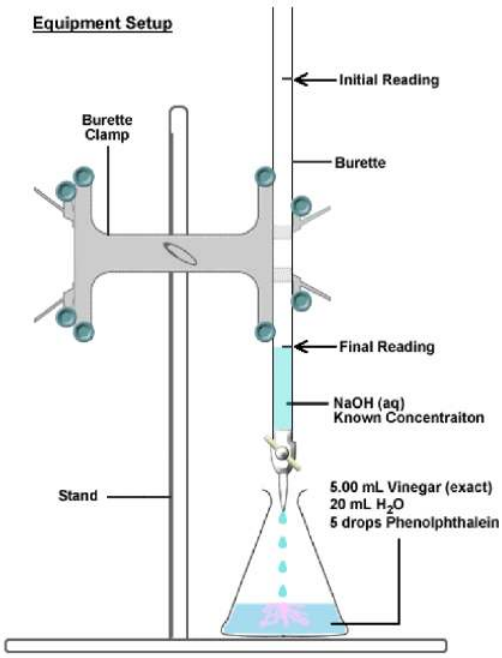
2. Rinse the inside of the burette with distilled water. Allow the distilled water to drain out through the tip in order to ensure that the tip is also rinsed.
3. Now rinse the burette with a small amount of NaOH (*aq*). To do this, add about 5-mL of NaOH (*aq*) to the burette, then twirl the burette on its side (over the sink) to rinse its entire inner surface. Then allow the NaOH (*aq*) to drain out through the tip.
4. Fill the burette with NaOH (*aq*) up to the top, between 0-mL and 5-mL. Use a funnel to do this carefully, below eye-level, and preferably over the sink. After this you will need to flush the tip of the burette – your instructor will show you how to do this. Now measure the volume at the level of the NaOH precisely, and record it as the "Initial Burette Reading" on your report. Also record the exact molarity of the NaOH (aq), which is labeled on the stock bottle.

Preparing the vinegar sample

5. The volumetric pipette used in this lab is designed to measure and transfer exactly 5.00 mL of solution. First, rinse the inside of the volumetric pipette with distilled water. Using the pipette bulb, draw the water into the pipette up above the 5-mL mark, then allow it to drain out through the tip. You may want to do this several times for practice. Then perform a final rinse, but this time use vinegar.
6. Now use the volumetric pipette to transfer 5.00-mL of vinegar into a clean 250-mL Erlenmeyer flask (see instructions on page 4). Record this volume of vinegar (precise to two decimal places) on your report. Then add about 20-mL of distilled water and 5 drops of phenolphthalein to this Erlenmeyer flask.

Performing the titration

7. Begin the titration by slowly adding NaOH (*aq*) from the burette to the vinegar in the Erlenmeyer flask. Swirl Erlenmeyer flask as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly in the flask as the base is added, but it will quickly disappear as the flask is swirled.
8. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the NaOH (*aq*) **drop by drop**. Eventually the addition of just one drop of NaOH (*aq*) will turn the solution in the Erlenmeyer flask a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. **Do not add any more** NaOH **(*aq*) at this point.** Measure this volume of NaOH (*aq*) precisely, and record it as the "Final Burette Reading" on your report. Then show the resulting solution in the flask to your instructor so s/he can record the final color on your report form.
9. Refill your burette with NaOH (*aq*), and then repeat this procedure for a second sample of vinegar, and then a third sample of vinegar. You do not need to flush the tip of the burette again. Note that if you use less than 25-mL of NaOH (*aq*) for the second titration, you do not need to refill the burette for the third titration; also that you will need to clean out and re-use one of your Erlenmeyer flasks for the third titration. You and your partner should take turns performing these titrations.
10. When finished, dispose of your chemical waste as instructed.



Pipetting Instructions

1. Get the appropriate amount of the solution you wish to pipette in a clean, dry beaker. Never pipette directly out of the stock bottles of solution. This creates a contamination risk.
2. Insert the tip of the pipette into the beaker of solution so that it is about a quarter inch from the bottom. Be sure not to press the tip against the bottom of the container.
3. If you are right handed, hold the pipette in your right hand, leaving your index finger free to place over the top of the pipette. With your left hand, squeeze the pipette bulb. Press it firmly over the top of the pipette, but DO NOT INSERT THE PIPET DEEP INTO THE BULB!
4. Release the pressure on the bulb and allow the solution to be drawn up into the pipette until it is above the volume mark. Do not allow the solution to be sucked into the bulb itself.
5. Quickly remove the bulb and place your index finger firmly over the top of the pipette. Then remove the pipette tip from the beaker of solution.
6. Slowly roll your finger to one side and allow the liquid to drain until the bottom of the meniscus is aligned with the volume mark. With practice you will be able to lower the liquid very, very slowly.
7. When the bottom of the meniscus is even with the volume mark, press your index finger firmly on the top of the pipette so no liquid leaks out. Touch the tip once to the side of the beaker to remove any hanging drops.
8. To transfer the solution, place the tip of the pipette against the wall of the receiving container at a slight angle. Then allow the liquid to drain from the pipette.
9. When the solution stops flowing, touch the pipette once to the side of the receiving container to remove any hanging drops. DO NOT blow out the remaining solution. The pipette has been calibrated to deliver the appropriate amount of solution with some remaining in the tip.

Calculations

Molarity of Acetic Acid in Vinegar

- First, using the known molarity of the NaOH (*aq*) and the volume of NaOH (*aq*) required to reach the equivalence point, calculate the moles of NaOH used in the titration.
- From this mole value (of NaOH), obtain the moles of $\mathrm{HC_2H_3O_2}$ in the vinegar sample, using the mole-to-mole ratio in the balanced equation.
- Finally, calculate the molarity of acetic acid in vinegar from the moles of $\mathrm{HC_2H_3O_2}$ and the volume of the vinegar sample used.

Mass Percent of Acetic Acid in Vinegar

- First, convert the moles of $HC_2H_3O_2$ in the vinegar sample (previously calculated) to a mass of $HC_2H_3O_2$, via its molar mass.
- Then determine the total mass of the vinegar sample from the vinegar volume and the vinegar density. Assume that the vinegar density is 1.000 g/mL (= to the density of water).
- Finally, calculate the mass percent of acetic acid in vinegar from the mass of $HC_2H_3O_2$ and the mass of vinegar.

Pre-laboratory Assignment: Titration of Vinegar

1. In this lab, you will perform a titration using sodium hydroxide and acetic acid (in vinegar). Write the balanced neutralization reaction that occurs between sodium hydroxide and acetic acid.
2. Specialized equipment is needed to perform a titration.

- Consider the sodium hydroxide reactant.
  - Name the specialized device the sodium hydroxide is placed in.
  - Is the concentration of the sodium hydroxide known or unknown?
  - Is sodium hydroxide the analyte or the titrant?
- Consider the acetic acid reactant.
  - What type of flask is the acetic acid placed in?
  - What volume of acetic acid is used?
  - What specialized device is used to obtain this precise volume?
  - Is the acetic acid the analyte or the titrant?

3. You will add sodium hydroxide to the acetic acid until all the acetic acid is consumed. This is a special point in the titration called the _____ point.
4. An indicator solution is used to indicate when all the acetic acid has been consumed and that the reaction in complete.

- What is the name of the indicator solution?
- Is this indicator mixed with sodium hydroxide or acetic acid?
- How exactly does the indicator let you know when the reaction is complete?

Lab Report: Titration of Vinegar

Experimental Data

	Trial 1	Trial 2	Trial 3
Initial Buret Reading			
Final Buret Reading			
Volume of NaOH (aq) used			
Molarity of NaOH (aq) used			
Volume of Vinegar used			
Color at equivalence point – to be recorded by your instructor			

Data Analysis

Write the balanced equation for the neutralization reaction between aqueous sodium hydroxide and acetic acid.

The Molarity of Acetic Acid in Vinegar

Use your two best sets of results (with the palest pink equivalence points) along with the balanced equation to determine the molarity of acetic acid in vinegar. Show all work for each step in the spaces provided.

Data used ⇒	Trial _____	Trial _____
Moles of NaOH used in titration		

Data used ⇒	Trial _____	Trial _____
Moles of $HC_2H_3O_2$ neutralized in vinegar sample		
Molarity of $HC_2H_3O_2$ in vinegar		
Average Molarity		

The Mass Percent of Acetic Acid in Vinegar

Use your two best sets of results along with calculated values in the previous table to determine the mass percent of acetic acid in vinegar. Show all work for each step in the spaces provided.

Data used ⇒	Trial _____	Trial _____
Mass of $HC_2H_3O_2$ in vinegar sample		
Mass of vinegar sample (assume density = 1.00 g/mL)		
Mass Percent of $HC_2H_3O_2$ in vinegar		
Average Mass Percent		

Questions

1. What was the purpose of the phenolphthalein indicator in this experiment? Be specific.
2. Suppose you added 40 mL of water to your vinegar sample instead of 20 mL. Would the titration have required more, less or the same amount of $NaOH$ (*aq*) for a complete reaction? Explain.
3. Consider a 0.586 M aqueous solution of barium hydroxide, $Ba(OH)_2$ (*aq*).

- How many grams of $Ba(OH)_2$ are dissolved in 0.191 dL of 0.586 M $Ba(OH)_2$ (*aq*)?
- How many individual hydroxide ions (OH^{-1}) are found in 13.4 mL of 0.586 M $Ba(OH)_2$ (*aq*)?
- What volume (in L) of 0.586 M $Ba(OH)_2$ (*aq*) contains 0.466 ounces of $Ba(OH)_2$ dissolved in it?

4. If 16.0 mL of water are added to 31.5 mL of 0.586 M $Ba(OH)_2$ (*aq*), what is the new solution molarity?
5. Suppose you had titrated your vinegar sample with barium hydroxide instead of sodium hydroxide:

$$Ba(OH)_2(aq) + 2\,HC_2H_3O_2(aq) \longrightarrow Ba(C_2H_3O_2)_2(aq) + 2\,H_2O(l) \qquad (12.10.4)$$

- What volume (in mL) of 0.586 M $Ba(OH)_2$ (*aq*) must be added to a 5.00 mL sample of vinegar to reach the equivalence point? Use your average vinegar molarity (see page 1) in this calculation.

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12.11: Lab 11 - Types of Chemical Reactions

Objectives

- To perform and observe the results of a variety of chemical reactions.
- To become familiar with the observable signs of chemical reactions.
- To identify the products formed in chemical reactions and predict when a reaction will occur.
- To write balanced equations for the reactions studied.
- To use the results from the single replacement reactions to devise a partial activity series.

Matter undergoes three kinds of change: physical, chemical, and nuclear. While the composition of a chemical substance is not altered by physical changes (such as freezing and evaporation), chemical changes, or reactions, result in the formation of new substances when bonds are formed and/or broken. Some relatively simple but common types of chemical reactions are illustrated in this experiment. Examples and descriptions of each reaction type appear in the following section.

Chemical reactions are often accompanied by observable changes as they occur. These include:

- Color change.
- Formation of a precipitate—noted as the formation of a cloudy solution, formation of a gel, or
- an obvious solid.
- Evolution of a gas—noted as bubbling in the solution.
- Appearance or disappearance of distinct separation between two or more liquids.
- Evolution of heat—noted as a temperature increase.
- Absorption of heat—noted as a temperature decrease.
- Plating out of one metal on another.
- Decomposition, pitting, or the disappearance of a solid metal

One or more of these changes may occur in the reactions that are performed in this experiment.

Reaction Types

Please note that some reactions may be classified as more than one type of reaction; for example, combination and decomposition reactions that involve elemental substances are also oxidation-reduction reactions.

Combination Reactions (also called Synthesis Reactions) occur when two or more substances, elements or compounds, combine to form one new substance. For example, hydrogen and oxygen gases combine to give water:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l) \tag{12.11.1}$$

Decomposition Reactions occur when a compound breaks apart to yield two or more new substances. As an example, potassium chlorate decomposes when heated to yield potassium chloride and oxygen gas. Potassium chlorate is one of the ingredients used in match heads. This reaction occurs as the match ignites, causing the match head to be surrounded by an oxygen atmosphere.

$$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g) \tag{12.11.2}$$

Displacement Reactions (also called Single Replacement Reactions) involve the displacement of one element in a compound by another element. Several examples of displacement reactions are given below.

- Displacement of one metal by another:

$$Pb(s) + Cu(NO_3)_2(aq) \longrightarrow Cu(s) + Pb(NO_3)_2(aq) \tag{12.11.3}$$

- Displacement of hydrogen gas from an acid by a metal:

$$Mg(s) + 2\,HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g) \tag{12.11.4}$$

- Displacement of hydrogen gas from water by a metal:

$$2\,K(s) + 2\,H_2O(l) \longrightarrow 2\,KOH(aq) + H_2(g) \tag{12.11.5}$$

- Displacement of one halogen by another:

$$Cl_2(aq) + NaBr(aq) \longrightarrow NaCl(aq) + Br_2(aq) \quad (12.11.6)$$

Displacement reactions are also classified as oxidation-reduction reactions. For example, in the first reaction given above, elemental lead is oxidized to lead(II) and copper is reduced from copper(II) to elemental copper. Two electrons are transferred from lead to copper in this process:

$$Pb(s) \longrightarrow Pb^{2+}(aq) + 2\,e^- \quad \text{oxidation of lead} \quad (12.11.7)$$

$$Cu^{2+}(aq) + 2\,e^- \longrightarrow Cu(s) \quad \text{reduction of copper} \quad (12.11.8)$$

The ability of one metal to displace another depends on their relative ease of oxidation—a more active metal (one that is more easily oxidized) displaces a less active metal. In the first reaction above, lead is more active than copper. The relative activities of metals can be tabulated in an activity series, ranking the metals by relative ease of oxidation. A metal that displaces hydrogen gas from acid is more active than hydrogen. A metal that displaces hydrogen gas from acid, but not from water, is less active than one that can displace hydrogen from both acids and water. The ease with which a substance is oxidized is quantified as its standard oxidation potential; you will learn more about this in the second semester of General Chemistry.

Exchange Reactions (also called Double Replacement or Metathesis Reactions) occur when two compounds that form ions in solution react by switching ion partners. Thus, these reactions have the general form:

$$AB + CD \longrightarrow AD + CB \quad (12.11.9)$$

One of three conditions must be met for these reactions to occur: (1) the formation of an insoluble ionic compound, observed as a precipitate, (2) the formation of a gas, or (3) the production of water from hydroxide and hydrogen ions (an example of an acid-base neutralization). In each case one of the products results by combining two ions that are removed from the solution by the reaction.

Precipitation Reactions occur when aqueous solutions of two ionic compounds are mixed and the ions combine to make a compound that is insoluble in water (the precipitate). For example, sodium phosphate can be used in an exchange reaction to precipitate calcium ions out of hard water as calcium phosphate, a reaction that is used in some commercial water softeners.

$$3\,CaCl_2(aq) + 2\,Na_3PO_4(aq) \longrightarrow Ca_3(PO_4)_2(s) + 6\,NaCl(aq) \quad (12.11.10)$$

The solubility behavior of the ions that you will be using is summarized in the following table:

General Rule	Important Exceptions
Group IA and Ammonium compounds are soluble.	None
Nitrates and Acetates are soluble.	None
Chlorides, Bromides and Iodides are soluble.	The halides of silver, lead(II), and mercury(I) are insoluble.
Sulfates are soluble.	The sulfates of lead(II), barium, and calcium are insoluble.
Carbonates, Chromates and Phosphates are insoluble.	The carbonates, chromates and phosphates of Group IA and ammonium are soluble.
Sulfides are insoluble.	The sulfides of Group IA, ammonium, calcium, and barium are soluble.
Hydroxides are insoluble.	The hydroxides of Group IA and barium are soluble. The other Group IIA hydroxides are moderately soluble to insoluble.

Gas Forming Reactions typically go to completion because one or more of the products are removed from the reaction vessel via the formation of a gas, which leaves the reaction mixture as bubbles. Gases thus produced include hydrogen sulfide, sulfur dioxide, carbon dioxide and ammonia.

Hydrogen sulfide, H_2S, is formed by the direct combination of an acid (source of H^+) and the sulfide ion:

$$Na_2S(aq) + 2\,HCl(aq) \longrightarrow H_2S(g) + 2\,NaCl(aq) \quad (12.11.11)$$

Sulfur dioxide, SO_2, is formed by the decomposition of sulfurous acid, which is initially formed in a reaction between an acid and the sulfite ion:

$$Na_2SO_3(s) + 2\,HCl(aq) \longrightarrow H_2SO_3(aq) + 2\,NaCl(aq) \quad (12.11.12)$$

$$H_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) \quad (12.11.13)$$

$$Na_2SO_3(s) + 2\,HCl(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,NaCl(aq) \quad (12.11.14)$$

Carbon dioxide, CO_2, is formed by the decomposition of carbonic acid, which is initially formed in a reaction between an acid and the carbonate ion:

$$Na_2CO_3(s) + 2\,HCl(aq) \longrightarrow H_2CO_3(aq) + 2\,NaCl(aq) \quad (12.11.15)$$

$$H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) \quad (12.11.16)$$

$$Na_2CO_3(s) + 2\,HCl(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,NaCl(aq) \quad (12.11.17)$$

Ammonia, NH_3, is formed from the combination of ammonium and hydroxide ions:

$$NH_4Cl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) + NH_3(g) \quad (12.11.18)$$

Neutralization Reactions (also called Acid-Base Reactions) involve the transfer of a proton (H^+) from the acid to the base. Water is always a product when the base contains the hydroxide ion (see example below). Some of the gas-forming reactions may also be classified as neutralization reactions.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) \quad (12.11.19)$$

Procedure

Materials and Equipment

Solids: Mg, $CuSO_4 \cdot 5\,H_2O$, Ca, Cu, Zn, $NaHCO_3$
Solutions: 6 M HCl, 6 M NaOH, 6 M H_2SO_4, 1 M NH_4NO_3, and 0.1 M solutions of $CuSO_4$, $ZnSO_4$, $AgNO_3$, NaCl, $Ni(NO_3)_2$, $Pb(NO_3)_2$, and K_2CrO_4

Equipment: crucible tongs, one large test tube, two small test tubes, ten small test tubes, test tube holder, test tube rack, 100-mL beaker, red litmus paper, Bunsen Burner
Supplies for Instructor Demonstrations: CaO solution (prepared in advance by stockroom), sucrose, 18 M H_2SO_4, distilled water, Na, means of cutting Na and removing from its storage vessel, two 100-mL beakers, straw, red and blue litmus paper, glass stirring rod

Safety

- Do not stare directly at the magnesium when it burns as the light can hurt your eyes.
- Do not touch metals with your hands.

Instructions for Performing each Reaction

Perform each of the following reactions except those that are to be demonstrated by your instructor. You will use 1 mL of solution in many of the reactions; estimate this by drops (typically 12-15 drops from the reagent bottle dispenser) or by measuring 1 mL once in your graduated cylinder and then transferring it to a test tube to determine how far it fills the test tube. Record your observations on the data page as you complete each reaction. Make sure that you observe the results of every reaction even if you didn't actually mix the chemicals yourself. Then write a balanced equation for each reaction. Be sure to include the states of all compounds in your equations (solid, liquid, aqueous, or gas). If no reaction occurs write the words "no reaction" (or NR) instead of the products in your balanced equation and indicate why your think there was no reaction. Unless otherwise indicated dispose of all waste in the waste container provided. **Do not put metal strips in the sink.**

Part A: Combination Reactions

1. *Instructor Demonstration:* Pour about 35 mL of a clear saturated solution containing calcium oxide into a 100 mL beaker. Allow the solution to stand for about 15 minutes. Observe. Use a straw to blow bubbles into the solution for a few seconds. Observe the solution again
2. Hold a small strip of magnesium metal (used in flashbulbs and fireworks) in your crucible tongs and ignite the metal in the hot portion of a burner flame. Don't forget to note the color and composition of the residue left on the tongs.

Part B: Decomposition Reactions

1. *Instructor Demonstration:* Perform this reaction in the fume hood. Fill a 100 mL beaker about one-third full of granulated sugar (sucrose, $C_{12}H_{22}O_{11}$). Add about 20 mL of concentrated (18 M) sulfuric acid and stir until mixed well. Continue stirring until the mixture darkens. Observe. Do not touch the reaction products or the beaker with your hands; use a stirring rod to guide the solid product that forms.
2. Place a small amount (an amount that will fit on the end of a spatula) of solid copper(II) sulfate pentahydrate in a medium test tube. Use a test tube holder to hold the tube at about a 45° angle and heat in a burner flame for a few minutes, remembering not to point the tube at anyone in the room. Note any changes in the appearance of the solid and anything else that appears in the test tube. Allow the solid to cool and add a few drops of water. Observe. Dispose of the copper compound in the waste container.

Part C: Displacement Reactions

Use 1 mL of each solution unless otherwise specified. For reactions involving metals, use just one piece of metal. Do not put the metal pieces in the sink. If no discernable initial change is noted, let the reaction mixture stand for at least five to ten minutes before observing again. Not all of the combinations will yield observable reactions. Repeat the reaction if there is any doubt about whether a reaction occurred or not.

1. *Instructor Demonstration:* Cautiously add a small piece of sodium metal to water. Test the resulting solution with red and blue litmus papers (red litmus paper will turn blue in the presence of a base; blue litmus will turn red in the presence of an acid).
2. Calcium metal and water (15 mL) in a large test tube
3. Zinc metal and water
4. Copper metal and 6 M hydrochloric acid
5. Zinc metal and 6 M hydrochloric acid
6. Zinc metal and 0.1 M copper(II) sulfate
7. Copper metal and 0.1 M zinc sulfate
8. Copper metal and 0.1 M silver nitrate

Part D: Exchange Reactions

Use 1 mL of each solution unless otherwise specified. Be sure to mix the solutions well.

1. 0.1 M silver nitrate and 0.1 M sodium chloride
2. 0.1 M nickel(II) nitrate and three drops of 6 M sodium hydroxide
3. 0.1 M lead(II) nitrate and 0.1 M potassium chromate
4. 1 M ammonium nitrate and 6 M sodium hydroxide. Warm the test tube gently by passing it back and forth through a burner flame. Hold a strip of moistened red litmus paper in the tube without letting it come in contact with the sides of the tube and note any color changes to the paper. Remove the tube from the flame and quickly and cautiously note the smell.
5. Place 5 mL of 6 M hydrochloric acid in a 100 mL beaker. Carefully add several spatulas of solid sodium bicarbonate. Observe.
6. Combine about 5 mL each of 6 M sodium hydroxide and 6 M sulfuric acid in a large test tube. Mix with a stirring rod. Cautiously feel the outside of the test tube. If you cannot detect anything, make sure that you used the correct concentrations of acid and base. **Be very careful with the concentrated sulfuric acid; it is very caustic and can dissolve skin and clothing.**

Pre-laboratory Assignment: Types of Reactions

1. Many of the reactions use 1 mL of solution. How can you estimate this volume?
2. Do you need to dry the test tubes before using them for the reactions in this experiment? Why or why not?
3. The following reactions are performed, and the results are shown below. Use these results to determine the relative activities of the two elements involved in each reaction next to that reaction. Then place the elements gold, hydrogen, zinc and tin in an

activity series in order of decreasing activity.

$$\mathrm{Sn(s) + HCl(aq) \longrightarrow H_2(g) + SnCl_2(aq)} \qquad \_\_\_\_\_ > \_\_\_\_\_ \tag{12.11.20}$$

$$\mathrm{Au(s) + Sn(NO_3)_2(aq) \longrightarrow NR} \qquad \_\_\_\_\_ > \_\_\_\_\_ \tag{12.11.21}$$

$$\mathrm{Au(s) + HCl \longrightarrow NR} \qquad \_\_\_\_\_ > \_\_\_\_\_ \tag{12.11.22}$$

$$\mathrm{Zn(s) + Sn(NO_3)_2(aq) \longrightarrow Zn(NO_3)_2(aq) + Sn(s)} \qquad \_\_\_\_\_ > \_\_\_\_\_ \tag{12.11.23}$$

From most active (most easily oxidized) to least active:

_____>_____>_____>_____

Now use the above results to write products for the reactions below. Write NR if no reaction is expected.

$$\mathrm{Sn(s) + Zn(NO_3)_2(aq) \longrightarrow} \tag{12.11.24}$$

$$\mathrm{Zn(s) + Au(NO_3)_3(aq) \longrightarrow} \tag{12.11.25}$$

4. Suppose that each of the following pairs of aqueous solutions is combined. For those where a reaction is expected, write a balanced formula equation, with state labels, for the reaction that occurs. If no reaction is expected, indicate this and explain why no reaction is expected.

- Barium chloride + potassium sulfate
- Aluminum nitrate + sodium chloride
- Sodium hydroxide + phosphoric acid

Lab Report: Types of Reactions

Record your observations on these data pages as you perform each reaction. Write a balanced formula equation with state labels for each reaction. If no reaction occurs, follow the instructions in the Procedure.

Part A: Combination Reactions

1. *Calcium oxide and carbon dioxide (demonstration)*

Observations	
Reaction Equation	

2. *Magnesium and oxygen*

Observations	
Reaction Equation	

Part B: Decomposition Reactions

1. *Sucrose and sulfuric acid catalyst (demonstration)*

Observations	
Reaction Equation	

2. *Thermal decomposition of copper(II) sulfate pentahydrate,* $\mathrm{CuSO_4 \cdot 5\,H_2O}$

Observations	
Reaction Equation	

Part C: Displacement Reactions

In addition to providing observations and an equation for each reaction, use your results to determine the relative activities of the two elements involved in each reaction.

1. *Sodium and water (demonstration)*

Observations	
Reaction Equation	
Relative activities	

2. *Calcium and water*

Observations	
Reaction Equation	
Relative activities	

3. *Zinc and water*

Observations	
Reaction Equation	
Relative activities	

4. *Copper and hydrochloric acid*

Observations	
Reaction Equation	
Relative activities	

5. *Zinc and hydrochloric acid*

Observations	
Reaction Equation	
Relative activities	

6. *Zinc and copper(II) sulfate*

Observations	
Reaction Equation	
Relative activities	

7. *Copper and zinc sulfate*

Observations	
Reaction Equation	
Relative activities	

8. *Copper and silver nitrate*

Observations	
Reaction Equation	
Relative activities	

Arrange copper, silver, calcium, zinc, and hydrogen in an activity series from most active to least active on the basis of the results from the displacement reactions that you performed. Recall that a more active metal displaces a less active metal, a more active metal to is needed to displace hydrogen from water than to displace it from an acid, and that a metal that displaces hydrogen from acid is ranked as more active than hydrogen.

_____>_____>_____>_____>_____

most active (most easily oxidized) to least active

Part D: Exchange Reactions

1. *Silver nitrate and sodium chloride*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

2. *Nickel(II) nitrate and sodium hydroxide*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

3. *Lead(II) nitrate and potassium chromate*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

4. *Ammonium nitrate and sodium hydroxide*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

5. *Hydrochloric acid and sodium bicarbonate*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

6. *Sulfuric acid and sodium hydroxide*

Observations	
Reaction Equation (balanced Molecular)	
Net Ionic Equation	

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12.12: Lab 12 - Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions

Objectives

- To observe the solubility properties of various ionic compounds containing alkaline earth metal cations.
- To observe the relative abilities of the halogens to be reduced to halides, or act as oxidizing agents.
- To use the above observations to identify an unknown salt consisting of an alkaline earth metal cation and a halide anion.

Elements within a given column of the periodic table tend to have similar properties due to their similar valence electron configurations. Because of this, columns of elements are often referred to as “groups” or “families” of elements. These families include the alkali metals, alkaline earth metals, halogens, and noble gases. The physical and chemical properties of the elements within a given family tend to change gradually as one goes from one element in the column to the next. In this experiment the properties of elements in the alkaline earth metal and halogen families will be studied and this data used to identify an unknown salt consisting of an alkaline earth metal cation and a halide anion.

Group 2A: The Alkaline Earth Metals

The alkaline earth metals— beryllium, magnesium, calcium, strontium, barium, and radium—are all moderately reactive. Beryllium compounds are quite rare and often very poisonous and radium compounds are highly radioactive; thus, neither of these will be studied. Alkaline earth metals lose two electrons to make ions with a +2 charge and can thus be represented generically as M^{2+}. When solutions containing these cations are mixed with solutions containing anions such as CO_3^{2-} or SO_4^{2-}, ionic compounds of the general form MX will precipitate if the compound MX is insoluble under the reaction conditions used, as shown in the net ionic equation below.

$$M^{2+}(aq) + X^{2-}(aq) \longrightarrow MX(s) \qquad \text{if } MX \text{ is insoluble} \tag{12.12.1}$$

In this experiment M^{2+} = Ba^{2+}, Ca^{2+}, Mg^{2+}, or Sr^{2+}; and X^{2-} = SO_4^{2-}, CO_3^{2-}, $C_2O_4^{2-}$, or CrO_4^{2-}.

No precipitate will be observed if the compound MX is soluble.

When the solubilities of compounds containing various cations combined with a given anion are compared, a solubility trend that follows the order in the periodic table is expected. For example, for the solubilities of the sulfate salts, the solubility is expected either to increase or decrease as we go down the alkaline earth family. These solubility properties will be used to identify an unknown compound containing a Group 2A cation.

Group 7A: The Halogens

The elementary halogens are also relatively reactive. They include fluorine, chlorine, bromine, iodine, and astatine. We will not study astatine or fluorine since the former is radioactive and the latter is too reactive to be safe. Unlike the alkaline earth metals, the halogens tend to gain electrons to form anions, such as Cl^- and Br^-. Since they are reduced when this occurs, the halogens are oxidizing agents, species that tend to oxidize (remove electrons from) other species. Thus it is possible for some halogens (Cl_2, Br_2, I_2) to react with halide ions (Cl^-, Br^-, I^-) in a single replacement reaction. Taking X_2 to be a halogen, and Y^- to be a halide ion, the reaction would be as follows:

$$X_2 + 2\,Y^- \longrightarrow 2\,X^- + Y_2 \tag{12.12.2}$$

The reaction will only occur if X_2 is a better oxidizing agent than Y_2, since X_2 has to remove electrons from the Y^- ions. If Y_2 is a better oxidizing agent than X_2 then no reaction will occur. When comparing two or more halogens, the one that is a better oxidizing agent is considered to be “more active.” Solutions of halogens and halide ions will be combined to determine the relative oxidizing abilities of the halogens. These should show a trend as one goes from one halogen to the next in the Periodic Table.

Since the halogens have characteristic colors in non-polar organic solvents, such as hexane, while the halide ions are colorless we can use the color changes that occur (or don’t occur) to determine whether or not one halogen has displaced another in the above reaction. The relative oxidizing strengths of the halogens will be determined from the reactivity patterns. This data will then be used to determine the identity of the halide anion in an unknown compound.

Example 15.1

Suppose that an aqueous solution of bromine (Br_2) is mixed with hexane, which is less dense than and insoluble in water. The hexane layer will be the top layer due to the relative densities of hexane and water. The bromine is much more soluble in hexane than in water and goes into the hexane layer if you shake the mixture well, giving the hexane layer an orange color. Now suppose we add a solution containing chloride ion (Cl^-), to the bromine mixture and mix well. There are two possible results:

$$Br_2 + 2\,Cl^- \longrightarrow 2\,Br^- + Cl_2$$

If this reaction occurs the color of the hexane layer will change from orange (the color of Br_2 in hexane) to that of a solution of Cl_2 in hexane.

$$Br_2 + 2\,Cl^- \longrightarrow \text{No Reaction}$$

There will be no color change; the hexane layer will remain orange. This implies that bromine is not a better oxidizing agent than chlorine or, in other words, that chlorine is a better oxidizing agent than bromine.

In summary, no color change means that no reaction occurred and a color change means that a reaction occurred. In either event, the halogen that remains after the halogen and halide are mixed is the less active halogen.

It is important to keep in mind the difference between the halogen elements and the halide ions:

Halogens	Halide Ions
Bromine, Br_2	Bromide ion, Br^-
Chlorine, Cl_2	Chloride ion, Cl^-
Iodine, I_2	Iodide ion, I^-

The **halogens** are molecular substances, oxidizing agents, have odors, and are distinct colors. They are only slightly soluble in water and much more soluble in hexane.

The **halide ions** are soluble only in water, have no color or odor, and are not oxidizing agents. They do not dissolve in hexane.

Identification of a Salt Containing an Alkaline Earth Cation and a Halide

Given the observed solubility properties of the alkaline earth cations and the oxidizing ability of the halogens, it is possible to develop a systematic procedure for determining the identities of the alkaline earth cation and halide present in an unknown. You will devise such a procedure and then carry it out on an unknown substance containing an alkaline earth cation and a halide.

Procedure

Materials and Equipment

Reagents: 0.1 M $Ba(NO_3)_2$, $Ca(NO_3)_2$, $Mg(NO_3)_2$, and $Sr(NO_3)_2$; 1 M H_2SO_4, 1 M Na_2CO_3, 0.25 M $(NH_4)_2C_2O_4$, 1 M K_2CrO_4, 1 M acetic acid, bromine water (aqueous Br_2), chlorine water (freshly prepared is best), iodine water, 0.1 M NaCl, NaBr, and NaI; hexane

Equipment: four or more small test tubes, test tube rack, stirring rod, and cork stoppers(*) for test tubes

Safety

Hexane is flammable. Do not use it anywhere near an open flame. All waste must be poured into the waste container. Avoid breathing the halogen vapors. Don't use your finger to stopper the tubes. Notify your instructor if you spill any bromine solution; it can be cleaned up with sodium thiosulfate solution. All reaction mixtures should be placed in the waste container when you are done with them.

Part A: Relative Solubilities of Some Salts of the Alkaline Earth Metals

1. Each of the following reactions is performed in a small test tube, using about 1 mL (12-15 drops from the reagent bottle pipet) of each solution. The reaction mixtures must be mixed well by using a stirring rod or a method described by your instructor.

The stirring rod should be rinsed in a beaker of deionized water between uses. The test tubes must be washed and then rinsed them with distilled water before you use them and before you reuse them for another reaction.

2. The results are to be recorded in the data table provided, noting whether a precipitate forms, and any characteristics—as color, odor, amount, size of particles, and settling tendencies— that might distinguish it.

$$H_2SO_4 + Ba(NO_3)_2 \quad (12.12.3)$$

$$H_2SO_4 + Ca(NO_3)_2 \quad (12.12.4)$$

$$H_2SO_4 + Mg(NO_3)_2 \quad (12.12.5)$$

$$H_2SO_4 + Sr(NO_3)_2 \quad (12.12.6)$$

$$Na_2CO_3 + Ba(NO_3)_2 \quad (12.12.7)$$

$$Na_2CO_3 + Ca(NO_3)_2 \quad (12.12.8)$$

$$Na_2CO_3 + Mg(NO_3)_2 \quad (12.12.9)$$

$$Na_2CO_3 + Sr(NO_3)_2 \quad (12.12.10)$$

$$(NH_4)_2C_2O_4 + Ba(NO_3)_2 \quad (12.12.11)$$

$$(NH_4)_2C_2O_4 + Ca(NO_3)_2 \quad (12.12.12)$$

$$(NH_4)_2C_2O_4 + Mg(NO_3)_2 \quad (12.12.13)$$

$$(NH_4)_2C_2O_4 + Sr(NO_3)_2 \quad (12.12.14)$$

$$K_2CrO_4 + Ba(NO_3)_2 + \text{acetic acid} \quad (12.12.15)$$

$$K_2CrO_4 + Ca(NO_3)_2 + \text{acetic acid} \quad (12.12.16)$$

$$K_2CrO_4 + Mg(NO_3)_2 + \text{acetic acid} \quad (12.12.17)$$

$$K_2CrO_4 + Sr(NO_3)_2 + \text{acetic acid} \quad (12.12.18)$$

Part B: Relative Oxidizing Abilities of the Halogens

1. The color of the bromine in hexane is determined by placing 1 mL of bromine-saturated water in a small test tube, adding 1 mL of hexane, and mixing the contents well until the bromine color is mostly in the hexane layer. The color of the hexane layer should be noted at this point. Although the colors of the two layers may be different it is the color in the hexane layer that will be used throughout this experiment.
2. The colors of chlorine and iodine in hexane are determined as for bromine. Chlorine water is combined with hexane and then iodine water is combined with hexane, noting the color of the hexane layer in each case.
3. Each of the following reactions is to be performed in a small test tube by combining 1 mL of hexane with 1 mL of each of the two reagents listed. The contents should be mixed well and the color of the (upper) hexane layer noted. The test tubes must be washed and then rinsed with distilled water before you use them and before you reuse them for another reaction.

$$Br_2 + NaCl \quad (12.12.19)$$

$$Br_2 + NaBr \quad (12.12.20)$$

$$Br_2 + NaI \quad (12.12.21)$$

$$Cl_2 + NaCl \quad (12.12.22)$$

$$Cl_2 + NaBr \quad (12.12.23)$$

$$Cl_2 + NaI \quad (12.12.24)$$

$$I_2 + NaCl \quad (12.12.25)$$

$$I_2 + NaBr \quad (12.12.26)$$

$$I_2 + NaI \quad (12.12.27)$$

Part C: The Unknown

The solubility properties of the alkaline earths, the oxidizing properties of the halogens, and the methodology of this experiment will be used to devise a systematic procedure for determining the chemical formula of an unknown that contains one alkaline earth metal cation and one halide ion. A written procedure or flow chart may be required before an unknown is issued. Appropriate tests will be performed on the unknown so as to ascertain the cation and anion present. It may not be necessary to perform all of the reactions used in parts A and B to identify your unknown.

Clean Up

Reaction mixtures should be disposed of in the waste container.

Pre-Laboratory Assignment: Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions

1. Suppose that substances X, Y, and Z each potentially react with the anion (X^-, Y^-, or Z^-) of another species in the series. In solution X is yellow, Y is blue, and Z is red. The anions are all colorless.

When a solution of Z is reacted with X^- the color changes from red to yellow, meaning that the reaction $Z + X^- \longrightarrow Z^- + X$ has occurred.

- Which reactant was oxidized? _____
- Which reactant was reduced? _____
- What is the better oxidizing agent, Z or X? _____

When a solution of X is reacted with Y^-, the color changes from yellow to blue.

- Write an equation for the reaction that occurred, assuming it is the same type of reaction as above.
- Which reactant was oxidized? _____
- Which reactant was reduced? _____
- What is the better oxidizing agent, X or Y? _____

Rank X, Y, and Z in order of decreasing oxidizing ability.

_____ > _____ > _____

From best oxidizing agent to worst oxidizing agent

Would you expect a reaction when Y is combined with Z^-? Why or why not?

2. Hexane and water form two layers when they are combined. What layer is on the bottom and on the top? Why?
3. Silver ion reacts with chloride ion to form insoluble $AgCl$. Aluminum ion does not form a precipitate when reacted with chloride ions. Lead (II) ion reacts with chloride ion to form insoluble $PbCl_2$. $PbCl_2$ is soluble in hot water whereas $AgCl$ is not. Use this information to outline a procedure that would determine whether Ag^+, Pb^{2+}, or Al^{3+} is present in an unknown that contains one, and only one, of these three ions. Use the back of the page for your answer.

Lab Report: Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions

Part A: Relative Solubilities of Some Salts of the Group 2A Cations

Data and Observations

Indicate whether a precipitate forms (“ppt”) or there is no precipitate (“no ppt”). Note any distinguishing characteristics of the precipitates.

	1 M H_2SO_4	**1 M Na_2CO_3**	**0.25M $(NH_4)_2C_2O_4$**	**1 M K_2CrO_4 1 M Acetic Acid**
$Ba(NO_3)_2$				
$Ca(NO_3)_2$				
$Mg(NO_3)_2$				

$Sr(NO_3)_2$				

Questions and Conclusions

1. Consider the relative solubilities of the Group 2 cations in the various precipitating reagents. On the basis of the trends you observed, list the four alkaline earth metal ions in order based on the solubility of the salts formed—the one that forms the greatest number of soluble compounds is the most soluble and the one that forms the greatest number of precipitates is the least soluble

_____ > _____ > _____ > _____

From most soluble to least soluble

2. Is this order consistent with the data from all of the precipitation reagents? If not, why did you order them as you did?
3. Is this order consistent with a periodic trend? If no, how is it different? If yes, does solubility increase or decrease down the column?

Part B: Relative oxidizing powers of the halogens

Data and Observations

Indicate the color of the hexane layer after mixing hexane and each halogen in the space provided in the header row of the following table. Indicate the color of the hexane layer after addition of each halide reagent and the corresponding halogen that is present in the top portion of each box. If a reaction occurred, write a net ionic equation in the lower portion of the box. If no reaction occurred, write "NR" instead of a reaction equation

	Br_2 color _____	**Cl_2 color** _____	**I_2 color** _____
Cl^-	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:
Br^-	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:
I^-	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:	Color: _____ Halogen: _____ Reaction:

Questions and Conclusions

1. Rank the halogens in order of their decreasing oxidizing power: _____ > _____ > _____

From best oxidizing agent to worst oxidizing agent

2. Does the order above follow a periodic trend? If no, how is it different? If yes, does oxidizing ability ("activity') increase or decrease down the column?

Part C: The Unknown

ID Code: _____

Outline the procedure you will use to identify the ions present in the unknown.

Cation:

Anion:

Give the observations made when the above procedure was done. Explain how the data and observations lead to your conclusion.

Based on the work detailed above:

Cation Present _____

Anion Present _____

Formula of Unknown _____

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12.13: Lab 13 - Mole Ratios and Reaction Stoichiometry

Learning Objectives

- To experimentally determine the mole-to-mole ratios between the reactants and products in the following two double displacement "gas forming" reactions:

$$\text{sodium bicarbonate} + \text{hydrochloric acid} \longrightarrow \text{sodium chloride} + \text{carbon dioxide} + \text{water} \tag{12.13.1}$$

$$\text{sodium carbonate} + \text{hydrochloric acid} \longrightarrow \text{sodium chloride} + \text{carbon dioxide} + \text{water} \tag{12.13.2}$$

The easiest way to obtain the mole-to-mole ratios would be to simply balance the chemical equations for these reactions. This would be considered a theoretical approach to the problem. The balanced equations for reactions 12.13.1and 12.13.2are:

$$1\,\mathrm{NaHCO_3(s)} + 1\,\mathrm{HCl(aq)} \longrightarrow 1\,\mathrm{NaCl(aq)} + 1\,\mathrm{CO_2(g)} + 1\,\mathrm{H_2O(l)} \tag{12.13.3}$$

$$1\,\mathrm{Na_2CO_3(s)} + 2\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{NaCl(aq)} + 1\,\mathrm{CO_2(g)} + 1\,\mathrm{H_2O(l)} \tag{12.13.4}$$

In reaction 12.13.3, the balancing coefficients indicate that there is a 1:1 mole ratio between reactant $\mathrm{NaHCO_3}$ and product NaCl. This means that for every 1 mole of sodium bicarbonate that reacts, 1 mole of sodium chloride should be produced. However in reaction 12.13.4, the balancing coefficients indicate that there is a 1:2 mole ratio between reactant $\mathrm{Na_2CO_3}$ and product NaCl. In this case, for every 1 mole of sodium carbonate that reacts, 2 moles of sodium chloride should be produced.

To determine these mole-to-mole ratios experimentally, a quantitative analysis of both reactions is required. Specifically, a pre-weighed mass of sodium bicarbonate/carbonate will be allowed to react with a slight excess of hydrochloric acid. The sodium chloride product will then be carefully retrieved, dried and weighed at the end of the reaction. This mass of collected sodium chloride is called an *experimental yield*. Both the mass of sodium bicarbonate/carbonate reactant used and the mass of sodium chloride product can be converted to mole quantities via their respective molar masses. Finally, dividing both the reactant and product moles by the lower of the two values should yield the simplest whole number mole-to-mole ratio of reactant and product.

While an experimental product yield is obtained by actually performing a reaction in lab, a *theoretical yield* is the maximum mass of product that could be obtained from a reaction provided that no errors occur. Theoretical product yields can only be determined by performing a series of stoichiometric calculations. Thus, using this method, theoretical yields of sodium chloride will be calculated for reactions A and B. Once obtained, the *percent yield* of sodium chloride can be determined for both reactions, where

$$\text{Percent Yield} = \frac{\text{Experimental Yield}}{\text{Theoretical Yield}} \times 100 \tag{12.13.5}$$

Good experimental practices in the lab (with minimum error) generally result in a high percent yield, where the experimental yield closely approaches the theoretical yield.

Procedure

Materials and Equipment

Solid sodium bicarbonate ($\mathrm{NaHCO_3}$), solid sodium carbonate ($\mathrm{Na_2CO_3}$), 6 M hydrochloric acid (HCl), electronic balance, evaporating dish, watch glass (to fit as a cover for the evaporating dish), stand and ring clamp, wire gauze, dropper pipette, small beaker, and Bunsen burner.

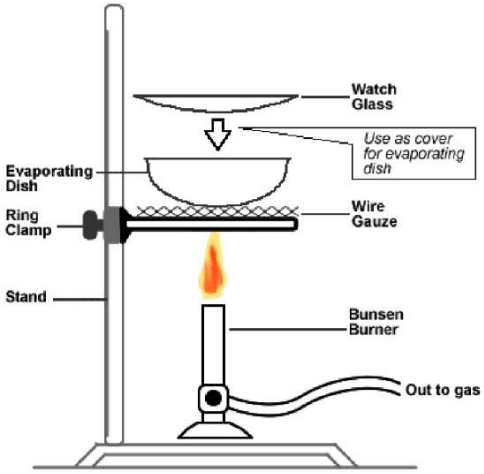
Safety

Be especially careful when handling the 6 M HCl (*aq*), as it can cause chemical burns to the skin. If any acid spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor. Acid spills may also be neutralized using the sodium bicarbonate solution by the sinks. Also, be sure to exercise appropriate caution when using the Bunsen burner and handling hot equipment.

Experimental Procedure

1. Measure and record the mass of your clean dry evaporating dish + watch glass (assembled together with the watch glass acting as a cover on top of the evaporating dish).

2. Carefully add 0.3 – 0.4 g of solid sodium bicarbonate ($NaHCO_3$) to the evaporating dish. ***Do not do this over the balance!*** Then measure and record the mass of the evaporating dish + watch glass + $NaHCO_3$.
3. Obtain about a 5-mL quantity of hydrochloric acid (HCl) in your small beaker. Then using your dropper pipette, add the HCl **drop by drop** to the sodium bicarbonate in the evaporating dish. The reaction will be evident by the bubbling that takes place. Gently mix the reactants after every 4-5 drops of HCl. Continue adding HCl until the bubbling stops and all of the $NaHCO_3$ is dissolved – this indicates that the reaction is complete.
4. Assemble the stand, ring clamp and wire gauze apparatus for heating as shown in the figure below. Cover the evaporating dish with the watch glass and place it on the wire gauze.



5. **Gently heat** the solution in the covered evaporating dish with a Bunsen burner flame in order to remove the water generated in the reaction (as well as any excess HCl present). The flame should be adjusted to a lower temperature and wafted under the evaporating dish constantly. Continue heating until the contents are **completely dry**. Note that the watch glass cover should also be dry!
6. After allowing the evaporating dish to cool to room temperature, measure and record the mass of the evaporating dish + watch glass + residue ($NaCl$).
7. Repeat steps 1 to 6 with a 0.3 – 0.4 g sample of sodium carbonate (Na_2CO_3).
8. The waste from this experiment may be disposed of in the sink.
9. *Analysis:* Experimental Mole-to-Mole Ratios – Convert the initial mass of sodium bicarbonate (or carbonate) reactant to moles (via its molar mass). In a similar manner, convert the final mass of collected sodium chloride product to moles (via its molar mass). Finally, obtain the simplest whole number mole-to-mole ratio by dividing both the reactant and product moles by the lower of the two values.
10. *Analysis:* Percent Yields – Calculate the theoretical yield of $NaCl$ for both reactions 12.13.3and 12.13.4via standard mass-to-mass stoichiometry. Use your masses of sodium bicarbonate/carbonate reactants weighed out in lab as the starting point and the mole ratios from the balanced equations for these calculations. Then determine your percent yield for each reaction using the calculated theoretical yields along with your experimental yields of $NaCl$, obtained in lab. See the Theory Section for the equation, if necessary.

Pre-laboratory Assignment: Mole Ratios and Reaction Stoichiometry

1. Write balanced equations for the two reactions you will perform in this lab.

- Reaction 12.13.3:
- Reaction 12.13.4:

2. Your goal in this lab is to experimentally verify the mole-to-mole ratios between a certain reactant and a certain product in both reactions.

- Identify the two substances in Reaction 12.13.4.
- What is this theoretical mole-to-mole ratio in Reaction 12.13.4?

3. In Reaction 12.13.3, you will react a pre-weighed sample of sodium bicarbonate with acid. In Reaction 12.13.4, you will use sodium carbonate instead of the bicarbonate.

- Name the "container" that you will perform both reactions in.
- What is the purpose of the watch glass?

4. How will you know when enough acid has been added to the sodium bicarbonate (in Reaction 12.13.3) or sodium carbonate (in Reaction 12.13.4), and that the reactions are complete?
5. After mixing the reactants together, you will then use a Bunsen burner to heat the contents of the reaction "container". When should you stop heating?
6. Once heating is complete, you will then weigh the substance that remains in the reaction "container".

- What is the name of this substance?
- The mass of this substance could be described as your (circle one): theoretical yield/experimental yield/percent yield

Lab Report: Mole Ratios and Reaction Stoichiometry

Reaction A: Sodium Bicarbonate and Hydrochloric Acid

Experimental Data

Mass of evaporating dish + watch glass	
Mass of evaporating dish + watch glass + sodium bicarbonate	
Mass of sodium bicarbonate used	
Mass of evaporating dish + watch glass + sodium chloride	
Mass of sodium chloride collected (experimental yield)	

Data Analysis

1. Use your data to determine the experimental mole-to-mole ratio between sodium bicarbonate and sodium chloride. Show your work for each step.

- Convert the mass of sodium bicarbonate used to moles.
- Convert the mass of sodium chloride collected to moles.
- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium bicarbonate and sodium chloride.

Simplest mole ratio *before* rounding:

_____moles $NaHCO_3$: _____moles $NaCl$

Simplest whole number mole ratio *after* rounding:

_____moles $NaHCO_3$: _____moles $NaCl$

2. Determine your percent yield of sodium chloride in reaction 12.13.3 Show your work for each step.

- Write the balanced equation for reaction 12.13.3– the reaction between sodium bicarbonate and hydrochloric acid.
- Using mass-to-mass stoichiometry, calculate the theoretical yield of $NaCl$ for reaction 12.13.3 Use your initial mass of sodium bicarbonate reactant as a starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.
- Calculate your percent yield of sodium chloride product.

Reaction B: Sodium Carbonate and Hydrochloric Acid

Experimental Data

Mass of evaporating dish + watch glass	
Mass of evaporating dish + watch glass + sodium carbonate	
Mass of sodium carbonate used	

Mass of evaporating dish + watch glass + sodium chloride	
Mass of sodium chloride collected (experimental yield)	

Data Analysis

1. Use your data to determine the experimental mole-to-mole ratio between sodium carbonate and sodium chloride. Show your work for each step.

- Convert the mass of sodium carbonate used to moles.
- Convert the mass of sodium chloride collected to moles.
- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium carbonate and sodium chloride.

Simplest mole ratio *before* rounding:

_____moles Na_2CO_3: _____moles $NaCl$

Simplest whole number mole ratio after rounding:

_____moles Na_2CO_3: _____moles $NaCl$

2. Determine your percent yield of sodium chloride in reaction 12.13.4 Show your work for each step.

- Write the balanced equation for reaction 12.13.4– the reaction between sodium carbonate and hydrochloric acid.
- Using mass-to-mass stoichiometry, calculate the theoretical yield of $NaCl$ for reaction 12.13.4 Use your initial mass of sodium carbonate reactant as a starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.
- Calculate your percent yield of sodium chloride product.

3. Is your percent yield here for reaction 12.13.4greater than or less than 100%? Give one possible source of error that could explain the percent yield you obtained.

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12.14: Lab 14 - Properties of Hydrates

Objectives

- Identification of hydrates in a group of compounds
- Investigation of the properties of hydrates
- Determination of the number moles of water of hydration in a hydrate

Water, the most common chemical on earth, can be found in the atmosphere as water vapor. Some chemicals, when exposed to water in the atmosphere, will reversibly either adsorb it onto their surface or include it in their structure forming a complex in which water generally bonds with the cation in ionic substances. The water present in the latter case is called *water of hydration* or *water of crystallization*. Common examples of minerals that exist as hydrates are gypsum ($\mathrm{CaSO_4 \cdot 2\,H_2O}$), Borax ($\mathrm{Na_3B_4O_7 \cdot 10\,H_2O}$) and Epsom salts ($\mathrm{MgSO_4 \cdot 7\,H_2O}$). Hydrates generally contain water in stoichiometric amounts; hydrates' formulae are represented using the formula of the anhydrous (non-water) component of the complex followed by a dot then the water ($\mathrm{H_2O}$) preceded by a number corresponding to the ratio of $\mathrm{H_2O}$ moles per mole of the anhydrous component present. They are typically named by stating the name of the anhydrous component followed by the Greek prefix specifying the number of moles of water present then the word *hydrate* (example: $\mathrm{MgSO_4 \cdot 7\,H_2O}$: magnesium sulfate heptahydrate).

Properties of Hydrates

It is generally possible to remove the water of hydration by heating the hydrate. Le Chatelier's principle predicts that an addition of heat to an endothermic reaction (heat is a "reactant") will shift the reaction to the right (product side). Heating will shift the equation of dehydration below to the right since it is an endothermic reaction. The residue obtained after heating, called the anhydrous compound, will have a different structure and texture and may have a different color than the hydrate.

Example:

$$\underbrace{\mathrm{CuSO_4 \cdot 5\,H_2O(s)}}_{\text{Deep Blue}} \xrightarrow{\Delta} \underbrace{\mathrm{CuSO_4(s)}}_{\text{Ashy White}} + \mathrm{5\,H_2O(g)} \tag{12.14.1}$$

$$\underbrace{\mathrm{CuSO_4(s)}}_{\text{Ashy White}} \xrightarrow{\mathrm{H_2O(l)}} \underbrace{\mathrm{CuSO_4(aq)}}_{\text{Deep Blue}} \tag{12.14.2}$$

Any anhydrous compound from a hydrate generally has the following properties:

- Highly soluble in water
- When dissolved in water, the anhydrous compound will have a color similar to that of the original hydrate even if it had changed color going from the hydrate to the anhydrous compound.

Most hydrates are stable at room temperature. However, some spontaneously lose water upon standing in the atmosphere, they are said to be *efflorescent*.

Other compounds can spontaneously absorb water from the surrounding atmosphere, they are said to be *hygroscopic*. Some hygroscopic substances, such as $\mathrm{P_2O_5}$ and anhydrous $\mathrm{CaCl_2}$, are widely used to "dry" liquids and gases (see experiment on the Molecular Weight of $\mathrm{CO_2}$); they are referred to as desiccants. Other hygroscopic substances, such as solid NaOH, absorb so much water from the atmosphere that they dissolve in this water, these substances are said to be *deliquescent*. Some compounds like carbohydrates release water upon heating by decomposition of the compound rather than by loss of the water of hydration. These compounds are not considered true hydrates as the hydration process is not reversible.

Formula of a Hydrate ($\mathrm{Anhydrous\ Solid \cdot \mathit{x}H_2O}$)

The formula of a hydrate can be determined by dehydrating a known mass of the hydrate, then comparing the masses of the original hydrate and the resulting anhydrous solid. The mass of water evaporated is obtained by subtracting the mass of the anhydrous solid from the mass of the original hydrate (12.14.3):

$$m_{\mathrm{H_2O}} = m_{\text{Hydrate}} - m_{\text{Anhydrous Solid}} \tag{12.14.3}$$

From the masses of the water and anhydrous solid and the molar mass of the anhydrous solid (the formula of the anhydrous solid will be provided), the number of moles of water and moles of the anhydrous solid are calculated as shown below (12.14.4, 12.14.5):

$$n_{\mathrm{H_2O}} = \frac{m_{\mathrm{H_2O}}}{MM_{\mathrm{H_2O}}} \tag{12.14.4}$$

$$n_{\text{Anhydrous Solid}} = \frac{m_{\text{Anhydrous Solid}}}{MM_{\text{Anhydrous Solid}}} \tag{12.14.5}$$

In order to determine the formula of the hydrate, [$\text{Anhydrous Solid} \cdot x\mathrm{H_2O}$], the number of moles of water per mole of anhydrous solid (x) will be calculated by dividing the number of moles of water by the number of moles of the anhydrous solid (Equation 12.14.6).

$$x = \frac{n_{\mathrm{H_2O}}}{n_{\text{Anhydrous Solid}}} \tag{12.14.6}$$

Procedure

Materials and Equipment

Solutions: Nitric Acid (6M)
Solids: Nickel (II) chloride, Cobalt (II) chloride, Sucrose, Calcium Carbonate, Barium chloride, Sodium tetraborate, Potassium chloride, Sodium sulfate hydrate, Iron (III) chloride, Potassium aluminum sulfate, Calcium chloride, Copper sulfate, Unknown.

Materials: Tests tubes (small or medium size), watch glasses, crucible* and cover*, crucible tongs, clay triangle.

Safety

Use caution when heating the crucible and cover. A hot crucible looks like a cold one, avoid direct contact with the crucible, clay triangle and ring stand until you are sure they are cooled. Use crucible tongs when cleaning the crucible with concentrated nitric acid.

Part A: Reversibility of hydration (optional, by Instructor)

In this section we will demonstrate the dehydration and re-hydration of cobalt (II) chloride hexahydrate. When heated gently, the red burgundy $\mathrm{CoCl_2 \cdot 6\,H_2O}$ will decomposes into the violet $\mathrm{CoCl_2 \cdot 2\,H_2O}$ then to the blue anhydrous $\mathrm{CoCl_2}$. When this anhydrous compound is dissolved in water it will go back to the original red burgundy color.

1. In an evaporating dish, gently heat a small amount (0.3 – 0.5 g) of $\mathrm{CoCl_2 \cdot 6\,H_2O}$ crystals until its color changes to violet then to blue.
2. When this color change appears to be complete, add 3 to 5 mL of water and observe the color of the dissolved substance.
3. Then reheat the solution to dryness.

Part B: Hygroscopic and Efflorescent Solids

In this section you will observe the changes in the physical properties of compounds, including wetness, color, structure, texture and mass. You should then decide if the compound is hygroscopic, efflorescent or neither using the change in the mass of the substance.

1. On an analytical balance, weigh a pea-sized sample of each of the compounds below on separate clean and dry watch glasses. Record the values as initial masses of containers and samples. Label and place all samples at the same location in the room, well out of the way so they won't be spilled.
2. After one hour, note any change in the physical appearance of each sample. Weigh the samples and record the masses as final masses. Calculate the change in mass for each sample. A substance is classified as efflorescent if its mass decreases by 0.005 g or more; and it is classified as hygroscopic if its mass increases by 0.005 g or more.

Compounds to be tested: $\mathrm{Na_2SO_4 \cdot 10\,H_2O}$, $\mathrm{FeCl_3}$, $\mathrm{KAl(SO_4)_2}$, $\mathrm{CaCl_2}$, $\mathrm{CuSO_4}$.

Part C: Identification of hydrates

In this section you will try to determine through the testing of a series of compounds, which ones are true hydrates. Some compounds may possess some of the properties of hydrates without being true hydrates. For a compound to be a true hydrate, it has to show all properties of true hydrates, including evolution of water upon heating, solubility of its anhydrous residue in water and reversibility in the color of the residue back to the color of the hydrate when dissolved in water.

1. For each of the chemical compounds below, place a pea-sized (≈30 mg) amount of the compound (just enough to cover the bottom of the test tube) in a dry test tube and note its color.
2. Heat the test tube and note any condensation that may appear at the mouth of the test tube as evidence of dehydration, note the color of the residue.
3. Let the residue cool down (put the test tube in a beaker not on a plastic test tube rack) then try to dissolve in about 3 mL of water (about 1/3 of the small test tube), warming gently if necessary to dissolve the residue (dissolve only substances that have shown condensation). Note the color of the dissolved residue.

If the compound possesses all three of the above-mentioned properties, it is a true hydrate; if at least one of them is not present, the compound is not.

Compounds to be tested: Nickel (II) chloride, Cobalt (II) chloride, Sucrose, Calcium Carbonate, Barium chloride, Sodium tetraborate, Potassium chloride.

Part D: Determination of the formula of a hydrate

In this section we will determine the number of moles of water present per mole of anhydrous solid in a given hydrate.

1. Using crucible tongs, clean a porcelain crucible and its cover using concentrated nitric acid (6 M). Pour the used nitric acid in the waste container provided. Rinse the crucible and its cover with distilled water.
2. Set the crucible with its cover slightly open on a clay triangle and heat strongly for at least 10 minutes. Allow the crucible to cool down to room temperature (do not set the hot crucible on the bench top). Weigh the crucible with its cover to the nearest 0.001 g.
3. Making sure to handle the crucible and its cover with clean tongs, add about 1 g (weighed to the nearest 0.001 g) of the unknown hydrate. Record the mass of the crucible, cover and sample. Heat the content of the crucible with its cover slightly open to allow the water of hydration to escape, first gently (about 10 minutes), then strongly (about 5 minutes).
4. Center the crucible's cover and let it cool down to room temperature. Weigh and record the mass of the cooled crucible with its cover and content (anhydrous residue).
5. Save the residue and perform your calculations. If the results of your calculations suggest that you have some water left in the residue, reheat your sample for an additional 5 minutes, allow it to cool down and weigh it again.

When you have completed the experiment, dissolve all your heated residues in water, put all your solids and liquids in the waste crock.

Pre-laboratory Assignment: Properties of Hydrates

1. A student trying to determine if a white solid is a true hydrate heats the sample and finds that there is evolution of water, that the residue obtained is soluble in water and that the solution is colorless. Is this a true hydrate? Explain.
2. 1 A student is given a cobalt (II) chloride hydrate. He weighs a clean and dry crucible with its cover and records a mass of 18.456 g. He then weighs the sample in the crucible and cover and obtains a mass of 19.566 g. He heats the sample, allows it to cool to room temperature and reweighs it to obtain a mass of 19.062 g. In the process, the sample's color changed from red-burgundy to blue.

- Mass of hydrate:
- Mass of anhydrous $CoCl_2$:
- Mass of water driven off:
- Moles of water:
- Moles of anhydrous $CoCl_2$:
- Moles of water per mole of $CoCl_2$:
- Formula of hydrate:

3. Why did the color of the hydrate change?

4. What color would you expect to see when this student dissolves the blue residue in water at the end of the experiment?

Lab Report: Properties of Hydrates

Part A: Reversibility of Hydration (Optional)

Record your observations:

Part B: Hygroscopic and Efflorescent Solids

	Substance	Initial mass of container and sample	Final mass of container and sample	Change in mass	Observations on structure, texture , wetness, etc. ...	Conclusion
1	$CaCl_2$					
2	$Na_2SO_4 \cdot 10\,H_2O$					
3	$KAl(SO_4)_2 \cdot 12\,H_2O$					
4	$CuSO_4$					
5	$FeCl_3$					

Part C: Hydrates

	Substance	Initial color	Water upon heating (Y/N)	Color Residue	Residue soluble (Y/N)	Color of residue dissolved	Hydrate (Y/N)
1	Nickel (II) chloride						
2	Cobalt (II) chloride						
3	Sucrose						
4	Calcium carbonate						
5	Barium chloride						
6	Sodium tetraborate						
7	Potassium chloride						

Part D: Determination of the formula of a hydrate

Data

1. Mass of crucible and cover:
2. Mass of crucible, cover and solid hydrate:
3. Mass of crucible, cover and anhydrous solid:

Calculations

1. Mass of hydrate:
2. Mass of anhydrous solid:
3. Mass of water lost:

4. Formula of anhydrous solid (from Instructor):
5. Molar mass of anhydrous solid:
6. Moles of H_2O present in the hydrate:
7. Moles of anhydrous solid present:
8. Ratio of moles H_2O:Anhydrous solid = x:
9. Formula of hydrate [$\text{Anhydrous solid} \cdot x H_2O$]:
10. Name this compound:
11. Unknown ID:

Questions:

1. Did the compound(s) that appeared wet in section B lose or gain water? Explain what may have happened.
2. What will be the effect, on the mass of the residue, of overheating the hydrate so that the compound decomposes. Will this likely lead to a higher or lower value of x than the actual value?
3. What will be the effect, on the mass of the residue, of not heating the hydrate enough to drive off all the water of hydration in the hydrate. Will this likely lead to a higher or lower value of x than the actual value?

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12.15: Lab 15 - The Properties of Oxygen Gas

Objectives

- To generate (and collect) oxygen gas via the decomposition of hydrogen peroxide.
- To investigate the properties of oxygen, particularly as an agent of combustion.

Oxygen is one of the most abundant elements on this planet. Our atmosphere is 21% free elemental oxygen. Oxygen is also extensively combined in compounds in the earths crust, such as water (89%) and in mineral oxides. Even the human body is 65% oxygen by mass.

Free elemental oxygen occurs naturally as a gas in the form of diatomic molecules, O_2 (*g*). Oxygen exhibits many unique physical and chemical properties. For example, oxygen is a colorless and odorless gas, with a density greater than that of air, and a very low solubility in water. In fact, the latter two properties greatly facilitate the collection of oxygen in this lab. Among the unique chemical properties of oxygen are its ability to support respiration in plants and animals, and its ability to support combustion.

In this lab, oxygen will be generated as a product of the decomposition of hydrogen peroxide. A catalyst is used to speed up the rate of the decomposition reaction (without being consumed by the reaction), which would otherwise be too slow to use as a source of oxygen. The particular catalyst used here is active yeast. 

The reaction for generating oxygen gas from hydrogen peroxide is shown below:

$$2\,H_2O_2(aq) \xrightarrow{catalyst} 2\,H_2O(l) + O_2(g) \tag{12.15.1}$$

$$\text{hydrogen peroxide} \xrightarrow{catalyst} \text{water} + \text{oxygen} \tag{12.15.2}$$

The oxygen gas produced will be collected in bottles by a method known as the downward displacement of water (see figure below). Once collected, several tests will be performed in order to investigate the role of oxygen in a variety of combustion reactions.

A combustion reaction is commonly referred to as “burning”. During a combustion reaction, oxygen reacts chemically with the substance being burned. Note that since our atmosphere is roughly 21% oxygen, many substances readily burn in air. Both oxygen and the substance being burned (the reactants) are consumed during the combustion reaction, while new substances (the products) and heat energy are generated. Since heat is produced, this is an exothermic reaction.

Combustion reactions can be described by the following chemical equation:

$$\text{Substance being burned} + \text{Oxygen} \longrightarrow \text{Products} + \text{Heat} \tag{12.15.3}$$

The actual products of a combustion reaction depend on what substance is burned and how much oxygen is present. In general, however, when a pure element burns in oxygen the product is called an oxide. An oxide is a compound containing both the element and oxygen chemically combined together.

Some examples of element combustion are shown below. Several such reactions will be performed using the oxygen gas collected in this lab.

Combustion of an Element can be described using the following chemical equation:

$$\text{Element} + \text{Oxygen} \longrightarrow \text{Oxide of Element} + \text{Heat} \tag{12.15.4}$$

Two examples are:

$$C(s) + O_2(g) \longrightarrow CO_2(g) + \text{Heat} \tag{12.15.5}$$

$$\text{Carbon} + \text{Oxygen} \longrightarrow \text{Carbon Dioxide} + \text{Heat} \tag{12.15.6}$$

$$2\,Hg(l) + O_2(g) \longrightarrow 2\,HgO(s) + \text{Heat} \tag{12.15.7}$$

$$\text{Mercury} + \text{Oxygen} \longrightarrow \text{Mercury (II) Oxide} + \text{Heat} \tag{12.15.8}$$

Procedure

Materials and Equipment

Materials: hydrogen peroxide solution, active yeast, wooden splints, candle, sulfur, steel wool, magnesium ribbon, aluminum pellets, 6 M hydrochloric acid

Equipment: 250-mL Erlenmeyer flask, five wide-mouth bottles, four glass 'cover' plates, pneumatic trough, "stopper + thistle tube + tubing" apparatus, utility clamp, ring stand, deflagration spoon, crucible tongs, small beaker, medium test tube and plastic test tube rack.

Safety

Exercise caution and/or use gloves when using the hydrogen peroxide (H_2O_2) and the hydrochloric acid (HCl) as they can cause chemical burns and skin irritation. If either of these chemicals comes into contact with your skin, immediately rinse with water for a minimum of fifteen minutes and notify your instructor.

Also, do not look directly at the burning magnesium. In addition to being very bright, it emits harmful UV radiation that could cause damage to the retina of the eye.

Personal protective equipment (PPE) required: lab coat, safety goggles, closed-toe shoes, gloves

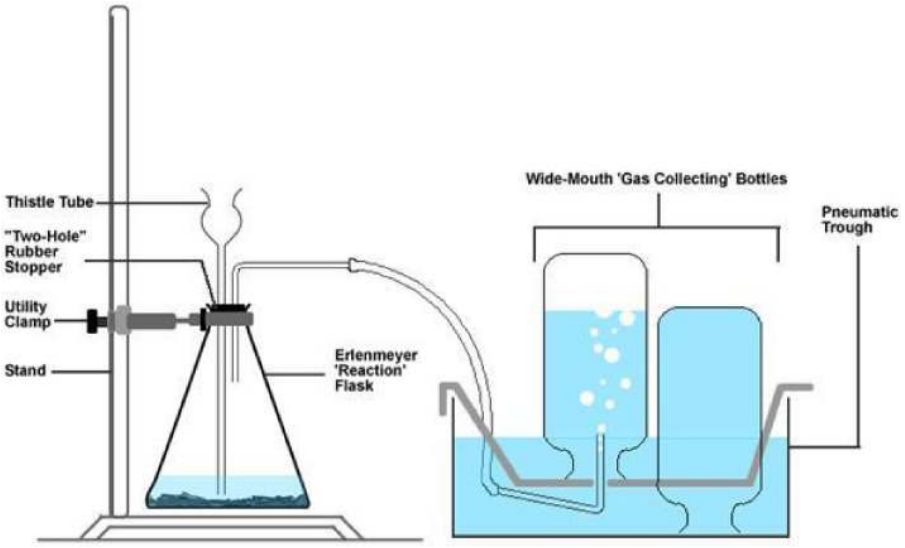
Part A: Generating and Collecting Oxygen Gas

1. Obtain the following equipment:

- 250-mL Erlenmeyer flask (locker)
- "two-hole stopper + thistle tube + glass tubing + rubber tubing" apparatus (stockroom)
- five wide mouth 'gas-collecting' bottles (under sink)
- four glass 'cover' plates (front desk)
- pneumatic trough (under sink) filled with water to 1⁄2 inch above the metal shelf

2. Fill four of the five wide-mouth bottles **to the brim** with water (the fifth will be used later). Then gently slide a glass plate over the mouth of each bottle. *Make sure that there are no air bubbles at the top of the glass plate.*
3. While holding the glass plate, gently invert a bottle and lower it into the water in the trough. Remove the glass plate when the mouth of the bottle is below the water level in the trough. Repeat this for all four bottles. Place the glass plates aside, as they will be used later.
4. Place one gas-collecting bottle on the metal shelf. Make sure that the mouth of the bottle is below the water level.
5. Now focus on your reaction vessel, the Erlenmeyer flask. Add a pea-sized amount of active yeast (the catalyst) to the flask, followed by approximately 25-mL of distilled water.
6. Finally, assemble all your equipment together as demonstrated by your instructor, or as shown in the figure below. Make sure that

- the end of the thistle tube is completely covered with water at the bottom of the flask,
- the end of the glass tubing running from the Erlenmeyer flask is inserted under the opening in the bottom of the metal shelf into the gas-collecting bottle (which is full of water),
- the Erlenmeyer flask is stabilized with a utility clamp



7. Obtain about 30-mL of hydrogen peroxide (H_2O_2) in your smallest beaker. Then carefully add about 10-mL of this H_2O_2 through the thistle tube. The generation of oxygen gas should begin immediately. If at any time the rate of the reaction in the Erlenmeyer flask appears to slow down, add another 10-mL portion of H_2O_2.
8. The oxygen produced will fill the inverted bottle by displacing the water in it. This is because oxygen does not dissolve in water, due to its low solubility. When the first bottle is completely filled with gas, place the second bottle on the metal rack in its place and allow it to fill in a like manner. Repeat this for the third and fourth bottles.
9. As soon as each bottle is completely filled, remove it by placing a glass plate under the bottle's mouth while under water, then lifting the bottle and plate from the pneumatic trough. Place the bottle on the lab bench **mouth up** and *do not remove the glass plate*. Since oxygen is denser than air, it sinks to the bottom of the flask and will not readily leak out the top.
10. Using masking tape, label each bottle of gas in the order they are collected: Bottle #1, Bottle #2, Bottle #3 and Bottle #4. Label the fifth unused bottle "Empty Bottle".
11. Once all four bottles are filled with oxygen, do not add any more H_2O_2 to the Erlenmeyer flask. Set it aside and allow the reaction to go to completion. At the end of the lab, the chemicals remaining in the reaction flask and any unused H_2O_2 must be disposed of in the labeled container in the hood. In the meantime, proceed to Part B.

Part B: The Properties of Oxygen Gas

Test 1: Combustion of wood

Light a wooden split, and then blow it out. While it is still glowing, place it in the empty wide- mouth bottle (air-filled only). Record your observations. Now re-light the same wooden split, and again blow it out. Quickly insert the splint into Bottle #1 (oxygen-filled) while it is still glowing. Record your observations.

Test 2: Combustion of candle wax

Place a tealight candle on a glass plate and light it. Lower the empty bottle over the candle. Measure and record the number of seconds that the candle continues to burn. Then re-light the candle and lower Bottle #2 over it. Again, measure and record the number of seconds that the candle continues to burn.

Test 3: Combustion of sulfur

This test must be performed in the hood. Take the empty bottle and Bottle #3 to the hood your instructor directs you to. Place a small lump of sulfur in a deflagrating spoon. Light the Bunsen burner in the hood, and heat the sulfur in the spoon. The sulfur will first melt, then burn with an almost invisible blue flame. Insert the spoon with the burning sulfur in the empty bottle and record your observations. Then insert it in Bottle #3, and again record your observations. When finished, extinguish the burning sulfur in a beaker of water.

Test 4: Combustion of iron

Pour about 30-mL of tap water into Bottle #4 and replace the glass plate quickly. Take a loose (frayed out) 2 or 3 centimeter piece of steel wool and hold it in a Bunsen burner flame for a very brief instant with your crucible tongs (it will glow red). Then

immediately lower the steel wool into Bottle #4. Record your observations. Repeat with the empty bottle and record your observations.

Test 5: Combustion of hydrogen

This test must be performed in air only. (Note: The hydrogen burned in this test must be first generated by a reaction between aluminum and hydrochloric acid.) To a medium test tube, add aluminum pellets, followed by about 3-mL of hydrochloric acid. Bubbles should begin to appear as hydrogen gas is produced. Place the test tube in the plastic test tube rack. After 30-60 seconds have elapsed, light a wooden splint. Do not blow it out. Hold the burning splint to the mouth of test tube (where the hydrogen gas is being evolved) and record your observations.

Test 6: Combustion of magnesium

This test is an instructor demonstration. It must be performed in air only. Hold a 1-inch piece of magnesium metal in a Bunsen burner flame with your crucible tongs until it ignites (in air). Record your observations, remembering not to look directly at the burning magnesium!

Lab Report: The Properties of Oxygen Gas

Part A: Generating and Collecting Oxygen Gas

1. Write the equation for the reaction used to generate oxygen gas.

- Word Equation:
- Balanced Chemical Equation:

2. What is the *name* of the catalyst used in this reaction?

- What is the purpose of this catalyst? (There are two parts to this answer.)

3. In addition to oxygen, what other substance is produced by this reaction? What happens to this substance?

4. Two notable physical properties of oxygen are its low solubility in water and a density greater than air.

- Which one of these properties allows the oxygen gas to be collected via the displacement of water? Explain.
- Which one of these properties allows the oxygen gas collected to be stored in the bottles mouth up? Explain.

5. What percentage of the air we breathe is oxygen gas?

Part B: The Properties of Oxygen Gas

Test	Observations
Test 1	*Observations*
Glowing splint in empty bottle	
Glowing splint in Bottle #1	
Test 2	*Observations*
Burning candle in empty bottle	
Burning candle in Bottle #2	
Test 3	*Observations*
Burning sulfur in empty bottle	
Burning sulfur in Bottle #3	
Test 4	*Observations*
Glowing steel in empty bottle	
Glowing steel in Bottle #4	
Test 5	*Observations*

Test	Observations
Burning hydrogen in air	
Test 6	*Observations*
Burning magnesium in air	

Analysis of Combustion Results

1. Why did you perform each test in two separate bottles (air-filled and oxygen-filled)?
2. Are the combustion reactions of oxygen exothermic or endothermic? Support your answer with experimental evidence from the tests you performed.
3. Consider your results for the first four tests you performed. In which bottles, air-filled or oxygen-filled, did the combustion reactions occur more vigorously? Why?
4. Consider your Test 2 results. Although the candle burns for a longer period of time in one bottle, it eventually goes out in both the empty bottle and Bottle #2. Why does it go out?
5. When an element burns in oxygen gas, the product is called an oxide.

- The wood in the splint consists mostly of carbon. The combustion of carbon produces carbon dioxide, CO_2. Write the equation for the combustion of wood (carbon).
  - Balanced Chemical Equation:
- The combustion of sulfur produces sulfur dioxide, SO_2. Write the equation for the combustion of sulfur.
  - Balanced Chemical Equation:
- The combustion of hydrogen produces water, H_2O. Write the equation for the combustion of hydrogen.
  - Balanced Chemical Equation:
- Steel wool consists mostly of iron. The combustion of iron produces iron(III) oxide, Fe_2O_3. Write the equation for the combustion of steel wool (iron).
  - Balanced Chemical Equation:

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12.16: Lab 16 - The Dumas Method

Objectives

- To determine the molecular mass of an unknown volatile liquid using the Dumas method and the ideal gas law.

In the early 19th century, Jean-Baptiste Dumas, a distinguished French chemist, created a relatively simple method for determining the molecular mass of a volatile substance. In this experiment we will use a modified version of his technique to determine the molecular mass of an unknown volatile liquid.

The density of a gas is given by the ideal-gas equation as,

$$D_{gas} = \frac{m}{V} = \frac{PM}{RT} \tag{12.16.1}$$

where M is the molecular mass of the gas. Solving for molecular mass, we obtain:

$$M = \frac{mRT}{PV} \tag{12.16.2}$$

Thus, the molecular mass of a gas can be determined by measuring the **temperature**, **pressure**, **mass**, and the **volume** of a substance in its gaseous phase.

In this experiment we shall use a 125-mL Erlenmeyer flask in place of the glass bulb used by Dumas. A small sample of an unknown volatile liquid will be placed in the flask and the liquid vaporized by immersing the flask in a hot water bath. A piece of aluminum foil will be used to seal the flask and a tiny pinhole made in the foil to allow excess vapor to escape. The **temperature** of the gas will be determined by measuring the temperature of the water bath surrounding the flask and assuming that the gas and the water bath are in thermal equilibrium. Because an excess initial quantity of liquid is used, the volume of vapor produced from the liquid is greater than the volume of the flask. Upon heating, the vapor that is created initially pushes the air out of the flask and then the vapor begins exiting the flask through the pinhole until the pressure inside the flask is equal to the atmospheric pressure. Thus, the **pressure** of the gas can be determined by measuring the air pressure in the laboratory using a barometer. After the last of the liquid is vaporized the flask is removed from the hot water bath and the vapor inside the flask is allowed to cool and recondense. The **mass** of the vapor is measured by weighing the condensate remaining in the flask. The **volume** of the vapor is equal to the volume of the flask. To determine the volume of the flask we will weigh the flask after filling it to the brim with deionized water and then measure the temperature of the water. From the measured mass of the water and the known density of water at a given temperature, we can determine its volume, and thus the volume of the flask and the vapor. Three trials will be performed to determine the average molecular mass of the unknown liquid. A Table of densities of water at various temperatures is given below:

Density of Liquid Water

Temperature (°C)	Density (g/mL)
15	0.9991
16	0.9989
17	0.9988
18	0.9986
19	0.9984
20	0.9982
21	0.9980
22	0.9978
23	0.9975
24	0.9973

Temperature (°C)	Density (g/mL)
25	0.9970
26	0.9968
30	0.9957

Information from the CRC *Handbook of Chemistry and Physics*, 64th ed., 1983-4

Procedure

Materials and Equipment

You will need the following additional items for this experiment: second utility clamp split stopper, ring stand, hot plate, aluminum foil, rubber band, 1 small dithizone dye crystal, 2-3 boiling chips

SAFETY

Students should wear goggles. All heating should be done in the fume hood.

WASTE DISPOSAL: All chemicals used must go in the proper waste container for disposal.

Experimental Set-up and Procedure

1. Using a hot plate, a ring stand, two utility clamps, your 600-mL beaker, a 125-mL Erlenmeyer flask, a split stopper and your thermometer, assemble the apparatus shown in the Figure below inside one of the laboratory fume hoods. Use the split stopper to help clamp the thermometer. Be certain there is sufficient water in the 600-mL beaker to surround the bottom and neck of the 125-mL Erlenmeyer flask, but no so much that it overflows upon immersion of the flask. You do not need the foil and rubber band attached at this time. To assist in observation of the unknown liquid it is a good idea to mount the flask at a slight angle, as shown. Take special care that none of the water enters the flask when placing it in the beaker.



2. Again note that the level of the water in the hot water bath should be almost to the brim of the beaker when the flask is inserted, so that it covers as much of the neck of the clamped Erlenmeyer flask as possible. When you have the basic set-up accomplished, remove the flask. Place several boiling chips in the water bath and begin heating the bath. You may want to keep a squirt bottle of water handy to help maintain the level of water in the bath between trials.

3. Fill the 125-mL Erlenmeyer flask to the brim with deionized water and weigh the flask being careful not to spill any of the water out. Be sure the outside of the flask is completely dry. Record this mass. Using your thermometer measure the temperature of the water in the flask. Record this temperature.
4. Dry the 125-mL Erlenmeyer flask thoroughly inside and out using either an acetone rinse or by heating the flask over a Bunsen burner and then allowing it to cool to room temperature (your instructor will tell you which procedure to use). *If you use acetone be certain that none of the waste acetone goes down the sink; all chemical waste should be placed in the proper waste container.* Place the dry empty flask on the digital balance and record its mass.
5. Subtract the mass of the dry empty flask from the mass of the flask filled with water to determine the mass of water in the flask. Using Table 1 in the background section of this experiment determine the density of the water at the temperature recorded. Using these values determine the volume of your flask. Record these values.
6. If there is a barometer in your laboratory room, use it to measure the atmospheric pressure and record this value on your report form. Otherwise your instructor should provide this value. Using your forceps (tweezers), add a single crystal – just a single speck about the size of a grain of pepper – of the dry dithizone dye to the flask. Cover the flask mouth with a small piece of aluminum foil and secure in place with a rubber band. (Keep the aluminum foil cap as short as possible, so that it does not extend down more than a few centimeters below the position of the rubber band; this will help keep water from getting trapped between the foil and the flask.) Using a push-pin, make a tiny pinhole in the foil cap. Measure and record the mass of the dry flask containing the dithizone crystal with the foil and rubber band in place using the digital balances.
7. Obtain an unknown sample from your instructor and record the unknown number on your report form (your instructor may also require you to attach the label from the unknown to your report).

Trials 1 – 3

1. Bring the water in the bath to a gentle boil. Careful remove the foil cap so as not to tear it and add approximately 2 mL of your unknown liquid to the flask. Swirl to dissolve the dithizone dye in the unknown liquid. Secure the cap to the flask and carefully immerse the flask in the boiling water bath. Clamp the neck of the flask to the support at a slight angle as shown on the previous page using your utility clamp and lower it so that as much of the neck of the flask as possible is submerged in the boiling water bath. It is extremely important that no water gets inside the flask and that the water in the bath is already boiling when you immerse the flask. Mount the thermometer in the bath, so it is as close to the unknown as possible.
2. The unknown is a volatile liquid that boils at a temperature below that of water. Consequentially, the liquid inside the 125-mL Erlenmeyer flask will begin to boil when immersed in the boiling water bath. Watch the liquid carefully. After driving out all of the air in the flask the vapor from the unknown liquid should begin to exit the pinhole in the foil. The unknown liquid should vaporize completely (as evidenced by the disappearance of the color of the dissolved dithizone dye and the dye precipitating out on the wall of the flask). Keep the flask in the boiling water for three minutes after the last of the unknown liquid vaporizes to ensure that the vapor is in complete thermal equilibrium with the boiling water in the bath. Record the temperature of the boiling water indicated on your thermometer.
3. Carefully remove the 125-mL flask from the boiling water bath, being certain not to get any water or water vapor into the flask. (The vapor inside the flask will recondense as it cools and so you should observe some liquid form at the bottom of the flask after removing it from the water bath.) Dry the outside of the flask thoroughly using a paper towel and allow it to stand away from any heat source until it reaches room temperature. Make sure there are no water droplets on the outside of the aluminum cap. Reweigh the flask (with the condensed unknown liquid inside and the aluminum cap and rubber band in place) and record this mass.
4. Remove the cap, add another 2 mL or so of your unknown liquid. Repeat the steps above twice more for a total of three trials beginning at the step following the heading, “Trials 1 – 3” above. (There is no need to rinse and dry the flask or add more dithizone dye between trials.)
5. Finally, discard the aluminum foil, return the rubber band and equipment your borrowed, and pour the liquid in the flask into the chemical waste container. You may need to use a small amount of your remaining unknown solution or acetone to rinse out all of the dithizone dye.
6. Using the data you collected and the equations presented in the background section of this experiment determine the average molecular mass of your unknown liquid and then answer the questions on the report form.

Pre-laboratory Assignment: Determination of Molar Mass via the Dumas Method

A student performs the Dumas bulb experiments and collects the following data:

- Mass of empty flask: 31.022 g

- Mass of flask filled with water: 161.175 g
- Temperature of water in flask: 22.7°C
- Atmospheric pressure: 759.1 mmHg

She then performs the experiment with her unknown liquid and records the following data:

- Temperature of boiling water bath: 100.1°C
- Mass of condensate after boiling: 0.306 g

Use these data to find the molecular mass of her unknown liquid. Show all work below.

Lab Report: Determination of Molar Mass via the Dumas Method

Data

Unknown Number:	
Mass of the flask filled to the brim with water:	
Temperature of the water inside the flask:	
Mass of dry empty flask:	
Mass of water in flask (*difference of flask + water and empty flask*):	
Density of water at the recorded temperature (from Table 1):	
Volume of flask: (*determined from your measured mass of water and its density*)	
Atmospheric Pressure (from barometer):	
Mass of the dry flask, dithizone crystal, foil, and rubber band:	

Data:

Trial One:	
Temperature of boiling water bath:	
Mass of flask, dithizone, foil, rubber band, and condensate:	
Mass of condensate (subtract mass of flask, foil, etc. from above):	
Trial Two:	
Temperature of boiling water bath:	
Mass of flask, dithizone, foil, rubber band, and condensate:	
Mass of condensate (subtract mass of flask, foil, etc. from above):	
Trial Three:	
Temperature of boiling water bath:	
Mass of flask, dithizone, foil, rubber band, and condensate:	
Mass of condensate (subtract mass of flask, foil, etc. from above):	

Results

Determine the molecular mass of your unknown liquid for each of your three trials. Show all of your calculations below. Record the average molecular mass of your unknown liquid below.

Record the average molecular mass of your unknown liquid: _____ Unknown number: _____

Questions

Indicate how each of the following possible sources of error would be expected to affect your calculated molecular mass. (Indicate both the directions and, when possible, the approximate magnitude of the errors.) Justify each of your answers with supporting rationale or equations.

1. Air was not completely swept out of the flask by the unknown liquid after it had completely vaporized, so that the gas in the flask was a mixture of air and unknown.
2. The volume of the flask at 100°C was greater than that at room temperature because of expansion of the glass. (The expansion of the glass is 1.4×10^{-5} deg^{-1}. That is, the volume of the flask increases by 1.4×10^{-3} percent per degree rise in temperature.)
3. The vapor was not an ideal gas. (Hint: Remember that the temperature of the vapor was not far above the boiling point of the unknown liquid. This fact should tell you the direction of the deviation from ideality.)
4. Although the unknown liquid completely vaporized, the gas in the bulb never reached the temperature of the boiling water.

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12.17: Lab 17 - The Molecular Weight of Carbon Dioxide

Objectives

- To determine the molecular weight (molar mass) of carbon dioxide based on measurements of the pressure, temperature, volume and mass of a sample of the gas.
- To compare the experimental molecular weight to the theoretical molecular weight of carbon dioxide.

From the ideal gas law, $PV = nRT$, it is easy to show that the density of an ideal gas is:

$$D_{gas} = \frac{m}{V} = \frac{PM_w}{RT} \tag{12.17.1}$$

where M_w is the molecular weight (molar mass) of the gas.

This can be solved to obtain:

$$M_w = \frac{mRT}{PV} \tag{12.17.2}$$

Thus, the molecular weight (or molar mass) of a gas can be determined by measuring the **temperature**, **pressure**, **volume** and the **mass** of a sample of the gas.

In this experiment the molecular weight of carbon dioxide will be determined. The carbon dioxide gas is produced by the reaction of calcium carbonate with hydrochloric acid.

$$\mathrm{CaCO_3(s) + 2\,HCl(aq) \longrightarrow CO_2(g) + CaCl_2(aq) + H_2O(l)} \tag{12.17.3}$$

The carbon dioxide gas will be collected in an Erlenmeyer flask that is covered with aluminum foil. The foil and the fact that carbon dioxide gas is denser than air inhibit the rate at which the carbon dioxide mixes with the outside air.

The **temperature** of the gas will be measured by briefly placing a thermometer underneath the foil.

Since the flask is not sealed the **pressure** of the gas in the flask will be equal to the pressure in the laboratory. The laboratory pressure will be measured with a mercury barometer.

The **volume** of the gas is equal to the volume of the flask. The volume of the flask will be measured by filling it to the brim with deionized water. From the measured mass of the water and the known density of water the volume of the water, and thus the volume of the flask, will be determined. A table of densities of water at various temperatures is given below.

Density of Liquid Water

page2image2224 **Temperature (°C)**	page2image3168 **Density (g/mL)**
15	0.9991
16	0.9989
17	0.9988
18	0.9986
19	0.9984
20	0.9982
21	0.9980
22	0.9978
23	0.9975
24	0.9973
25	0.9970

Temperature (°C)	Density (g/mL)
26	0.9968
30	0.9957

Information from the CRC *Handbook of Chemistry and Physic*s, 64th ed., 1983-4

The **mass** of gas in the flask can be determined by weighing the gas filled flask and then subtracting the mass of the empty flask. In most experiments the mass of an "empty" container is actually the mass of the container plus the air it holds. This is generally not a problem since the mass of the air is relatively small compared to the mass of a liquid or solid sample and can therefore be neglected. However when weighing a gas, not accounting for the mass of the air in the container would be a serious systematic error. The mass of air in the flask will be estimated from the volume of the flask and the known density of dry air. A table of dry air densities is given below. It is reasonable to approximate the air as being dry since the mole fraction of water vapor in air is very small (even in air that is saturated with water vapor the mole fraction of water vapor is only 0.03). The mass of the empty flask will then be calculated by subtracting the calculated mass of dry air from the measured mass of the flask filled with air. After the flask has been filled with carbon dioxide, it will be weighed, and the mass of the gas will be determined by subtracting the calculated mass of the truly empty flask.

Density of Dry Air

Temperature (°C)	Density (g/L)		
--	P = 750 torr.	P = 760 torr.	P = 770 torr.
17	1.201	1.217	1.233
18	1.197	1.213	1.229
19	1.193	1.209	1.225
20	1.189	1.205	1.221
21	1.185	1.201	1.216
22	1.181	1.197	1.212
23	1.177	1.193	1.208
24	1.173	1.189	1.204
25	1.169	1.185	1.200

Information from the CRC *Handbook of Chemistry and Physics*, 64th ed., 1983-4

Procedure

Materials and Equipment

The following items must be obtained from the stockroom.

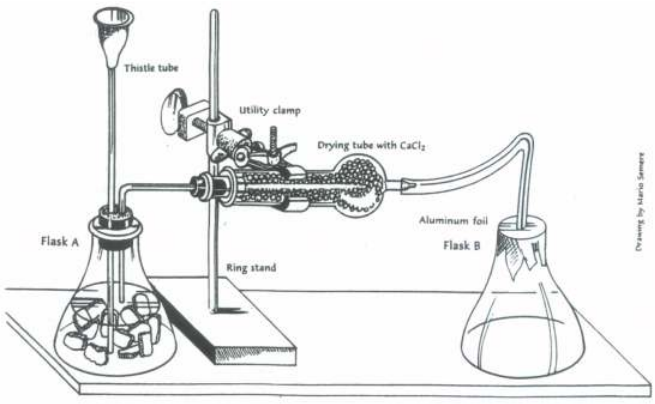
- A straight piece of glass tubing connected to a piece of rubber tubing
- A two holed rubber stopper with a thistle tube inserted into one hole and a piece of bent glass tubing (connected to a piece of rubber tubing) inserted into the other hole
- A $CaCl_2$ drying tube
- A piece of aluminum foil that is approximately 15 cm by 15 cm and two 250-mL Erlenmeyer flasks are also needed. **Make sure that the rubber stoppers will fit tightly into the mouth of each Erlenmeyer flask.**

Safety

6 M hydrochloric acid is capable of causing serious chemical burns and blindness and should be handled with care. If the acid comes in contact with your skin immediately rinse the affected area with water for several minutes. As always, you should be

wearing your safety goggles.

1. Obtain one of the 250 mL Erlenmeyer flasks. This will be designated as flask B. Make sure that the flask is both clean and dry. Record the mass of the flask and the foil to the nearest 0.001 g using the analytical balance. Note that it is mass (flask + foil + air) that is being measured here. Also record the temperature of the air in the flask.
2. Place about 25 grams of calcium carbonate into the other 250 mL Erlenmeyer flask designated as flask A. Add about 10 mL of water, or enough to cover the chips completely. Insert the two-holed rubber stopper (with the thistle tube and the bent tube) into flask A, making sure that the thistle tube is adjusted so that it is beneath the water but not touching the bottom of flask. The stopper must fit tightly into the flask. Also obtain about 25 mL of dilute (6 M) hydrochloric acid, HCl, in a small, labeled beaker.
3. Assemble the apparatus shown in figure 1. Insert the straight glass tube (the one that is attached to the flexible tubing but not attached to the one-holed rubber stopper) into flask B by placing it between the foil and the flask and pressing the foil against it to hold it in place. Be careful to fold and shape the foil only as much as necessary since it is fragile and will easily tear. Attach the flexible end of the tubing to the drying tube. The small rubber stopper attached to the bent glass tubing should be inserted into the other end of the drying tube.



Apparatus for Generation and Collection of Carbon Dioxide Gas

4. When all is ready pour 5–10 mL of the hydrochloric acid into the top of the thistle tube and allow it to run through the tube and into flask A. The reaction should begin immediately as evidenced by gas evolution. Allow the reaction to continue for at least 20 minutes to displace all of the air in flask B with carbon dioxide gas. During this time pay attention to what is happening in flask A; if the gas evolution ceases, add additional HCl solution through the thistle tube. After the 20 minutes remove the tube from flask B (keep the foil in place) and **immediately** weigh the flask containing carbon dioxide on the analytical balance to the nearest 0.001 g.
5. Reassemble the apparatus and allow gas evolution to continue and flow into flask B for an additional 15 minutes. Again, weigh flask B (with the foil). The two masses (before and after the 15 minutes) should agree closely (to within 0.005 g). If the mass has increased by more than 0.005 g reassemble the apparatus and continue to collect carbon dioxide gas for an additional 5 minutes then reweigh the flask. Continue this procedure until successive masses agree to within 0.005 g (or until a decrease in mass is observed). Next, measure and record the temperature of the carbon dioxide in the flask. Use the barometer to read the atmospheric pressure.
6. Finally, in order to determine the volume of flask B, fill this flask with deionized water to the brim, wipe any water from the outside of the flask and weigh it, with the foil, to the nearest 0.01 g using the **triple beam balance**. Next, measure and record the temperature of the water. The volume of water in the flask can now be calculated from the mass and density of water in the flask at the temperature recorded.

Cleanup

Neutralize the contents of flask A and any unused HCl by adding solid sodium bicarbonate until the solution no longer fizzes upon addition of additional bicarbonate. **Remember that goggles should be worn at all times in the laboratory**. This step is best performed in the sink as the flask tends to overflow. Decant off the liquid, leaving any solid marble chips behind. Rinse the chips with water and decant off the water. Pour the chips into the labeled recovery beaker.

Pre-laboratory Assignment: Molecular Weight of a Gas

1. Write a balanced equation including state symbols for the reaction that takes place in flask A.
2. A flask holds a total of 402.5 g of water at 20.4°C. What is the volume of the flask?
3. Assuming that the pressure is 758.4 torr and the temperature is 24.6°C, calculate the mass of dry air in the flask.
4. Assuming the mass of the flask filled with dry air is 125.383 g, calculate the mass of the empty flask.
5. Assuming the mass of the flask filled with CO_2 is 125.557g , the temperature is 23.2°C, and the pressure is 758.4 torr, calculate the molecular weight of CO_2.
6. Calculate the actual molecular weight of CO_2 and determine the percent error.

Lab Report: Molecular Weight of a Gas

Data and Observations

Mass of foil and flask filled with air	g
Temperature of dry air in flask	°C
Barometric pressure	mm Hg (torr) = atm
Density of dry air at experimental temperature and pressure	g/L = g/mL
Mass of foil and flask filled with CO_2, after first twenty minutes	g
Mass of foil and flask filled with CO_2, after second fifteen minutes	g
Mass of foil and flask filled with CO_2, after additional five min (if needed)	g
Temperature of CO_2 in flask	°C = K
Mass of flask filled with water	g (with foil)
Temperature of water	°C
Density of water at above temperature	g/mL

Calculations and Analysis

- The volume of the flask is equal to the volume of water in the flask. This volume is calculated from the mass and the density of the water. The mass of the water is the difference between the $\text{mass}_{\text{(flask + foil + water)}}$ and the $\text{mass}_{\text{(empty flask + foil)}}$. However, since the mass of the water is much larger than the mass of air displaced by the water the mass of water can be approximated as the difference between the $\text{mass}_{\text{(flask + foil + water)}}$ (#9 on Data Sheet) and the $\text{mass}_{\text{(flask + foil + air)}}$ (#1 on Data Sheet). The density of water at different temperatures is tabulated above.
- The mass of dry air in the flask is calculated from the volume of the flask and the density of dry air at the recorded temperature and pressure. The density of dry air at different combinations of temperature and pressure is tabulated above.
- The mass of the empty flask is equal to the difference between the $\text{mass}_{\text{(flask + foil + air)}}$ (#1 on Data Sheet) and the calculated mass of dry air.
- The mass of carbon dioxide in the flask is determined by subtracting the mass of the empty flask from the largest value of $\text{mass}_{\text{(flask + foil + carbon dioxide)}}$ (the largest value of #6 or #7 on the Data Sheet).
- The molecular weight of carbon dioxide is calculated using Equation 12.17.2
- The percentage error can be calculated using the following formula.

$$\text{Percentage Error} = \frac{|(\text{Experimental Value}) - (\text{Theoretical Value})|}{(\text{Theoretical Value})} \times 100 \qquad (12.17.4)$$

For each calculation, show your work and put a box around each answer.

1. Volume of flask

2. Mass of dry air in flask
3. Mass of empty flask
4. Mass of carbon dioxide in flask
5. Experimental molecular weight
6. Percentage Error

Questions and Conclusions

1. Because the mass of the flask with water exceeds the amount that can be weighed on the digital balances, the triple-beam-balance is used, which only weighs to ±0.01 g. Why doesn't using this balance affect the number of significant digits to which the experimental molecular weight of carbon dioxide can be reported?
2. Why can we neglect the mass of the air in the flask when obtaining the mass of the water in the flask (as we do when weighing most things) but not neglect the mass of the air when we calculate the mass of carbon dioxide in the flask?
3. Suggest a possible source of error that would account for any discrepancy between your experimentally determined molecular weight and the theoretical molecular weight.

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12.18: Lab 18 - Calorimetry and Hess's Law

Objectives

- To experimentally measure the ΔH values of two reactions using the technique of constant pressure calorimetry.
- To apply these ΔH values in a Hess's Law calculation to determine the enthalpy of combustion of a metal.

The combustion of a metal in oxygen produces the corresponding metal oxide as the only product. Such reactions are exothermic and release heat. For example, the combustion of iron releases 1651 kJ of heat energy for every four moles of iron burned:

$$4\,\mathrm{Fe(s)} + 3\,\mathrm{O_2(g)} \longrightarrow 2\,\mathrm{Fe_2O_3(s)} \qquad \Delta H_1 = -1651kJ \tag{12.18.1}$$

Since it is difficult to measure the enthalpy of combustion of a metal directly, in this lab it will be determined indirectly by applying Hess's Law of Heat Summation. Hess's Law states that the enthalpy change of an overall process is equal to the sum of the enthalpy changes of its individual steps.

Hess's Law Example 12.1 :

Determine ΔH for the target reaction $2\,\mathrm{NO_2(g)} + \frac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{N_2O_5(g)}$ given the following information,

$$\text{Reaction A} \qquad \mathrm{N_2O_5(g)} \longrightarrow 2\,\mathrm{NO(g)} + \tfrac{3}{2}\mathrm{O_2(g)} \qquad \Delta H_A = +223.7kJ \tag{12.18.2}$$

$$\text{Reaction B} \qquad \mathrm{NO_2(g)} \longrightarrow \mathrm{NO(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \qquad \Delta H_B = -57.1kJ \tag{12.18.3}$$

Solution

Reactions A and B have to be carefully manipulated before they can be summed to produce the target reaction. Reaction A must be reversed, causing a sign change to ΔH_A. Reaction B must be multiplied by a factor of 2, causing ΔH_B to be multiplied by 2. Only then will they yield the target equation when added together:

$$2\,\mathrm{NO(g)} + \tfrac{3}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{N_2O_5(g)} \qquad \Delta H = -(+223.7) = -223.7kJ \tag{12.18.4}$$

$$+ \quad (2\,\mathrm{NO_2(g)} \longrightarrow 2\,\mathrm{NO(g)} + \mathrm{O_2(g)}) \qquad \Delta H = 2 \times (-57.1) = -114.2kJ \tag{12.18.5}$$

$$2\,\mathrm{NO_2(g)} + \tfrac{1}{2}\mathrm{O_2(g)} \longrightarrow \mathrm{N_2O_5(g)} \qquad \text{Target} \tag{12.18.6}$$

Thus, $\Delta H_{\text{Target}} = -223.7 + (-114.2) = -337.9kJ$

In order to use Hess's Law to find the heat of combustion of a metal, it is first necessary to obtain reaction enthalpies (ΔH values) for equations that can be summed together appropriately. To accomplish this, two reactions will be studied in this lab. In one reaction, a given metal will react with hydrochloric acid producing hydrogen and the metal chloride. In the other reaction, the corresponding metal oxide will react with hydrochloric acid producing water and the metal chloride. For example, the reactions involving iron and iron(III) oxide are as follows:

$$2\,\mathrm{Fe(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{H_2(g)} \qquad \Delta H_2 \tag{12.18.7}$$

$$\mathrm{Fe_2O_3(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{H_2O(l)} \qquad \Delta H_3 \tag{12.18.8}$$

Since both reactions are exothermic, the heat released (q) will be absorbed into the surrounding reaction mixture. As long as the reactions are performed in an insulated container (such as a coffee cup calorimeter) there will be negligible heat exchange with the container walls or outside air. By monitoring the temperature of the reaction mixture when specific quantities of reactants are used, the amount of heat (in J) released by these reactions can be determined by applying the equation:

$$\text{heat released by reaction}(-q_{\text{reaction}}) = \text{heat absorbed by reaction mixture}(+q_{\text{mixture}}) = (m \times c \times \Delta T)_{\text{mixture}} \tag{12.18.9}$$

Here m is the total mass of the reaction mixture (in g), ΔT is the maximum temperature change that occurs during the reaction (in °C), and c is the specific heat capacity of the mixture (in J/g°C). Note that since the reactions occur in aqueous solution, it is reasonable to substitute the specific heat capacity of water (= 4.184 J/g°C) for the specific heat capacity of the mixture.

Recall that at constant pressure (the conditions of this experiment), the heat released by the reaction equals the reaction enthalpy:

$$q_P = \Delta H \tag{12.18.10}$$

Since the heat released by each reaction is proportional to the amount of metal/metal oxide used, ΔH_2 and ΔH_3 can be easily calculated per gram or mole of metal/metal oxide used. It should be noted that reactions 12.18.7and 12.18.8by themselves still cannot be summed to

produce Reaction 12.18.1. Another reaction is required:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l) \qquad \Delta H_4 \tag{12.18.11}$$

ΔH for this reaction (the formation of water from its elements) must be obtained from tabulated thermodynamic data in the textbook. Finally, the reactions 12.18.7, 12.18.8 and 12.18.11 and their enthalpies may be summed together according to Hess's Law to determine the enthalpy of combustion of the given metal 12.18.1.

Procedure

Materials and Equipment

Mg (*s*), MgO (*s*), Zn (*s*), ZnO (*s*), Al (*s*), Al_2O_3 (*s*), 1 M HCl (*aq*), 6 M HCl (*aq*), coffee cup calorimeter with lid*, thermometer*, looped stirring rod*, slotted stopper*, 100-mL graduated cylinder, 50-mL beaker, utility clamp, stand, electronic balance, and wash bottle.

*Items with an asterisk must be checked out from the stockroom.

Safety

Data Acquisition

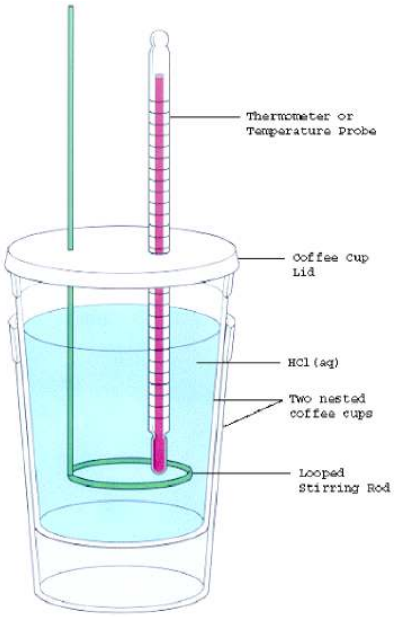
Instead of a thermometer, some sections may use a data acquisition system (laptop computer, Vernier® interface, temperature probe, and LoggerPro® software) to directly monitor temperature changes over time. Detailed instructions for setting up this system will be provided at the beginning of your lab period. Please note that your experimental procedure will still be the same regardless of the method used to monitor temperature.

The Heat of Combustion of a Metal/Metal Oxide

1. You will be assigned a specific metal/metal oxide pair to investigate by your instructor. Record their identities on your report form. Note that you will perform the following procedure for a total of four times, twice with the metal, then twice with the metal oxide.
2. The table below indicates the quantities of reactants to be used for each metal/metal oxide combination. Note that the reactions involving Zn and Al require the concentrated 6 M acid.

Mg / MgO	Zn / ZnO	Al / Al_2O_3
0.15 g Mg, 25 mL 1 M HCl	0.40 g Zn, 25 mL 6 M HCl	0.15 g Al, 25 mL 6 M HCl
0.25 g MgO, 25 mL 1 M HCl	0.60 g ZnO, 25 mL 6 M HCl	0.75 g Al_2O_3, 25 mL 6 M HCl

3. Use an electronic balance to weigh your empty, dry calorimeter (the two nested Styrofoam® cups). Remove it from the balance, then pour approximately 25-mL of HCl (*aq*) into it and weigh it again. Record these masses on your report (the difference is the mass of HCl (*aq*) used).
4. Now weigh an empty, dry 50-mL beaker. Remove it from the balance, then add the recommended mass of your assigned metal to it and weigh it again. Record these masses on your report (the difference is the mass of metal used).
5. Assemble your equipment as shown in the figure below. The thermometer (or temperature probe) and the stirring rod must be inserted through the holes in the calorimeter lid. The thermometer bulb should be immersed in the acid, but not touch the bottom of the calorimeter. Clamp the thermometer in place using the slotted stopper and utility clamp.
6. Measure the temperature of the HCl in the calorimeter (while covered with the lid). Once thermal equilibrium is established, record the temperature. Next, carefully add the metal sample to the acid. Quickly replace the lid and monitor the temperature change until the reaction is complete. Stir the mixture continuously with the stirring rod as the reaction occurs. Record the maximum temperature achieved by the mixture. Note that the mixture first warms up as the reaction occurs, but will then gradually cool as heat is lost to the surroundings. However, as Styrofoam is a poor conductor of heat this cooling will occur slowly. Thus it will be very easy for you to identify the maximum temperature.
7. When finished, dispose of your chemical waste as directed by your instructor. Then rinse the calorimeter, thermometer and stirring rod thoroughly with distilled water, dry, and repeat the experiment again. Once you have completed both trials with the metal, perform your two trials using the metal oxide using the identical procedure.



Note that the thermometer must be clamped in place using the slotted stopper and utility clamp/stand. You may also want to place the nested cups in a medium beaker for extra stability.

Pre-laboratory Assignment: Calorimetry and Hess's Law

1. Show your work for all calculations. When 1.104 grams of iron metal are mixed with 26.023 grams of hydrochloric acid in a coffee cup calorimeter, the temperature rises from 25.2 °C to a maximum of 33.5 °C. The reaction that occurs is given below.

$$2\,\mathrm{Fe(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{H_2(g)} \tag{12.18.12}$$

- Determine the amount of heat (in J) absorbed by the reaction mixture. Assume that the specific heat capacity of the mixture is the same as the specific heat capacity of water.
- How much heat (in J) was released by the reaction that occurred?
- Is this reaction exothermic or endothermic? Is ΔH reaction positive or negative?
- Under constant pressure conditions (as used in this experiment), the heat released by the reaction equals the reaction enthalpy, $q_{\text{released}} = \Delta H_{\text{reaction}}$. Determine $\Delta H_{\text{reaction}}$ in Joules per gram of metal used (J/g).
- Determine $\Delta H_{\text{reaction}}$ in kilojoules per mole of metal used (kJ/mol).
- Determine $\Delta H_{\text{reaction}}$ in kilojoules for the balanced reaction equation provided (kJ).

2. Consider the following three reactions:

$$2\,\mathrm{Fe(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{H_2(g)} \qquad \Delta H_A \tag{12.18.13}$$

$$\mathrm{Fe_2O_3(s)} + 6\,\mathrm{HCl(aq)} \longrightarrow 2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{H_2O(l)} \qquad \Delta H_B \tag{12.18.14}$$

$$2\,\mathrm{H_2(g)} + \mathrm{O_2(g)} \longrightarrow 2\,\mathrm{H_2O(l)} \qquad \Delta H_C \tag{12.18.15}$$

Show how these equations must be summed together according to Hess's Law to determine ΔH for the combustion of iron (target equation shown below). Also show clearly how the ΔH values of each of the three reactions must be manipulated to determine the enthalpy of combustion of iron.

$$4\,\mathrm{Fe(s)} + 3\,\mathrm{O_2(g)} \longrightarrow 2\,\mathrm{Fe_2O_3(s)} \qquad \Delta H = ? \tag{12.18.16}$$

3. Using tabulated ΔH_f^o values in the text, determine the enthalpy change (in kJ) that occurs during the formation of water from its elements:

$$2\,\mathrm{H_2(g)} + \mathrm{O_2(g)} \longrightarrow 2\,\mathrm{H_2O(l)} \qquad \Delta H = ? \tag{12.18.17}$$

Note that this value (and the equation) will be used in your data analysis for this lab.

Lab Report: Calorimeter and Hess's Law

Metal + HCl Reaction

Experimental Data

Assigned Metal:

	Trial 1	Trial 2
Mass of dry, empty calorimeter		
Mass of calorimeter plus HCl		
Mass of HCl used		
Mass of dry, empty beaker		
Mass of beaker plus metal		
Mass of metal used		
Initial (equilibrium) temperature of HCl		
Final (maximum) temperature of mixture		

Data Analysis

1. Write the balanced equation for the reaction between your assigned metal and HCl. All balancing coefficients should be whole numbers.
2. Complete the table below with the results of your calculations.

	Trial 1	Trial 2
Total mass of mixture, m		
Temperature change of mixture, ΔT		
SHC of mixture, c (use SHC of water)		
Heat absorbed by mixture, in J		
Heat released by reaction, in J		
ΔH_{rxn} in J/g of metal used		
ΔH_{rxn} in kJ/mol of metal used		
ΔH_{rxn} in kJ for the rxn as balanced in 12.18.1		
Average ΔH_{rxn} in kJ		

3. Show your work for the following calculations using your Trial 1 data only:

- Heat absorbed by mixture, in J
- Heat released by reaction, in J
- ΔH_{rxn} in J/g of metal used
- ΔH_{rxn} in kJ/mol of metal used
- ΔH_{rxn} in kJ for reaction as balanced in 12.18.1

4. Is this reaction exothermic or endothermic? What is your experimental evidence supporting this? Is ΔH_{rxn} positive or negative?

Metal Oxide + HCl Reaction

Experimental Data

Assigned Metal Oxide:

	Trial 1	Trial 2
Mass of dry, empty calorimeter		

	Trial 1	Trial 2
Mass of calorimeter plus HCl		
Mass of HCl used		
Mass of dry, empty beaker		
Mass of beaker plus metal oxide		
Mass of metal oxide used		
Initial (equilibrium) temperature of HCl		
Final (maximum) temperature of mixture		

Data Analysis

1. Write the balanced equation for the reaction between your assigned metal oxide and HCl. All balancing coefficients should be whole numbers.
2. Complete the table below with the results of your calculations.

	Trial 1	Trial 2
Total mass of mixture, m		
Temperature change of mixture, ΔT		
SHC of mixture, c (use SHC of water)		
Heat absorbed by mixture, in J		
Heat released by reaction, in J		
ΔH_{rxn} in J/g of metal used		
ΔH_{rxn} in kJ/mol of metal used		
ΔH_{rxn} in kJ for the rxn as balanced in 12.18.1		
Average ΔH_{rxn} in kJ		

Note: You are not required to show your work for the calculations you performed to complete the above table.

Enthalpy of Combustion of a Metal

1. Write the balanced equation for the combustion of your assigned metal. All balancing coefficients should be whole numbers.
2. Using Hess's Law, determine the enthalpy of combustion of your assigned metal as balanced above. To do this, you will need the balanced thermochemical equations for the two reactions studied in this lab, plus the balanced thermochemical equation for the formation of water from its elements. Clearly show how these three equations (and their reaction enthalpies) must be combined to give the target combustion reaction.
3. Calculate the theoretical value for the enthalpy of combustion of your assigned metal, using the tabulated ΔH_f^o values from the text.
4. Determine the percent error in your experimentally determined value for the enthalpy of combustion.

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Index

A

accuracy
1.6: Measurement Uncertainty, Accuracy, and Precision
atom
1.3: Phases and Classification of Matter
Atomic Emission Spectroscopy (AES)
12.6: Lab 6 - Atomic Emission

C

calorimetry
12.18: Lab 18 - Calorimetry and Hess's Law
Celsius
1.5: Measurements
chemical change
1.4: Physical and Chemical Properties
chemical property
1.4: Physical and Chemical Properties
Chemistry
1.2: Chemistry in Context
compound
1.3: Phases and Classification of Matter
conversion factors
1.7: Mathematical Treatment of Measurement Results
cubic centimeter
1.5: Measurements
cubic meter
1.5: Measurements

D

density
1.5: Measurements
12.3: Lab 3 - The Densities of Solutions and Solids
dimensional analysis
1.7: Mathematical Treatment of Measurement Results
Dumas Bulb
12.16: Lab 16 - The Dumas Method

E

electromagnetic radiation
12.5: Lab 5 - Flame Tests of Metal Cations
element
1.3: Phases and Classification of Matter
equivalence point
12.10: Lab 10 - Titration of Vinegar
exact number
1.6: Measurement Uncertainty, Accuracy, and Precision
extensive property
1.4: Physical and Chemical Properties

F

Fahrenheit
1.7: Mathematical Treatment of Measurement Results
flame tests
12.5: Lab 5 - Flame Tests of Metal Cations

G

gas
1.3: Phases and Classification of Matter

H

Hess' Law
12.18: Lab 18 - Calorimetry and Hess's Law
heterogeneous mixture
1.3: Phases and Classification of Matter
homogeneous mixture
1.3: Phases and Classification of Matter
hydrate
12.14: Lab 14 - Properties of Hydrates
hypothesis
1.2: Chemistry in Context

I

intensive property
1.4: Physical and Chemical Properties
international system of units
1.5: Measurements

K

Kelvin
1.5: Measurements
kilogram
1.5: Measurements

L

law
1.2: Chemistry in Context
law of conservation of matter
1.3: Phases and Classification of Matter
length
1.5: Measurements
liquid
1.3: Phases and Classification of Matter
liter
1.5: Measurements

M

macroscopic domain
1.2: Chemistry in Context
mass
1.3: Phases and Classification of Matter
Meter
1.5: Measurements
microscopic domain
1.2: Chemistry in Context
milliliter
1.5: Measurements
mixture
1.3: Phases and Classification of Matter
molecule
1.3: Phases and Classification of Matter

N

nomenclature
12.7: Lab 7 - Chemical Nomenclature Dry Lab

O

oxygen
12.15: Lab 15 - The Properties of Oxygen Gas

P

physical change
1.4: Physical and Chemical Properties
physical property
1.4: Physical and Chemical Properties
plasma
1.3: Phases and Classification of Matter
polarity
12.9: Lab 9 - VSEPR Theory and Shapes of Molecules Dry Lab
polyatomic ions
12.7: Lab 7 - Chemical Nomenclature Dry Lab
precision
1.6: Measurement Uncertainty, Accuracy, and Precision
pure substance
1.3: Phases and Classification of Matter

Q

qualitative analysis
12.12: Lab 12 - Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions

R

rounding
1.6: Measurement Uncertainty, Accuracy, and Precision

S

scientific method
1.2: Chemistry in Context
second
1.5: Measurements
SI Units
1.5: Measurements
significant figures
1.6: Measurement Uncertainty, Accuracy, and Precision
solid
1.3: Phases and Classification of Matter
symbolic domain
1.2: Chemistry in Context

T

theory
1.2: Chemistry in Context
titration
12.10: Lab 10 - Titration of Vinegar

U

uncertainty
1.6: Measurement Uncertainty, Accuracy, and Precision
unit
1.5: Measurements

V

vinegar
12.10: Lab 10 - Titration of Vinegar
volume
1.5: Measurements
VSEPR
12.9: Lab 9 - VSEPR Theory and Shapes of Molecules Dry Lab

W

weight

1.3: Phases and Classification of Matter

Y

yield

12.13: Lab 13 - Mole Ratios and Reaction Stoichiometry

Glossary

Sample Word 1 | Sample Definition 1

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    - 1.1: Introduction - *CC BY-SA 4.0*
    - 1.2: Chemistry in Context - *CC BY-SA 4.0*
    - 1.3: Phases and Classification of Matter - *CC BY-SA 4.0*
    - 1.4: Physical and Chemical Properties - *CC BY-SA 4.0*
    - 1.5: Measurements - *CC BY-SA 4.0*
    - 1.6: Measurement Uncertainty, Accuracy, and Precision - *CC BY-SA 4.0*
    - 1.7: Mathematical Treatment of Measurement Results - *CC BY-SA 4.0*
    - 1.8: Key Terms - *Undeclared*
    - 1.9: Key Equations - *Undeclared*
  - 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*
  - 3: Electronic Structure and Periodic Properties - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Energy - *CC BY 4.0*
    - 3.3: The Bohr Model - *CC BY 4.0*
    - 3.4: Development of Quantum Theory - *CC BY 4.0*
    - 3.5: Electronic Structure of Atoms (Electron Configurations) - *CC BY 4.0*
    - 3.6: Periodic Variations in Element Properties - *CC BY 4.0*
    - 3.7: The Periodic Table - *Undeclared*
    - 3.8: Ionic and Molecular Compounds - *Undeclared*
    - 3.9: Key Terms - *Undeclared*
    - 3.10: Key Equations - *Undeclared*
    - 3.11: Summary - *Undeclared*
  - 4: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 4.1: Introduction - *CC BY 4.0*
    - 4.2: Ionic Bonding - *CC BY 4.0*
    - 4.3: Covalent Bonding - *CC BY 4.0*
    - 4.4: Chemical Nomenclature - *CC BY 4.0*
    - 4.5: Lewis Symbols and Structures - *CC BY 4.0*
    - 4.6: Formal Charges and Resonance - *CC BY 4.0*
    - 4.7: Molecular Structure and Polarity - *CC BY 4.0*
    - 4.8: Key Terms - *Undeclared*
    - 4.9: Key Equations - *Undeclared*
    - 4.10: Summary - *Undeclared*
  - 5: Advanced Theories of Covalent Bonding - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Valence Bond Theory - *CC BY 4.0*
    - 5.3: Hybrid Atomic Orbitals - *CC BY 4.0*
    - 5.4: Multiple Bonds - *CC BY 4.0*
    - 5.5: Molecular Orbital Theory - *CC BY 4.0*
    - 5.6: Key Terms - *Undeclared*
    - 5.7: Key Equations - *Undeclared*
    - 5.8: Summary - *Undeclared*
  - 6: Composition of Substances and Solutions - *CC BY 4.0*
    - 6.1: Introduction - *Undeclared*
    - 6.2: Formula Mass - *CC BY 4.0*

- 6.3: Determining Empirical and Molecular Formulas - *CC BY 4.0*
- 6.4: Molarity - *CC BY 4.0*
- 6.5: Other Units for Solution Concentrations - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*

7: Stoichiometry of Chemical Reactions - *CC BY 4.0*
- 7.1: Introduction - *CC BY 4.0*
- 7.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 7.3: Classifying Chemical Reactions - *CC BY 4.0*
- 7.4: Reaction Stoichiometry - *CC BY 4.0*
- 7.5: Reaction Yields - *CC BY 4.0*
- 7.6: Quantitative Chemical Analysis - *CC BY 4.0*
- 7.7: Key Terms - *Undeclared*
- 7.8: Key Equations - *Undeclared*
- 7.9: Summary - *Undeclared*

8: Gases - *CC BY 4.0*
- 8.1: Introduction - *Undeclared*
- 8.2: Gas Pressure - *CC BY 4.0*
- 8.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 8.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 8.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 8.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 8.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 8.8: Key Terms - *Undeclared*
- 8.9: Key Equations - *Undeclared*
- 8.10: Summary - *Undeclared*

9: Thermochemistry - *CC BY 4.0*
- 9.1: Introduction - *CC BY 4.0*
- 9.2: Energy Basics - *CC BY 4.0*
- 9.3: Calorimetry - *CC BY 4.0*
- 9.4: Enthalpy - *CC BY 4.0*
- 9.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 9.6: Key Terms - *Undeclared*
- 9.7: Key Equations - *Undeclared*
- 9.8: Summary - *Undeclared*

10: Liquids and Solids - *Undeclared*
- 10.1: Introduction - *CC BY 4.0*
- 10.2: Intermolecular Forces - *CC BY 4.0*
- 10.3: Properties of Liquids - *CC BY 4.0*
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- 10.9: Key Equations - *Undeclared*
- 10.10: Summary - *Undeclared*

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- 11.6: Colloids - *CC BY 4.0*
- 11.7: Key Terms - *Undeclared*
- 11.8: Key Equations - *Undeclared*
- 11.9: Summary - *Undeclared*

12: Lab Manual - *Undeclared*
- 12.1: Lab 1 - Using Excel for Graphical Analysis of Data - *Undeclared*
- 12.2: Lab 2 - Measurements in the Laboratory - *CC BY-NC 4.0*
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- 12.10: Lab 10 - Titration of Vinegar - *CC BY-NC 4.0*
- 12.11: Lab 11 - Types of Chemical Reactions - *CC BY-NC 4.0*
- 12.12: Lab 12 - Using Periodic Properties to Identify Group 2A Cations and Group 7A Anions - *CC BY-NC 4.0*
- 12.13: Lab 13 - Mole Ratios and Reaction Stoichiometry - *CC BY-NC 4.0*
- 12.14: Lab 14 - Properties of Hydrates - *CC BY-NC 4.0*
- 12.15: Lab 15 - The Properties of Oxygen Gas - *CC BY-NC 4.0*
- 12.16: Lab 16 - The Dumas Method - *CC BY-NC 4.0*
- 12.17: Lab 17 - The Molecular Weight of Carbon Dioxide - *CC BY-NC 4.0*
- 12.18: Lab 18 - Calorimetry and Hess's Law - *CC BY-NC 4.0*

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    - 1.1: Introduction - *CC BY-SA 4.0*
    - 1.2: Chemistry in Context - *CC BY-SA 4.0*
    - 1.3: Phases and Classification of Matter - *CC BY-SA 4.0*
    - 1.4: Physical and Chemical Properties - *CC BY-SA 4.0*
    - 1.5: Measurements - *CC BY-SA 4.0*
    - 1.6: Measurement Uncertainty, Accuracy, and Precision - *CC BY-SA 4.0*
    - 1.7: Mathematical Treatment of Measurement Results - *CC BY-SA 4.0*
    - 1.8: Key Terms - *Undeclared*
    - 1.9: Key Equations - *Undeclared*
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    - 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*
  - 3: Electronic Structure and Periodic Properties - *CC BY 4.0*
    - 3.1: Introduction - *Undeclared*
    - 3.2: Electromagnetic Energy - *CC BY 4.0*
    - 3.3: The Bohr Model - *CC BY 4.0*
    - 3.4: Development of Quantum Theory - *CC BY 4.0*
    - 3.5: Electronic Structure of Atoms (Electron Configurations) - *CC BY 4.0*
    - 3.6: Periodic Variations in Element Properties - *CC BY 4.0*
    - 3.7: The Periodic Table - *Undeclared*
    - 3.8: Ionic and Molecular Compounds - *Undeclared*
    - 3.9: Key Terms - *Undeclared*
    - 3.10: Key Equations - *Undeclared*
    - 3.11: Summary - *Undeclared*
  - 4: Chemical Bonding and Molecular Geometry - *CC BY 4.0*
    - 4.1: Introduction - *CC BY 4.0*
    - 4.2: Ionic Bonding - *CC BY 4.0*
    - 4.3: Covalent Bonding - *CC BY 4.0*
    - 4.4: Chemical Nomenclature - *CC BY 4.0*
    - 4.5: Lewis Symbols and Structures - *CC BY 4.0*
    - 4.6: Formal Charges and Resonance - *CC BY 4.0*
    - 4.7: Molecular Structure and Polarity - *CC BY 4.0*
    - 4.8: Key Terms - *Undeclared*
    - 4.9: Key Equations - *Undeclared*
    - 4.10: Summary - *Undeclared*
  - 5: Advanced Theories of Covalent Bonding - *CC BY 4.0*
    - 5.1: Introduction - *CC BY 4.0*
    - 5.2: Valence Bond Theory - *CC BY 4.0*
    - 5.3: Hybrid Atomic Orbitals - *CC BY 4.0*
    - 5.4: Multiple Bonds - *CC BY 4.0*
    - 5.5: Molecular Orbital Theory - *CC BY 4.0*
    - 5.6: Key Terms - *Undeclared*
    - 5.7: Key Equations - *Undeclared*
    - 5.8: Summary - *Undeclared*
  - 6: Composition of Substances and Solutions - *CC BY 4.0*
    - 6.1: Introduction - *Undeclared*
    - 6.2: Formula Mass - *CC BY 4.0*

- 6.3: Determining Empirical and Molecular Formulas - *CC BY 4.0*
- 6.4: Molarity - *CC BY 4.0*
- 6.5: Other Units for Solution Concentrations - *CC BY 4.0*
- 6.6: Key Terms - *Undeclared*
- 6.7: Key Equations - *Undeclared*
- 6.8: Summary - *Undeclared*

7: Stoichiometry of Chemical Reactions - *CC BY 4.0*

- 7.1: Introduction - *CC BY 4.0*
- 7.2: Writing and Balancing Chemical Equations - *CC BY 4.0*
- 7.3: Classifying Chemical Reactions - *CC BY 4.0*
- 7.4: Reaction Stoichiometry - *CC BY 4.0*
- 7.5: Reaction Yields - *CC BY 4.0*
- 7.6: Quantitative Chemical Analysis - *CC BY 4.0*
- 7.7: Key Terms - *Undeclared*
- 7.8: Key Equations - *Undeclared*
- 7.9: Summary - *Undeclared*

8: Gases - *CC BY 4.0*

- 8.1: Introduction - *Undeclared*
- 8.2: Gas Pressure - *CC BY 4.0*
- 8.3: Relating Pressure, Volume, Amount, and Temperature- The Ideal Gas Law - *CC BY 4.0*
- 8.4: Stoichiometry of Gaseous Substances, Mixtures, and Reactions - *CC BY 4.0*
- 8.5: Effusion and Diffusion of Gases - *CC BY 4.0*
- 8.6: The Kinetic-Molecular Theory - *CC BY 4.0*
- 8.7: Non-Ideal Gas Behavior - *CC BY 4.0*
- 8.8: Key Terms - *Undeclared*
- 8.9: Key Equations - *Undeclared*
- 8.10: Summary - *Undeclared*

9: Thermochemistry - *CC BY 4.0*

- 9.1: Introduction - *CC BY 4.0*
- 9.2: Energy Basics - *CC BY 4.0*
- 9.3: Calorimetry - *CC BY 4.0*
- 9.4: Enthalpy - *CC BY 4.0*
- 9.5: Strengths of Ionic and Covalent Bonds - *Undeclared*
- 9.6: Key Terms - *Undeclared*
- 9.7: Key Equations - *Undeclared*
- 9.8: Summary - *Undeclared*

10: Liquids and Solids - *Undeclared*

- 10.1: Introduction - *CC BY 4.0*
- 10.2: Intermolecular Forces - *CC BY 4.0*
- 10.3: Properties of Liquids - *CC BY 4.0*
- 10.4: Phase Transitions - *CC BY 4.0*
- 10.5: Phase Diagrams - *CC BY 4.0*
- 10.6: The Solid State of Matter - *CC BY 4.0*
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- 10.8: Key Terms - *Undeclared*
- 10.9: Key Equations - *Undeclared*
- 10.10: Summary - *Undeclared*

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- 11.1: Introduction - *CC BY 4.0*
- 11.2: The Dissolution Process - *CC BY 4.0*
- 11.3: Electrolytes - *CC BY 4.0*
- 11.4: Solubility - *CC BY 4.0*
- 11.5: Colligative Properties - *CC BY 4.0*
- 11.6: Colloids - *CC BY 4.0*
- 11.7: Key Terms - *Undeclared*
- 11.8: Key Equations - *Undeclared*
- 11.9: Summary - *Undeclared*

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- 12.1: Lab 1 - Using Excel for Graphical Analysis of Data - *Undeclared*
- 12.2: Lab 2 - Measurements in the Laboratory - *CC BY-NC 4.0*
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