INTRODUCTORY CHEMISTRY



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Chemistry 108, Introductory Chemistry, San Ramon Campus of Diablo Valley College

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TABLE OF CONTENTS

Licensing

1: What is Chemistry?

1.1: Scope of Chemistry

1.2: What is Matter?

1.3: Classifying Matter According to Its Composition

1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas

1.5: Differences in Matter- Physical and Chemical Properties

1.6: Changes in Matter - Physical and Chemical Changes

1.7: The Scientific Method - How Chemists Think

1.8: A Beginning Chemist - How to Succeed

2: The Numerical Side of Chemistry

2.1: Taking Measurements

2.2: Significant Figures - Writing Numbers to Reflect Precision

2.3: Scientific Notation - Writing Large and Small Numbers

2.4: Significant Figures in Calculations

2.5: The Basic Units of Measurement

2.6: Temperature - Random Motion of Molecules and Atoms

2.7: Density

2.8: Problem Solving and Unit Conversions

2.9: Metric Unit Conversions

2.10: Solving Multi-step Conversion Problems

2.11: Units Raised to a Power

2.12: Derived Units

3: The Evolution of Atomic Theory

3.1: Cutting Aluminum until you get Atoms

3.2: Indivisible - The Atomic Theory

3.3: Law of Conservation of Mass

3.4: Law of Definite Proportions

3.5: Law of Definite Proportions - Calculating Mass Percent

3.6: The Nuclear Atom

3.7: The Properties of Protons, Neutrons, and Electrons

3.8: Elements- Defined by Their Number of Protons

3.9: Isotopes - When the Number of Neutrons Varies

3.10: Atomic Mass - The Average Mass of an Element's Atoms

3.11: Looking for Patterns - The Periodic Table

3.12: Ions - Losing and Gaining Electrons

3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character

4: The Modern Model of the Atom

4.1: Electromagnetic Spectrum

4.2: Wavelength and Frequency Calculations

4.3: Energy and Frequency Calculations

4.4: The Bohr Model - Atoms with Orbits

4.5: The Quantum-Mechanical Model- Atoms with Orbitals

4.6: Electron Configurations

4.7: Electron Configuration Shortcut - Noble Gas Configurations

4.8: Valence Electrons

4.9: Electron Configurations of Cations

4.10: Electron Configurations of Transition Metal Cations

4.11: Electron Configurations of Anions

4.12: Octet Rule

4.13: Ionic Bonding

5: Chemical Bonding and Nomenclature

5.1: A Molecular View of Elements and Compounds

5.2: Representing Valence Electrons with Dots

5.3: Lewis Structures of Ionic Compounds- Electrons Transferred

5.4: Lewis Structures of Covalent Compounds- Electrons Shared

5.5: Rules for Drawing Lewis Structures

5.6: Resonance - Equivalent Lewis Structures for the Same Molecule

5.7: Bond Polarity

5.8: Chemical Nomenclature

5.9: Writing Formulas for Ionic Compounds

5.10: Naming Ionic Compounds

5.11: Naming Covalent Compounds

5.12: Naming Acids

5.13: Nomenclature Summary

6: The Shape of Molecules

6.1: VSEPR Theory

6.2: Molecule Polarity

7: Intermolecular Forces

7.1: Interactions between Molecules

7.2: Properties of Liquids and Solids

7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"

7.4: Types of Crystalline Solids

8: Chemical Reactions

8.1: Evidence of a Chemical Reaction

8.2: The Chemical Equation

8.3: How to Write Balanced Chemical Equations
8.4: Classifying Chemical Reactions
8.5: Aqueous Solutions and Solubility
8.6: Precipitation Reactions
8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations
8.8: Acid–Base Reactions

9: Stoichiometry and the Mole

9.1: Avogadro's Number
9.2: Conversions Between Moles and Particles
9.3: Molar Mass
9.4: Conversions Between Moles and Mass
9.5: Conversions Between Mass and Number of Particles
9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"
9.7: Stoichiometry
9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!
9.9: Mole-Mass and Mass-Mass Calculations
9.10: Yields
9.11: Limiting Reagents
9.12: Mass Percent Composition from a Chemical Formula
9.13: Calculating Empirical Formulas
9.14: Calculating Molecular Formulas

10: Gases

10.1: Kinetic Molecular Theory- A Model for Gases
10.2: What is Gas Pressure?
10.3: Boyle’s Law - Pressure and Volume
10.4: Charles’s Law- Volume and Temperature
10.5: Gay-Lussac's Law- Temperature and Pressure
10.6: Avogadro’s Law- Volume and Moles
10.7: The Combined Gas Law
10.8: The Ideal Gas Law
10.9: Molar Volume of a Gas at STP
10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!
10.11: Gas Stoichiometry
10.12: Dalton's Law of Partial Pressures

11: Solutions

11.1: Solutions - Homogeneous Mixtures
11.2: Solutions of Solids Dissolved in Water
11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
11.4: Solution Dilution (M1V1 = M2V2)
11.5: Calculating Solution Concentration- Mass Percent

11.6: Solution Stoichiometry

12: Chemical Equilibrium

12.1: The Idea of Dynamic Chemical Equilibrium

12.2: The Equilibrium Constant, Keq

12.3: Heterogeneous Equilibria

12.4: Calculations with Equilibrium Constants

12.5: Le Chatelier's Principle

13: Electrolytes, Acids and Bases

13.1: Electrolytes

13.2: Acids- Properties and Examples

13.3: Bases- Properties and Examples

13.4: Strong and Weak Acids and Bases

13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory

13.6: Autodissociation of Water and Kw

13.7: The pH and pOH Scales

13.8: Acid-Base Reactions

13.9: Acid–Base Titrations

13.10: Buffers- Solutions that Resist pH Change

Index

Glossary

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1: WHAT IS CHEMISTRY?

1.1: Scope of Chemistry

1.2: What is Matter?

1.3: Classifying Matter According to Its Composition

1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas

1.5: Differences in Matter- Physical and Chemical Properties

1.6: Changes in Matter - Physical and Chemical Changes

1.7: The Scientific Method - How Chemists Think

1.8: A Beginning Chemist - How to Succeed

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1.1: Scope of Chemistry



Figure 1.1.1 (Public Domain; User:Phanton/Wikipedia via https://commons.wikimedia.org/wiki/F...tive_Soaps.jpg)

How does soap relate to chemistry?

Chemistry affects every area of our lives. Here is just one example of chemistry in action – the making of soaps. Soap was once made by boiling animal fat in ashes – the product was hard on the skin and not very pleasant to use. Today, soap manufacture involves complicated chemical processes to provide a wide variety of soaps for different skin types. Colors and odors can be custom-made for that individual experience.

"What is this made of? How can we produce this material quickly and at a low cost? Will this product harm us or help us?"—these are all questions that can be answered using the science of chemistry.

What is Chemistry?

Science is a general term used to describe the principled, rigorous study of the natural world. Many interconnected disciplines fall under this broader concept. For example, physics is the study of motion and forces. Biology is the study of living things. Geology is the study of the Earth and the rocks and minerals of which it is comprised. Chemistry is the study of the composition of matter and the changes that matter undergoes. Matter is anything that has mass and takes up space. Virtually everything around us is matter, including both living and nonliving things. Chemistry affects nearly everything we see and every action we take. Chemistry explains why milk that is left in the refrigerator for too long turns sour. Chemistry explains why certain pollutants called chlorofluorocarbons have done lasting damage to the ozone layer of our planet. Chemistry explains why the leaves of deciduous trees turn from green in the summer to various shades of red and yellow in the autumn (Figure below).

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Figure 1.1.2: Chemical reactions in the leaves of deciduous trees cause them to change color from green to red, orange, or yellow before dropping to the ground. (Public Domain; User:Jongleur100/Wikimedia Commons via https://commons.wikimedia.org/wiki/F...untry_lane.jpg)

Chemistry touches every area of our lives. The medicines we take, the food we eat, the clothes we wear—all of these materials and more are, in some way or another, products of chemistry.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Chemists look at the world in two ways, often simultaneously. The two worlds of the chemist are the **macroscopic** world and the **microscopic** world. Macroscopic refers to substances and objects that can be seen, touched, and measured directly. Microscopic refers to the small particles that make up all matter. Chemists must observe matter and do experiments macroscopically; then make generalizations and propose explanations that are microscopic in nature. For example, anyone can observe the physical change in appearance that occurs as an iron object, such as a tractor, is left out in the elements and gradually turns to rust. However, a chemist looking at the rusting tractor considers the individual atoms that make up the iron, and how they are changing as a result of exposure to oxygen in the air, and water from rain. Throughout the study of chemistry, there is often a switch back and forth between the macroscopic and microscopic worlds.

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Figure 1.1.3: Rusting artillery shells. (the US Marine Corps, Source: http://commons.wikimedia.org/wiki/File:USMC-060507-M-7799R-003.jpg; License: Public Domain)

Summary

- Chemistry is the study of matter and the changes it undergoes.
- Chemistry considers both macroscopic and microscopic information.

Review

1. Give two examples of chemistry in your everyday life.
2. What is the macroscopic world?
3. What is the microscopic world?

Explore More

1. Read the label on a prepared food product (for example: bread, cereal, dessert). List all of the ingredients in the product. Look up each ingredient on the Internet and write down what that material is doing in the food product.
2. Select your favorite hobby or activity. List all of the items you use in that activity or hobby. For each item, find out how chemistry has contributed to the creation or better operation of that item.

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1.2: What is Matter?

Learning Objectives

- Define matter and explain how it is composed of building blocks known as "atoms".

We are all familiar with matter. The definition of **Matter** is anything that has mass and volume (takes up space). For most common objects that we deal with every day, it is fairly simple to demonstrate that they have mass and take up space. You might be able to imagine, however, the difficulty for people several hundred years ago to demonstrate that air had mass and volume. Air (and all other gases) are invisible to the eye, have very small masses compared to equal amounts of solids and liquids, and are quite easy to compress (change volume). Without sensitive equipment, it would have been difficult to convince people that gases are matter. Today, we can measure the mass of a small balloon when it is deflated and then blow it up, tie it off, and measure its mass again to detect the additional mass due to the air inside. The mass of air, under room conditions, that occupies a one quart jar is approximately 0.0002 pounds. This small amount of mass would have been difficult to measure in times before balances were designed to accurately measure very small masses. Later, scientists were able to compress gases into such a small volume that the gases turned into liquids, which made it clear that gases are matter.

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Figure 1.2.1: Everything from an ant, to a truck, to the earth, and even the entire galaxy is composed of matter. Images used with permission from Wikipedia (CC_SA-BY-3.0; credit High Contrast).

Even though the universe consists of "things" as wildly different as ants and galaxies, the matter that makes up all of these "things" is composed of a very limited number of building blocks. These building blocks are known as **atoms**, and so far, scientists have discovered or created a grand total of 118 different types of atoms. Scientists have given a name to each different type of atom. A substance that is composed of only one type of atom is called an element. At this point, what should amaze you is that all forms of matter in our universe are made with only 118 different building blocks. In some ways, it's sort of like cooking a gourmet, five-course meal using only three ingredients! How is it possible? To answer that question, you have to understand the ways in

which different elements are put together to form matter.

The most important method that nature uses to organize atoms into matter is the formation of **molecules**. Molecules are groups of two or more atoms that have been bonded together. There are millions of different ways to bond atoms together, which means that there are millions of different possible molecules. Each of these molecules has its own set of chemical properties, and it's these properties with which chemists are most concerned. You will learn a lot more about atoms and molecules, including how they were discovered, in a later part of the textbook.

Summary

All matter has mass and occupies space. All physical objects are made of matter. Matter itself is composed of tiny building blocks known as "atoms". There are only 118 different types of atoms known to man. Frequently, atoms are bonded together to form "molecules".

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1.3: Classifying Matter According to Its Composition

Learning Objectives

- Explain the difference between a pure substance and a mixture.
- Explain the difference between an element and a compound.
- Explain the difference between a homogeneous mixture and a heterogeneous mixture.

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 1.3.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

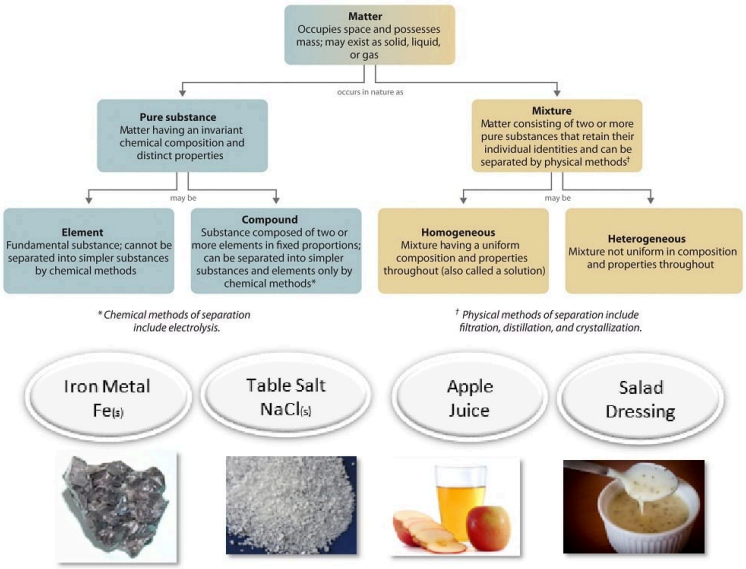


Figure 1.3.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

A **homogeneous** mixture is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. The amount of salt in the salt water can vary from one sample to another. All solutions are considered homogeneous because the dissolved material is present in the same amount throughout the solution.

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components

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of the soup.

Phase

A phase is any part of a sample that has a uniform composition and properties. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly, but instead form two separate layers. Each of the layers is called a phase.

Example 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. filtered tea
b. freshly squeezed orange juice
c. a compact disc
d. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
e. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
   B) Because the composition of the solution is uniform throughout, it is ***a homogeneous mixture***.
b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
   B) Because its composition is not uniform throughout, orange juice is ***a heterogeneous mixture***.
c. **A)** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence, a compact disc is not chemically pure.
   B) The regions of different composition indicate that a compact disc is ***a heterogeneous mixture.***
d. **A)** Aluminum oxide is a single, chemically ***pure compound***.
e. **A)** Selenium is one of the known ***elements***.

Exercise 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. white wine
b. mercury
c. ranch-style salad dressing
d. table sugar (sucrose)

Answer a: available at the end of this section.

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Example 1.3.2

How would a chemist categorize each example of matter?

a. saltwater
b. soil
c. water
d. oxygen

Solution

a. Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a homogeneous mixture, or a solution.
b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
c. Water is a substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a substance, is an element.

Exercise 1.3.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a: available at the end of this section.

Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

- Element: a substance that is made up of only one type of atom.
- Compound:a substance that is made up of more than one type of atom bonded together.
- Mixture: a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties.

Contributions & Attributions

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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

Answers

Exercise 1.3.1

homogeneous mixture (solution)
element
heterogeneous mixture

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compound

Exercise 1.3.2

a homogeneous mixture (solution), assuming it is filtered coffee

element

heterogeneous mixture

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1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas

Learning Objectives

- To describe the solid, liquid and gas phases.

Water can take many forms. At low temperatures (below $0^{\circ}\mathrm{C}$), it is a solid. When at "normal" temperatures (between $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$), it is a liquid. While at temperatures above $100^{\circ}\mathrm{C}$, water is a gas (steam). The state that water is in depends upon the temperature. Each state has its own unique set of physical properties. Matter typically exists in one of three states: **solid**, **liquid**, or **gas**.

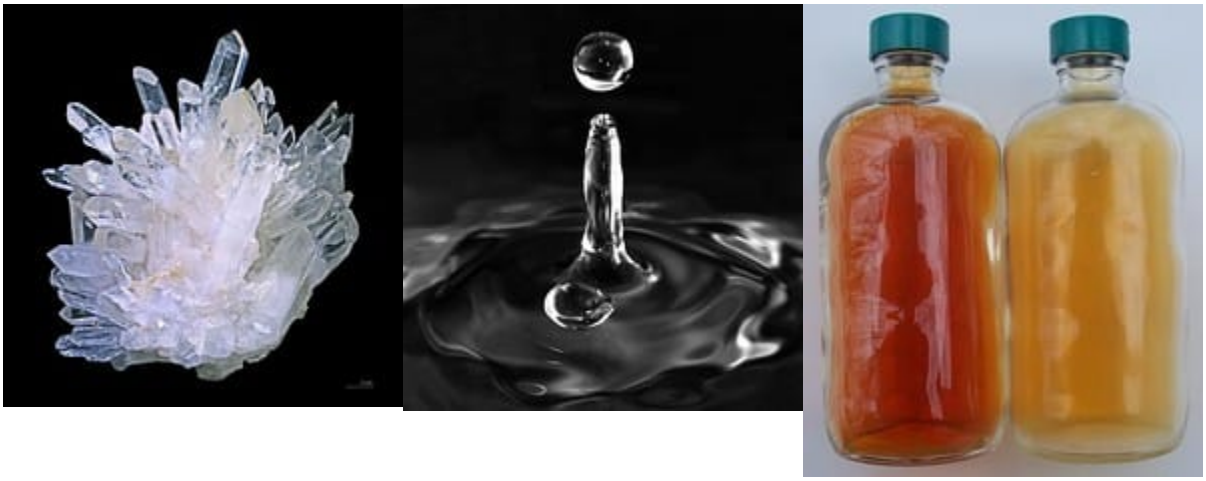


Figure 1.4.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure 1.4.2 shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

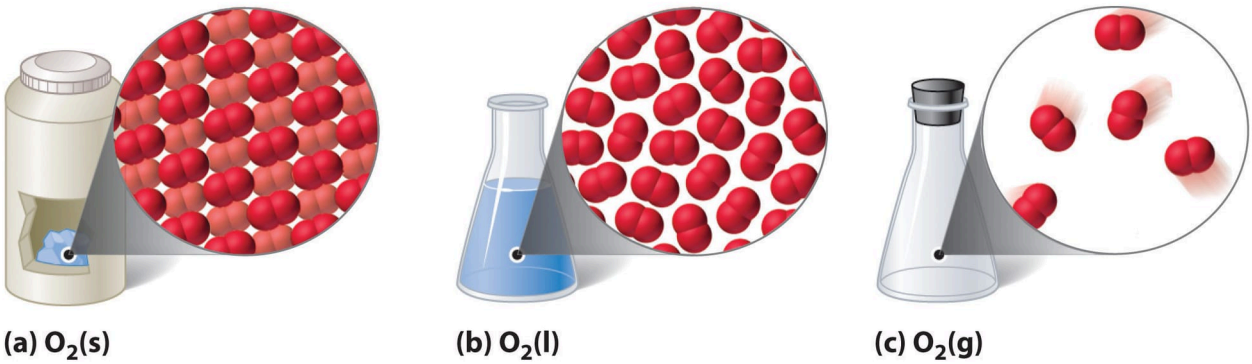
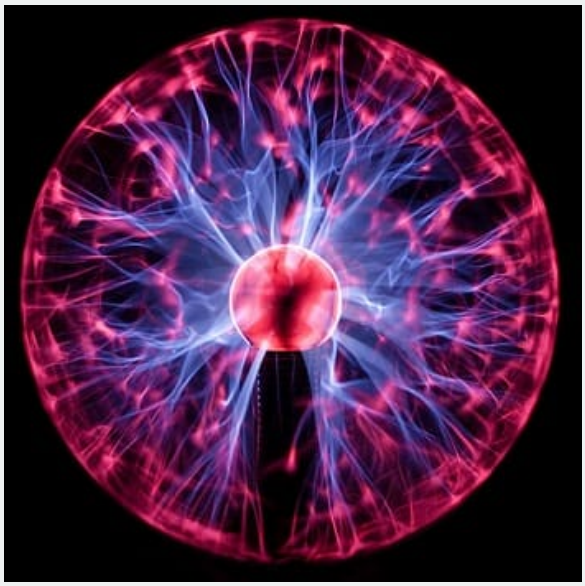


Figure 1.4.2: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

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Plasma: A Fourth State of Matter

Technically speaking, a fourth state of matter called plasma exists, but it does not naturally occur on earth, so we will omit it from our study here.



A plasma globe operating in a darkened room. (CC BY-SA 3.0; Chocolateoak).

Solids

In the solid state, the individual particles of a substance are in fixed positions with respect to each other because there is not enough thermal energy to overcome the intermolecular interactions between the particles. As a result, solids have a definite shape and volume. Most solids are hard, but some (like waxes) are relatively soft. Many solids composed of ions can also be quite brittle.

Solids are defined by the following characteristics:

- Definite shape (rigid)
- Definite volume
- Particles vibrate around fixed axes

If we were to cool liquid mercury to its freezing point of $-39^{\circ}\mathrm{C}$, and under the right pressure conditions, we would notice all of the liquid particles would go into the solid state. Mercury can be solidified when its temperature is brought to its freezing point. However, when returned to room temperature conditions, mercury does not exist in solid state for long, and returns back to its more common liquid form.

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure $1.4.3$. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

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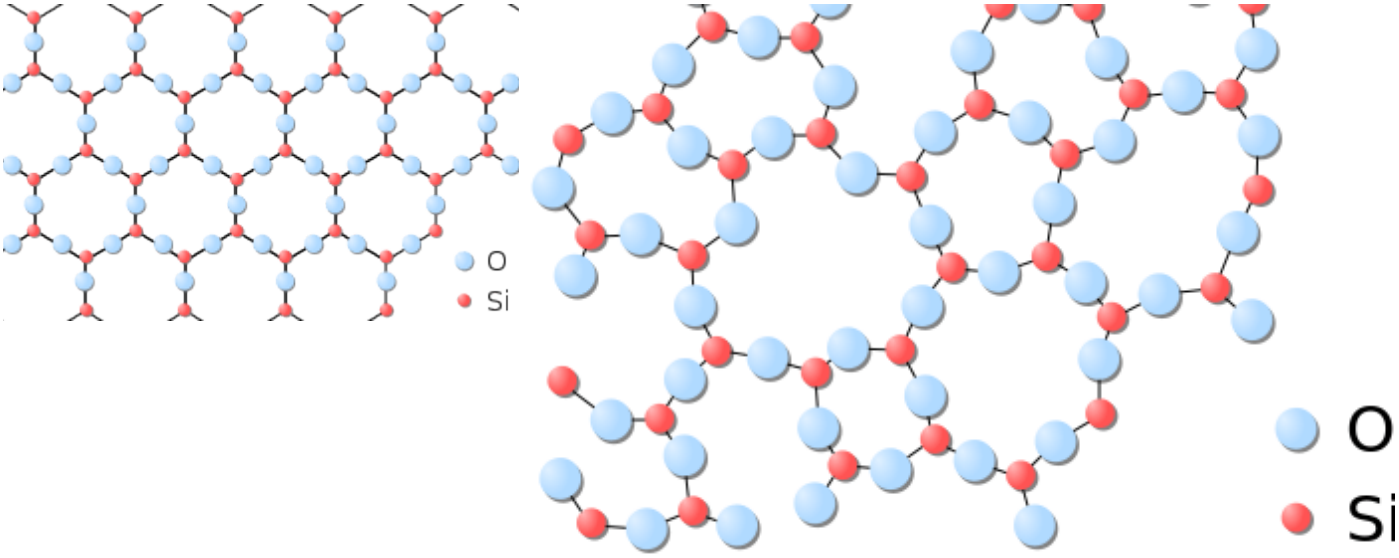


Figure 1.4.3: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

If the particles of a substance have enough energy to partially overcome intermolecular interactions, then the particles can move about each other while remaining in contact. This describes the liquid state. In a liquid, the particles are still in close contact, so liquids have a definite volume. However, because the particles can move about each other rather freely, a liquid has no definite shape and takes a shape dictated by its container.

Liquids have the following characteristics:

- No definite shape (takes the shape of its container).
- Has definite volume.
- Particles are free to move over each other, but are still attracted to each other.

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension)—a property that all liquids exhibit. Mercury has a relatively high surface tension, which makes it very unique. Here you see mercury in its common liquid form.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 1.4.1: *Mercury boiling to become a gas.*

If we heat liquid mercury to its boiling point of $357^{\circ}\mathrm{C}$ under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. This describes the gas state, which we will consider in more detail elsewhere. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either. The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Gases have the following characteristics:

- No definite shape (takes the shape of its container)
- No definite volume
- Particles move in random motion with little or no attraction to each other
- Highly compressible

Table 1.4.1: Characteristics of the Three States of Matter

Characteristics	Solids	Liquids	Gases
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

Example 1.4.1

What state or states of matter does each statement, describe?

a. This state has a definite volume, but no definite shape.
b. This state has no definite volume.
c. This state allows the individual particles to move about while remaining in contact.

Solution

a. This statement describes the liquid state.
b. This statement describes the gas state.
c. This statement describes the liquid state.

Exercise 1.4.1

What state or states of matter does each statement describe?

a. This state has individual particles in a fixed position with regard to each other.
b. This state has individual particles far apart from each other in space.
c. This state has a definite shape.

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Summary

- Three states of matter exist—solid, liquid, and gas.
- Solids have a definite shape and volume.
- Liquids have a definite volume, but take the shape of the container.
- Gases have no definite shape or volume.

Answers

Exercise 1.4.1

solid

Exercise 1.4.1

gas

Exercise 1.4.1

solid

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1.5: Differences in Matter- Physical and Chemical Properties

Learning Objectives

To separate physical from chemical properties.

All matter has physical and chemical properties. Physical properties are characteristics that scientists can measure without changing the composition of the sample under study, such as mass, color, and volume (the amount of space occupied by a sample). Chemical properties describe the characteristic ability of a substance to react to form new substances; they include its flammability and susceptibility to corrosion. All samples of a pure substance have the same chemical and physical properties. For example, pure copper is always a reddish-brown solid (a physical property) and always dissolves in dilute nitric acid to produce a blue solution and a brown gas (a chemical property).

Physical Property

A *physical property* is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets, a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C. The average densities of some common substances are in Table 1.5.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table 1.5.1: Densities of Common Substances

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Hardness helps determine how an element (especially a metal) might be used. Many elements are fairly soft (silver and gold, for example) while others (such as titanium, tungsten, and chromium) are much harder. Carbon is an interesting example of hardness. In graphite, (the "lead" found in pencils) the carbon is very soft, while the carbon in a diamond is roughly seven times as hard.

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Figure 1.5.1: Pencil (left) and Diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

Melting and boiling points are somewhat unique identifiers, especially of compounds. In addition to giving some idea as to the identity of the compound, important information can be obtained about the purity of the material.

Chemical Properties

Chemical properties of matter describe its potential to undergo some chemical change or reaction by virtue of its composition. The elements, electrons, and bonds that are present give the matter potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, after studying chemistry for some time, you should be able to look at the formula of a compound and state some chemical property. For example, hydrogen has the potential to ignite and explode given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.



Figure 1.5.2: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure 1.5.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of

chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 1.5.2: Contrasting Physical and Chemical Properties

Physical Properties	Chemical Properties
Gallium metal melts at 30 °C.	Iron metal rusts.
Mercury is a very dense liquid.	A green banana turns yellow when it ripens.
Gold is shiny.	A dry piece of paper burns.

Example 1.5.1

Which of the following is a chemical property of iron?

a. Iron corrodes in moist air.
b. Density = 7.874 g/cm^3
c. Iron is soft when pure.
d. Iron melts at 1808 K.

Solution

"Iron corrodes in moist air" is the only chemical property of iron from the list.

Exercise 1.5.1A

Which of the following is a physical property of matter?

a. corrosiveness
b. pH (acidity)
c. density
d. flammability

Answer available at the end of this section.

Exercise 1.5.1B

Which of the following is a chemical property?

a. flammability
b. melting point
c. boiling point
d. density

Answer available at the end of this section.

Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change. 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.

Answers

Exercise $1.5.1A$

c

Exercise $1.5.1B$

a

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1.6: Changes in Matter - Physical and Chemical Changes

Learning Objectives

- Label a change as chemical or physical.
- List evidence that can indicate a chemical change occurred.

Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are classified as either physical or chemical changes. Chemists learn a lot about the nature of matter by studying the changes that matter can undergo. Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.

Physical Change

Physical changes are changes in which no bonds are broken or formed. This means that the same types of compounds or elements that were there at the beginning of the change are there at the end of the change. Because the ending materials are the same as the beginning materials, the properties (such as color, boiling point, etc.) will also be the same. Physical changes involve moving molecules around, but not changing them. Some types of physical changes include:

- Changes of state (changes from a solid to a liquid or a gas and vice versa).
- Separation of a mixture.
- Physical deformation (cutting, denting, stretching).
- Making solutions (special kinds of mixtures).

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. Melting is an example of a physical change. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. When liquid water is heated, it changes to water vapor. However, even though the physical properties have changed, the molecules are exactly the same as before. We still have each water molecule containing two hydrogen atoms and one oxygen atom covalently bonded. When you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of the pennies or the nickels—you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you started tearing is still paper when you are done. Again, this is an example of a physical change.

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Figure 1.6.1: Ice melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; however, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.

Chemical Change

Chemical changes occur when bonds are broken and/or formed between molecules or atoms. This means that one substance with a certain set of properties (such as melting point, color, taste, etc) is turned into a different substance with different properties. Chemical changes are frequently harder to reverse than physical changes.

One good example of a chemical change is burning a candle. The act of burning paper actually results in the formation of new chemicals (carbon dioxide and water) from the burning of the wax. Another example of a chemical change is what occurs when natural gas is burned in your furnace. This time, on the left there is a molecule of methane, $\mathrm{CH_4}$, and two molecules of oxygen, $\mathrm{O_2}$; on the right are two molecules of water, $\mathrm{H_2O}$, and one molecule of carbon dioxide, $\mathrm{CO_2}$. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.

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Figure 1.6.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

We can't actually see molecules breaking and forming bonds, although that's what defines chemical changes. We have to make other observations to indicate that a chemical change has happened. Some of the evidence for chemical change will involve the energy changes that occur in chemical changes, but some evidence involves the fact that new substances with different properties are formed in a chemical change.

Observations that help to indicate chemical change include:

- Temperature changes (either the temperature increases or decreases).
- Light given off.
- Unexpected color changes (a substance with a different color is made, rather than just mixing the original colors together).
- Bubbles are formed (but the substance is not boiling—you made a substance that is a gas at the temperature of the beginning materials, instead of a liquid).
- Different smell or taste (do not taste your chemistry experiments, though!).
- A solid forms if two clear liquids are mixed (look for *floaties*—technically called a precipitate).

Example 1.6.1

Label each of the following changes as a physical or chemical change. Give evidence to support your answer.

a. Boiling water.
b. A nail rusting.
c. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
d. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

a. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $\mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2O}\,(g)$
b. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $\mathrm{Fe} + \mathrm{O_2} \rightarrow \mathrm{Fe_2O_3}$
c. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
d. Chemical: the formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

Exercise 1.6.1

Label each of the following changes as a physical or chemical change.

a. A mirror is broken.
b. An iron nail corroded in moist air
c. Copper metal is melted.
d. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

Separating Mixtures Through Physical Changes

Homogeneous mixtures (solutions) can be separated into their component substances by physical processes that rely on differences in some physical property, such as differences in their boiling points. Two of these separation methods are distillation and crystallization. Distillation makes use of differences in volatility, a measure of how easily a substance is converted to a gas at a given temperature. A simple distillation apparatus for separating a mixture of substances, at least one of which is a liquid. The most volatile component boils first and is condensed back to a liquid in the water-cooled condenser, from which it flows into the receiving flask. If a solution of salt and water is distilled, for example, the more volatile component, pure water, collects in the receiving flask, while the salt remains in the distillation flask.

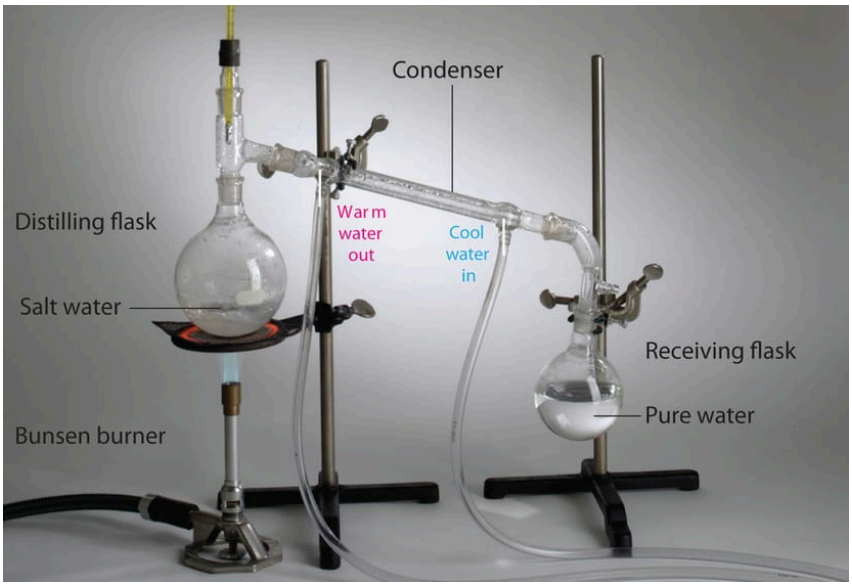


Figure 1.6.3: *The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.*

Parts of a distillation setup: Bunsen burner, salt water in distilling flask, condenser with cool water in and warm water out, pure water in receiving flask

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as brandy or whiskey. This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States.

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Another example for using physical properties to separate mixtures is filtration (Figure 1.6.4). Filtration is any mechanical, physical or biological operation that separates solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.

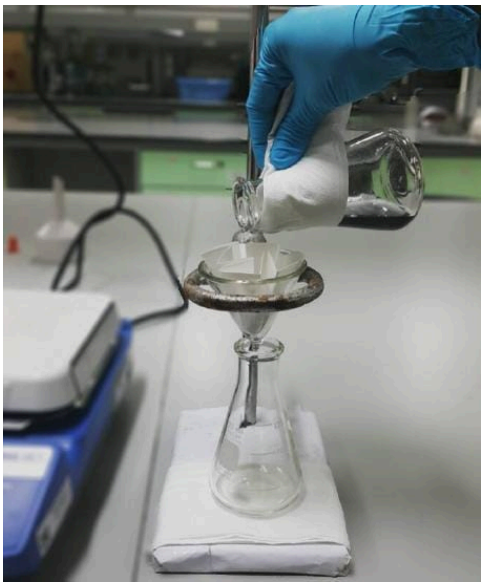


Figure 1.6.4: *Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).*

Summary

- Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.
- Physical changes are changes that do not alter the identity of a substance.
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change has occurred include color change, temperature change, light given off, formation of bubbles, formation of a precipitate, etc.

Contributions & Attributions

-

- Boundless (www.boundless.com)

Answers

Exercise 1.6.1
physical change

Exercise 1.6.1
chemical change

Exercise 1.6.1
physical change

Exercise 1.6.1
chemical change

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1.7: The Scientific Method - How Chemists Think

Learning Objectives

- Identify the components of the scientific method.

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments; which leads to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.7.1).

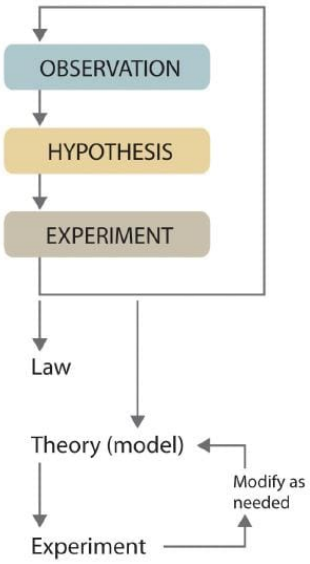


Figure 1.7.1: The Steps in the Scientific Method.

Step 1: Make observations

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: "the outside air temperature is cooler during the winter season," "table salt is a crystalline solid," "sulfur crystals are yellow," and "dissolving a penny in dilute nitric acid forms a blue solution and a brown gas." **Quantitative observations** are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: "the melting point of crystalline sulfur is 115.21° Celsius," and "35.9 grams of table salt—the chemical name of which is sodium chloride—dissolve in 100 grams of water at 20° Celsius." For the question of the dinosaurs' extinction, the initial observation was quantitative: iridium concentrations in sediments dating to 66 million years ago were 20–160 times higher than normal.

Step 2: Formulate a hypothesis

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a hypothesis, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience alternating periods of light and darkness corresponding to observed movements of the sun, moon, clouds, and shadows is consistent with either one of two hypotheses:

a. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
b. The sun revolves around Earth every 24 hours.

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Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either support or refute it.

Step 3: Design and perform experiments

After a hypothesis has been formed, scientists conduct experiments to test its validity. Experiments are systematic observations or measurements, preferably made under controlled conditions—that is—under conditions in which a single variable changes.

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether or not the original hypothesis is valid. If the hypothesis is valid, the scientist can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified and requires further experimentation.

Step 5: Development into a law and/or theory

More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a law, a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply states what happens; it does not address the question of why.

One example of a law, the law of definite proportions, which was discovered by the French scientist Joseph Proust (1754–1826), states that a chemical substance always contains the same proportions of elements by mass. Thus, sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine, in this case 39.34% sodium and 60.66% chlorine by mass, and sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass.

Whereas a law states only what happens, a theory attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time unless a major experimental error is discovered. In contrast, a theory, by definition, is incomplete and imperfect, evolving with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure 1.7.1 at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

Example 1.7.1

Classify each statement as a law, a theory, an experiment, a hypothesis, an observation.

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
d. When 10 g of ice were added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
e. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is a possible explanation for the origin of birds, so it is a hypothesis.
c. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
d. The temperature is measured before and after a change is made in a system, so these are observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

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Exercise 1.7.1

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter to ever play professional basketball.
e. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer a available at the end of this section.

Summary

The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. The key steps in the scientific method include the following:

- Step 1: Make observations.
- Step 2: Formulate a hypothesis.
- Step 3: Test the hypothesis through experimentation.
- Step 4: Accept or modify the hypothesis.
- Step 5: Develop into a law and/or a theory.

Answers

Exercise 1.7.1

experiment
law
theory
hypothesis
observation

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1.8: A Beginning Chemist - How to Succeed

Examples of the practical applications of chemistry are everywhere (Figure 1.8.1). Engineers need to understand the chemical properties of the substances needed to design biologically compatible implants for joint replacements; or to design roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Archeology and paleontology rely on chemical techniques to date bones and artifacts and identify their origins. Although law is not normally considered a field related to chemistry, forensic scientists use chemical methods to analyze blood, fibers, and other evidence as they investigate crimes. In particular, DNA matching—comparing biological samples of genetic material to see whether they could have come from the same person—has been used to solve many high-profile criminal cases as well as clear innocent people who have been wrongly accused or convicted. Forensics is a rapidly growing area of applied chemistry. In addition, the proliferation of chemical and biochemical innovations in industry is producing rapid growth in the area of patent law. Ultimately, the dispersal of information in all the fields in which chemistry plays a part requires experts who are able to explain complex chemical issues to the public through television, print journalism, the Internet, and popular books.

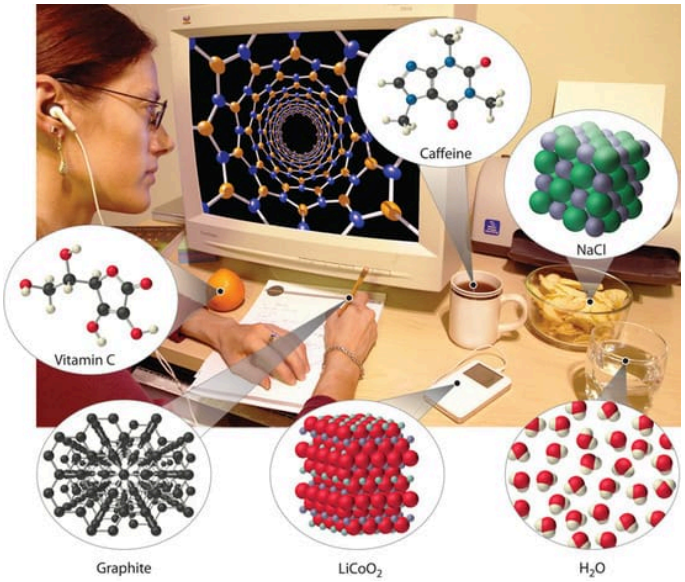


Figure 1.8.1: Chemistry in Everyday Life. Although most people do not recognize it, chemistry and chemical compounds are crucial ingredients in almost everything we eat, wear, and use.

Chemical compounds in everyday life: Vitamin C, graphite, lithium cobalt oxide, caffeine, sodium chloride, water

Hopefully at this point you are fully convinced of how important and useful the study of chemistry can be. You may, however, still be wondering exactly *what* it is that a chemist does. Chemistry is the study of matter and the changes that matter undergoes. In general, chemists are interested in both characteristics that you can test and observe, like a chemical's smell or color, and characteristics that are far too small to see, like what the oxygen you breathe in or the carbon dioxide you breath out looks like under a microscope 1,000 times more powerful than any existing in the world today.

Wait a minute... how can a chemist *know* what oxygen and carbon dioxide look like under a microscope that doesn't even exist? What happened to the scientific method? What happened to relying on *observations* and *careful measurements*? In fact, because chemists *can't* see the underlying structure of different materials, they have to rely on the scientific method even more! Chemists are a lot like detectives. Suppose a detective is trying to solve a murder case—what do they do? Obviously, the detective starts by visiting the site of the crime and looking for evidence. If the murderer has left enough clues behind, the detective can piece together a theory explaining what happened.

Even though the detective wasn't at the crime scene when the crime was committed and didn't actually see the murderer kill the victim, with the right evidence, the detective can be pretty sure of how the crime took place. It is the same with chemistry. When chemists go into the laboratory, they collect evidence by making measurements. Once chemists have collected enough clues from the properties that they can observe, they use that evidence to piece together a theory explaining the properties that they cannot observe—the properties that are too small to see.

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What kinds of properties do chemists actually measure in the laboratory? Well, you can probably guess a few. Imagine that you go to dinner at a friend's house and are served something that you don't recognize, what types of observations might you make to determine exactly *what* you've been given? You might smell the food. You might note the color of the food. You might try to decide whether the food is a liquid or a solid because if it's a liquid, it's probably soup or a drink. The temperature of the food could be useful if you wanted to know whether or not you had been served ice cream! You could also pick up a small amount of food with your fork and try to figure out how much it weighs—a light dessert might be something like an angel cake, while a heavy dessert is probably a pound cake. The quantity of food you have been given might be a clue too. Finally, you might want to know something about the food's texture—is it hard and granular like sugar cubes, or soft and easy to spread, like butter?

Believe it or not, the observations you are likely to make when trying to identify an unknown food are very similar to the observations that a chemist makes when trying to learn about a new material. Chemists rely on smell, color, state (whether it is a solid or liquid or gas), temperature, volume, mass (which is related to weight—as will be discussed in a later section), and texture. There is, however, one property possibly used to learn about a food, but that should *definitely* not be used to learn about a chemical—taste!

In the sections on the Atomic Theory, you will see exactly how measurements of certain properties helped early scientists to develop theories about the chemical structure of matter on a scale much smaller than they could ever hope to see. You will also learn how these theories, in turn, allow us to make predictions about new materials that humankind has not yet created.

The video below gives you some important tips on how to study chemistry in this class. With practice, you too can learn to think like a chemist, and you may even enjoy it!

YouTube Video

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Video 1.8.1: How To Study Chemistry.

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2: THE NUMERICAL SIDE OF CHEMISTRY

2.1: Taking Measurements

2.2: Significant Figures - Writing Numbers to Reflect Precision

2.3: Scientific Notation - Writing Large and Small Numbers

2.4: Significant Figures in Calculations

2.5: The Basic Units of Measurement

2.6: Temperature - Random Motion of Molecules and Atoms

2.7: Density

2.8: Problem Solving and Unit Conversions

2.9: Metric Unit Conversions

2.10: Solving Multi-step Conversion Problems

2.11: Units Raised to a Power

2.12: Derived Units

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2.1: Taking Measurements

Learning Objectives

- Express quantities properly, using a number and a unit.

A coffee maker's instructions tell you to fill the coffee pot with 4 cups of water and to use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure 2.1.1); the nurse is also measuring.



Figure 2.1.1: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GFDL; Pia von Lützau).

Chemists measure the properties of matter and express these measurements as quantities. A quantity is an amount of something and consists of a number and a unit. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5. If you ask a friend how far they walk from home to school, and the friend answers "12" without specifying a unit, you do not know whether your friend walks 12 kilometers, 12 miles, 12 furlongs, or 12 yards. *Both a number and a unit must be included to express a quantity properly.*

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. The next two sections examine the rules for expressing numbers.

Example 2.1.1

Identify the number and the unit in each quantity.

a. one dozen eggs
b. 2.54 centimeters
c. a box of pencils
d. 88 meters per second

Solution

a. The number is one, and the unit is a dozen eggs.
b. The number is 2.54, and the unit is centimeter.
c. The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.
d. The number is 88, and the unit is meters per second. Note that in this case the unit is actually a combination of two units: meters and seconds.

Key Take Away

- Identify a quantity properly with a number and a unit.

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2.2: Significant Figures - Writing Numbers to Reflect Precision

Learning Objectives

- Identify the number of significant figures in a reported value.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length at $2.553\ \text{cm}$ as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

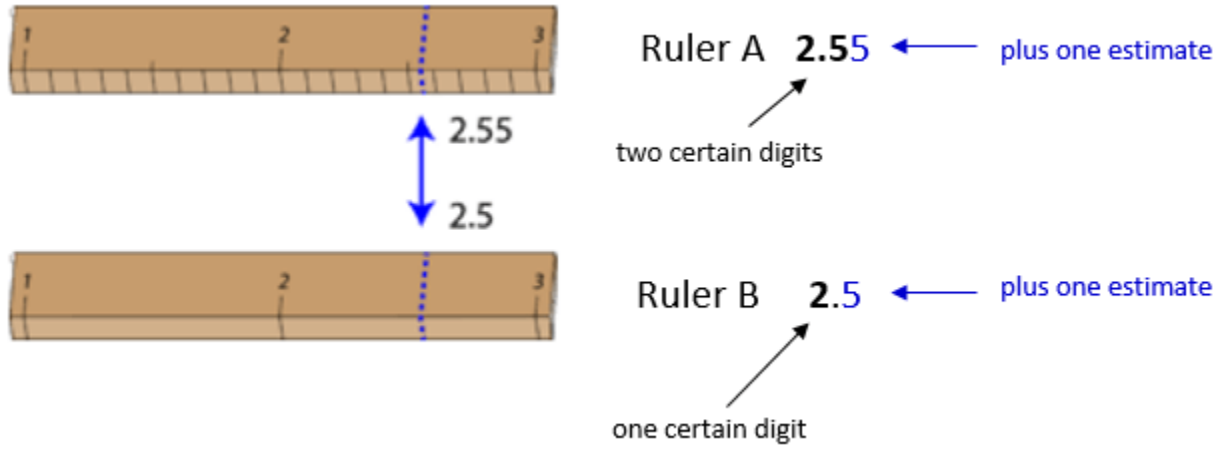


Figure 2.2.1: Measurement with two different rulers. Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest $0.1\ \text{g}$, other highly sensitive balances are capable of measuring to the nearest $0.001\ \text{g}$ or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 2.2.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

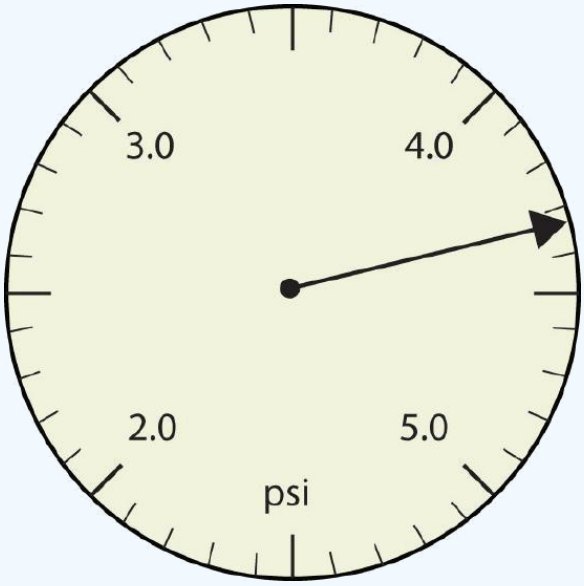
With either ruler, it is clear that the length of the object is between 2 and $3\ \text{cm}$. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as $2.5\ \text{cm}$. However, another person may judge that the measurement is $2.4\ \text{cm}$ or perhaps $2.6\ \text{cm}$. While the 2 is known for certain, the value of the tenths digit is uncertain.

The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as $2.55\ \text{cm}$. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be $2.54\ \text{cm}$ or $2.56\ \text{cm}$. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

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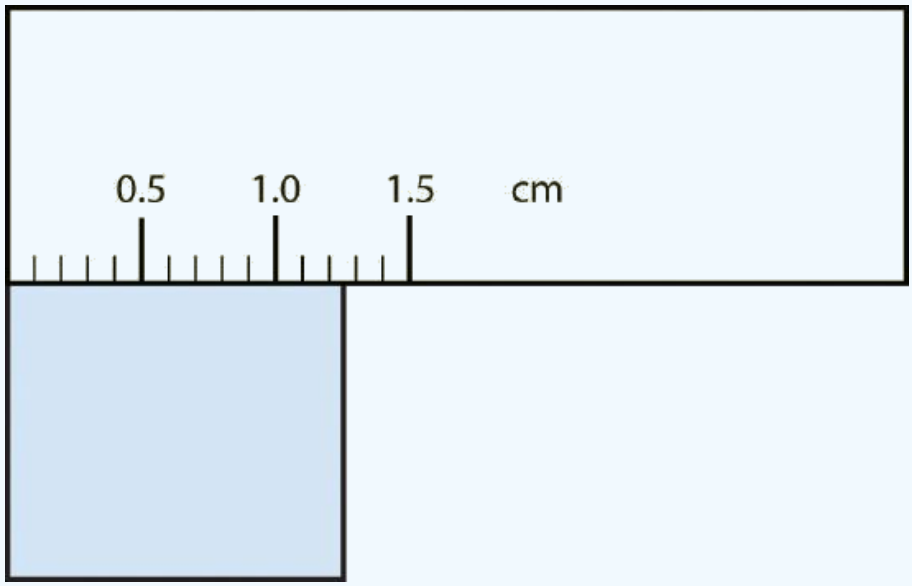
Example 2.2.1: Reporting Measurements to the Proper Number of Significant Figures

Use each diagram to report a measurement to the proper number of significant figures.



a.

b.



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

Solutions

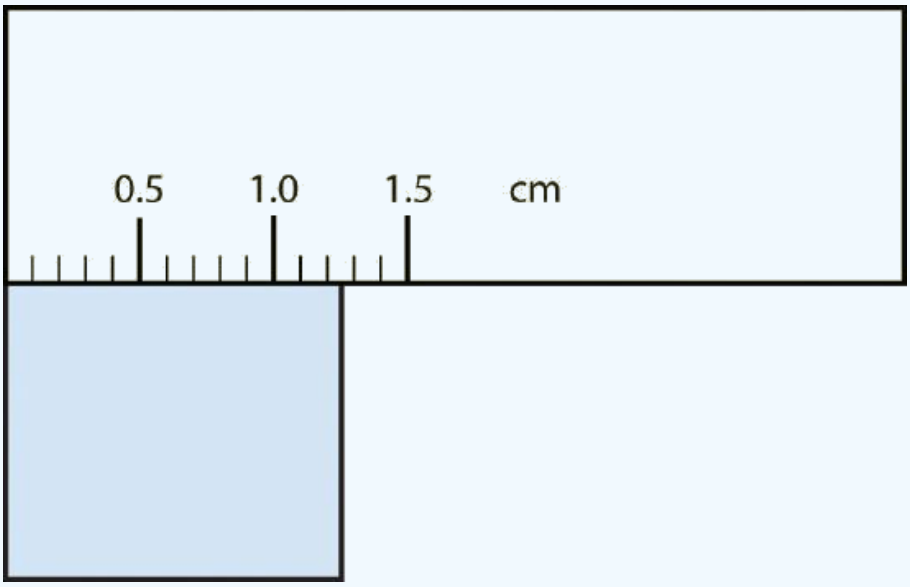
Solutions to Example 2.3.1

	Explanation	Answer
a.	The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it's at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. The symbol psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire. The measurement is reported to three significant figures.	4.33 psi

	Explanation	Answer
b.	The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle's width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. The measurement is reported to three significant figures.	1.25 cm

Exercise 2.2.1

What would be the reported width of this rectangle?



Answer available at the end of this section.

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 2.2.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.

Rule	Examples
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous. ◦ 1.4×10^3 has two significant figures. ◦ 1.40×10^3 three significant figures. ◦ 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000...). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

Example 2.2.2

Give the number of significant figures in each. Identify the rule for each.

a. 5.87
b. 0.031
c. 52.90
d. 00.2001
e. 500
f. 6 atoms

Solution

Solution to Example 2.3.2

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d	The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e	The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f	The 6 is a counting number. A counting number is an exact number.	6, infinite

Exercise 2.2.2

Give the number of significant figures in each.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people
e. 3,800 g

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Answer e available at the end of this section.

Accuracy and Precision

Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrates the concepts of accuracy and precision.

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YouTube Video

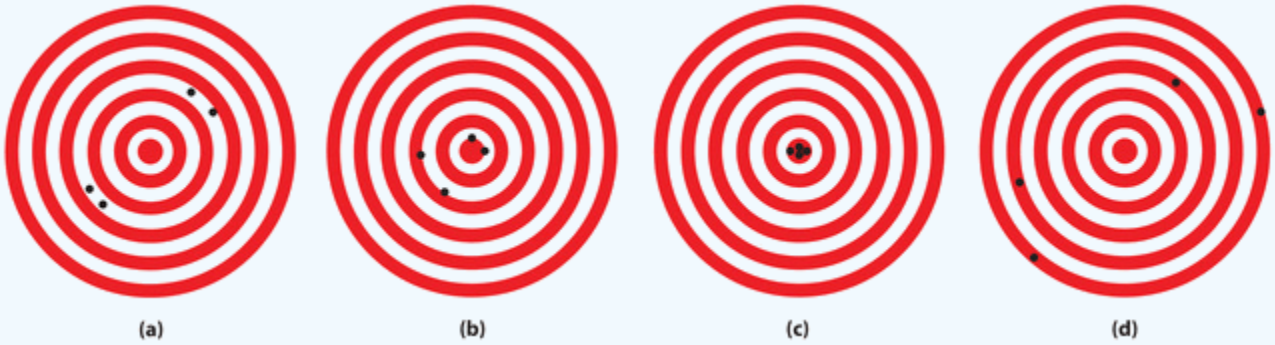
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Video 2.2.1: Difference between precision and accuracy.

Example 2.2.3

The following archery targets show marks that represent the results of four sets of measurements.



Which target shows

a. a precise, but inaccurate set of measurements?
b. a set of measurements that is both precise and accurate?
c. a set of measurements that is neither precise nor accurate?

Solution

a. Set **(a)** is precise, but inaccurate.
b. Set **(c)** is both precise and accurate.
c. Set **(d)** is neither precise nor accurate.

Summary

Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool. Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

Answers

Exercise 2.2.1
1.25 cm

Exercise 2.2.2
three significant figures

Exercise 2.2.2
four significant figures

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Exercise 2.2.2

four significant figures

Exercise 2.2.2

infinite (exact number)

Exercise 2.2.2

Ambiguous, could be two, three or four significant figures.

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2.3: Scientific Notation - Writing Large and Small Numbers

Learning Objectives

- Express a large number or a small number in scientific notation.
- Carry out arithmetical operations and express the final answer in scientific notation

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. In scientific notation, these numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 (1 ≤ N < 10), and n is a positive or negative integer ($10^0 = 1$). The number 10 is called the base because it is this number that is raised to the power n. Although a base number may have values other than 10, the base number in scientific notation is always 10.

A simple way to convert numbers to scientific notation is to move the decimal point as many places to the left or right as needed to give a number from 1 to 10 (N). The magnitude of n is then determined as follows:

- If the decimal point is moved to the left n places, n is positive.
- If the decimal point is moved to the right n places, n is negative.

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example $2.3.1$.

Example $2.3.1$: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

a. 637.8
b. 0.0479
c. 7.86
d. 12,378
e. 0.00032
f. 61.06700
g. 2002.080
h. 0.01020

Solution

Solutions to Example 2.2.1

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, n = 2.	6.378×10^2

	Explanation	Answer
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, n = –2.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^0 = 1$.)	7.86×10^{0}
d	Because the decimal point was moved four places to the left, n = 4.	1.2378×10^{4}
e	Because the decimal point was moved four places to the right, n = –4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^{1}
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^{3}
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example $2.3.2$ illustrates how to do this.

Example $2.3.2$: Expressing Sums and Differences in Scientific Notation

Carry out the appropriate operation and then express the answer in scientific notation.

a. $(1.36 \times 10^2) + (4.73 \times 10^3)$
b. $(6.923 \times 10^{-3}) - (8.756 \times 10^{-4})$

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) =$ $(1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) =$ $(0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87×10^3

	Explanation	Answer
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474×10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.047×10^{-3}

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example 2.3.3.

Example 2.3.3: Expressing Products and Quotients in Scientific Notation

Perform the appropriate operation and express your answer in scientific notation.

a. $(6.022 \times 10^{23})(6.42 \times 10^{-2})$
b. $\dfrac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$
c. $\dfrac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

Solution to Example 2.2.3

	Explanation	
a	In multiplication, we add the exponents: $(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	3.87×10^{22}

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	Explanation	
b	In division, we subtract the exponents: $\frac{1.67\times10^{-24}}{9.12\times10^{-28}} =$ $\frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^4$ In converting 0.183×10^4 to scientific notation, n has become more negative by 1 because the value of N has increased.	1.83 $\times 10^3$
c	This problem has both multiplication and division: $\frac{(6.63\times10^{-34})(6.0\times10)}{(8.52\times10^{-2})} =$ $\frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$	4.7 $\times 10^{-31}$

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2.4: Significant Figures in Calculations

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement $207.518\ \text{m}$. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table $2.4.1$.

Table $2.4.1$: Rounding examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Notice that the more rounding that is done, the less reliable the figure is. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

It is important to be aware of significant figures when you are mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238... to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers that have only three significant figures each? When performing mathematical operations, there are two rules for limiting the number of significant figures in an answer—one rule is for addition and subtraction, and one rule is for multiplication and division.

> In operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the **least precise** operation. An answer is no more precise than the least precise number used to get the answer.

Multiplication and Division

For multiplication or division, the rule is to count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

$$\underbrace{38.65}_{\text{4 sig figs}} \times \underbrace{105.93}_{\text{5 sig figs}} = \underbrace{4{,}094.1945}_{\text{reduce to 4 sig figs}}$$

The final answer, limited to four significant figures, is 4,094. The first digit dropped is 1, so we do not round up.

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 450 has two significant figures and would be written in scientific notation as 4.5×10^2, whereas 450.0 has four significant figures and would be written as 4.500×10^2. In scientific notation, all significant figures are listed explicitly.

Example 2.4.1

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 23.096 × 90.300
b. 125 × 9.000

Solution

a

Explanation	Answer
The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	2.0856×10^3

b

Explanation	Answer
The calculator gives 1,125 as the answer, but we limit it to three significant figures.	1.13×10^3

Addition and Subtraction

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.71, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

$$\begin{array}{r} 1.2 \\ \underline{4.41} \\ 5.61 \end{array}$$

↑ limit final answer to the tenths column: 5.6

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We drop the last digit—the 1—because it is not significant to the final answer.

The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater, and rounded down if the first dropped digit is less than 5.

$$\begin{array}{r} 77.2 \\ \underline{10.46} \\ 87.66 \end{array}$$

↑ limit final answer to the tenths column and round up: 87.7

Example 2.4.2

a. 13.77 + 908.226
b. 1,027 + 611 + 363.06

Solution

a

Explanation	Answer
The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place, we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$

b

Explanation	Answer
The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2,001.06 = 2.001 \times 10^3$

Exercise 2.4.2

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 217 ÷ 903
b. 13.77 + 908.226 + 515
c. 255.0 − 99
d. 0.00666 × 321

Answer a: available at the end of this section.

Remember that calculators do not understand significant figures. *You* are the one who must apply the rules of significant figures to a result from your calculator.

Calculations Involving Multiplication/Division and Addition/Subtraction

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain at least one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.

YouTube Video

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Video 2.4.1: Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).

YouTube Video

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Video 2.4.2: https://www.youtube.com/watch?v=__csP0NtlGI

In the worked examples in this text, we will often show the results of intermediate steps in a calculation. In doing so, we will show the results to only the correct number of significant figures allowed for that step, in effect treating each step as a separate calculation. This procedure is intended to reinforce the rules for determining the number of significant figures, but in some cases it may give a final answer that differs in the last digit from that obtained using a calculator, where all digits are carried through to the last step.

Example 2.4.3

a. 2(1.008 g) + 15.99 g
b. 137.3 s + 2(35.45 s)
c. $\frac{118.7g}{2} - 35.5g$

Solution

a.

Explanation	Answer
2(1.008 g) + 15.99 g = **Perform multiplication first.** 2 (1.008 g *4 sig figs*) = **2.016 g** *4 sig figs* The number with the least number of significant figures is 1.008 g; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Then, perform the addition.** **2.016 g** *thousandths place* + 15.99 g *hundredths place (least precise)* = 18.006 g **Round the final answer.** Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)

b.

Explanation	Answer
137.3 s + 2(35.45 s) = **Perform multiplication first.** 2(35.45 s *4 sig figs*) = **70.90 s** *4 sig figs* The number with the least number of significant figures is 35.45; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Then, perform the addition.** 137.3 s *tenths place (least precise)* + **70.90 s** *hundredths place* = 208.20 s **Round the final answer.** Round the final answer to the tenths place based on 137.3 s.	208.2 s

c.

Explanation	Answer
$\frac{118.7g}{2} - 35.5g =$ **Perform division first.** $\frac{118.7g}{2}$ *4 sig figs* = **59.35 g** *4 sig figs* The number with the least number of significant figures is 118.7 g; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Perform subtraction next.** **59.35 g** *hundredths place* – 35.5 g *tenths place (least precise)* = 23.85 g **Round the final answer.** Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

Exercise 2.4.3

Complete the calculations and report your answers using the correct number of significant figures.

a. 5(1.008s) - 10.66 s
b. 99.0 cm+ 2(5.56 cm)

Answer a available at the end of this section.

Summary

- Rounding
  - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1 (e.g. 2.9699 rounded to three significant figures is 2.97).
  - If the number to be dropped is less than 5, there is no change (e.g. 4.00443 rounded to four significant figures is 4.004).
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the number with the fewest significant figures.
- The rule in addition and subtraction is that the answer is given the same number of decimal places as the term with the fewest decimal places.

Answers

Exercise 2.4.2

$0.240 = 2.40 \times 10^{-1}$
$1,437 = 1.437 \times 10^{3}$
$156 = 1.56 \times 10^{2}$
$2.14 = 2.14 \times 10^{0}$

Exercise 2.4.3

-5.62 s
110.2 cm

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2.5: The Basic Units of Measurement

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

How long is a yard? It depends on whom you ask and when you asked the question. Today we have a standard definition of the yard, which you can see marked on every football field. If you move the ball ten yards, you get a first down and it does not matter whether you are playing in Los Angeles, Dallas, or Green Bay. But at one time that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. Of course, the problem there is simple: new king, new distance (and then you have to re-mark all of those football fields).

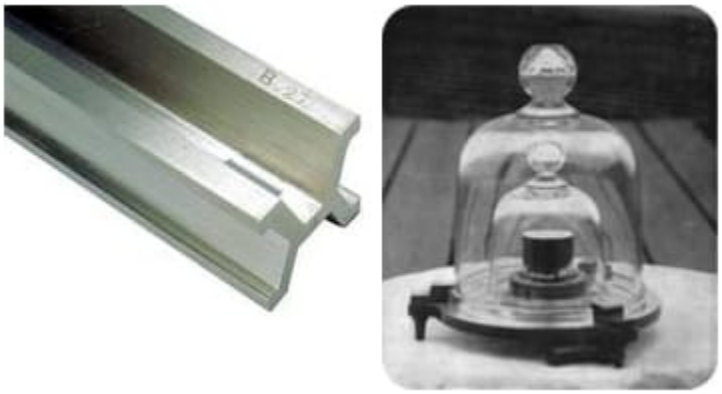


Figure 2.5.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 2.5.1.

Table 2.5.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Unfortunately, the Metric System is Not Ubiquitous

The map below shows the adoption of the SI units in countries around the world. The United States has legally adopted the metric system for measurements, but does not use it in everyday practice. Great Britain and much of Canada use a combination of metric and imperial units.

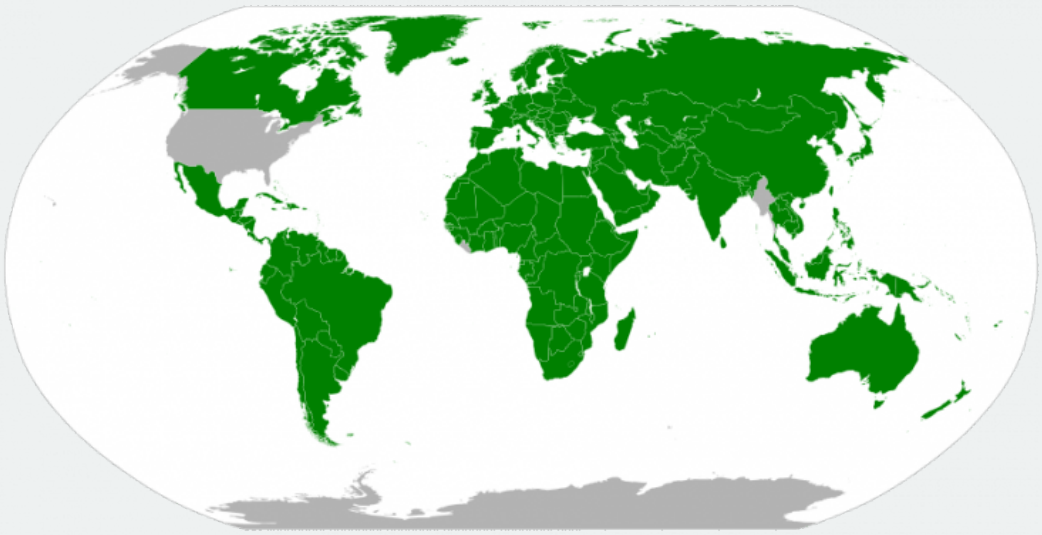


Figure 2.5.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great". Table 2.5.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 2.5.2: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\mathrm{Gm}) = 10^9\ \mathrm{m}$
mega	M	1,000,000	1 megameter $(\mathrm{Mm}) = 10^6\ \mathrm{m}$
kilo	k	1,000	1 kilometer $(\mathrm{km}) = 1,000\ \mathrm{m}$
hecto	h	100	1 hectometer $(\mathrm{hm}) = 100\ \mathrm{m}$
deka	da	10	1 dekameter $(\mathrm{dam}) = 10\ \mathrm{m}$
		1	1 meter (m)
deci	d	1/10	1 decimeter $(\mathrm{dm}) = 0.1\ \mathrm{m}$
centi	c	1/100	1 centimeter $(\mathrm{cm}) = 0.01\ \mathrm{m}$
milli	m	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001\ \mathrm{m}$

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Prefix	Unit Abbreviation	Meaning	Example
micro	μ	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6}$ m
nano	n	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9}$ m
pico	p	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12}$ m

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. We use "m" for meter and not "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So we would write 3.5 milliliters as 3.5 mL.

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs 6.5 kg, this is easier than saying it weighs 6500 g or 0.65 dag. All three are correct, but the kg units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

Example 2.5.1: Unit Abbreviations

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kiloliter
b. microsecond
c. decimeter
d. nanogram

Solutions

Solutions to Example 2.5.1

	Explanation	Answer
a	The prefix kilo means “1,000 ×,” so 1 kL equals 1,000 L.	kL
b	The prefix micro implies 1/1,000,000th of a unit, so 1 µs equals 0.000001 s.	µs
c	The prefix deci means 1/10th, so 1 dm equals 0.1 m.	dm
d	The prefix nano means 1/1000000000, so a nanogram is equal to 0.000000001 g.	ng

Exercise 2.5.1

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kilometer
b. milligram
c. nanosecond
d. centiliter

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

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Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

Answers

Exercise 2.5.1

km

Exercise 2.5.1

mg

Exercise 2.5.1

ns

Exercise 2.5.1

cL

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2.6: Temperature - Random Motion of Molecules and Atoms

Learning Objectives

- Identify the different between temperature and heat.
- Recognize the different scales used to measure temperature

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.



Figure 2.6.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.



Figure 2.6.2: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The

German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32^{\circ}\mathrm{F}$ and the boiling point as $212^{\circ}\mathrm{F}$. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$ respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".

The Kelvin Scale

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of $0\ \mathrm{K}$, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is $273.15\ \mathrm{K}$, while the boiling point is $373.15\ \mathrm{K}$. Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales where temperatures are referred to as "degrees F" or "degrees C", we simply designate temperatures in the Kelvin scale as kelvins.

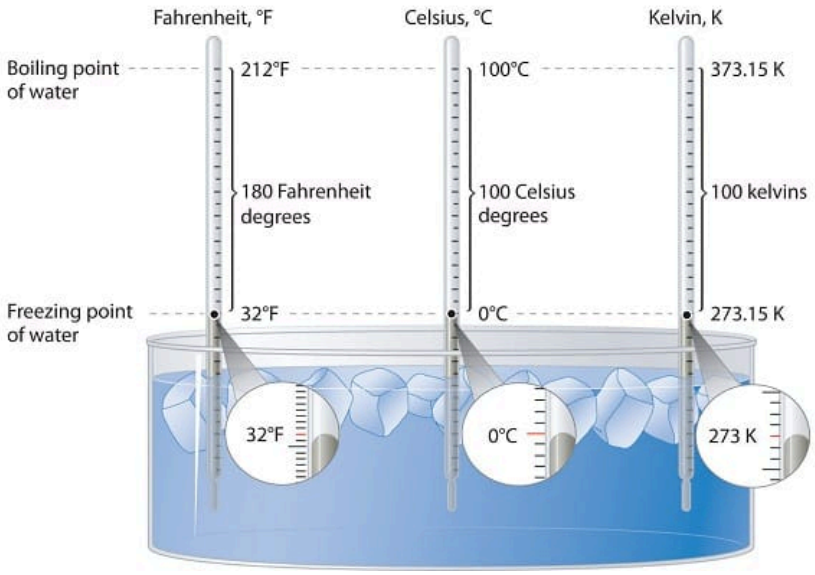


Figure 2.6.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Converting Between Scales

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in °C}) + 273.15 = T\,(\text{in K}) \tag{3.10.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in °C}) \tag{3.10.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which

the numerical value is the same on both the Fahrenheit and Celsius scales: –40°C = –40°F. The relationship between the scales is as follows:

$$°C = \frac{(°F - 32)}{1.8} \tag{3.10.3}$$

$$°F = 1.8 \times (°C) + 32 \tag{3.10.4}$$

Example 2.6.1: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 3.10.3:

$$°C = \frac{(103.5°F - 32)}{1.8} \tag{2.6.1}$$

$$= 39.7\,°C \tag{2.6.2}$$

Converting from Celsius to Kelvin requires the use of Equation 3.10.1:

$$K = 39.7\,°C + 273.15 \tag{2.6.3}$$

$$= 312.9\,K \tag{2.6.4}$$

Exercise 2.6.1

Convert each temperature to °C and °F.

a. the temperature of the surface of the sun (5800 K)
b. the boiling point of gold (3080 K)
c. the boiling point of liquid nitrogen (77.36 K)

Answer (a) available at the end of this section.

Answer (b) available at the end of this section.

Answer (c) available at the end of this section.

Summary

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).

Contributions & Attributions

Answers

Exercise 2.6.1
5527 K, 9980 °F

Exercise 2.6.1
2807 K, 5084 °F

Exercise 2.6.1
-195.79 K, -320.42 °F

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2.7: Density

Learning Objectives

- Define density.
- Use density as a conversion factor.

Density (ρ) is a physical property found by dividing the mass of an object by its volume. Regardless of the sample size, density is always constant. For example, the density of a pure sample of tungsten is always 19.25 grams per cubic centimeter. This means that whether you have one gram or one kilogram of the sample, the density will never vary. The equation, as we already know, is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

or just

$$\rho = \frac{m}{V} \tag{2.7.1}$$

Based on this equation, it's clear that density can, and does, vary from element to element and substance to substance due to differences in the relationship of mass and volume. Pure water, for example, has a density of 0.998 g/cm^3 at 25° C. The average densities of some common substances are in Table 2.7.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Table 2.7.1: Densities of Common Substances

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Density can be measured for all substances—solids, liquids and gases. For solids and liquids, density is often reported using the units of g/cm^3. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

Example 2.7.1: Ethyl Alcohol

Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g

Solution

This is a direct application of Equation 2.7.1:

$$\rho = \frac{23.71002\,g}{30.2\,mL} = 0.785\,g/mL$$

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Exercise 2.7.1

a. Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.
b. If you have a 2.130 mL sample of acetic acid with mass 0.002234 kg, what is the density in kg/L?

Answer a available at the end of this section.

Density as a Conversion Factor

Conversion factors can also be constructed for converting between different kinds of units. For example, density can be used to convert between the mass and the volume of a substance. Consider mercury, which is a liquid at room temperature and has a density of 13.6 g/mL. The density tells us that 13.6 g of mercury have a volume of 1 mL. We can write that relationship as follows:

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

$$\frac{13.6\ g}{1\ mL} = 1$$

and

$$\frac{1\ mL}{13.6\ g} = 1$$

Which one do we use? It depends, as usual, on the units we need to cancel and introduce. For example, suppose we want to know the mass of 2.0 mL of mercury. We would use the conversion factor that has milliliters on the bottom (so that the milliliter unit cancels) and grams on top, so that our final answer has a unit of mass:

$$2.0\ \cancel{\text{mL}} \times \frac{13.6\ \text{g}}{1\ \cancel{\text{mL}}} = 27.2\ \text{g} = 27\ \text{g}$$

In the last step, we limit our final answer to two significant figures because the volume quantity has only two significant figures; the 1 in the volume unit is considered an exact number, so it does not affect the number of significant figures. The other conversion factor would be useful if we were given a mass and asked to find volume, as the following example illustrates.

Density can be used as a conversion factor between mass and volume.

Example 2.7.2: Mercury Thermometer Steps for Problem Solving

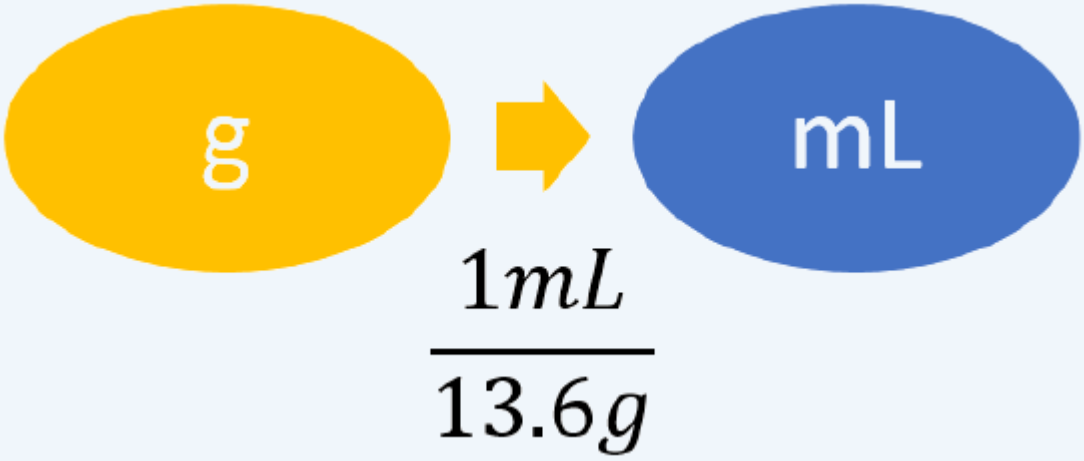
A mercury thermometer for measuring a patient's temperature contains 0.750 g of mercury. What is the volume of this mass of mercury?

Solution

Solution to Example 2.9.2

Steps for Problem Solving	**Unit Conversion**

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Identify the "given" information and what the problem is asking you to "find."	Given: 0.750 g Find: mL
List other known quantities.	13.6 g/mL (density of mercury)
Prepare a concept map.	
Calculate.	$0.750\ \cancel{g} \times \frac{1\ \text{mL}}{13.6\ \cancel{g}} = 0.055147...\ \text{mL} \approx 0.0551\ \text{mL}$ We have limited the final answer to three significant figures.

Exercise 2.7.2

What is the volume of 100.0 g of air if its density is 1.3 g/L?

Answer available at the end of this section.

Summary

- Density is defined as the mass of an object divided by its volume.
- Density can be used as a conversion factor between mass and volume.

Answers

Exercise 2.7.1

$0.274\ kg/L$

$1.049\ kg/L$

Exercise 2.7.2

$77\ L$

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2.8: Problem Solving and Unit Conversions

Learning Objectives

- To convert a value reported in one unit to a corresponding value in a different unit using conversion factors.

During your studies of chemistry (and physics also), you will note that mathematical equations are used in many different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Algebra skills become very important here!

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, or a more complex one can convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with (Table 2.8.1).

Table 2.8.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table 2.8.1. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions in the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds = 1 minute).

If you learned the SI units and prefixes described, then you know that 1 cm is 1/100th of a meter.

$$1\ \text{cm} = \frac{1}{100}\ \text{m} = 10^{-2}\text{m}$$

or

$$100\ \text{cm} = 1\ \text{m}$$

Suppose we divide both sides of the equation by 1m (both the number *and* the unit):

$$\frac{100\ \text{cm}}{1\ \text{m}} = \frac{1\ \text{m}}{1\ \text{m}}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the

right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100\text{ cm}}{1\text{ m}} = \frac{1000\text{ mm}}{1\text{ m}} = \frac{1 \times 10^6 \mu\text{m}}{1\text{ m}} = 1$$

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units.

Performing Dimensional Analysis

Dimensional analysis is amongst the most valuable tools that physical scientists use. Simply put, it is the conversion between an amount in one unit to the corresponding amount in a desired unit using various conversion factors. This is valuable because certain measurements are more accurate or easier to find than others. The use of units in a calculation to ensure that we obtain the final proper units is called *dimensional analysis*.

Here is a simple example. How many centimeters are there in 3.55 m? Perhaps you can determine the answer in your head. If there are 100 cm in every meter, then 3.55 m equals 355 cm. To solve the problem more formally with a conversion factor, we first write the quantity we are given, 3.55 m. Then we multiply this quantity by a conversion factor, which is the same as multiplying it by 1. We can write 1 as $\frac{100\text{ cm}}{1\text{ m}}$ and multiply:

$$3.55\text{ m} \times \frac{100\text{ cm}}{1\text{ m}}$$

The 3.55 m can be thought of as a fraction with a 1 in the denominator. Because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out:

$$\frac{3.55\ \cancel{\text{m}}}{1} \times \frac{100\text{ cm}}{1\ \cancel{\text{m}}}$$

The final step is to perform the calculation that remains once the units have been canceled:

$$\frac{3.55}{1} \times \frac{100\text{ cm}}{1} = 355\text{ cm}$$

In the final answer, we omit the 1 in the denominator. Thus, by a more formal procedure, we find that 3.55 m equals 355 cm. A generalized description of this process is as follows:

quantity (in old units) × conversion factor = quantity (in new units)

You may be wondering why we use a seemingly complicated procedure for a straightforward conversion. In later studies, the conversion problems you encounter *will not always be so simple*. If you master the technique of applying conversion factors, you will be able to solve a large variety of problems.

In the previous example, we used the fraction $\frac{100\text{ cm}}{1\text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1\text{ m}}{100\text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

$$3.55\text{ m} \times \frac{1\text{ m}}{100\text{ cm}} = 0.0355\frac{\text{m}^2}{\text{cm}}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 2.8.1 shows a concept map for constructing a proper conversion.

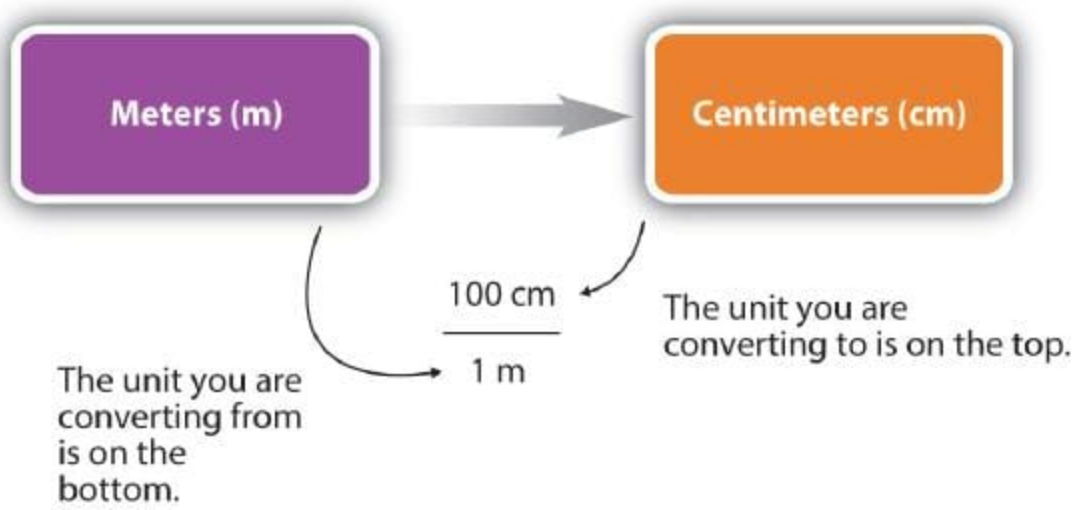


Figure 2.8.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Converting meters to centimeters: the unit you are converting is on the bottom of the fraction, and the unit to convert to is on top

General Steps in Performing Dimensional Analysis

1. Identify the "**given**" information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to "**find**"? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

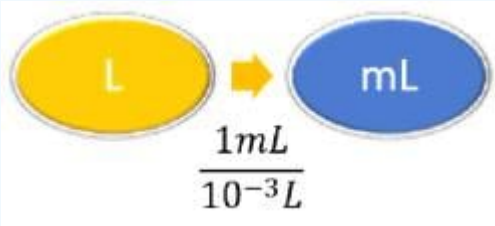
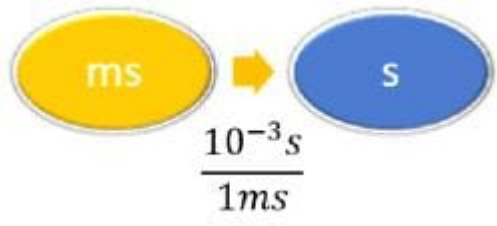
Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

- Numbers in conversion factors based on prefix changes, such as kilograms to grams, are *not* considered in the determination of significant figures in a calculation because the numbers in such conversion factors are exact.
- Exact numbers are defined or counted numbers, not measured numbers, and can be considered as having an infinite number of significant figures. (In other words, 1 kg is exactly 1,000 g, by the definition of kilo-.)
- Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact.
- In contrast, conversion factors that come from measurements (such as density, as we will see shortly) or that are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

Example 2.8.1

Steps for Problem Solving for Example 2.6.1 and 2.6.2

	Example 2.8.1	**Example** 2.8.2
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\ mL = 10^{-3} L$	$1\ ms = 10^{-3} s$
Prepare a concept map and use the proper conversion factor.		
Cancel units and calculate.	$4.7\cancel{L} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{L}} = 4,700\ \text{mL}$ or $4.7\cancel{L} \times \frac{1,000\ \text{mL}}{1\ \cancel{L}} = 4,700\ \text{mL}$ or 4.7 x 10^3 2SF, not ambiguous	$18\ \cancel{\text{ms}} \times \frac{10^{-3}\ \text{s}}{1\ \cancel{\text{ms}}} = 0.018\ \text{s}$ or $18\ \cancel{\text{ms}} \times \frac{1\ \text{s}}{1,000\ \cancel{\text{ms}}} = 0.018\ \text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

Exercise 2.8.1

Perform each conversion.

a. 101,000 ns to seconds
b. 32.08 kg to grams
c. 1.53 grams to cg

Answer a: available at the end of this section.

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce the appropriate units.

Answers

Exercise 2.8.1

$1.01000x10^{-4}s$

$3.208x10^{4}g$

$1.53x10^{2}cg$

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2.9: Metric Unit Conversions

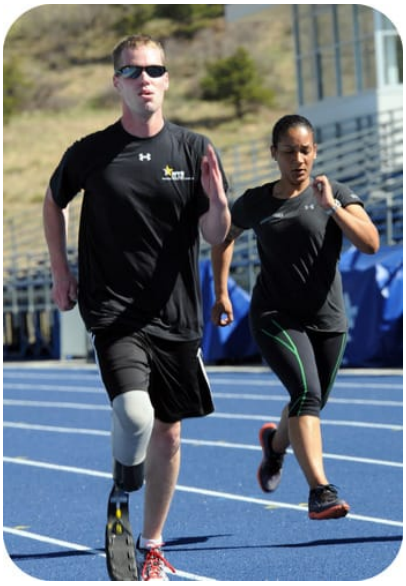


Figure 2.9.1 (Public Domain; the US Army via Wikipedia)

How can a number of track laps be converted to a distance in meters?

You are training for a 10-kilometer run by doing laps on a 400-meter track. You ask yourself how many times you need to run around the track in order to cover ten kilometers (...more than you realize). By using dimensional analysis, you can easily determine the number of laps needed to cover the $10\ \text{km}$ distance.

Metric Unit Conversions

The metric system's many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Example 2.9.1

A particular experiment for science class requires $120\ \text{mL}$ of a solution. The science teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires $120\ \text{mL}$
- $1\ \text{L} = 1000\ \text{mL}$

Unknown

Since each experiment requires $120\ \text{mL}$ of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get $4800\ \text{mL}$ of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate.

$$4800\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 4.8\ \text{L}$$

Note that the conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3:Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Two-Step Metric Unit Conversions

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (un-prefixed) unit, followed by a conversion to the desired unit. The following example will illustrate this method.

Example 2.9.2

Convert $4.3\ \text{cm}$ to μm.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1\ \text{m} = 100\ \text{cm}$
- $1\ \text{m} = 10^6\ \mu\text{m}$

Unknown

You may need to consult a table for the multiplication factor represented by each metric prefix. First convert cm to m, then convert m to μm.

Step 2: Calculate.

$$4.3\ \text{cm} \times \frac{1\ \text{m}}{100\ \text{cm}} \times \frac{10^6\ \mu\text{m}}{1\ \text{m}} = 43,000\ \mu\text{m}$$

Each conversion factor is written so that the unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

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Summary

- Dimensional analysis can be used to carry out metric unit conversions.

Review

1. Perform the following conversions.
   a. 0.074 km to m
   b. 24,600 µg to g
   c. 4.9×10^{7} µg to kg
   d. 84 dm to mm

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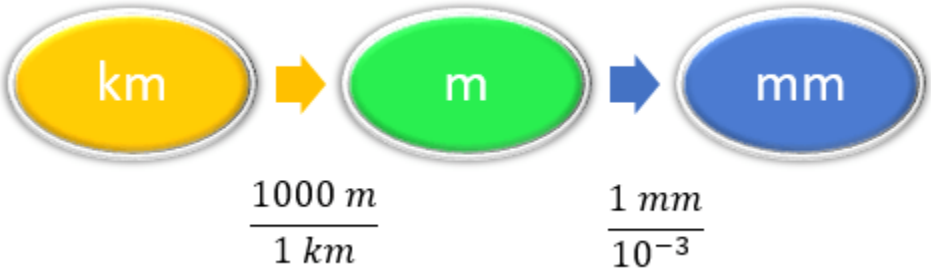
2.10: Solving Multi-step Conversion Problems

Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence. We first convert the given amount in km to the base unit, which is meters. We know that 1,000 m =1 km.

Then we convert meters to mm, remembering that $1\ \text{mm} = 10^{-3}\ \text{m}$.

Concept Map



Convert kilometers to meters to millimeters: use conversion factors 1000 meters per 1 kilometer and 1 millimeter per 0.001 meter

Calculation

$$54.7\ \cancel{\text{km}} \times \frac{1{,}000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1\ \text{mm}}{10^{-3}\ \cancel{\text{m}}} = 54{,}700{,}000\ \text{mm}$$
$$= 5.47 \times 10^{7}\ \text{mm}$$

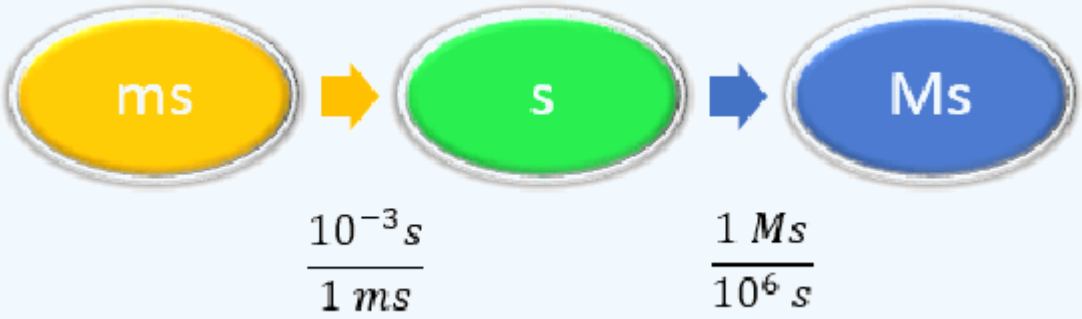
In each step, the previous unit is canceled and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.

Example 2.10.1: Unit Conversion

Convert 58.2 ms to megaseconds in one multi-step calculation.

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 58.2 ms Find: Ms

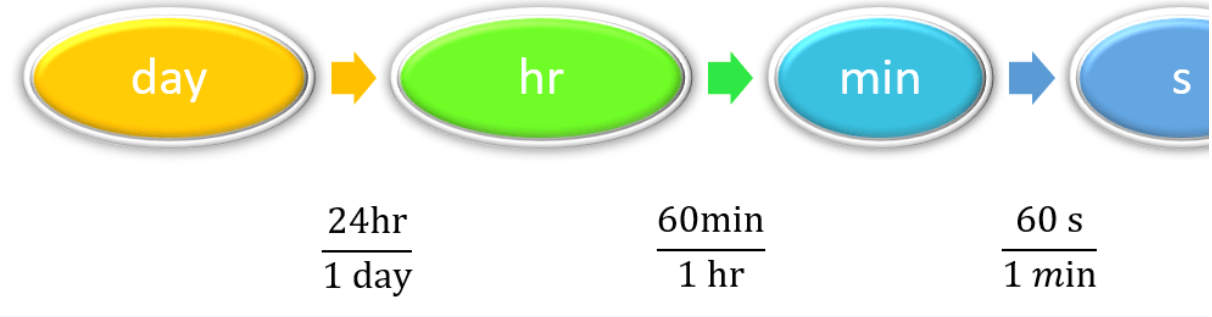
Steps for Problem Solving	Unit Conversion
List other known quantities	$1ms = 10^{-3}s$ $1Ms = 10^{6}s$
Prepare a concept map.	 $\frac{10^{-3}s}{1\ ms}$ $\frac{1\ Ms}{10^{6}\ s}$ Convert milliseconds to seconds to microseconds: use conversion factors 0.001 second per millisecond and 1 microsecond per 1 million seconds
Calculate.	$58.2\ \cancel{ms} \times \frac{10^{-3}\ \cancel{s}}{1\ \cancel{ms}} \times \frac{1\ \text{Ms}}{1,000,000\ \cancel{s}} = 0.0000000582\ \text{Ms}$ $= 5.82 \times 10^{-8}\ \text{Ms}$ Neither conversion factor affects the number of significant figures in the final answer.

Example 2.10.2: Unit Conversion

How many seconds are in a day?

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 1 day Find: s

Steps for Problem Solving	Unit Conversion
List other known quantities.	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds
Prepare a concept map.	 $\frac{24\text{hr}}{1\text{ day}}$ $\frac{60\text{min}}{1\text{ hr}}$ $\frac{60\text{ s}}{1\ min}$ Convert day to hour to minute to second: use conversion factors 24 hours per day, 60 minutes per hour, and 60 seconds per
Calculate.	$1\text{ d} \times \frac{24\text{ hr}}{1\text{ d}} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ s}}{1\text{ min}} = 86,400\text{ s}$

Exercise 2.10.1

Perform each conversion in one multi-step calculation.

a. 43.007 ng to kg
b. 1005 in to ft
c. 12 mi to km

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school. Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities. Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone's health.



A woman consulting with a pharmacist. (Public Domain; Rhoda Baer via National Cancer Institute, an agency that is part of the National Institutes of Health.)

Summary

In multi-step conversion problems, the previous unit is canceled for each step and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.

Answers

Exercise 2.10.1

$4.3007 \times 10^{-11} kg$

Exercise 2.10.1

$83.75\ ft$

Exercise 2.10.1

$19\ km$

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2.11: Units Raised to a Power

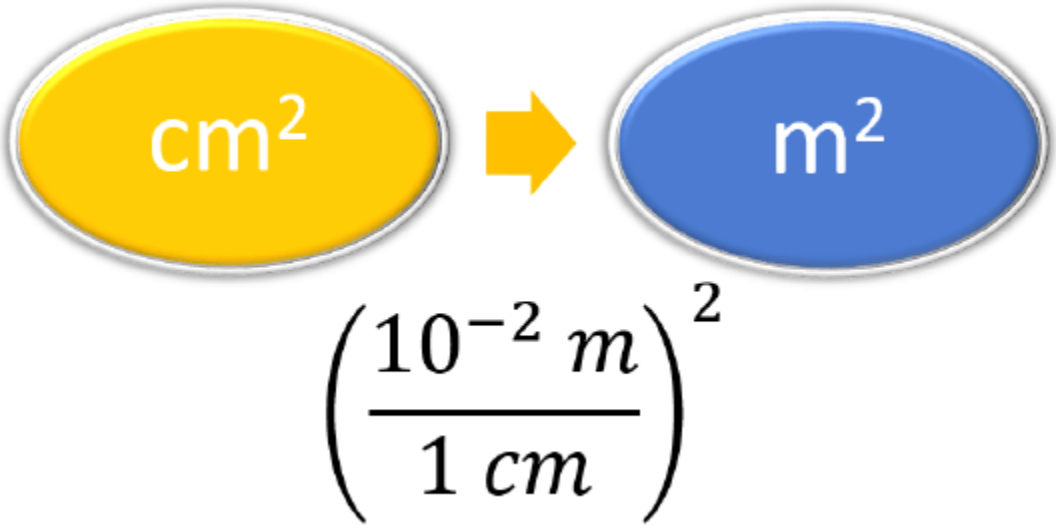
Learning Objectives

- To convert a value reported in one unit raised to a power of 10, to a corresponding value in a different unit raised to the same power of 10, using conversion factors.

Conversion factors for area and volume can also be produced by the dimensional analysis method. Just remember that if a quantity is raised to a power of 10, both the number and the unit must be raised to the same power of 10. For example, to convert $1500\ \text{cm}^2$ to m^2, we need to start with the relationship between centimeter and meter. We know that 1 cm = 10^{-2} m or 100 cm =1 m, but since we are given the quantity in 1500 cm^2, then we have to use the relationship:

$$1\,cm^2 = (10^{-2}\,m)^2 = 10^{-4}\,m^2$$

CONCEPT MAP



To convert centimeters squared to meters squared, use the conversion factor 0.01 meters per 1 centimeter, squared overall

CALCULATION

$$1500\ \cancel{\text{cm}}^2 \times \left(\frac{10^{-2}\ \text{m}}{1\ \cancel{\text{cm}}}\right)^2 = 0.15\ \text{m}^2$$

or

$$1500\ \cancel{\text{cm}}^2 \times \left(\frac{1\ \text{m}}{100\ \cancel{\text{cm}}}\right)^2 = 0.15\ \text{m}^2$$

or

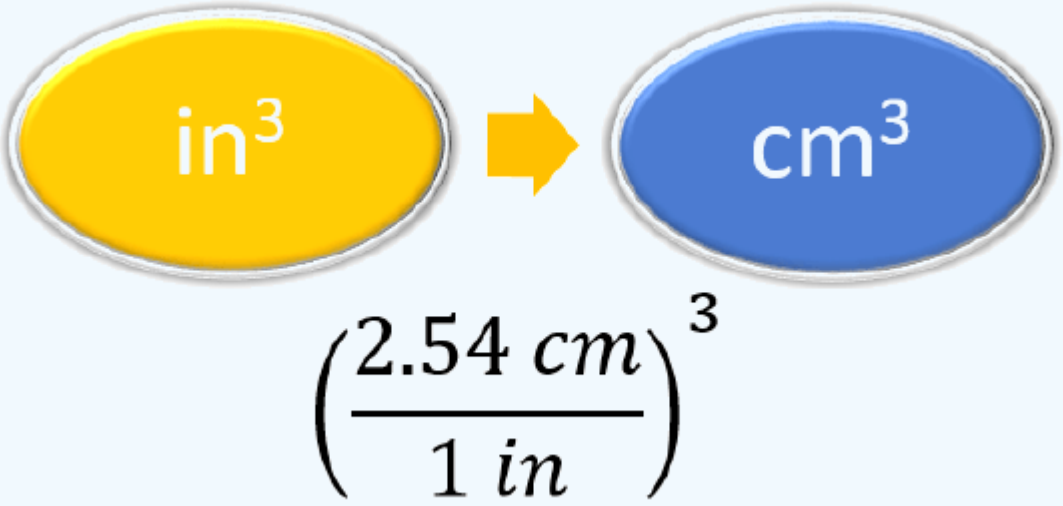
$$1500 \ \cancel{\text{cm}^2} \times \frac{1 \ \text{m}^2}{10,000 \ \cancel{\text{cm}^2}} = 0.15 \ \text{m}^2$$

Example 2.11.1: Volume of a Sphere

What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm^3)?

Solution

Solution for Example 2.8.1

Steps for Problem Solving	What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm^3)?
Identify the "given" information and what the problem is asking you to "find."	Given: radius = 4.30 in Find: cm^3 (volume)
Determine other known quantities.	Volume of a sphere: V = $\frac{4}{3} \times \pi \times r^3$ = $\frac{4}{3} \times 3.1416 \times (4.3\underline{0} in)^3$ = $33\underline{3}.04 in^3$
Prepare a concept map.	 To convert inches cubed to centimeters cubed, use conversion factor 2.54 centimeters per 1 inch, cubed overall
Calculate.	$33\underline{3}.04 \ \cancel{in^3} \left(\frac{2.54 cm}{1 \ \cancel{in}}\right)^3 = 5.46 \times 10^3 cm^3$

Steps for Problem Solving	What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm^3)?
Think about your result.	A centimeter is a smaller unit than an inch, so the answer in cubic centimeters is larger than the given value in cubic inches.

Exercise 2.11.1

Lake Tahoe has a surface area of 191 square miles. What is the area in square km (km^2)?

Answer available at the end of this section.

Answers

Exercise 2.11.1

495 km^2

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2.12: Derived Units



Figure 2.12.1 (Public Domain; User:Joegrimes/Wikipedia via Commons Wikimedia, Farm landly sander [commons.wikimedia.org])

How has farming evolved?

As farming becomes more expensive and less profitable (at least for small farms), farmers might decide to sell their land to builders who want to erect commercial or residential properties. In order to sell, an accurate property title is needed. The dimensions of the farm must be determined, and the acreage calculated from those dimensions.

Dimensional Analysis and Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. We have already discussed volume and energy as two examples of derived units. Some others are listed in the table below:

Derived SI Units

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Area	A	square meter	m^2	$\text{length} \times \text{width}$
Volume	V	cubic meter	m^3	$\text{length} \times \text{width} \times \text{height}$
Density	D	kilograms/cubic meter	$\mathrm{kg/m}^3$	\(\frac{\text{mass}}{\text{volume}}\0
Concentration	c	moles/liter	$\mathrm{mol/L}$	$\frac{\text{amount}}{\text{volume}}$
Speed (velocity)	v	meters/second	$\mathrm{m/s}$	$\frac{\text{length}}{\text{time}}$
Acceleration	a	meters/second/second	$\mathrm{m/s}^2$	$\frac{\text{speed}}{\text{time}}$
Force	F	newton	N	$\text{mass} \times \text{acceleration}$
Energy	E	joule	J	$\text{force} \times \text{length}$

Using dimensional analysis with derived units requires special care. When units are squared or cubed as with area or volume, the conversion factors themselves must also be squared. Shown below is the conversion factor for cubic centimeters and cubic meters.

$$\left(\frac{1\text{ m}}{100\text{ cm}}\right)^3 = \frac{1\text{ m}^3}{10^6\text{ cm}^3} = 1$$

Because a cube has 3 sides, each side is subject to the conversion of $1\ \text{m}$ to $100\ \text{cm}$. Since 100 cubed is equal to 1 million (10^6), there are $10^6\ \text{cm}^3$ in $1\ \text{m}^3$. Two convenient volume units are the liter (which is equal to a cubic decimeter) and the milliliter, which is equal to a cubic centimeter. The conversion factor would be:

$$\left(\frac{1\text{ dm}}{10\text{ cm}}\right)^3 = \frac{1\text{ dm}^3}{1000\text{ cm}^3} = 1$$

There are thus $1000\ \text{cm}^3$ in $1\ \text{dm}^3$, which is the same thing as saying there are $1000\ \text{mL}$ in $1\ \text{L}$.

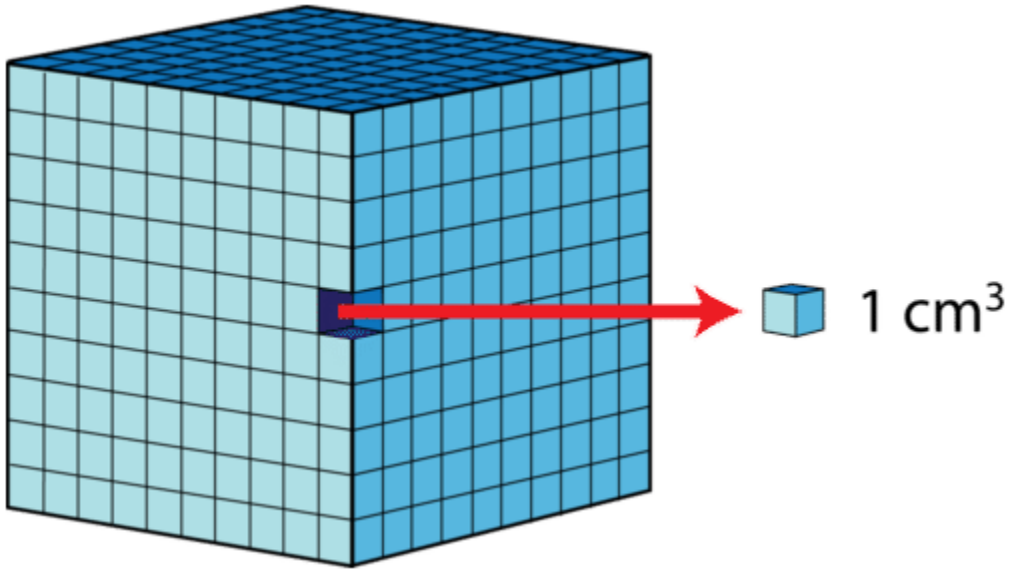


Figure 2.12.2: There are $1000\ \text{cm}^3$ in $1\ \text{dm}^3$. Since a cm^3 is equal to a mL and a dm^3 is equal to a L, we can say that there are $1000\ \text{mL}$ in $1\ \text{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 2.12.1

Convert $3.6 \times 10^8\ \text{mm}^3$ to mL.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1\ \text{m} = 1000\ \text{mm}$
- $1\ \text{mL} = 1\ \text{cm}^3$
- $1\ \text{m} = 100\ \text{cm}$

Unknown

This problem requires multiple steps and the technique for converting with derived units. Simply proceed one step at a time: mm^3 to m^3 to $\text{cm}^3 = \text{mL}$.

Step 2: Calculate.

$$3.6\text{ mm}^3 \times \left(\frac{1\text{ m}}{1000\text{ mm}}\right)^3 \times \left(\frac{100\text{ cm}}{1\text{ m}}\right)^3 \times \frac{1\text{ mL}}{1\text{ cm}^3} = 0.0036\text{ mL}$$

Numerically, the steps are to divide 3.6 by 10^9, followed by multiplying by 10^6. You may find that you can shorten the problem by a step by first determining the conversion factor from mm to cm, and using that instead of first converting to m. There are $10\ \text{mm}$ in $1\ \text{cm}$.

$$3.6\ \text{mm}^3 \times \left(\frac{1\ \text{cm}}{10\ \text{mm}}\right)^3 \times \frac{1\ \text{mL}}{1\ \text{cm}^3} = 0.0036\ \text{mL}$$

In this case, 3.6/1000 gives the same result of 0.0036.

Step 3: Think about your result.

Cubic millimeters are much smaller than cubic centimeters, so the final answer is much less than the original number of mm^3.

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Summary

- A derived unit is a unit that results from a mathematical combination of SI base units.
- Calculations involving derived units follow the same principles as other unit conversion calculations.

Review

1. What is a derived unit?
2. Convert 0.00722 km^3 to m^3
3. Convert 129 cm^3 to L
4. Convert 4.9 × 10^5 μm^3 to mm^3

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3: THE EVOLUTION OF ATOMIC THEORY

3.1: Cutting Aluminum until you get Atoms

3.2: Indivisible - The Atomic Theory

3.3: Law of Conservation of Mass

3.4: Law of Definite Proportions

3.5: Law of Definite Proportions - Calculating Mass Percent

3.6: The Nuclear Atom

3.7: The Properties of Protons, Neutrons, and Electrons

3.8: Elements- Defined by Their Number of Protons

3.9: Isotopes - When the Number of Neutrons Varies

3.10: Atomic Mass - The Average Mass of an Element's Atoms

3.11: Looking for Patterns - The Periodic Table

3.12: Ions - Losing and Gaining Electrons

3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character

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3.1: Cutting Aluminum until you get Atoms

Take some aluminum foil. Cut it in half. Now there are two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil. It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can this exercise be taken, at least in theory? Can one continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC by Democritus and other ancient Greek philosophers (Figure 3.1.1).

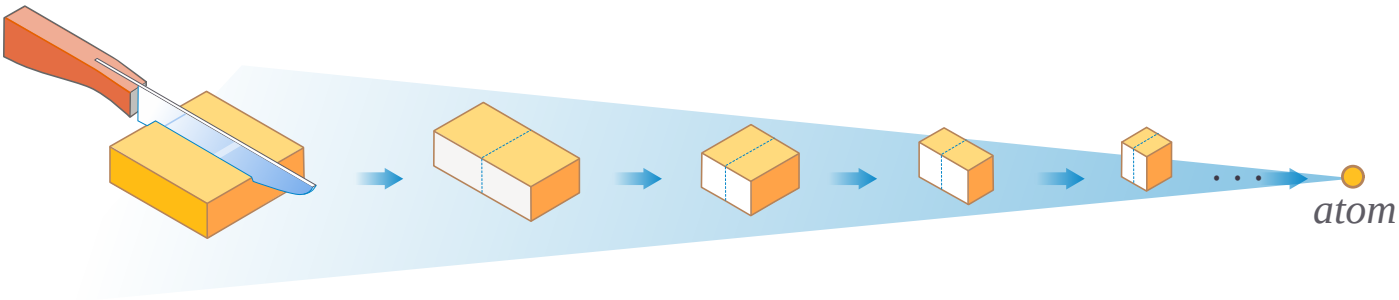


Figure 3.1.1: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya via LibreTexts)

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called diatomic molecules. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

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3.2: Indivisible - The Atomic Theory

Learning Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

You learned earlier that all matter in the universe is made out of tiny building blocks called atoms. All modern scientists accept the concept of the atom, but when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus and the Greek Philosophers

About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge, entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable "thing." One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon—tiny, indivisible, solid objects—make up all matter in the universe.



Figure 3.2.1: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

Democritus then reasoned that changes occur when the many atomos in an object were reconnected or recombined in different ways. Democritus even extended this theory, suggesting that there were different varieties of atomos with different shapes, sizes, and masses. He thought, however, that shape, size, and mass were the only properties differentiating the different types of atomos. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves, but rather, resulted from the different ways in which the atomos were combined and connected to one another.

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of

discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they are known today) was fully appreciated.

Greeks: "All Thought and No Action"

Greek philosophers were "all thought and no action" and did not feel the need to test their theories with reality. In contrast, Dalton's efforts were based on experimentation and testing ideas against reality.

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure $3.2.1$), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory were as follows:

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.

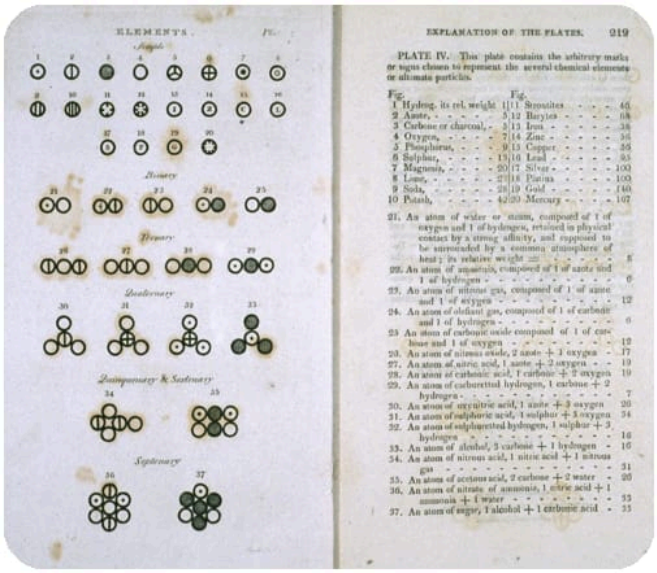


Figure 3.2.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video 3.2.1, control of individual atoms can be use used create animations.

YouTube Video

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Video 3.2.1: A Boy And His Atom: The World's Smallest Movie.

A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos." However, other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.
- Dalton's Atomic Theory is the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom. The general tenets of this theory are:
  - All matter is composed of extremely small particles called atoms.
  - Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
  - Atoms cannot be subdivided, created, or destroyed.
  - Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
  - In chemical reactions, atoms are combined, separated, or rearranged.

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3.3: Law of Conservation of Mass

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Look at Figure 3.3.1 below. It shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose you had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose you had been able to measure the oxygen used by the fire and the gases produced by the fire. What would you find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

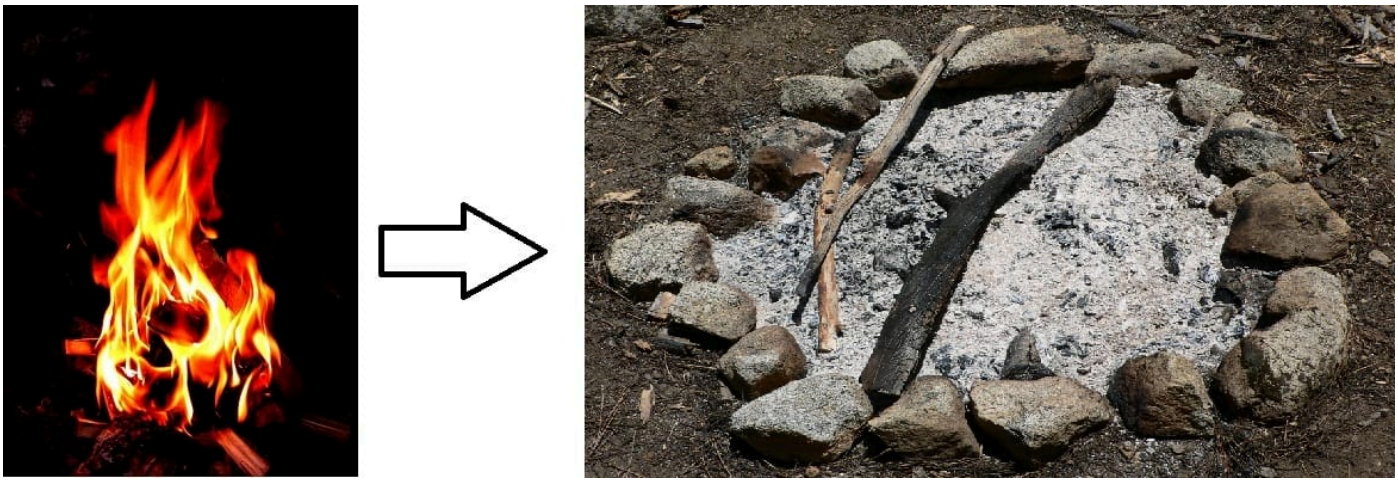


Figure 3.3.1: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). (CC BY-SA 2.5; Einar Helland Berger for fire and Walter Siegmund for ash).

Law of Conservation of Mass

The law of conservation of mass was created in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases equals the original mass of the charcoal and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. A reactant is the chemical reaction of two or more elements to make a new substance, and a product is the substance that is formed as the result of a chemical reaction (Video 3.3.1). Matter and its corresponding mass may not be able to be created or destroyed, but can change forms to other substances like liquids, gases, and solids.

YouTube Video

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Video 3.3.1: This is a nice little demonstration showing the Conservation of Mass in action.

If you witness a 300 kg tree burn to the ground, there are only ashes left after the burn, and all of them together weigh 10 kg. It may make you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only thing left that you can see is the 10 kg of ash. If you know the law of conservation of mass, then you know that the other 290 kg

has to go somewhere, because it has to equal the mass of the tree before it burnt down.

Example 3.3.1

If heating 10.0 grams of calcium carbonate ($CaCO_3$) produces 4.4 g of carbon dioxide (CO_2) and 5.6 g of calcium oxide (CaO), show that these observations are in agreement with the law of conservation of mass.

Solution

$$\text{Mass of the reactants} = \text{Mass of the products}$$
$$10.0\,\text{g of } CaCO_3 = 4.4\,\text{g of } CO_2 + 5.6\,\text{g of CaO}$$
$$10.0\,\text{g of reactant} = 10.0\,\text{g of products}$$

Because the mass of the reactant is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

Exercise 3.3.1

Potassium hydroxide (KOH) readily reacts with carbon dioxide (CO_2) to produce potassium carbonate (K_2CO_3) and water (H_2O). How many grams of potassium carbonate are produced if 224.4 g of KOH reacts with 88.0 g of CO_2? The reaction also produces 36.0 g of water.

Answer available at the end of this section.

The Law is also applicable to both chemical and physical changes. For example, if you have an ice cube that melts into a liquid and you heat that liquid up, it becomes a gas. It will appear to have disappeared, but is still there.

Summary

- Burning and other changes in matter do not destroy matter.
- The mass of matter is always the same before and after the changes occur.
- The law of conservation of mass states that matter cannot be created or destroyed.

Contributions & Attributions

- Binod Shrestha (University of Lorraine)

Answers

Exercise 3.3.1

276.4 g of potassium carbonate

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3.4: Law of Definite Proportions



Figure 3.4.1 (Public Domain; Staff Sgt. Christopher Flurry, U.S. Marine Corps via Commons Wikimedia, File:USMC-110609-M-FK922-004.jpg [commons.wikimedia.org])

We use electricity for many purposes—from cooking, to powering our televisions, to charging our cell phones. Wherever we travel in the United States, we want electricity to be available. What we also want (although we usually don't think about it) is for the electricity supply to be the same wherever we go. We want the same voltage (110 volts for the U.S.) to come from the outlet to whatever we plug in. If the voltage is less, the system will not work. If it is more, the equipment will be damaged. We want a definite amount of voltage—no more and no less.

Law of Definite Proportions

The discovery that mass was always conserved in chemical reactions was soon followed by the **law of definite proportions**, which states that a given chemical compound always contains the same elements in the exact same proportions by mass. As an example, any sample of pure water contains 11.19% hydrogen and 88.81% oxygen by mass. It does not matter where the sample of water came from or how it was prepared. Its composition, like that of every other compound, is fixed.



Figure 3.4.2: Water. (CC BY 2.0; Jenn Durfey via Flickr Photos, Water Faucet [www.flickr.com])

Another example is carbon dioxide. This gas is produced from a variety of reactions, often by the burning of materials. The structure of the gas consists of one atom of carbon and two atoms of oxygen. Carbon dioxide production is of interest in many areas, from the amount we breathe to the amount of the gas produced by burning wood or fossil fuels. By knowing the exact

composition of carbon dioxide, we can make predictions as to the effects of different chemical processes.



Figure 3.4.3: Carbon dioxide is produced during the burning process. (Public Domain; Patsy Lynch, FEMA via Commons Wikimedia, FEMA Photograph by Patsy Lynch taken on 04-17-2006 in Missouri [commons.wikimedia.org])

Summary

- The law of definite proportions states that a given chemical compound always contains the same elements in the exact same proportions by mass.

Review

1. State the law of definite proportions.
2. Will the composition of water molecules vary depending on its source?
3. Why is this law important?

Explore More

Use the resource below to answer the questions that follow.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

1. How many hydrogen atoms are there in a molecule of water?
2. How many oxygen atoms are there in a molecule of water?
3. What is the mass ratio of hydrogen to oxygen in a molecule of water?
4. Will the mass ratio of hydrogen to oxygen change depending on the size of the sample? For example if the sample size was 10 L or 10,000 L.

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3.5: Law of Definite Proportions - Calculating Mass Percent

Learning Objectives

- Determine percent composition of each element in a compound based on mass.

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter reveals that one serving size is considered to be $32\ \text{g}$. The label also gives the masses of various types of compounds that are present in each serving. One serving contains $7\ \text{g}$ of protein, $15\ \text{g}$ of fat, and $3\ \text{g}$ of sugar. By calculating the fraction of protein, fat, or sugar in one serving size of peanut butter and converting to percent values, we can determine the composition of peanut butter on a percent by mass basis.

Percent Composition

Chemists often need to know what elements are present in a compound and in what percentage. The **percent composition** is the percent by mass of each element in a compound. It is calculated in a similar way to that of the composition of the peanut butter.

$$\%\ \text{by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

Example 3.5.1: Percent Composition from Mass Data

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a $20.00\ \text{g}$ sample of the compound is decomposed, $16.07\ \text{g}$ of zinc remains. Determine the percent composition of the compound.

Solution

Solutions to Example 6.6.1

Steps for Problem Solving	When a $20.00\ \text{g}$ sample of the zinc-and-oxygen compound is decomposed, $16.07\ \text{g}$ of zinc remains. Determ percent composition of the compound.
Identify the "given" information and what the problem is asking you to "find."	Given : Mass of compound = 20.00 g Mass of Zn = 16.07 g Find: % Composition (% Zn and %O)
List other known quantities.	Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compour the percent by mass. Mass of oxygen = 20.00 g - 16.07 g = 3.93 g O

Steps for Problem Solving	When a $20.00\ \text{g}$ sample of the zinc-and-oxygen compound is decomposed, $16.07\ \text{g}$ of zinc remains. Deterr percent composition of the compound.
Cancel units and calculate.	$$\%\ \text{Zn} = \frac{16.07\ \text{g Zn}}{20.00\ \text{g}} \times 100\% = 80.35\%\ \text{Zn}$$ $$\%\ \text{O} = \frac{3.93\ \text{g O}}{20.00\ \text{g}} \times 100\% = 19.65\%\ \text{O}$$ Calculate the percent by mass of each element by dividing the mass of that element by the mass of the compc multiplying by 100%.
Think about your result.	The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound zinc.

Exercise 3.5.1

Sulfuric acid, H_2SO_4 is a very useful chemical in industrial processes. If 196.0 g of sulfuric acid contained 64.0g oxygen and 4.0 g of hydrogen, what is the percent composition of the compound?

Answer available at the end of this section.

Summary

- Processes are described for calculating the percent composition of a compound based on mass.

Answers

Exercise 3.5.1

2.04% H, 32.65% S, and 65.3% O

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3.6: The Nuclear Atom

Learning Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" model of the atom and the evidence for it.
- Draw a diagram of Thomson's "plum pudding" model of the atom and explain why it has this name.
- Describe Rutherford's gold foil experiment and explain how this experiment altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.



Figure 3.6.1: J. J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom.

Electrons and Plums

The electron was discovered by J. J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This was a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or similar to a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges, and the electron would be pulled back more

toward the center region of the atom.

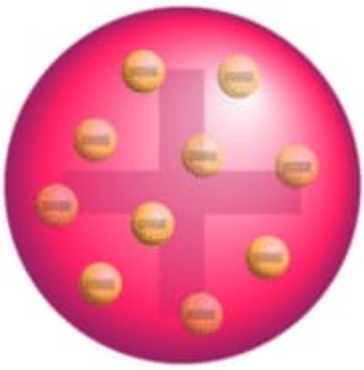


Figure 3.6.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871-1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Atoms and Gold

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

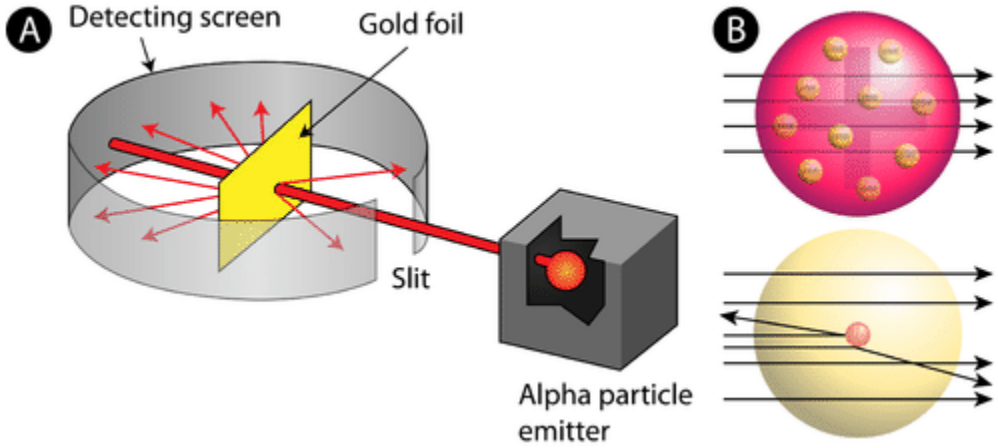


Figure 3.6.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the

positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupy the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Atomic Nucleus

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. The figure below shows these parts of the atom.

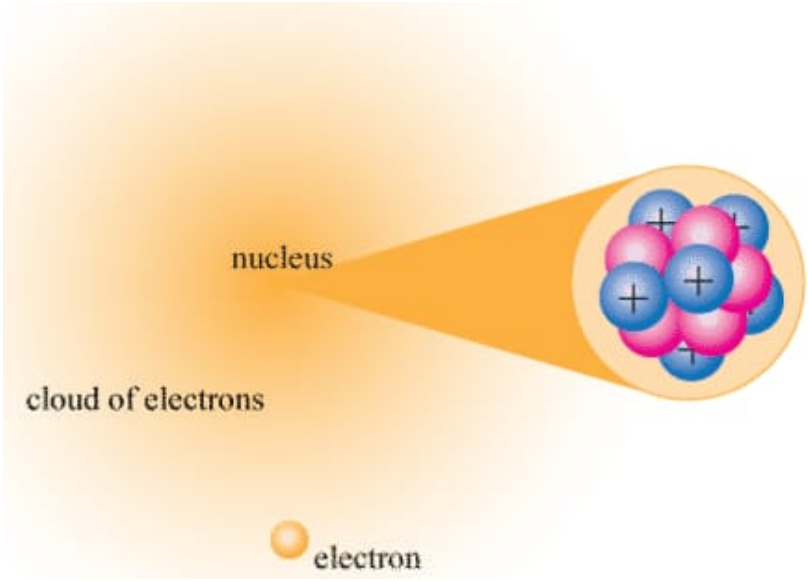


Figure 3.6.4: The nuclear atom.

The nucleus of the atom is extremely small. Its radius is only about 1/100,000 of the total radius of the atom. Electrons have virtually no mass, but protons and neutrons have a lot of mass for their size. As a result, the nucleus has virtually all the mass of an atom. Given its great mass and tiny size, the nucleus is very dense. If an object the size of a penny had the same density as the nucleus of an atom, its mass would be greater than 30 million tons!

Holding it all Together

Particles with opposite electric charges attract each other. This explains why negative electrons orbit the positive nucleus. Particles with the same electric charge repel each other. This means that the positive protons in the nucleus push apart from one another. So why doesn't the nucleus fly apart? An even stronger force—called the **strong nuclear force**—holds protons and neutrons together in the nucleus.

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

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3.7: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it was not entirely accurate, because contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to us: protons and neutrons. We already learned that J. J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus, called a **proton**. There is also a third subatomic particle, known as a neutron.

Electrons

Electrons are one of three main types of particles that make up atoms. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particle called leptons. All leptons have an electric charge of -1 or 0. Electrons are extremely small. The mass of an electron is only about 1/2000 the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of -1, which is equal but opposite to the charge of a proton, which is $+1$. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract one another, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. However, this is an incorrect perspective, as quantum mechanics demonstrates that electrons are more complicated.



Figure 3.7.1: Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

A proton is one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one $(+1)$ and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom.

Neutrons

Atoms of all elements—except for most atoms of hydrogen—have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge—they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they are bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 3.7.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	–1	outside the nucleus
neutron	n^0	1	1	0	inside the nucleus

Table $3.7.1$ gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they are neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons

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(approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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3.8: Elements- Defined by Their Number of Protons

Learning Objectives

- Define atomic number.
- Define mass number.
- Determine the number of protons, neutrons, and electrons in an atom.

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as H for hydrogen, C for carbon, or He for helium. However, it would more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.



Figure 3.8.1: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 3.8.1). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (Z). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 3.8.1: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many

electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, 79 electrons as well.

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table $3.8.1$). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table $3.8.1$).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units ($2\ \text{amu}$ for the protons + $2\ \text{amu}$ for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called isotopes and are discussed in detail later on).

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

An atom's mass number is very easy to calculate, provided that you know the number of protons and neutrons in an atom.

Example 4.5.1

What is the mass number of an atom of helium that contains 2 neutrons?

Solution

$(\text{number of protons}) = 2$ (Remember that an atom of helium always has 2 protons.)

$(\text{number of neutrons}) = 2$

$\text{mass number} = (\text{number of protons}) + (\text{number of neutrons})$

$$\text{mass number} = 2 + 2 = 4$$

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are O for oxygen, Zn for zinc, and Fe for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table $3.8.2$.

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Table 3.8.2: Symbols and Latin Names for Elements

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

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3.9: Isotopes - When the Number of Neutrons Varies

Learning Objectives

- Explain what isotopes are and how an isotope affects an element's atomic mass.
- Determine the number of protons, electrons, and neutrons of an element with a given mass number.

All atoms of the same element have the same number of protons, but some may have different numbers of neutrons. For example, all carbon atoms have six protons, and most have six neutrons as well. But some carbon atoms have seven or eight neutrons instead of the usual six. Atoms of the same element that differ in their numbers of neutrons are called **isotopes**. Many isotopes occur naturally. Usually one or two isotopes of an element are the most stable and common. Different isotopes of an element generally have the same physical and chemical properties because they have the same numbers of protons and electrons.

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 3.9.1. Most hydrogen atoms have just one proton, one electron, and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

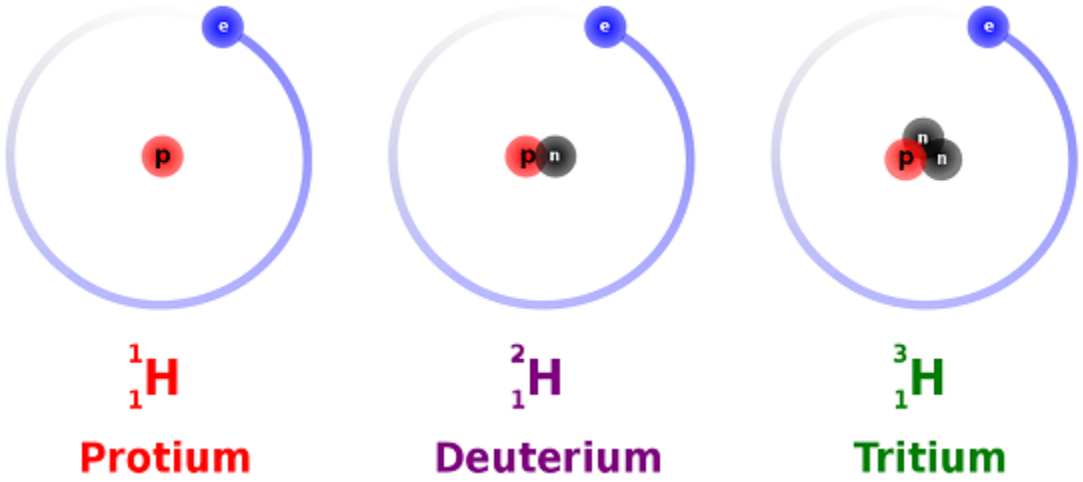


Figure 3.9.1: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

For most elements other than hydrogen, isotopes are named for their mass number. For example, carbon atoms with the usual 6 neutrons have a mass number of 12 (6 protons + 6 neutrons = 12), so they are called carbon-12. Carbon atoms with 7 neutrons have an atomic mass of 13 (6 protons + 7 neutrons = 13). These atoms are the isotope called carbon-13.

Example 3.9.1: Lithium Isotopes

a. What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons?
b. What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons?

Solution

A lithium atom contains 3 protons in its nucleus irrespective of the number of neutrons or electrons.

a.

$$\begin{aligned} \text{atomic number} &= (\text{number of protons}) = 3 \\ &\quad (\text{number of neutrons}) = 3 \end{aligned}$$

$$\begin{aligned} \text{mass number} &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number} &= 3 + 3 \\ &= 6 \end{aligned}$$

b.

$$\begin{aligned} \text{atomic number} &= (\text{number of protons}) = 3 \\ &\quad (\text{number of neutrons}) = 4 \end{aligned}$$

$$\begin{aligned} \text{mass number} &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number} &= 3 + 4 \\ &= 7 \end{aligned}$$

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons, and as an isotope with 4 neutrons, but it doesn't exist as an isotope with 2 neutrons or as an isotope with 5 neutrons.

Stability of Isotopes

Atoms need a certain ratio of neutrons to protons to have a stable nucleus. Having too many or too few neutrons relative to protons results in an unstable, or radioactive, nucleus that will sooner or later break down to a more stable form. This process is called radioactive decay. Many isotopes have radioactive nuclei, and these isotopes are referred to as radioisotopes. When they decay, they release particles that may be harmful. This is why radioactive isotopes are dangerous and why working with them requires special suits for protection. The isotope of carbon known as carbon-14 is an example of a radioisotope. In contrast, the carbon isotopes called carbon-12 and carbon-13 are stable.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundance"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element—chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below:



Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

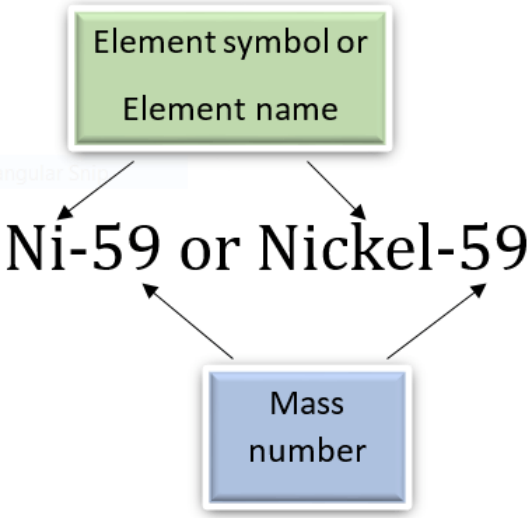
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\text{Ni}$$

$$^{238}_{92}\text{U}$$

In the nickel nucleus represented above, the atomic number 28 indicates that the nucleus contains 28 protons, and therefore, it must contain 31 neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons, as all uranium nuclei do; and this particular uranium nucleus has 146 neutrons.

Another way of representing isotopes is by adding a hyphen and the mass number to the chemical name or symbol. Thus the two nuclei would be Nickel-59 or Ni-59 and Uranium-238 or U-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) are given to the side of the name.



Example 3.9.2: Potassium-40

How many protons, electrons, and neutrons are in an atom of $^{40}_{19}\text{K}$?

Solution

$$\text{atomic number} = (\text{number of protons}) = 19$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 19$$

The mass number, 40, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 40 - 19 = 21.$$

Example 3.9.3: Zinc-65

How many protons, electrons, and neutrons are in an atom of zinc-65?

Solution

$$\text{number of protons} = 30$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 30$$

The mass number, 65, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 65 - 30 = 35$$

Exercise 3.9.3

How many protons, electrons, and neutrons are in each atom?

a. $^{60}_{27}\text{Co}$
b. Na-24
c. $^{45}_{20}\text{Ca}$
d. Sr-90

Answer a: available at the end of this section.

Summary

- The number of protons is always the same in atoms of the same element.
- The number of neutrons can be different, even in atoms of the same element.
- Atoms of the same element that contain the same number of protons, but different numbers of neutrons, are known as **isotopes**.
- Isotopes of any given element all contain the same number of protons, so they have the same atomic number (for example, the atomic number of helium is always 2).
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

Answers

Exercise 3.9.3

27 protons, 27 electrons, 33 neutrons
11 protons, 11 electrons, 13 neutrons
20 protons, 20 electrons, 25 neutrons
38 protons, 38 electrons, 52 neutrons

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3.10: Atomic Mass - The Average Mass of an Element's Atoms

Learning Objectives

- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units.

Calculating Atomic Mass

You can calculate the atomic mass (or average mass) of an element provided you know the **relative abundance** (the fraction of an element that is a given isotope), the element's naturally occurring isotopes, and the masses of those different isotopes. We can calculate this by the following equation:

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots$$

Look carefully to see how this equation is used in the following examples.

Example 3.10.1: Boron Isotopes

Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and mass of $10\ \text{amu}$. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of $11\ \text{amu}$. What is the atomic mass of boron?

Solution

Boron has two isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.20$ (Write all percentages as decimals), $\text{mass}_1 = 10$
- Isotope 2: $\%_2 = 0.80$, $\text{mass}_2 = 11$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)\,(10) + (0.80)\,(11)$$

$$\text{Atomic mass} = 10.8\,\text{amu}$$

The mass of an average boron atom, and thus boron's atomic mass, is $10.8\ \text{amu}$.

Example 3.10.2: Neon Isotopes

Neon has three naturally occurring isotopes. In a sample of neon, 90.92% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of $19.99\ \mathrm{amu}$. Another 0.3% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of $20.99\ \mathrm{amu}$. The final 8.85% of the atoms are Ne-22, which is an isotope of neon with 12 neutrons and a mass of $21.99\ \mathrm{amu}$. What is the atomic mass of neon?

Solution

Neon has three isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.9092$ (write all percentages as decimals), $\text{mass}_1 = 19.99$
- Isotope 2: $\%_2 = 0.003$, $\text{mass}_2 = 20.99$
- Isotope 3: $\%_3 = 0.0885$, $\text{mass}_3 = 21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)\,(19.99) + (0.003)\,(20.99) + (0.0885)\,(21.99)$$

$$\text{Atomic mass} = 20.17\,\text{amu}$$

The mass of an average neon atom is $20.17\ \mathrm{amu}$

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol B) is 10.8, which is what we calculated in Example $3.10.1$, and the atomic mass of neon (symbol Ne) is 20.8, which is what we calculated in Example $3.10.2$. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

Exercise 3.10.1

Chlorine has two naturally occurring isotopes. In a sample of chlorine, 75.77% of the atoms are Cl-35, with a mass of $34.97\ \mathrm{amu}$. Another 24.23% of the atoms are Cl-37, with a mass of $36.97\ \mathrm{amu}$. What is the atomic mass of chlorine?

Answer available at the end of this section.

Summary

- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- An element's atomic mass can be calculated provided the relative abundance of the element's naturally occurring isotopes and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.

Answers

Exercise $3.10.1$

35.45 amu

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3.11: Looking for Patterns - The Periodic Table

Learning Objectives

- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

> Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.

The periodic table is one of the cornerstones of chemistry because it organizes all of the known elements on the basis of their chemical properties. A modern version is shown in Figure 3.11.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

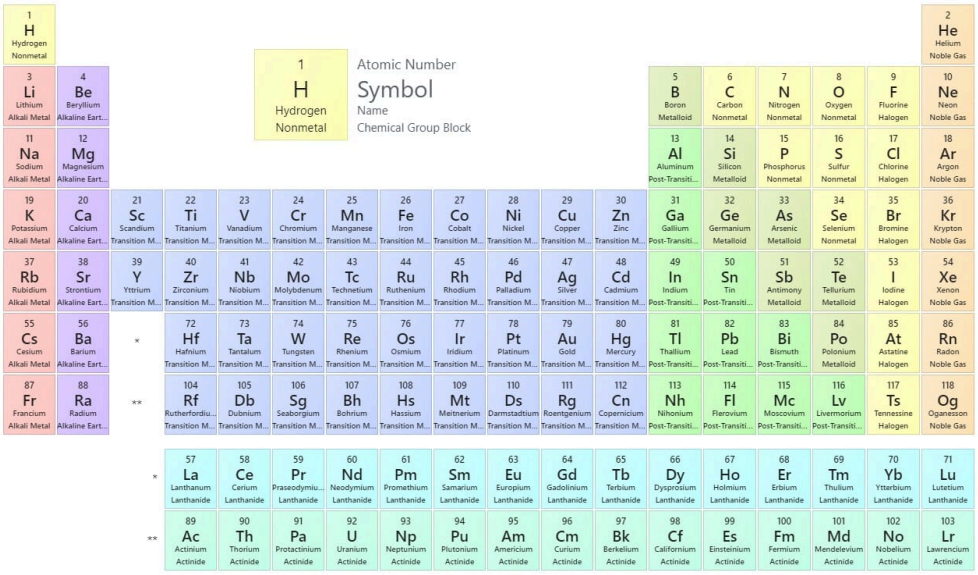


Figure 3.11.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

The word *halogen* comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.

To Your Health: Radon

Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building’s inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

Each row of elements on the periodic table is called a period. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 3.11.2. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 3.11.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

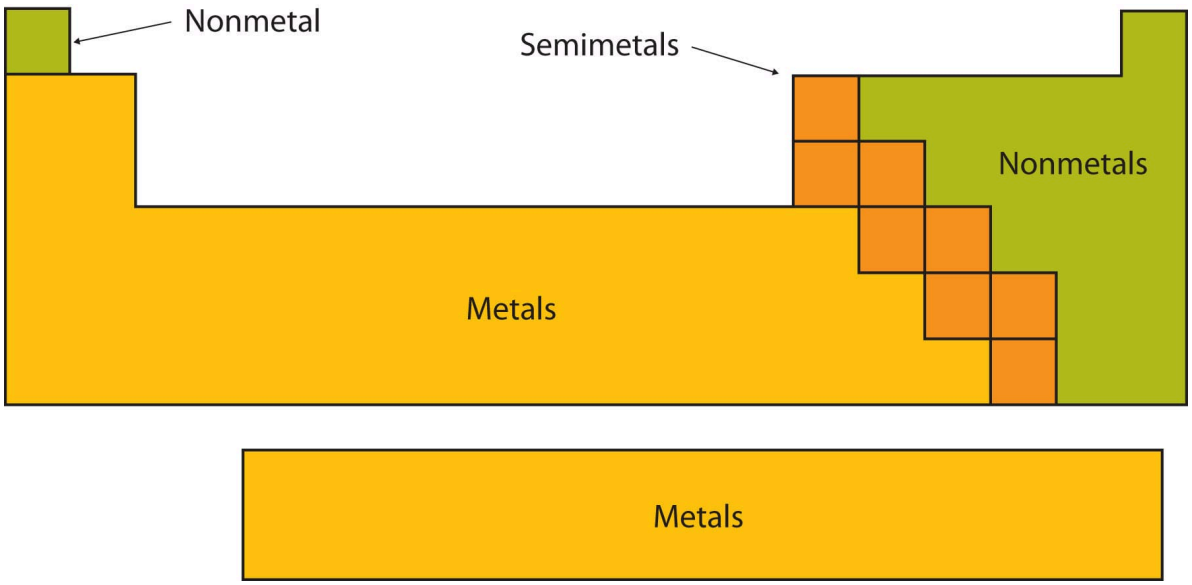


Figure 3.11.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

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Example 3.11.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

a. Se
b. Mg
c. Ge

Solution

a. In Figure 3.11.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

Exercise 3.11.1

Based on its location in the periodic table, do you expect indium to be a nonmetal, a metal, or a metalloid?

Answer available at the end of this section.

Another way to categorize the elements of the periodic table is shown in Figure 3.11.3. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

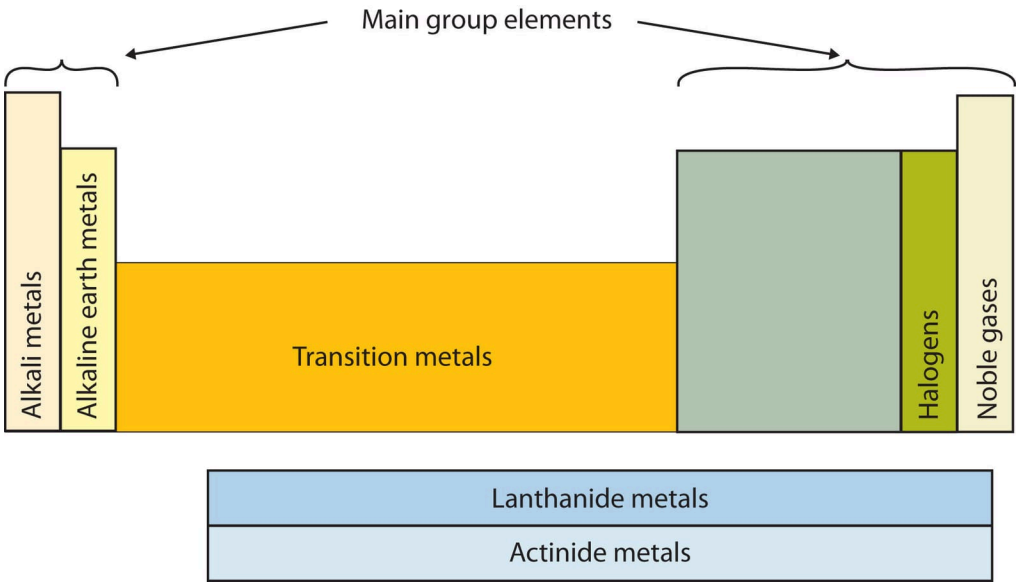


Figure 3.11.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

Group 2: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt forming," which reflects that all of the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

Example 3.11.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

a. Lithium is an alkali metal (Group 1)
b. Argon is a noble gas (Group 18)
c. Chlorine is a halogen (Group 17)

Exercise 3.11.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a: available at the end of this section.

Example 3.11.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. Li
b. Ar
c. Am
d. Fe

Solution

a. Lithium is a metal.
b. Argon is a non metal.
c. Americium is an inner transition metal.
d. Iron is a transition metal.

Exercise 3.11.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. F
b. U
c. Cu

Answer a: available at the end of this section.

Summary

The periodic table is an arrangement of the elements in order of increasing atomic number. Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods. Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.

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The elements can be broadly divided into metals, nonmetals, and semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
2. Sisler, Harry H. Electronic structure, properties, and the periodic law. New york; Reinhold publishing corporation, 1963.
3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

Contributions & Attributions

Answers

Exercise 3.11.1

Indium is a metal.

Exercise 3.11.2**: Groups**

Fluorine is a halogen (Group 17).

Calcium is a alkaline earth metal (Group 2).

Krypton is a noble gas (Group 18).

Exercise 3.11.3**: Classification of Elements**

Fluorine is a nonmetal.

Uranium is a metal (and a inner transition metal too).

Copper is a metal (and a transition metal too).

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3.12: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Atoms tend to lose, gain, or share some valance electrons, making bonds to acquire the electron configuration of the nearest noble gas, i.e., usually eight valence electrons for the main group elements, called an octet. Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while others lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called **cations**. Most metals become cations when they make ionic compounds.

Cations

A neutral sodium atom is likely to achieve an octet in its outermost shell by losing its one valence electron.

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-}$$

The cation produced in this way, $\mathrm{Na^+}$, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure $3.12.1$ is a graphical depiction of this process.

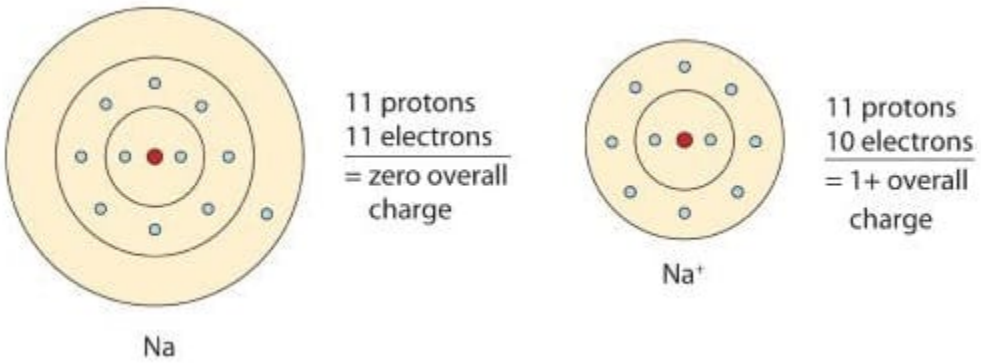


Figure $3.12.1$: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called **anions**. Most nonmetals become anions when they make ionic compounds.

A neutral chlorine atom has seven electrons in its outermost shell. Only one more electron is needed to achieve an octet in chlorine's valence shell. (In table salt, this electron comes from the sodium atom.)

$$\mathrm{e^-} + \mathrm{Cl} \longrightarrow \mathrm{Cl^-}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again,

the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure 3.12.2 is a graphical depiction of this process.

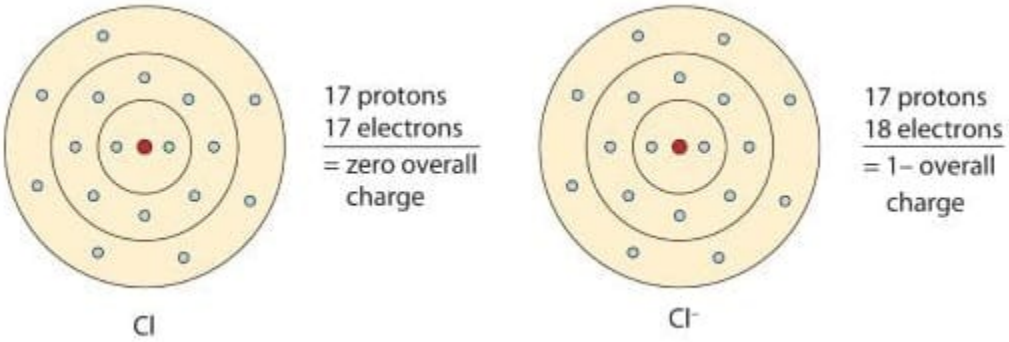


Figure 3.12.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1– charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall charge.

The names for positive and negative ions are pronounced CAT-eye-ons and ANN-eye-ons, respectively.

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1– charge. Figure 3.12.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written Ba^{2+}, not Ba^{+2}.

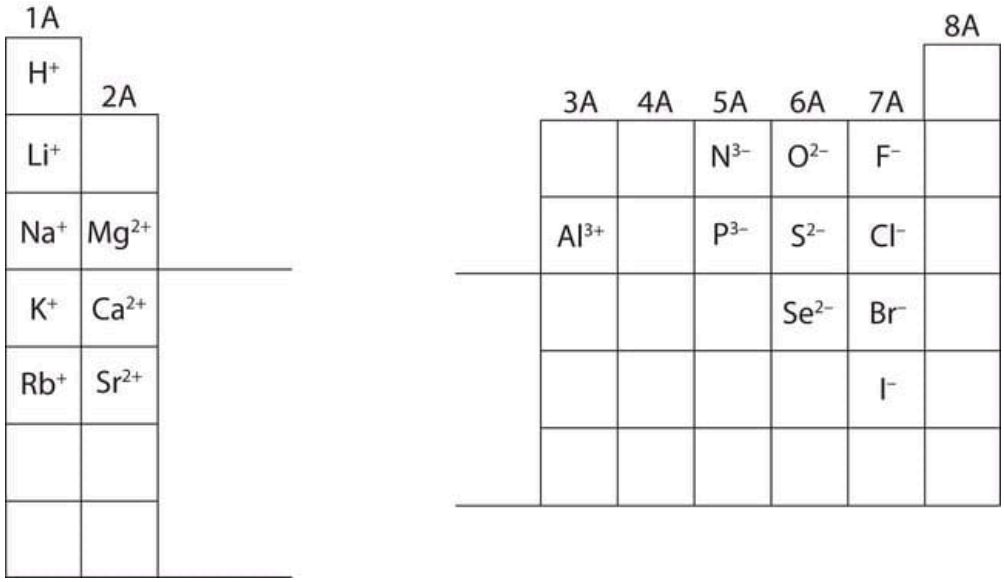


Figure 3.12.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

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3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character

Learning Objectives

- Be able to state how certain properties of atoms vary based on their relative position on the periodic table.

One of the reasons the periodic table is so useful is because its structure allows us to qualitatively determine how some properties of the elements vary versus their position on the periodic table. The variations of properties versus positions on the periodic table are called **periodic trends**. There is no other tool in science that allows us to judge relative properties of a class of objects like this, which makes the periodic table a very useful tool. Many periodic trends are general. There may be a few points where an opposite trend is seen, but there is an overall trend when considered across a whole row or down a whole column of the periodic table.

The first periodic trend we will consider is atomic radius. The **atomic radius** is an indication of the size of an atom. Although the concept of a definite radius of an atom is a bit fuzzy, atoms behave as if they have a certain radius. Such radii can be estimated from various experimental techniques, such as the x-ray crystallography of crystals.

As you go down a column of the periodic table, the atomic radii increase. This is because the valence electron shell is getting larger and there is a larger principal quantum number, so the valence shell lies physically farther away from the nucleus. This trend can be summarized as follows:

$$as \downarrow PT, atomic\ radius \uparrow$$

where PT stands for periodic table. Going across a row on the periodic table, left to right, the trend is different. Even though the valence shell maintains the same principal quantum number, the number of protons—and hence the nuclear charge—is increasing as you go across the row. The increasing positive charge casts a tighter grip on the valence electrons, so as you go across the periodic table, the atomic radii decrease. Again, we can summarize this trend as follows:

$$as \rightarrow PT, atomic\ radius \downarrow$$

Figure $3.13.1$ shows spheres representing the atoms of the *s* and *p* blocks from the periodic table to scale, showing the two trends for the atomic radius.

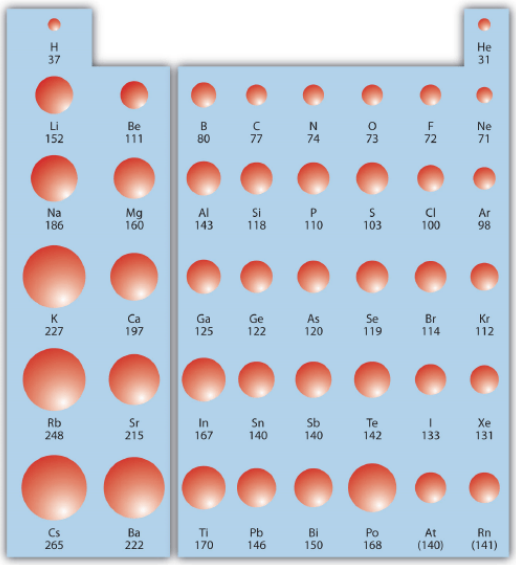


Figure 3.13.1: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

Example 3.13.1: Atomic Radii

Referring only to a periodic table and not to Figure 3.13.1, which atom is larger in each pair?

a. Si or S
b. S or Te

Solution

a. Si is to the left of S on the periodic table; it is larger because as you go across the row, the atoms get smaller.
b. S is above Te on the periodic table; Te is larger because as you go down the column, the atoms get larger.

Exercise 3.13.1: Atomic Radii

Referring only to a periodic table and not to Figure 3.13.1, which atom is smaller, Ca or Br?

Answer available at the end of this section.

Ionization energy (IE) is the amount of energy required to remove an electron from an atom in the gas phase:

$$A(g) \rightarrow A^+(g) + e^- \qquad \Delta H \equiv IE$$

IE is usually expressed in kJ/mol of atoms. It is always positive because the removal of an electron always requires that energy be put in (i.e., it is endothermic). IE also shows periodic trends. As you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE decreases) because the valence electron is farther away from the nucleus. Thus,

$$as \downarrow PT,\ IE \downarrow$$

However, as you go across the periodic table and the electrons get drawn closer in, it takes more energy to remove an electron; as a result, IE increases:

$$as \rightarrow PT,\ IE \uparrow$$

Figure 3.13.2 shows values of IE versus position on the periodic table. Again, the trend is not absolute, but the general trends going across and down the periodic table should be obvious.

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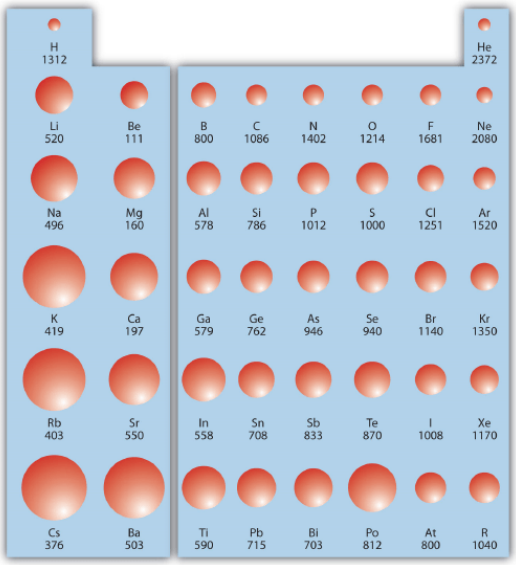


Figure 3.13.2: Ionization Energy on the Periodic Table. Values are in kJ/mol.

IE also shows an interesting trend within a given atom. This is because more than one IE can be defined by removing successive electrons (if the atom has them to begin with):

- First Ionization Energy (IE_1):

$$A(g) \rightarrow A^{+}(g) + e^{-}$$

- Second Ionization Energy (IE_2):

$$A^{+}(g) \rightarrow A^{2+}(g) + e^{-}$$

- Third Ionization Energy (IE_3):

$$A^{2+}(g) \rightarrow A^{3+}(g) + e^{-}$$

and so forth.

Each successive IE is larger than the previous because an electron is being removed from an atom with a progressively larger positive charge. However, IE takes a large jump when a successive ionization goes down into a new shell. For example, the following are the first three IEs for Mg, whose electron configuration is $1s^2 2s^2 2p^6 3s^2$:

- First Ionization Energy (IE_1) = 738 kJ/mol:

$$\mathrm{Mg\,(g)} \longrightarrow \mathrm{Mg^{+}\,(g)} + \mathrm{e^{-}}$$

- Second Ionization Energy (IE_2) = 1,450 kJ/mol:

$$\mathrm{Mg^{+}\,(g)} \longrightarrow \mathrm{Mg^{2+}\,(g)} + \mathrm{e^{-}}$$

- Third Ionization Energy (IE_3) = 7,734 kJ/mol:

$$\mathrm{Mg^{2+}\,(g)} \longrightarrow \mathrm{Mg^{3+}\,(g)} + \mathrm{e^{-}}$$

The second IE is twice the first, which is not a surprise: the first IE involves removing an electron from a neutral atom, while the second one involves removing an electron from a positive ion. The third IE, however, is over *five times* the previous one. Why is it so much larger? Because the first two electrons are removed from the 3*s* subshell, but the third electron has to be removed from the $n = 2$ shell (specifically, the 2*p* subshell, which is lower in energy than the $n = 3$ shell). Thus, it takes much more energy than just overcoming a larger ionic charge would suggest. It is trends like this that demonstrate that electrons within atoms are organized in groups.

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Example 3.13.2: Ionization Energies

Which atom in each pair has the larger first ionization energy?

a. Ca or Sr
b. K or K^+

Solution

a. Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE.
b. Because K^+ has a positive charge, it will be harder to remove another electron from it, so its IE is larger than that of K. Indeed, it will be significantly larger because the next electron in K^+ to be removed comes from another shell.

Exercise 3.13.2: Ionization Energies

Which atom has the lower ionization energy, C or F?

Answer available at the end of this section.

The opposite of IE is described by **electron affinity (EA),** which is the energy change when a gas-phase atom accepts an electron:

$$\mathrm{A\,(g) + e^- \longrightarrow A^-\,(g)} \qquad \Delta H \equiv EA$$

EA is also usually expressed in kJ/mol. EA also demonstrates some periodic trends, although they are less obvious than the other periodic trends discussed previously. Generally, as you go across the periodic table, EA increases its magnitude:

$$as \rightarrow PT,\ EA \uparrow$$

There is not a definitive trend as you go down the periodic table; sometimes EA increases, sometimes it decreases. Figure 3.13.3 shows EA values versus position on the periodic table for the *s*- and *p*-block elements. The trend is not absolute, especially considering the large positive EA values for the second column. However, the general trend going across the periodic table should be obvious.

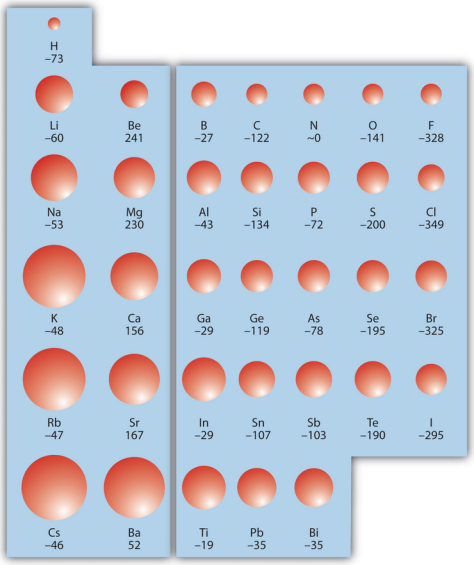


Figure 3.13.3: Electron Affinity on the Periodic Table. Values are in kJ/mol.

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Example 3.13.3: Electron Affinities

Predict which atom in each pair will have the highest magnitude of Electron Affinity.

a. C or F
b. Na or S

Solution

a. C and F are in the same row on the periodic table, but F is farther to the right. Therefore, F should have the larger magnitude of EA.
b. Na and S are in the same row on the periodic table, but S is farther to the right. Therefore, S should have the larger magnitude of EA.

Exercise 3.13.3: Electron Affinities

Predict which atom will have the highest magnitude of Electron Affinity: As or Br.

Answer available at the end of this section.

Metallic Character

The metallic character is used to define the chemical properties that metallic elements present. Generally, metals tend to lose electrons to form cations. Nonmetals tend to gain electrons to form anions. They also have a high oxidation potential—therefore they are easily oxidized and are strong reducing agents. Metals also form basic oxides; the more basic the oxide, the higher the metallic character.

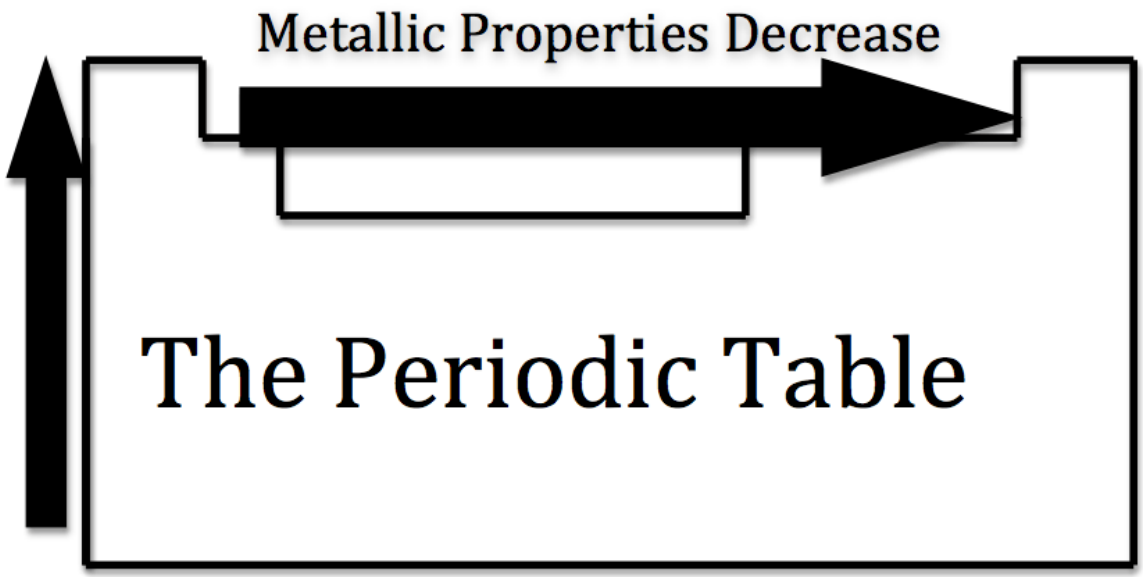


Figure 3.13.4: Courtesy of Jessica Thornton (UCD)

As you move across the table from left to right, the metallic character decreases, because the elements easily accept electrons to fill their valance shells. Therefore, these elements take on the nonmetallic character of forming anions. As you move up the table, the metallic character decreases, due to the greater pull that the nucleus has on the outer electrons. This greater pull makes it harder for the atoms to lose electrons and form cations.

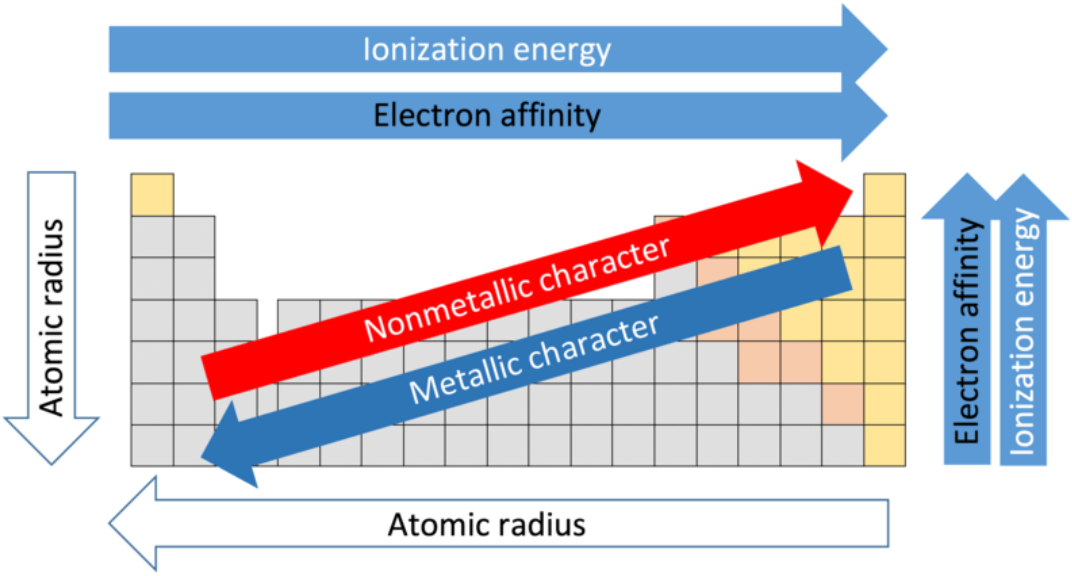
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Uses of the Periodic Properties of Elements

1. Predict greater or smaller atomic size and radial distribution in neutral atoms and ions.
2. Measure and compare ionization energies.
3. Compare electron affinities and electronegativities.
    - Predict redox potential.
    - Compare metallic character with other elements; ability to form cations.
    - Predict reactions that may or may not occur due to the trends.
    - Determine greater cell potential (sum of oxidation and reduction potential) between reactions.
    - Complete chemical reactions according to trends.

Summary

- Certain properties—notably atomic radius, ionization energies, and electron affinities—can be qualitatively understood by the positions of the elements on the periodic table. The major trends are summarized in the figure below.
- There are three factors that help in the prediction of the trends in the Periodic Table: number of protons in the nucleus, number of shells, and shielding effect.



Various periodic trends (CC BY-SA 4.0; Sandbh via Wikipedia)

Answers

Exercise 3.13.1**: Atomic Radii**

Br

Exercise 3.13.2**: Ionization Energies**

C

Exercise 3.13.3**: Electron Affinities**

Br

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4: THE MODERN MODEL OF THE ATOM

4.1: Electromagnetic Spectrum

4.2: Wavelength and Frequency Calculations

4.3: Energy and Frequency Calculations

4.4: The Bohr Model - Atoms with Orbits

4.5: The Quantum-Mechanical Model- Atoms with Orbitals

4.6: Electron Configurations

4.7: Electron Configuration Shortcut - Noble Gas Configurations

4.8: Valence Electrons

4.9: Electron Configurations of Cations

4.10: Electron Configurations of Transition Metal Cations

4.11: Electron Configurations of Anions

4.12: Octet Rule

4.13: Ionic Bonding

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4.1: Electromagnetic Spectrum



Figure 4.1.1 (Public Domain; Jon Sullivan / pdphoto.org via Wikipedia)

What are waves?

Waves can come in many sizes. Here we see a large wave crashing on the beach. Other waves can be very small and regular. We normally think of waves as being made of water, but there are forms of energy that take on the characteristics of waves. The idea of a wave has played a major role in our understanding of how the atom is put together and why it behaves the way it does.

Properties of Light

The nuclear atomic model proposed by Rutherford was a great improvement over previous models, but was still not complete. It did not fully explain the location and behavior of the electrons in the vast space outside of the nucleus. In fact, it was well known that oppositely charged particles attract one another. Rutherford's model did not explain why the electrons don't simply move toward, and eventually collide with, the nucleus. Experiments in the early twentieth century began to focus on the absorption and emission of light by matter. These studies showed how certain phenomena associated with light reveal insight into the nature of matter, energy, and atomic structure.

Wave Nature of Light

In order to begin to understand the nature of the electron, we first need to look at the properties of light. Prior to 1900, scientists thought light behaved solely as a wave. As we will see later, this began to change as new experiments demonstrated that light also has some of the characteristics of a particle. First, we will examine the wavelike properties of light.

Visible light is one type of **electromagnetic radiation**, which is a form of energy that exhibits wavelike behavior as it moves through space. Other types of electromagnetic radiation include gamma rays, x-rays, ultraviolet light, infrared light, microwaves, and radio waves. The figure below shows the **electromagnetic spectrum**, which is all forms of electromagnetic radiation. Notice that visible light makes up only a very, very small portion of the entire electromagnetic spectrum. All electromagnetic radiation moves through a vacuum at a constant speed of $2.998 \times 10^8 \text{ m/s}$. While the presence of air molecules slows the speed of light by a very small amount, we will use the value of $3.00 \times 10^8 \text{ m/s}$ as the speed of light in air.

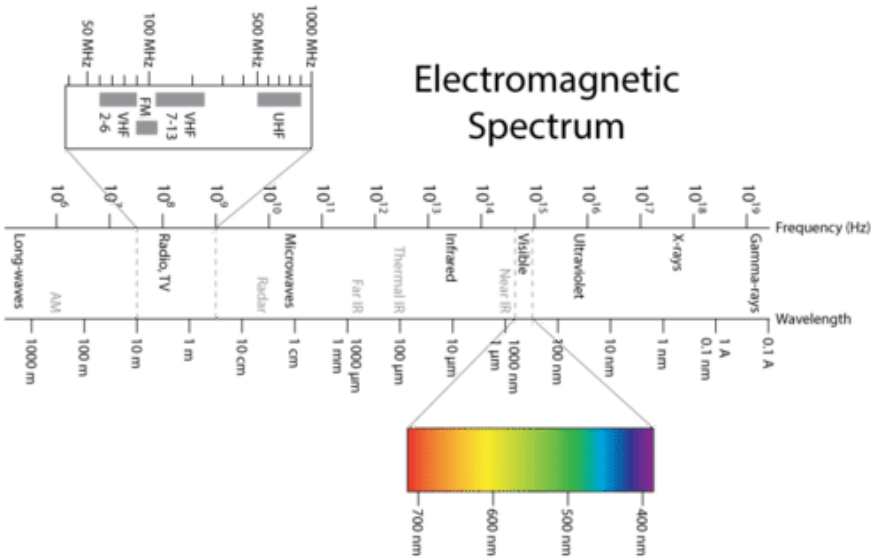


Figure 4.1.2: The electromagnetic spectrum encompasses a very wide range of wavelengths and frequencies. Visible light is only a very small portion of the spectrum with wavelengths from 400-700 nm. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

The figure above shows how the electromagnetic spectrum displays a wide variation in wavelength and frequency. Radio waves have wavelengths of as long as hundreds of meters, while the wavelength of gamma rays are on the order of 10^{-12} m. The corresponding frequencies range from 10^{6} to 10^{21} Hz. Visible light can be split into colors with the use of a prism (see below), yielding the visible spectrum of light. Red light has the longest wavelength and lowest frequency, while violet light has the shortest wavelength and highest frequency. Visible light wavelength ranges from about 400 to 700 nm with frequencies in the range of 10^{14} Hz.

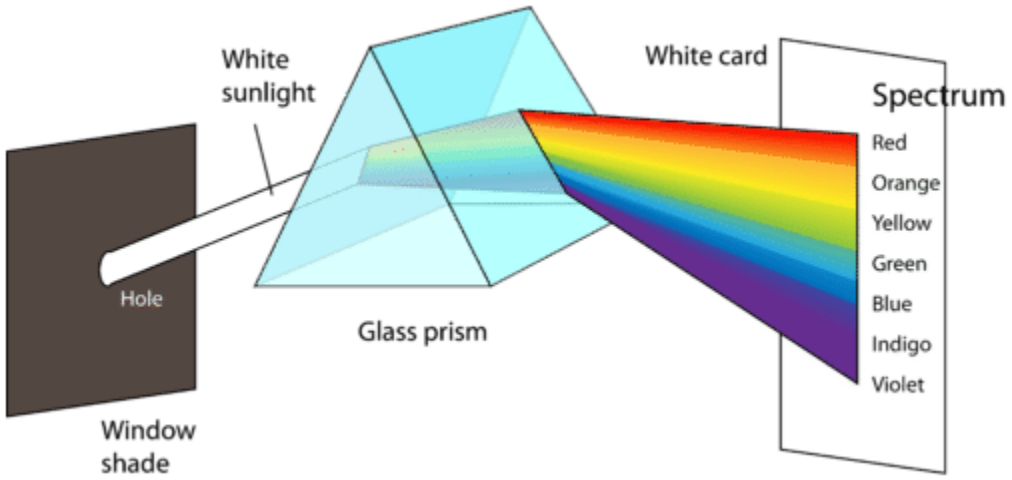


Figure 4.1.3: A small beam of white light is refracted (bent) as it passes through a glass prism. The shorter the wavelength of light, the greater is the refraction, so the light is separated into all of its colors. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Simulation

If we could see other wavelengths of energy, would they appear as colors? Use this simulation to explore the electromagnetic spectrum.

Summary

- Electromagnetic radiation is a form of energy.
- Visible light has wavelengths from $400\text{-}700\ \text{nm}$.
- The speed of light in air is $3.00 \times 10^8\ \text{m/s}$.

Review

1. What did Rutherford's nuclear atomic model not explain?
2. Prior to 1900, what did scientists believe about the nature of light?
3. What is visible light?
4. What is the range of wavelengths for visible light?

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4.2: Wavelength and Frequency Calculations



Figure 4.2.1 (CC BY 2.0; Piervincenzo Madeo (Flickr:piervix) via Flickr)

Do you enjoy going to the beach?

During the summer, almost everyone enjoys going to the beach. Beach-goers can swim, have picnics, and work on their tans. But if a beach-goer gets too much sun, they can burn. A particular set of solar wavelengths are especially harmful to the skin. This portion of the solar spectrum is known as UV B, with wavelengths of $280\text{-}320\ \text{nm}$. Sunscreens are effective in protecting skin against both the immediate skin damage and the long-term possibility of skin cancer.

Waves

Waves are characterized by their repetitive motion. Imagine a toy boat riding the waves in a wave pool. As the water wave passes under the boat, it moves up and down in a regular and repetitive fashion. While the wave travels horizontally, the boat only travels vertically up and down. The figure below shows two examples of waves.

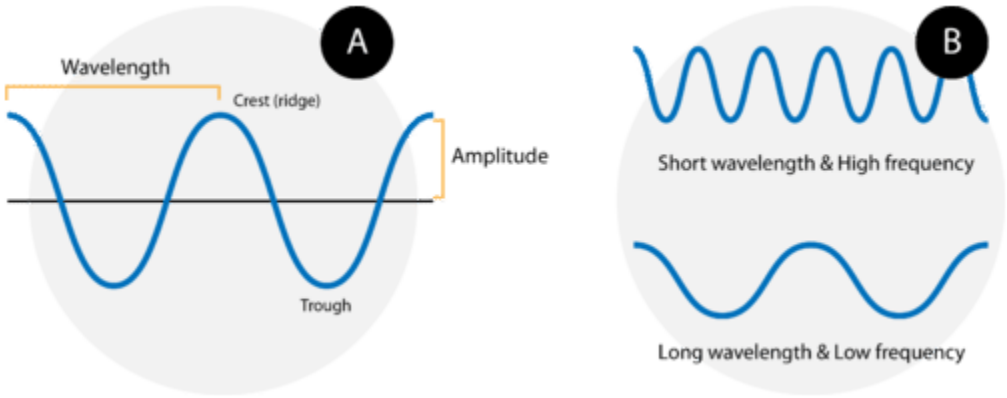


Figure 4.2.2: (A) A wave consists of alternating crests and troughs. The wavelength (λ) is defined as the distance between any two consecutive identical points on the waveform. The amplitude is the height of the wave. (B) A wave with a short wavelength (top) has a high frequency because more waves pass a given point in a certain amount of time. A wave with a longer wavelength (bottom) has a lower frequency. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

A wave cycle consists of one complete wave—starting at the zero point, going up to a wave **crest**, going back down to a wave

trough, and back to the zero point again. The **wavelength** of a wave is the distance between any two corresponding points on adjacent waves. It is easiest to visualize the wavelength of a wave as the distance from one wave crest to the next. In an equation, wavelength is represented by the Greek letter lambda (λ). Depending on the type of wave, wavelength can be measured in meters, centimeters, or nanometers $\left(1\ \mathrm{m} = 10^9\ \mathrm{nm}\right)$. The **frequency**, represented by the Greek letter nu (ν), is the number of waves that pass a certain point in a specified amount of time. Typically, frequency is measured in units of cycles per second or waves per second. One wave per second is also called a Hertz (Hz) and in SI units is a reciprocal second (s^{-1}).

Figure B above shows an important relationship between the wavelength and frequency of a wave. The top wave clearly has a shorter wavelength than the second wave. However, if you picture yourself at a stationary point watching these waves pass by, more waves of the first kind would pass by in a given amount of time. Thus the frequency of the first wave is greater than that of the second wave. Wavelength and frequency are therefore inversely related. As the wavelength of a wave increases, its frequency decreases. The equation that relates the two is:

$$c = \lambda\nu$$

The variable c is the speed of light. For the relationship to hold mathematically, if the speed of light is used in $\mathrm{m/s}$, the wavelength must be in meters and the frequency in Hertz.

Example 4.2.1

The color orange within the visible light spectrum has a wavelength of about $620\ \mathrm{nm}$. What is the frequency of orange light?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Wavelength $(\lambda) = 620\ \mathrm{nm}$
- Speed of light $(c) = 3.00 \times 10^8\ \mathrm{m/s}$
- Conversion factor $1\ \mathrm{m} = 10^9\ \mathrm{nm}$

Unknown

- Frequency

Convert the wavelength to m, then apply the equation $c = \lambda\nu$ and solve for frequency. Dividing both sides of the equation by λ yields:

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

Step 2: Calculate.

$$620\ \mathrm{nm} \times \left(\frac{1\ \mathrm{m}}{10^9\ \mathrm{nm}}\right) = 6.20 \times 10^{-7}\ \mathrm{m}$$

$$\nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8\ \mathrm{m/s}}{6.20 \times 10^{-7}} = 4.8 \times 10^{14}\ \mathrm{Hz}$$

Step 3: Think about your result.

The value for the frequency falls within the range for visible light.

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Summary

- All waves can be defined in terms of their frequency and intensity.
- $c = \lambda\nu$ expresses the relationship between wavelength and frequency.

Review

1. Define wavelength.
2. Define frequency.
3. What is the relationship between wavelength and frequency?

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4.3: Energy and Frequency Calculations

Every so often, you may hear a commercial or a news story with the words "quantum leap" in it. The quantum leap is supposed to be a major breakthrough—a big change, something extraordinarily large. The reality is far different. Instead of a big extravagant change, the "quantum" that scientists know about is a very small difference in the location of an electron around a nucleus—hardly an enormous shift at all.

Quantization of Energy

German physicist Max Planck (1858-1947) studied the emission of light by hot objects. You have likely seen a heated metal object glow an orange-red color (see below).



Figure 4.3.1: A heated object may glow different colors. The atoms in this piece of metal are releasing energy in discrete units called quanta. (Public Domain; Jack Delano, courtesy of the Library of Congress via Wikipedia)

Classical physics, which explains the behavior of large, everyday objects, predicted that a hot object would emit electromagnetic energy in a continuous fashion. In other words, every wavelength of light could possibly be emitted. Instead, what Planck found by analyzing the spectra was that the energy of the hot body could only be lost in small discrete units. A **quantum** is the minimum quantity of energy that can either be lost or gained by an atom. An analogy is that a brick wall can only undergo a change in height by units of one or more bricks, and not by any possible height. Planck showed that the amount of radiant energy absorbed or emitted by an object is directly proportional to the frequency of the radiation.

$$E = h\nu$$

In the equation, E is the energy, in joules, of a quantum of radiation, ν is the frequency, and h is a fundamental constant called **Planck's constant**. The value of Planck's constant is $h = 6.626 \times 10^{-34}\ \text{J} \cdot \text{s}$. The energy of any system must increase or decrease in units of $h \times \nu$. A small energy change results in the emission or absorption of low-frequency radiation, while a large energy change results in the emission or absorption of high-frequency radiation.

Example 4.3.1

What is the energy of a photon of green light with a frequency of $5.75 \times 10^{14}\ \text{Hz}$?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Frequency $(\nu) = 5.75 \times 10^{14}\ \text{Hz}$
- Planck's constant $(h) = 6.626 \times 10^{-34}\ \text{J} \cdot \text{s}$

Unknown

- energy (E)

Apply the equation $E = h\nu$ to solve for the energy.

Step 2: Calculate.

$$E = \left(6.626 \times 10^{-34}\ \text{J} \cdot \text{s}\right) \times \left(5.75 \times 10^{14}\ \text{Hz}\right) = 3.81 \times 10^{-19}\ \text{J}$$

Step 3: Think about your result.

While the resulting energy may seem very small, this is for only one photon of light. Visible quantities of light consist of huge quantities of photons. Recall that a hertz is equal to a reciprocal second, so the units agree in the equation.

Summary

- A quantum is the minimum quantity of energy that can be lost or gained by an atom.
- The value of Planck's constant is $h = 6.626 \times 10^{-34}\ \text{J} \cdot \text{s}$.

Review

1. What did classical physics predict about hot objects?
2. What did Planck find?
3. What is a quantum?

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4.4: The Bohr Model - Atoms with Orbits

Learning Objectives

- Define an energy level in terms of the Bohr model.
- Discuss how the Bohr model can be used to explain atomic spectra.

Electric light bulbs contain a very thin wire in them that emits light when heated. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances, but tungsten was chosen because the light it emits contains virtually every frequency and therefore, the light emitted by tungsten appears white. A wire made of some other element would emit light of some color that was not convenient for our uses. Every element emits light when energized by heating or passing electric current through it. Elements in solid form begin to glow when they are heated sufficiently, and elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs and is also the source of light in a fire.



Figure 4.4.1: Human/Need/Desire. Neon sculpture by Bruce Nauman (1983), who has been characterized as a conceptual artist.

Each Element Has a Unique Spectrum

The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Angstrom in 1868 and Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual frequencies that made up the light. The **emission spectrum** (or **atomic spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.



Figure 4.4.2: Atomic Emission Spectrum of Hydrogen.

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the color is spread out, we see that the hydrogen spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be

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defined using them.

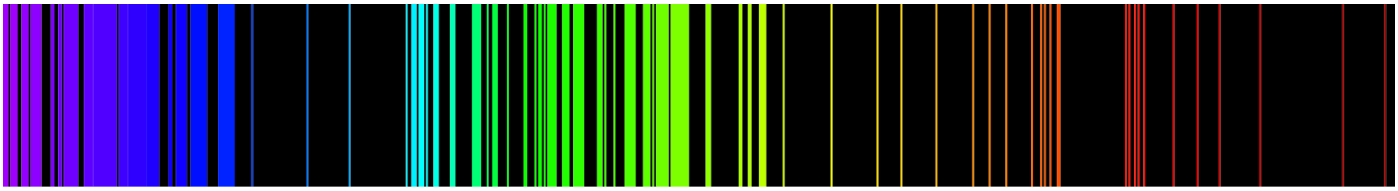


Figure 4.4.3: Atomic Emission Spectrum of Iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star, and after hearing that, wondered how scientists could know what elements were present in a place no one has ever been. Scientists determine what elements are present in distant stars by analyzing the light that comes from stars and finding the atomic spectrum of elements in that light. If the exact four lines that compose hydrogen's atomic spectrum are present in the light emitted from the star, that element contains hydrogen.

Bohr's Model of the Atom

By 1913, the concept of the atom had evolved from Dalton's indivisible spheres idea, to J. J. Thomson's plum pudding model, and then to Rutherford's nuclear atom theory. Rutherford, in addition to carrying out the brilliant experiment that demonstrated the presence of the atomic nucleus, also proposed that the electrons circled the nucleus in a planetary type motion. The solar system or planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system.



Figure 4.4.3: Niels Bohr with Albert Einstein at Paul Ehrenfest's home in Leiden (December 1925).

Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy, and would eventually crash into the nucleus. Furthermore, Rutherford's model was unable to describe how electrons give off light forming each element's unique atomic spectrum. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to be replaced.

In 1913, the Danish physicist Niels Bohr proposed a model of the electron cloud of an atom in which electrons orbit the nucleus and were able to produce atomic spectra. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light).

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Energy Levels

Bohr's key idea in his model of the atom is that electrons occupy definite orbitals that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbitals and it must have the precise energy required for that orbit. Orbits closer to the nucleus would require smaller amounts of energy for an electron, and orbits farther from the nucleus would require the electron to have a greater amount of energy. The possible orbits are known as **energy levels**. One of the weaknesses of Bohr's model was that he could not offer a reason why only certain energy levels or orbits were allowed.

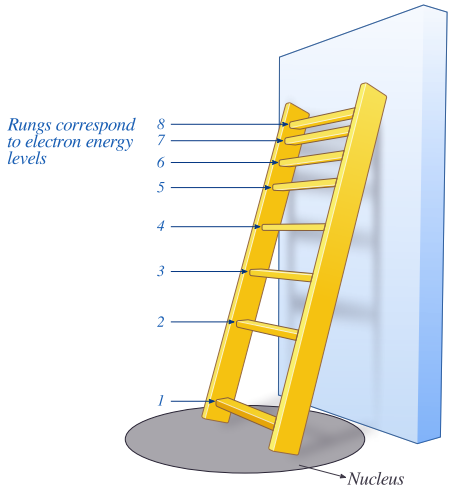


Figure 4.4.4: The energy levels of the electrons can be viewed as rungs on a ladder. Note that the spacing between rungs gets smaller at higher energies (CC BY-NC; Ümit Kaya)

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. It would be like a ladder that had rungs only at certain heights. The only way you can be on that ladder is to be on one of the rungs, and the only way you could move up or down would be to move to one of the other rungs. Suppose we had such a ladder with 10 rungs. Other rules for the ladder are that only one person can be on a rung in the normal state, and the ladder occupants must be on the lowest rung available. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all of the lower rungs are full. Bohr worked out rules for the maximum number of electrons that could be in each energy level in his model, and required that an atom in its normal state (ground state) had all electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, Bohr's model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.

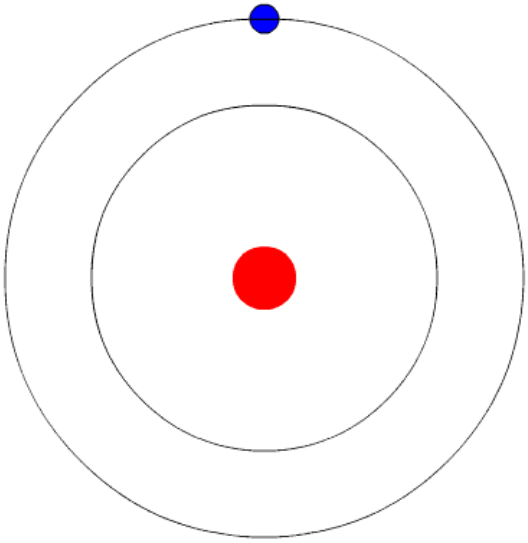


Figure 4.4.5: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; Kurzon).

Bohr's Model and Atomic Spectra

The evidence used to support Bohr's model came from the atomic spectra. He suggested that an atomic spectrum is made by the electrons in an atom moving energy levels. The electrons typically have the lowest energy possible, called the **ground state**. If the electrons are given energy (through heat, electricity, light, etc.) the electrons in an atom could absorb energy by jumping to a higher energy level, or **excited state**. The electrons then give off the energy in the form of a piece of light—called a **photon**—that they had absorbed, to fall back to a lower energy level. The energy emitted by electrons dropping back to lower energy levels will always be precise amounts of energy, because the differences in energy levels are precise. This explains why you see specific lines of light when looking at an atomic spectrum—each line of light matches a specific "step down" that an electron can take in that atom. This also explains why each element produces a different atomic spectrum. Because each element has different acceptable energy levels for its electrons, the possible steps each element's electrons can take differ from all other elements.

Summary

- Bohr's model suggests each atom has a set of unchangeable energy levels, and electrons in the electron cloud of that atom must be in one of those energy levels.
- Bohr's model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.

Vocabulary

- **Emission spectrum** (or **atomic spectrum**) - The unique pattern of light given off by an element when it is given energy.
- **Energy levels** - Possible orbits that an electron can have in the electron cloud of an atom.
- **Ground state** - To be in the lowest energy level possible.
- **Excited state** - To be in a higher energy level.
- **Photon** - A piece of electromagnetic radiation, or light, with a specific amount of energy.
- **Quantized** - To have a specific amount.

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4.5: The Quantum-Mechanical Model- Atoms with Orbitals

Learning Objectives

- Define quantum mechanics
- Differentiate between an orbit and an orbital.

How do you study something that seemingly makes no sense? We talk about electrons being in orbits and it sounds like we can tell where that electron is at any moment. We can draw pictures of electrons in orbit, but the reality is that we don't know exactly where they are. We are going to take a look at an area of science that even leaves scientists puzzled. When asked about quantum mechanics, Niels Bohr (who proposed the Bohr model of the atom) said: "Anyone who is not shocked by quantum theory has not understood it". Richard Feynman (one of the founders of modern quantum theory) stated: "I think I can safely say that nobody understands quantum theory." So, let's take a short trip into a land that challenges our everyday world.

Quantum Mechanics

The study of motion of large objects such as baseballs is called mechanics, or more specifically, classical mechanics. Because of the quantum nature of the electron and other tiny particles moving at high speeds, classical mechanics is inadequate to accurately describe their motion. **Quantum mechanics** is the study of the motion of objects that are atomic or subatomic in size and thus demonstrate wave-particle duality. In classical mechanics, the size and mass of the objects involved effectively obscures any quantum effects, so that such objects appear to gain or lose energies in any amounts. Particles whose motion is described by quantum mechanics gain or lose energy in small pieces called **quanta**.

One of the fundamental (and hardest to understand) principles of quantum mechanics is that the electron is both a particle and a wave. In the everyday macroscopic world of things we can see, something cannot be both. But this duality can exist in the quantum world of the submicroscopic on the atomic scale.

At the heart of quantum mechanics is the idea that we cannot accurately specify the location of an electron. All we can say is that there is a probability that it exists within this certain volume of space. The scientist Erwin Schrödinger developed an equation that deals with these calculations, which we will not pursue at this time.



Erwin Schrödinger.

Recall that in the Bohr model, the exact path of the electron was restricted to very well-defined circular **orbits** around the nucleus. An **orbital** is the quantum mechanical refinement of Bohr's orbit. In contrast to his concept of a simple circular orbit with a fixed radius, orbitals are mathematically derived regions of space with different *probabilities* of having an electron.

Summary

Quantum mechanics involves the study of material at the atomic level. This field deals with probabilities, since we cannot definitely locate a particle. Orbitals are mathematically derived regions of space with different *probabilities* of having an electron.

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4.6: Electron Configurations

Learning Objectives

- Relate the electron configurations of the elements to the shape of the periodic table.
- Determine the expected electron configuration of an element by its place on the periodic table.

Previously, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure 4.6.1. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.

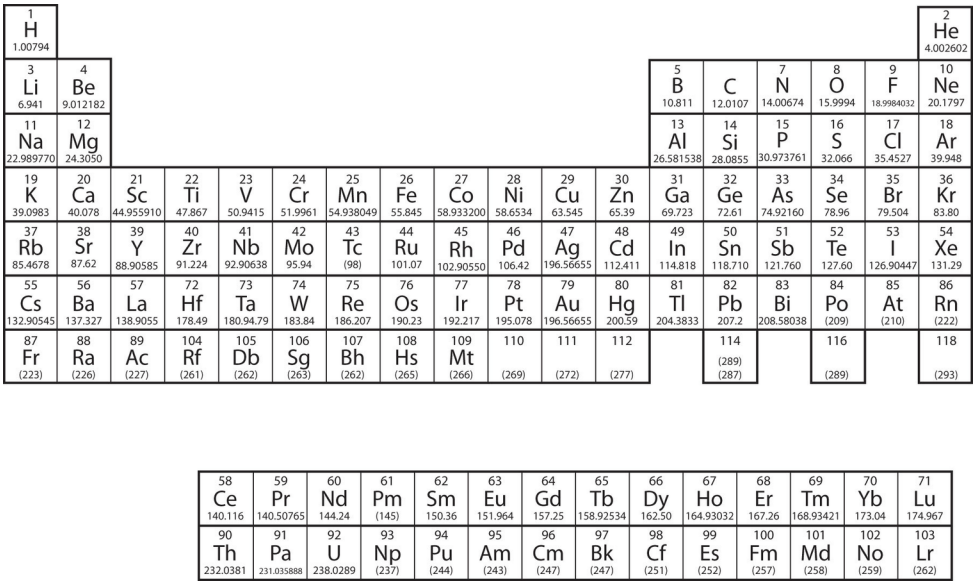


Figure 4.6.1: The Periodic Table

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations: the shape of the periodic table mimics the filling of the subshells with electrons.

> The shape of the periodic table mimics the filling of the subshells with electrons.

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure 4.6.2)

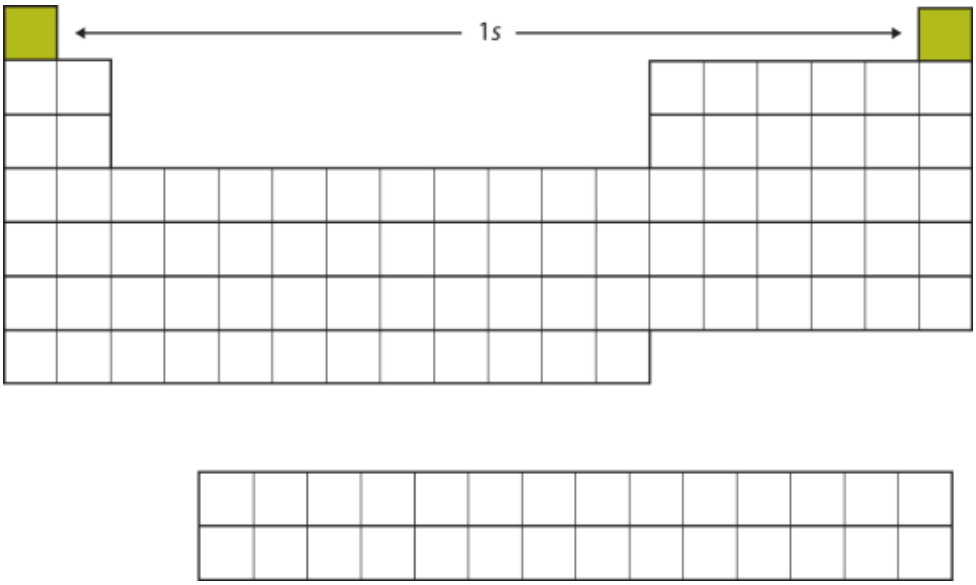


Figure 4.6.2: The 1s Subshell. H and He represent the filling of the 1s subshell.

The next two electrons, for Li and Be, would go into the 2s subshell. Figure 4.6.3 shows that these two elements are adjacent on the periodic table.

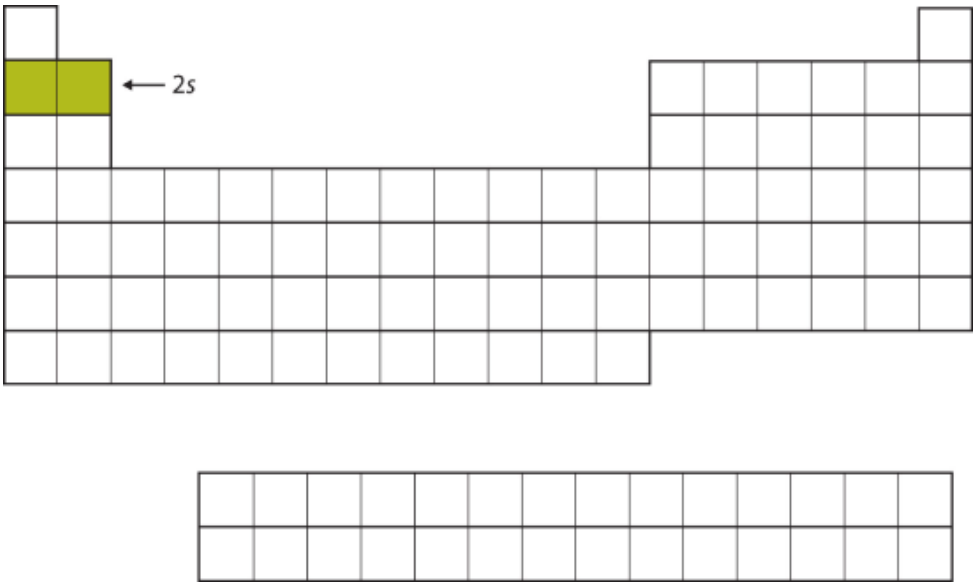


Figure 4.6.3: The 2s Subshell. In Li and Be, the 2s subshell is being filled.

For the next six elements, the 2*p* subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 4.6.4).

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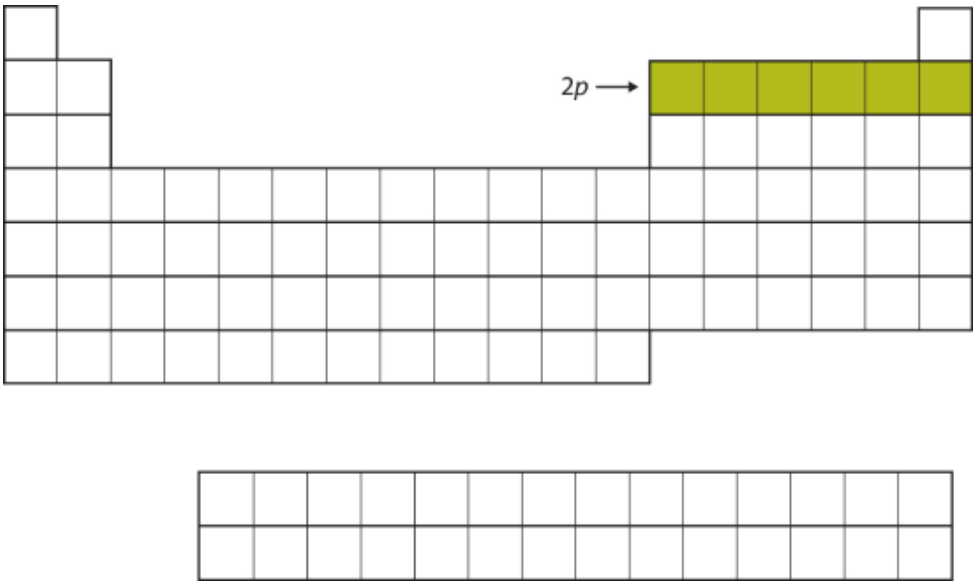


Figure 4.6.4: The 2*p* Subshell. For B through Ne, the 2*p* subshell is being occupied.

The next subshell to be filled is the 3*s* subshell. The elements when this subshell is being filled, Na and Mg, are back on the left side of the periodic table (Figure 4.6.5).

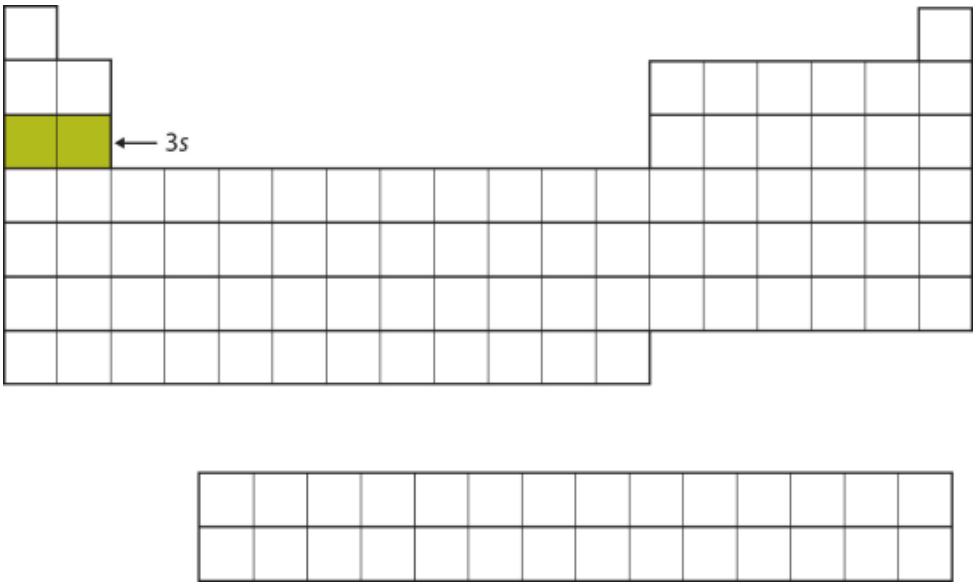


Figure 4.6.5: The 3*s* Subshell. Now the 3*s* subshell is being occupied.

Next, the 3*p* subshell is filled with the next six elements (Figure 4.6.6).

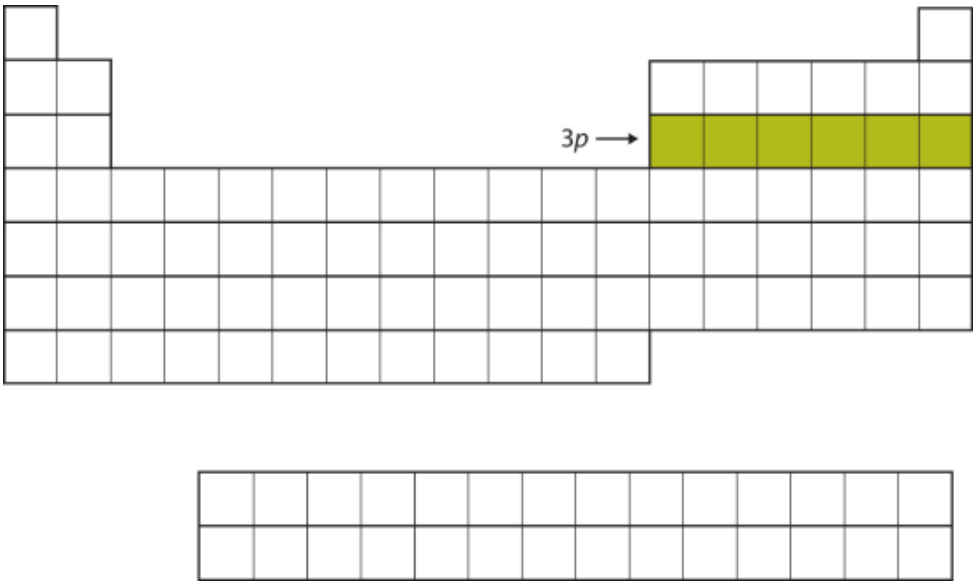


Figure 4.6.6: The 3*p* Subshell. Next, the 3*p* subshell is filled with electrons.

Instead of filling the 3*d* subshell next, electrons go into the 4s subshell (Figure 4.6.7).

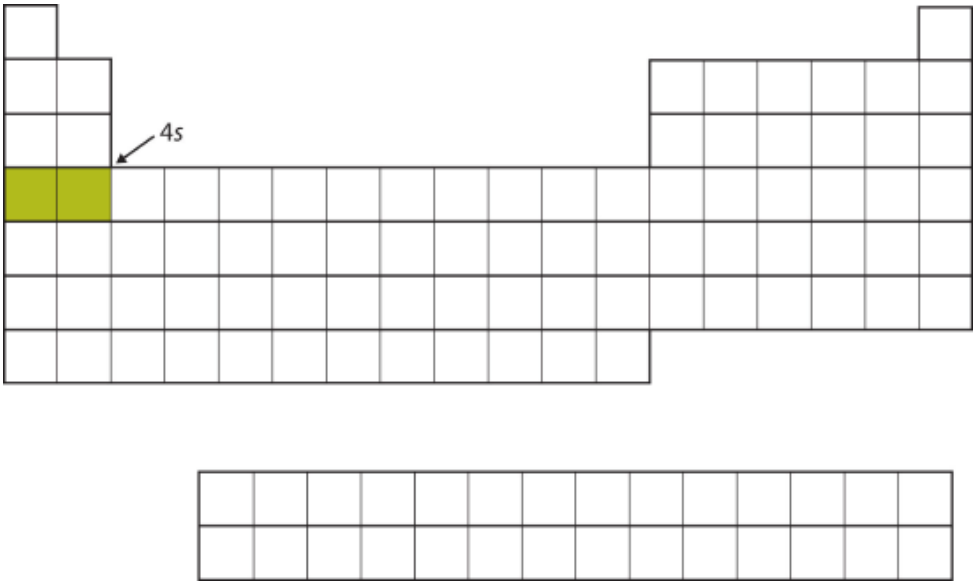


Figure 4.6.7: The 4s Subshell. The 4s subshell is filled before the 3*d* subshell. This is reflected in the structure of the periodic table.

After the 4s subshell is filled, the 3*d* subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure 4.6.8).

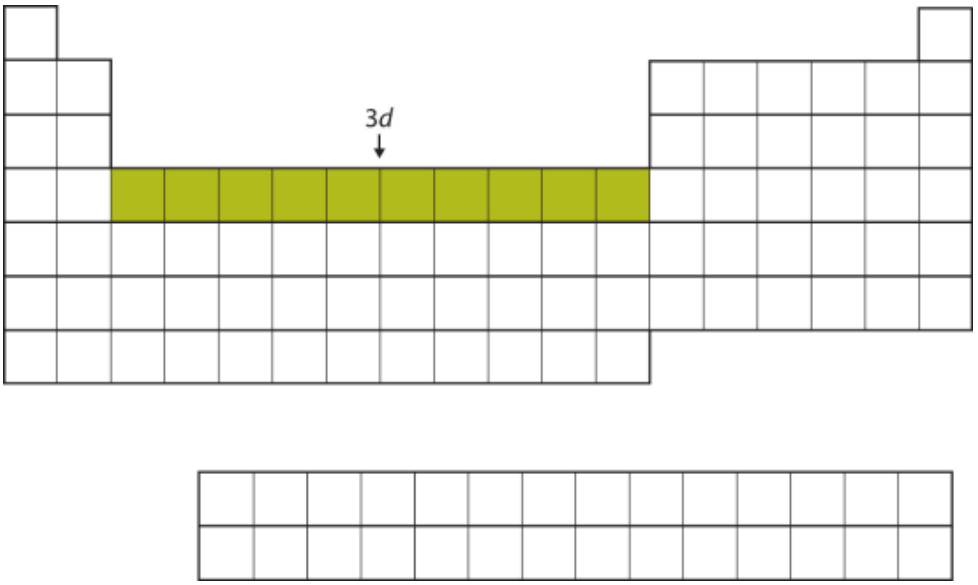


Figure 4.6.8: The 3*d* Subshell. The 3*d* subshell is filled in the middle section of the periodic table.

...And so forth. As we go across the rows of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the s subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** are the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 4.6.9 shows the blocks of the periodic table.

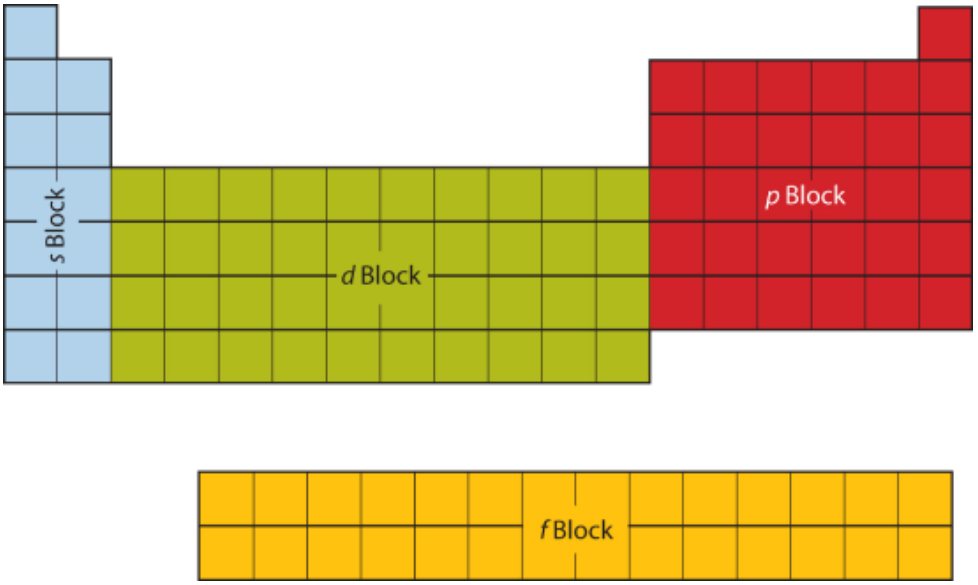


Figure 4.6.9: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

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Electrons, electron configurations, and the valence shell electron configuration highlighted.

H:	$1s^1$
Li:	$1s^2 2s^1$
Na:	$[Ne]3s^1$
K:	$[Ar]4s^1$
Rb:	$[Kr]5s^1$
Cs:	$[Xe]6s^1$

They all have a similar electron configuration in their valence shells: a single s electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry—*and they do*. The organization of electrons in atoms explains not only the shape of the periodic table, but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the s and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure 4.6.10. It is in the fourth column of the *p* block. This means that its electron configuration should end in a p^4 electron configuration. Indeed, the electron configuration of Se is $[Ar]4s^2 3d^{10} 4p^4$, as expected.

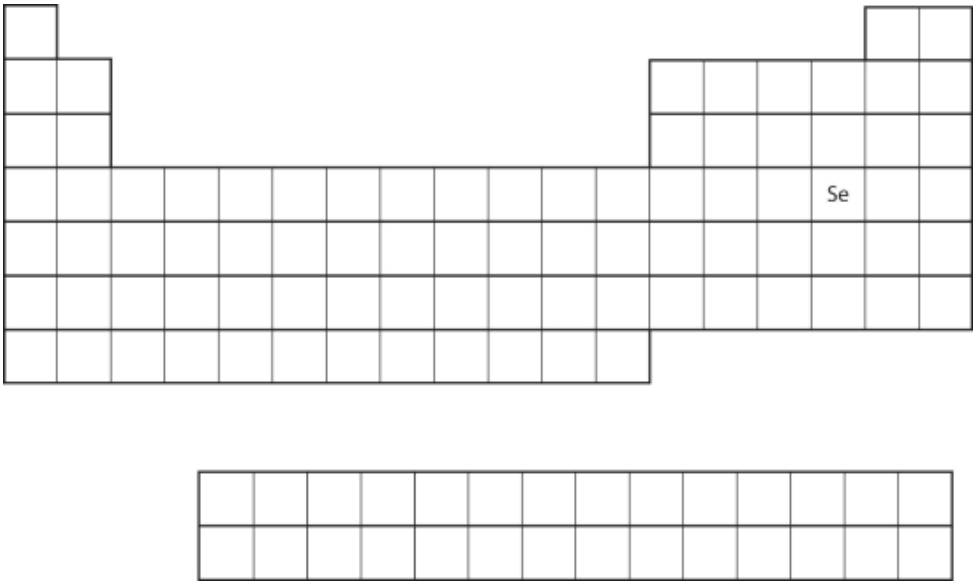


Figure 4.6.10: Selenium on the Periodic Table

Example 4.6.1: Predicting Electron Configurations

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure 4.6.11).

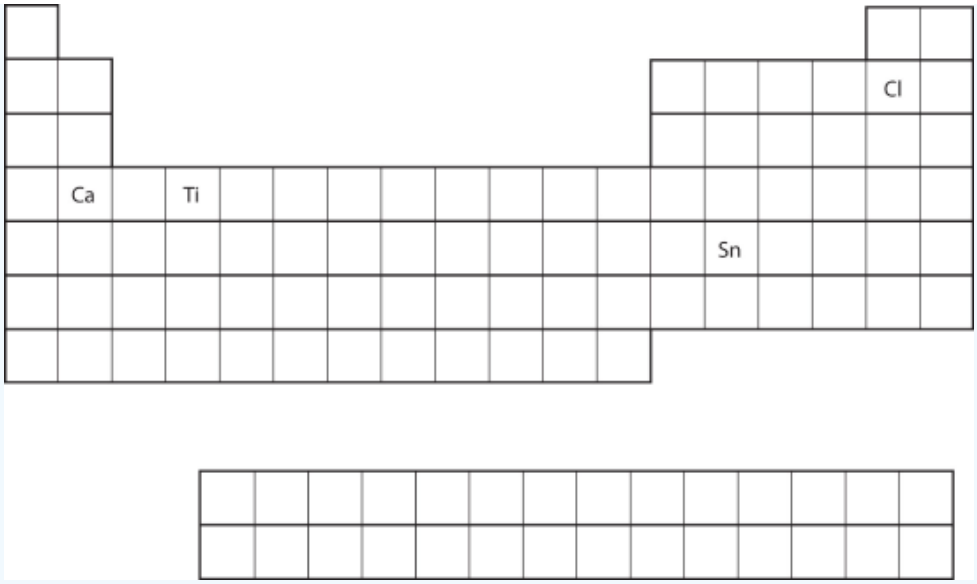


Figure 4.6.11: Various Elements on the Periodic Table

a. Ca
b. Sn

Solution

a. Ca is located in the second column of the *s* block. We expect that its electron configuration should end with s^2. Calcium's electron configuration is $[Ar]4s^2$.
b. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in p^2. Tin's electron configuration is $[Kr]5s^2 4d^{10} 5p^2$.

Exercise 4.6.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom. Figure 4.6.11.

a. Ti
b. Cl

Answer a available at the end of this section.

Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.

Answers

Exercise 4.6.1

$[Ar]4s^2 3d^2$

$[Ne]3s^2 3p^5$

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4.7: Electron Configuration Shortcut - Noble Gas Configurations



Figure 4.7.1 (Public Domain; User:Danbold/Wikipedia via Wikipedia)

How does it feel to be full after a meal?

Envision that you have nearly finished a great meal, but cannot put another bite in your mouth because there is no place for it to go. The noble gases have the same problem—there is no room for any more electrons in their outer shells. They are completely full and cannot handle any more.

Noble Gas Configuration

Sodium, element number 11, is the first element in the third period of the periodic table. Its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon $(Z = 10)$. This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. The elements that are found in the last column of the periodic table are an important group of elements called the noble gases. They are helium, neon, argon, krypton, xenon, and radon. A **noble gas configuration** of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons. So for sodium, we make the substitution of $[\mathrm{Ne}]$ for the $1s^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes $[\mathrm{Ne}]\, 3s^1$. Table 4.7.1 shows the noble gas configurations of the third period elements.

Table 4.7.1: Electron Configurations of Third-Period Elements

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Sodium	Na	11	$[\mathrm{Ne}]\, 3s^1$
Magnesium	Mg	12	$[\mathrm{Ne}]\, 3s^2$
Aluminum	Al	13	$[\mathrm{Ne}]\, 3s^2 3p^1$
Silicon	Si	14	$[\mathrm{Ne}]\, 3s^2 3p^2$
Phosphorus	P	15	$[\mathrm{Ne}]\, 3s^2 3p^3$
Sulfur	S	16	$[\mathrm{Ne}]\, 3s^2 3p^4$

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Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Chlorine	Cl	17	$[\mathrm{Ne}]\,3s^2 3p^5$
Argon	Ar	18	$[\mathrm{Ne}]\,3s^2 3p^6$

Again, the number of valence electrons increases from one to eight across the third period.

The fourth and subsequent periods follow the same pattern, except for the use of a different noble gas. Potassium has nineteen electrons, one more than the noble gas argon, so its configuration could be written as $[\mathrm{Ar}]\,4s^1$. In a similar fashion, strontium has two more electrons than the noble gas krypton, which would allow us to write its electron configuration as $[\mathrm{Kr}]\,5s^2$. All elements can be represented in this fashion.

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Summary

- The noble gas configuration system allows some shortening of the total electron configuration by using the symbol for the noble gas of the previous period as part of the pattern of electrons.

Review

1. What is the element represented by $[\mathrm{Ne}]\,3s^2 3p^2$?
2. What element has this electron configuration $[\mathrm{Ar}]\,3d^7 4s^2$?
3. What noble gas would be part of the electron configuration notation for Mn?
4. How would you write the electron configuration for Ba?

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4.8: Valence Electrons



Figure 4.8.1 (Public Domain; User:Chemicalinterest/Wikipedia via Wikipedia)

What makes a particular element very reactive and another element non-reactive?

A chemical reaction involves either electron removal, electron addition, or electron sharing. The path that a specific element will take in a reaction depends on where the electrons are in the atom and how many there are.

Valence Electrons

In the study of chemical reactivity, electrons in the outermost principal energy level are very important and so are given a special name. **Valence electrons** are the electrons in the highest occupied principal energy level of an atom. In the second period elements, the two electrons in the $1s$ sublevel are called **inner-shell electrons** and are not involved directly in the element's reactivity, or in the formation of compounds. Lithium has a single electron in the second principal energy level, and so we say that lithium has one valence electron. Beryllium has two valence electrons. How many valence electrons does boron have? Recognize that the second principal energy level consists of both the $2s$ and the $2p$ sublevels, and so the answer is three. In fact, the number of valence electrons goes up by one for each step across a period, until the last element is reached. Neon, with its configuration ending in $2s^2 2p^6$, has eight valence electrons.

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Summary

- Valence electrons are the outer-shell electrons of an atom.
- Valence electrons determine the reactivity of an atom.

Review

1. Define valence electron.
2. Define inner shell electron.
3. How many valence electrons are there in fluorine?
4. What are the 2s electrons in nitrogen?
5. How many inner shell electrons are there in beryllium?

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4.9: Electron Configurations of Cations



Figure 4.9.1 (Public Domain; User:Bugman/Wikipedia via Wikipedia)

How are cations formed?

In many parts of the country, the water contains high concentrations of minerals that stain clothes, build up deposits on bathtubs and water heaters, and create problems with soap foaming properly. This problem is caused by "hard water". The water contains excessive amounts of cations such as iron and calcium. These ions create a lot of problems in the water. Ion exchange resins can remove these minerals and clean up the water.

Cation Formation

Cations are the positive ions formed by the loss of one or more electrons. The most commonly formed cations of the representative elements are those that involve the loss of all of the valence electrons. Consider the alkali metal sodium (Na) . It has one valence electron in the third principal energy level. Upon losing that electron, the sodium ion now has an octet of electrons from the second principal energy level. The equation below illustrates this process.

$$\begin{array}{ll} \mathrm{Na} & \rightarrow \quad \mathrm{Na^+ + e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^1 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

The electron configuration of the sodium ion is now the same as that of the noble gas neon. The term **isoelectronic** refers to an atom and an ion of a different atom (or two different ions) that have the same electron configuration. The sodium ion is isoelectronic with the neon atom. Consider a similar process with magnesium and with aluminum:

$$\begin{array}{ll} \mathrm{Mg} & \rightarrow \quad \blacksquare + 2\mathrm{e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^2 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{ll} \mathrm{Al} & \rightarrow \quad \blacksquare + 3\mathrm{e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

In this case, the magnesium atom loses its two valence electrons in order to achieve the same noble gas configuration. The aluminum atom loses its three valence electrons. The $\mathrm{Mg^{2+}}$ ion, the $\mathrm{Al^{3+}}$ ion, the $\mathrm{Na^+}$ ion, and the $ceNe$ atom are all isoelectronic. For representative elements under typical conditions, three electrons are the maximum number that will be lost.

We can also show the loss of valence electron(s) with an electron dot diagram:

$$\mathrm{Na}\cdot \rightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$
$$\cdot\mathrm{Mg}\cdot \rightarrow \blacksquare + 2\mathrm{e}^{-}$$

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Summary

- Cations form when an atom loses one or more electrons.
- The resulting cation has the electron configuration of the noble gas atom in the row above it in the periodic table.

Review

1. What is a cation?
2. How many valence electrons does the sodium atom have?
3. Which atom is the sodium ion isoelectronic with?
4. How many electrons does magnesium lose to form the magnesium ion?

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4.10: Electron Configurations of Transition Metal Cations

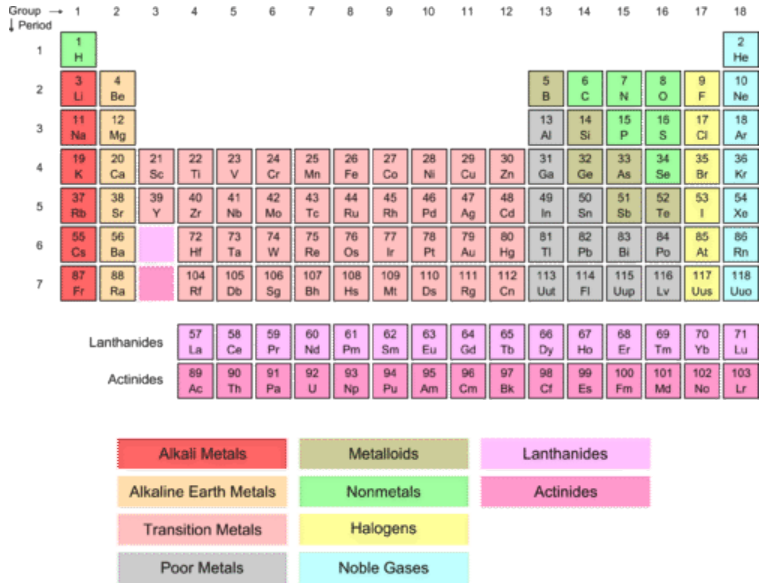


Figure 4.10.1 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Wikipedia)

How do transition metals form ions?

The transition metals are an interesting and challenging group of elements. They have perplexing patterns of electron distribution that do not always follow the electron filling rules. Predicting how they will form ions is also not always obvious.

Transition Metal Ions

Transition metals belong to the d block, meaning that the d sublevel of electrons is in the process of being filled with up to ten electrons. Many transition metals cannot lose enough electrons to attain a noble-gas electron configuration. In addition, the majority of transition metals are capable of adopting ions with different charges. Iron, which forms either the Fe^{2+} or Fe^{3+} ions, loses electrons as shown below.

$$\begin{array}{lcl} \mathrm{Fe} & \rightarrow & \blacksquare + 2\mathrm{e}^- \\ [\mathrm{Ar}]\ 3d^6\ 4s^2 & & [\mathrm{Ar}]\ 3d^6 \end{array}$$

$$\begin{array}{lcl} \mathrm{Fe} & \rightarrow & \blacksquare + 3\mathrm{e}^- \\ [\mathrm{Ar}]\ 3d^6\ 4s^2 & & [\mathrm{Ar}]\ 3d^5 \end{array}$$

According to the Aufbau process, the electrons fill the $4s$ sublevel before beginning to fill the $3d$ sublevel. However, the outermost s electrons are always the first to be removed in the process of forming transition metal cations. Because most transition metals have two valence electrons, the charge of $2+$ is a very common one for their ions. This is the case for iron above. A half-filled d sublevel (d^5) is particularly stable, which is the result of an iron atom losing a third electron.

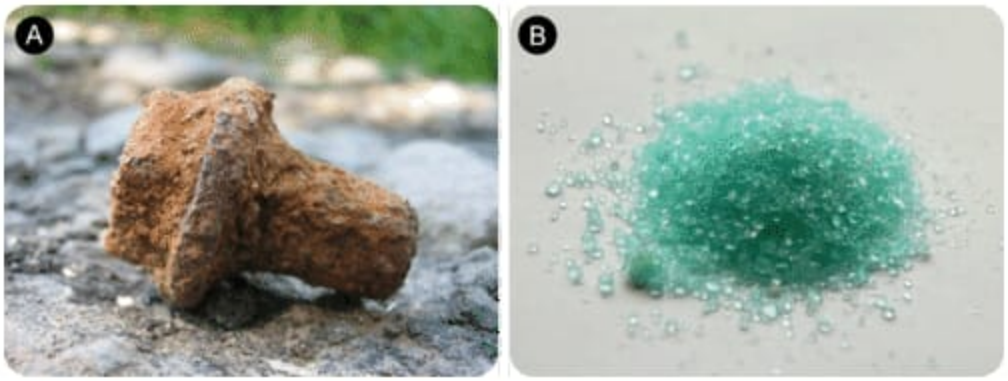


Figure 4.10.2: A. Rust is a complex combination of oxides of iron, among them iron (III) oxide, $\mathrm{Fe_2O_3}$. B. Iron (II) sulfate, $\mathrm{FeSO_4}$, has been known since ancient times as green vitriol and was used for centuries in the manufacture of inks. (CC by 2.0; Public Domain; (A) Paulnasca; (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Rust_screw.jpg; (B) http://commons.wikimedia.org/wiki/File:Iron%2528II%2529-sulfate-heptahydrate-sample.jpg)

Some transition metals that have relatively few d electrons may attain a noble-gas electron configuration. Scandium is an example.

$$\underset{[\mathrm{Ar}]\ 3d^1\ 4s^2}{\mathrm{Sc}} \rightarrow \underset{[\mathrm{Ar}]}{\blacksquare} + 3\mathrm{e}^-$$

Others may attain configurations with a full d sublevel, such as zinc and copper.

$$\underset{[\mathrm{Ar}]\ 3d^{10}\ 4s^2}{\mathrm{Zn}} \rightarrow \underset{[\mathrm{Ar}]\ 3d^{10}}{\blacksquare} + 2\mathrm{e}^-$$

$$\underset{[\mathrm{Ar}]\ 3d^{10}\ 4s^1}{\mathrm{Cu}} \rightarrow \underset{[\mathrm{Ar}]\ 3d^{10}}{\mathrm{Cu}^+} + \mathrm{e}^-$$

The resulting configuration above, with 18 electrons in the outermost principal energy level, is referred to as a **pseudo noble-gas electron configuration**. It gives particular stability to the Zn^{2+} and Cu^+ ions.

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Summary

- Transition metal ion formation is more complex than simple cation formation.
- Transition metal ions often involve rearrangements of both d and s electrons.

Review

1. What block do the transition metals fall in?
2. When writing the electron configuration for a transition metal, which sublevel (s, p, d, f) is filled first?
3. When forming transition metal ions, which sublevel (s, p, d, f) loses electrons first? What is the pseudo noble-gas electron configuration?

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4.11: Electron Configurations of Anions



Figure 4.11.1: A pressed glass open salt dish made in the early 1830s. (Public Domain; User:Daderot/Wikimedia Commons via Wikipedia)

How do you make chlorine safe to eat?

How do you transform a deadly gas into something you can sprinkle on your eggs and eat for breakfast? Chlorine in its free form is very dangerous if you breathe the fumes or come in contact with the gas. However, in reaction with sodium, sodium chloride forms as the sodium atom gives up an electron to chlorine, which accepts the electron to form the chloride anion.

Anions

Anions are the negative ions formed from the gain of one or more electrons. When nonmetal atoms gain electrons, they often do so until their outermost principal energy level achieves an octet. This process is illustrated below for the elements fluorine, oxygen, and nitrogen:

$$\begin{array}{lcl} \mathrm{F} + \mathrm{e}^- & \rightarrow & \mathrm{F}^- \\ 1s^2\ 2s^2\ 2p^5 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{lcl} \mathrm{O} + 2\mathrm{e}^- & \rightarrow & \blacksquare \\ 1s^2\ 2s^2\ 2p^4 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{lcl} \mathrm{N} + 3\mathrm{e}^- & \rightarrow & \blacksquare \\ 1s^2\ 2s^2\ 2p^3 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

All of these anions are isoelectronic with each other and with neon. They are also isoelectronic with the three cations from the previous section. Under typical conditions, three electrons are the maximum that will be gained in the formation of anions.

Outer electron configurations are constant within a group, so this pattern of ion formation repeats itself for Periods 3, 4, and following (see below).

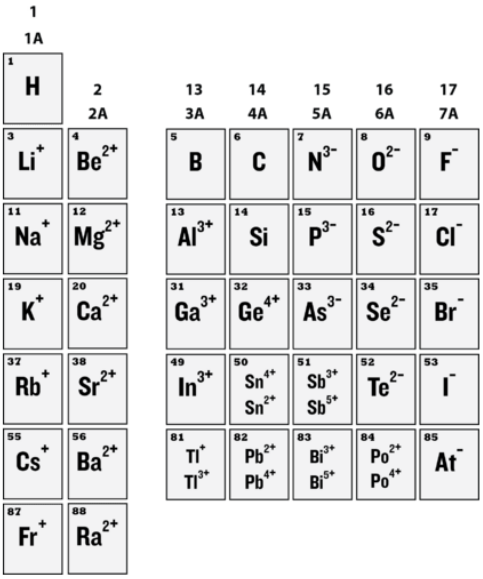


Figure 4.11.2: Ion charges. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

It is important not to misinterpret the concept of being isoelectronic. A sodium ion is very different from a neon atom, because the nuclei of the two contain different numbers of protons. One is an essential ion that is a part of table salt, while the other is an unreactive gas that is a very small part of the atmosphere. Likewise, sodium ions are very different than magnesium ions, fluoride ions, and all the other members of this isoelectronic series ([redacted], [redacted], F^-, Ne, Na^+, [redacted], [redacted]).

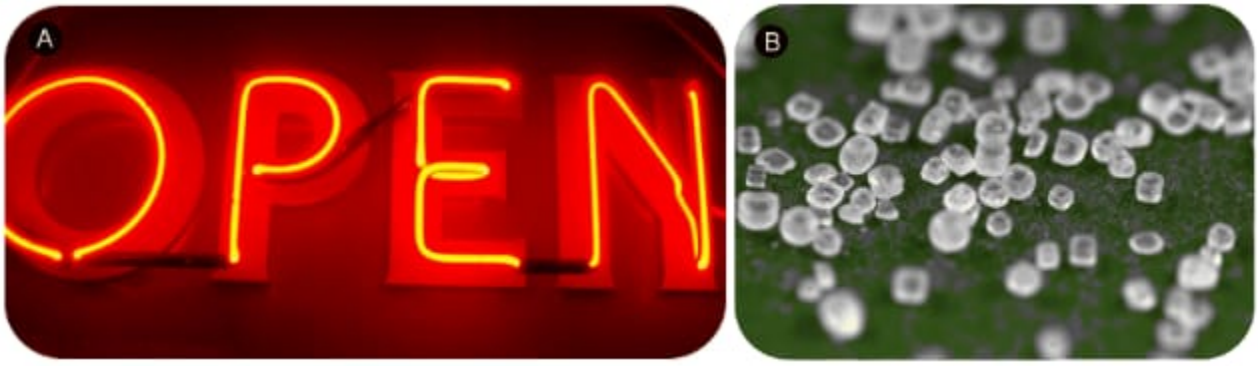


Figure 4.11.3: Neon gas (A) and sodium chloride crystals (B). Neon atoms and sodium ions are isoelectronic. Neon is a colorless and unreactive gas that glows a distinctive red-orange color in a gas discharge tube. Sodium ions are most commonly found in crystals of sodium chloride—ordinary table salt. (CC BY 2.0; (A) Andy Wright (Flickr: rightee); (B) Kevin Dooley;User:Bugman/Wikipedia via (A) http://www.flickr.com/photos/rightee/4356950; (B) http://www.flickr.com/photos/pagedooley/2769134850; http://commons.wikimedia.org/wiki/File:Ion_exchange_resin_beads.jpg)

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Summary

- Anions are negative ions formed by accepting electrons.
- The outermost principal energy level of an anion is usually an octet.

Review

1. What is an anion?
2. Write the electronic configurations for the chlorine atom and the chloride anion.
3. What does isoelectronic mean?

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4.12: Octet Rule



Figure 4.12.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How are electrons organized in atoms?

Graduations are exciting events in one's life. The processional, the ceremonies, even the speeches can be handled because this marks the end of a stage of life and the beginning of another one. The ceremony itself is challenging for those organizing it. There are just enough seats on the stage for the graduates. How do you avoid getting too many people in a row and not enough in the next row? Someone is stationed at the end of the row to count the students as they enter. Only so many are allowed to go into a row, and then you begin to fill the next row. Electrons in atoms behave the same way. There are rules that determine where electrons go in compounds.

Octet Rule

The noble gases are unreactive because of their electron configurations. The noble gas neon has the electron configuration of $1s^2\ 2s^2\ 2p^6$. It has a full outer shell and cannot incorporate any more electrons into the valence shell. The other noble gases have the same outer shell electron configuration, even though they have different numbers of inner-shell electrons.



Figure 4.12.2: Electron configuration of neon atom. (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

American chemist Gilbert Lewis (1875-1946) used this observation to explain the types of ions and molecules that are formed by other elements. He called his explanation the **octet rule**. The octet rule states that atoms tend to form compounds in ways that give them eight valence electrons, and thus the electron configuration of a noble gas. An exception to an octet of electrons is in the case of the first noble gas, helium, which only has two valence electrons. This primarily affects the element hydrogen, which forms stable compounds by achieving two valence electrons. Lithium, an alkali metal with three valence electrons, is also an exception to the octet rule. Lithium tends to lose one electron to take on the electron configuration of the nearest noble gas, helium, leaving it with two valence electrons.

There are two ways in which atoms can satisfy the octet rule. One way is by sharing their valence electrons with other atoms. The second way is by transferring valence electrons from one atom to another. Atoms of metals tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level. Atoms of nonmetals tend to gain electrons in order to fill their outermost principal energy level with an octet.

Watch the following video to learn how to use the octet rule to predict the charge of an ion.

YouTube Video

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View this YouTube Video online

Summary

- Atoms form compounds in ways that give them eight valence electrons.
- Metals tend to lose electrons to achieve this configuration.
- Nonmetals tend to gain electrons to achieve this configuration.

Review

1. What is the electron configuration of a noble gas?
2. Why is this configuration important?
3. What are the two ways in which atoms can satisfy the octet rule?
4. How do metals change to obey the octet rule?
5. How do nonmetals change to obey the octet rule?

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4.13: Ionic Bonding



Figure $4.13.1$ (Public Domain; Leon Brooks via Commons Wikimedia, Waneroo Beach [commons.wikimedia.org])

Does the sea really have salt in it?

We can get common table salt from several sources. It can be mined in the solid form in salt mines, or found as a solid in deposits. We can also get salt from the ocean, but it really does not exist as a salt when in solution. The sodium ions and chloride ions are dissolved, but not combined into a structure until all the water is removed.

Most of the rocks and minerals that make up the Earth's crust are composed of positive and negative ions held together by ionic bonding. An ionic compound is an electrically neutral compound consisting of positive and negative ions. You are very familiar with some ionic compounds, such as sodium chloride (NaCl). A sodium chloride crystal consists of equal numbers of positive sodium ions (Na^+) and negative chloride ions (Cl^-).

Ionic Bonds

Oppositely charged particles attract each other. This attractive force is often referred to as an **electrostatic force**. An **ionic bond** is the electrostatic force that holds ions together in an **ionic compound**. The strength of the ionic bond is directly dependent upon the quantity of the charges and inversely dependent on the distance between the charged particles. A cation with a $2+$ charge will make a stronger ionic bond than a cation with a $1+$ charge. A larger ion makes a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion.

YouTube Video

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Electron Dot Diagrams

We will use sodium chloride as an example to demonstrate the nature of the ionic bond and how it forms. As you know, sodium is a metal and loses its one valence electron to become a cation. Chlorine is a nonmetal and gains one electron in becoming an anion. Both achieve a noble-gas electron configuration. However, electrons cannot be simply "lost" to nowhere in particular. A more accurate way to describe what is happening is that a single electron is transferred from the sodium atom to the chlorine atom, as shown below.

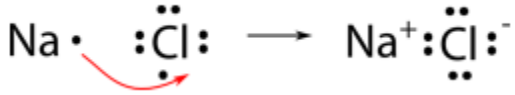


Figure 4.13.2 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The ionic bond is the attraction of the Na^{+} ion for the Cl^{-} ion. It is conventional to show the cation without dots around the symbol to emphasize that the original energy level that contained the valence electron is now empty. The anion is now shown with a complete octet of electrons.

For a compound such as magnesium chloride, it is not quite as simple. Because magnesium has two valence electrons, it needs to lose both to achieve the noble-gas configuration. Therefore, two chlorine atoms will be needed.

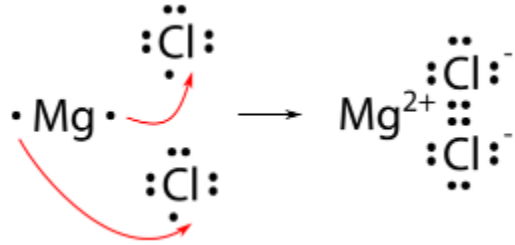


Figure 4.13.3 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The final formula for magnesium chloride is MgCl_2.

Summary

- An ionic compound contains positive and negative ions.
- An ionic bond is electrostatic in nature.
- Electron dot diagrams can be used to illustrate electron movements and ion formation.

Review

1. What is an ionic compound?
2. What is an ionic bond?
3. Which cation (Na^+ or Ca^{2+}) would form a stronger ionic bond with Cl^-?

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5: CHEMICAL BONDING AND NOMENCLATURE

5.1: A Molecular View of Elements and Compounds

5.2: Representing Valence Electrons with Dots

5.3: Lewis Structures of Ionic Compounds- Electrons Transferred

5.4: Lewis Structures of Covalent Compounds- Electrons Shared

5.5: Rules for Drawing Lewis Structures

5.6: Resonance - Equivalent Lewis Structures for the Same Molecule

5.7: Bond Polarity

5.8: Chemical Nomenclature

5.9: Writing Formulas for Ionic Compounds

5.10: Naming Ionic Compounds

5.11: Naming Covalent Compounds

5.12: Naming Acids

5.13: Nomenclature Summary

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5.1: A Molecular View of Elements and Compounds

Learning Objectives

- Classify substances as atomic elements, molecular elements, molecular compounds, or ionic compounds.

Atomic Elements

Most elements exist with **individual atoms** as their basic unit. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

Molecular Elements

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called **molecules**. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Table 5.1.1: Elements That Exist as Diatomic Molecules

Hydrogen, H	Oxygen	Nitrogen	Fluorine	Chlorine	Bromine	Iodine

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules—a molecule with only two atoms (Table 5.1.1). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2, N_2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 (Figure 5.1.1).

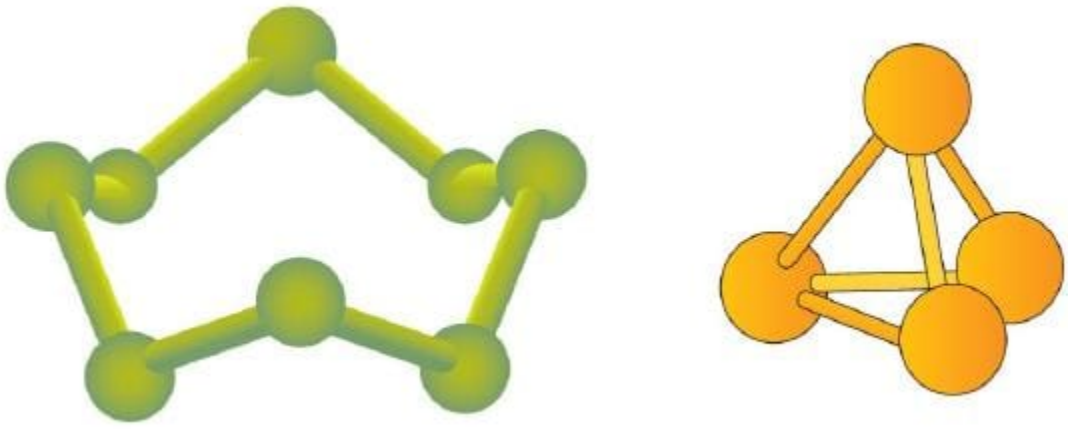


Figure 5.1.1: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 5.1.1 shows two examples of how molecules will be represented in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

Ionic Compounds

The elements in the periodic table are divided into specific groupings; the metals, the non-metals, the semi-metals, and so on. These groupings are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). The basic unit of ionic compounds is the **formula unit**.

Molecular Compounds

Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds. Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals) and silicon dioxide (SiO_2) will also be a covalently bound molecule (a semi-metal and a non-metal). The basic unit of molecular compounds is the **molecule**.

Example 5.1.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. Fe
b. PCl_3
c. LiBr
d. P_4
e. oxygen gas

Solution

a. **Fe** (iron) is an element that is represented with no subscript, so it is an **atomic element**.
b. **PCl_3** is made up of two nonmetals, so it is a **molecular compound**.
c. **LiBr** is made up of lithium, a metal, and bromine, a nonmetal, so it is an **ionic compound**.
d. **P_4** is a substance that is made up of four atoms of the same element, so it is a **molecular element**.
e. The formula for **oxygen gas** is **O_2** so it is a molecular element.

Exercise 5.1.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. I_2
b. He
c. H_2O
d. Al
e. CuCl

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

Answer e: available at the end of this section.

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Answers

Exercise 5.1.1
molecular element

Exercise 5.1.1
atomic element

Exercise 5.1.1
molecular compound

Exercise 5.1.1
atomic element

Exercise 5.1.1
ionic compound

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5.2: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, or a Lewis diagram, or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H}\cdot$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\dot{\mathbf{H}} \quad or \quad \cdot\mathbf{H} \quad or \quad \underset{\cdot}{\mathbf{H}}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He}\!:$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1s subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li}\cdot$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be}\!:$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\dot{\mathbf{B}}\!:$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\cdot\dot{\mathbf{C}}\!:$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\cdot\underset{\cdot}{\dot{\mathbf{N}}}\!:$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that each side has no more than two electrons.



Fluorine and neon have seven and eight dots, respectively:





With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the n = 3 shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

Example 5.2.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3s electrons:



2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the n = 4 shell, there are six electrons. Its electron dot diagram is as follows:



Exercise 5.2.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a available at the end of this section.

Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis electron dot diagrams for ions have less (for cations) or more (for anions) dots than the corresponding atom.

Answers

Exercise 5.2.1





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5.3: Lewis Structures of Ionic Compounds- Electrons Transferred

Learning Objectives

- State the octet rule.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something: many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the n = 2 shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\ddot{\underset{\cdot\,\cdot}{\mathbf{Cl}}}\mathbf{:}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \rightsquigarrow \cdot\ddot{\underset{\cdot\,\cdot}{\mathbf{Cl}}}\mathbf{:}$$

resulting in two ions—the Na^+ ion and the Cl^- ion:

$$\mathbf{Na}^+ \qquad \mathbf{:}\ddot{\underset{\cdot\,\cdot}{\mathbf{Cl}}}\mathbf{:}^-$$

$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Na}^+ + \mathbf{:}\ddot{\underset{\cdot\,\cdot}{\mathbf{Cl}}}\mathbf{:}^- \rightarrow Na^+Cl^- \;\; or \;\; NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

$$\mathbf{Mg}^{2+} + \left[:\ddot{\mathbf{O}}:\right]^{2-} \qquad Mg^{2+}O^{2-} \; or \; MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na}\cdot \rightharpoonup \cdot\ddot{\mathbf{O}}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



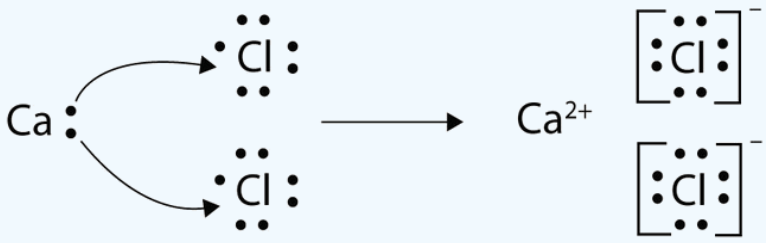
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

Example 5.3.1: Synthesis of Calcium Chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:



The oppositely charged ions attract each other to make $CaCl_2$.

Exercise 5.3.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

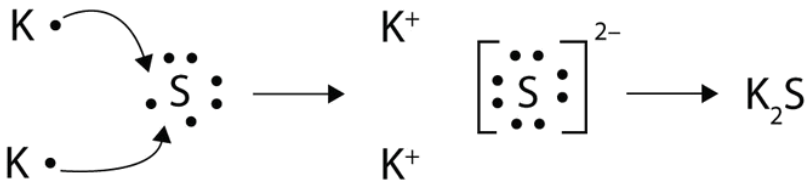
Answer available at the end of this section.

Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

Answers

Exercise 5.3.1



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5.4: Lewis Structures of Covalent Compounds- Electrons Shared

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

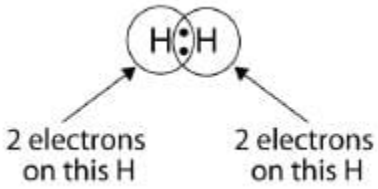
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1s subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H}$$

The two H atoms can share their electrons:

$$\mathbf{H:H}$$

We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:

H–H

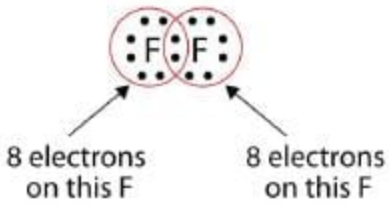
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



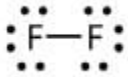
These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.



Note that each F atom has a complete octet around it now:



We can also write this using a dash to represent the shared electron pair:

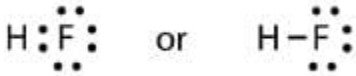


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

Example 5.4.1:

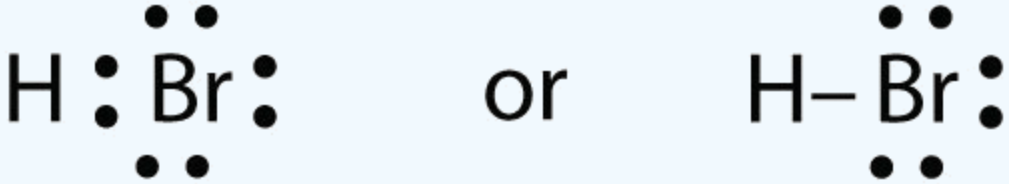
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:



The two atoms can share their unpaired electron:



Exercise 5.4.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer available at the end of this section.

When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far, and there are still some electrons that remain unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of C_2H_4. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two C atoms } = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four H atoms } = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

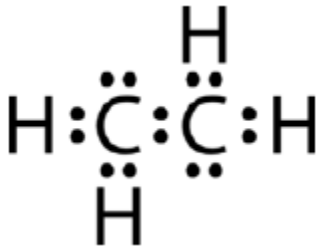


Figure 5.4.1: Incorrect dot structure of ethene. (CK12 License)

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

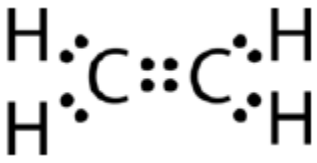


Figure 5.4.2: Correct dot structure for ethene. (CK12 License)

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as

shown in the figure below.

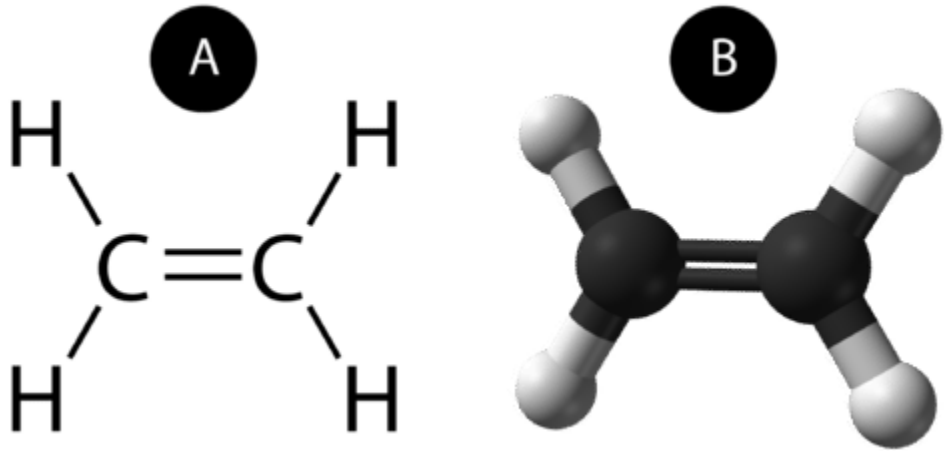


Figure 5.4.3: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

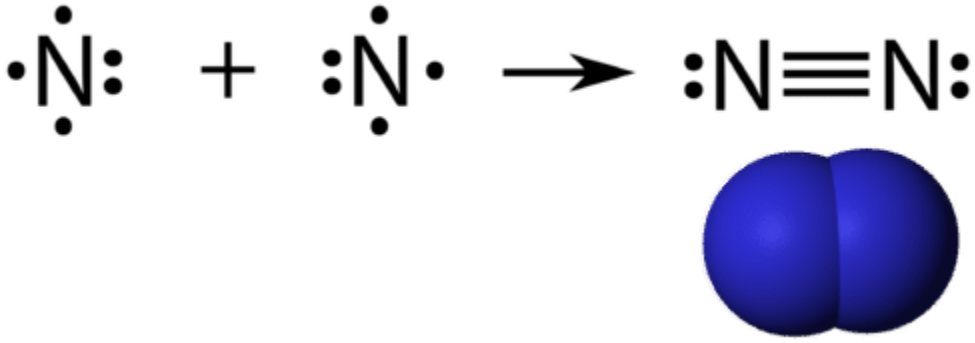


Figure 5.4.4: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributions & Attributions

-

- Anonymous by request

Answers

Exercise 5.4.1



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5.5: Rules for Drawing Lewis Structures

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

  For CO_3^{2-}, for example, we add two electrons to the total because of the –2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 5.5.1: Water

Write the Lewis Structure for H_2O.

Solution

Solutions to Example 10.4.1

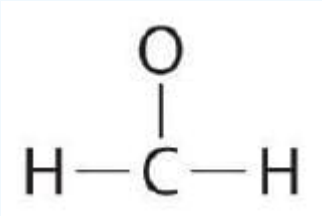
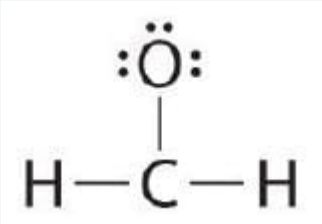
Steps for Writing Lewis Structures	Example 5.5.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.
2. Arrange the atoms to show specific connections.	**H O H** Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives **H -O- H** with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

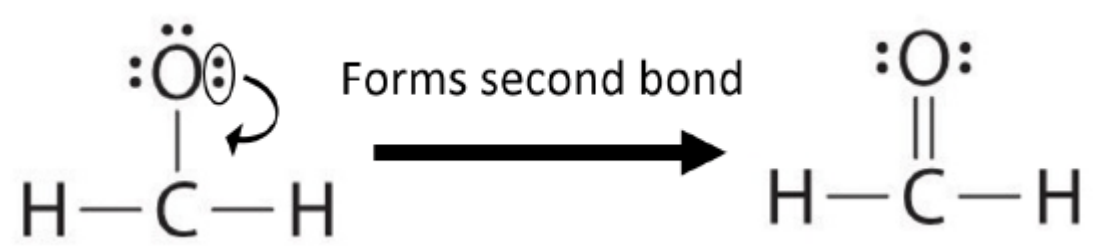
Example 5.5.2

Write the Lewis structure for the CH_2O molecule

Solution

Solutions to Example 10.4.2

Steps for Writing Lewis Structures	Example 5.5.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of bonded atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.

Steps for Writing Lewis Structures	Example 5.5.2
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 5.5.1

Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO2 available at the end of this section.

The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

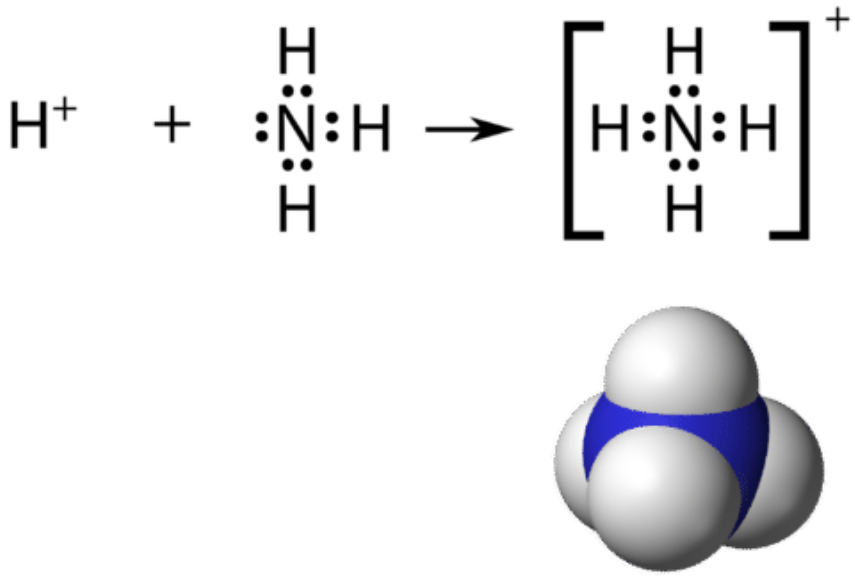


Figure 5.5.3: The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

1 N atom $= 5$ valence electrons

4 H atoms $= 4 \times 1 = 4$ valence electrons

subtract 1 electron for the $1+$ charge of the ion

total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

Exercise 5.5.2

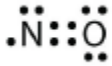
Draw the Lewis electron dot structure for the sulfate ion.

Answer (CK12 License) available at the end of this section.

Exceptions to the Octet Rule (BC Campus)

As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

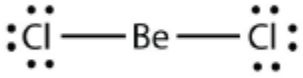
There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:



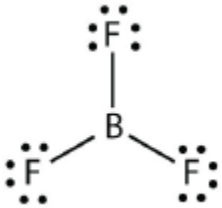
Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable

compound, it is very chemically reactive, as are most other odd-electron compounds.

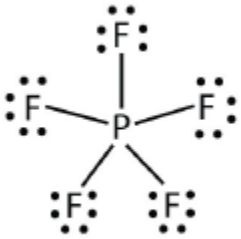
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 5.5.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

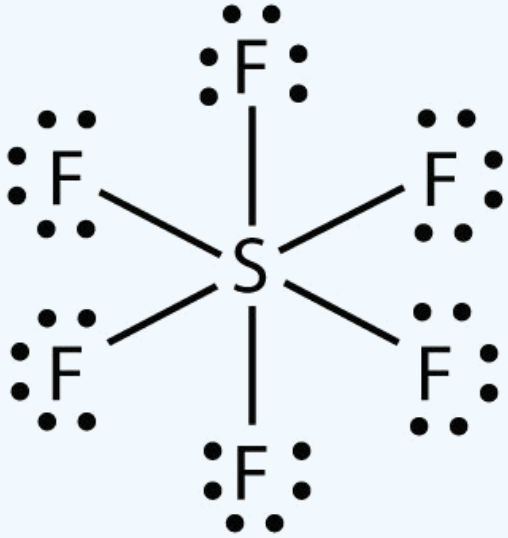
a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 5.5.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

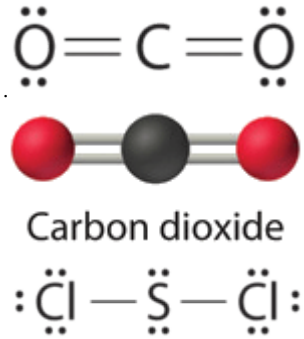
Answer available at the end of this section.

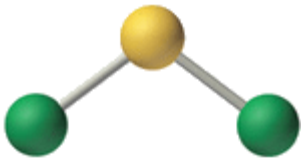
Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

Answers

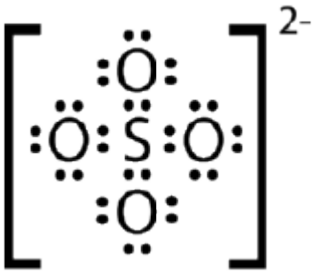
Exercise 5.5.1





Sulfur dichloride

Exercise 5.5.2



Exercise 5.5.3**: Xenon Difluoride**



The Xe atom has an expanded valence shell with more than eight electrons around it.

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5.6: Resonance - Equivalent Lewis Structures for the Same Molecule

Learning Objectives

- Explain the concept of resonance and how it works with within molecules.

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone (O_3) molecule in Figure $5.6.1$. There are a total of 18 electrons in the structure and so the following two structures are possible.

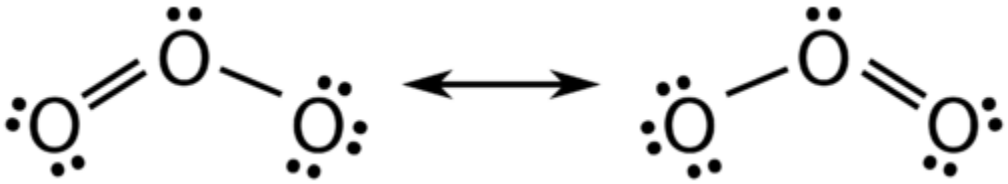


Figure $5.6.1$: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left $(5.6.1)$ can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $O{-}O$ single bond and a double bond.

Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure $5.6.2$).



Figure $5.6.2$: "Half-bond" model of ozone molecule. This is a better description of the electronic structure of ozone than either of the resonance structures in Figure $5.6.1$.

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- in Figure $5.6.3$.

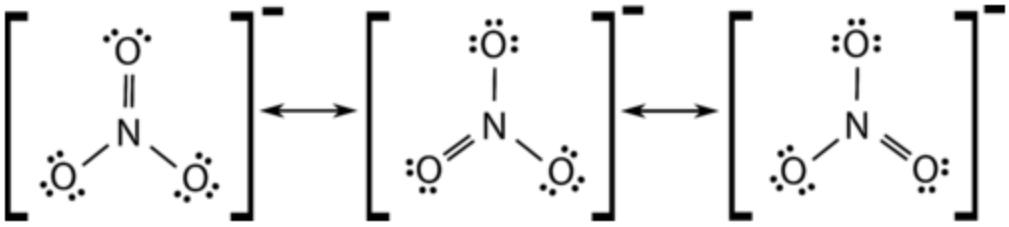


Figure 5.6.3: Resonance structure of nitrate anion.

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

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5.7: Bond Polarity



Figure 5.7.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What makes people share?

Have you ever spent time with someone you really didn't like? You had nothing in common with them and did not want to have anything to do with them. On the other hand, there are people that you enjoy being with. You have a lot in common and like to share with them. Atoms work the same way. If there are strong differences in their attraction of electrons, one atom gets the electrons and the other atom loses them. If they are similar, they share electrons to form a covalent bond.

Bond Polarity

Electronegativity is defined as the ability of an atom to attract electrons when the atoms are in a compound. Electronegativities of elements are shown in the periodic table below.

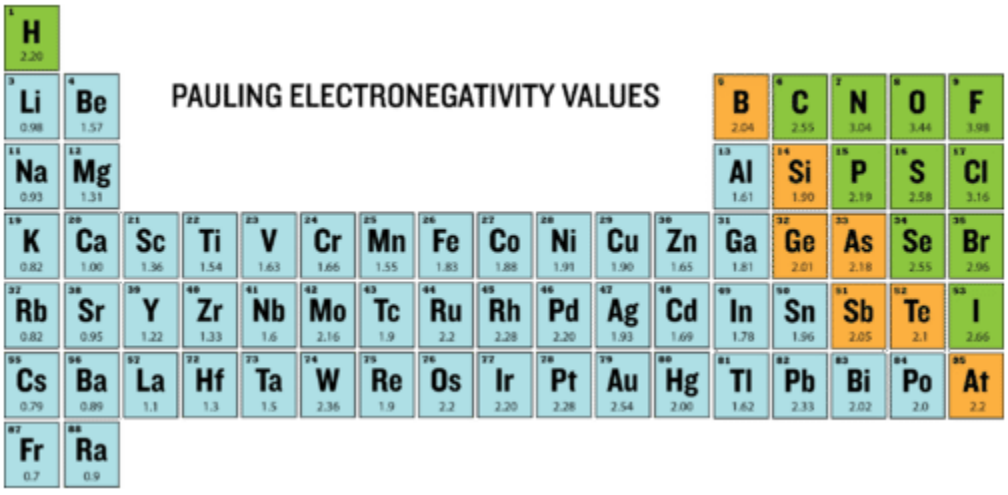


Figure 5.7.2: Electronegativities of elements. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The degree to which a given bond is ionic or covalent is determined by calculating the difference in electronegativity between the two atoms involved in the bond.

As an example, consider the bond that occurs between an atom of potassium and an atom of fluorine. Using the table, the difference in electronegativity is equal to $4.0 - 0.8 = 3.2$. Since the difference in electronegativity is relatively large, the bond between the two atoms is ionic. Since the fluorine atom has a much larger attraction for electrons than the potassium atom does, the valence electron from the potassium atom is completely transferred to the fluorine atom. The diagram below shows how difference in electronegativity relates to the ionic or covalent character of a chemical bond.

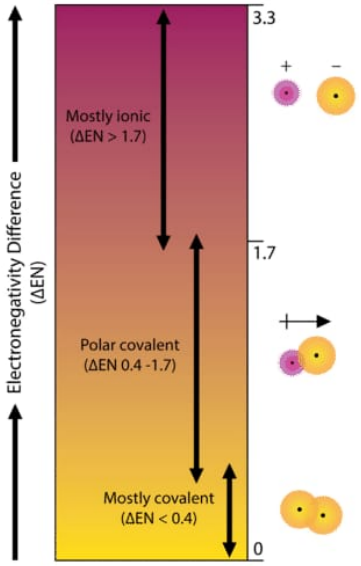


Figure 5.7.3: Bond type is predicted on the difference in electronegativity of the two elements involved in the bond. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.7 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

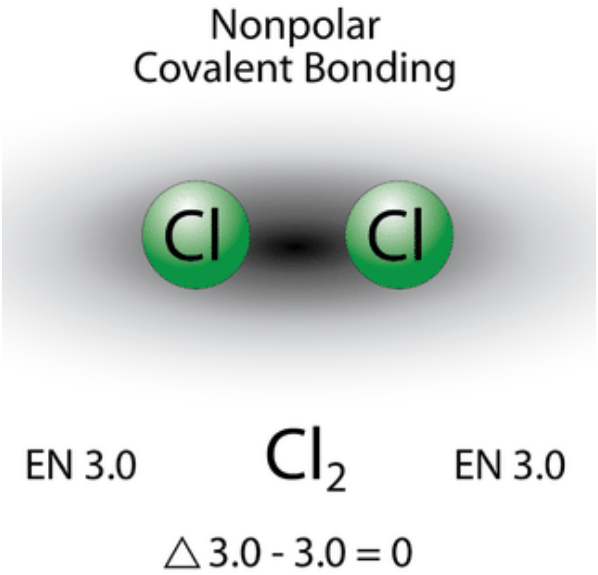


Figure 5.7.4: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the

Cl_2 molecule is symmetrical. Note that molecules in which the electronegativity difference is very small (<0.4) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta EN = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.4 and 1.7 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons, and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

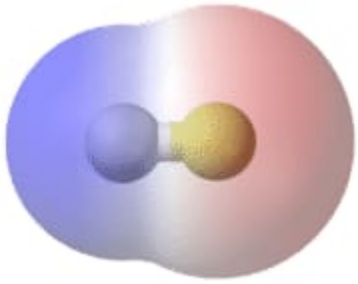


Figure 5.7.5: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Commons Wikimedia, Hydrogen fluoride elpot transparent 3D balls [commons.wikimedia.org])

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

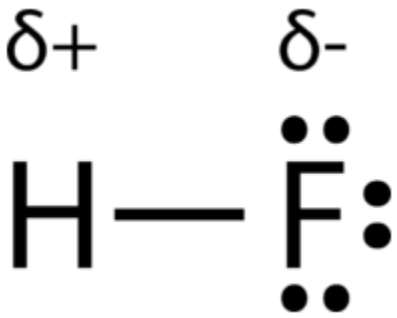


Figure 5.7.6: Use of δ to indicate partial charge. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

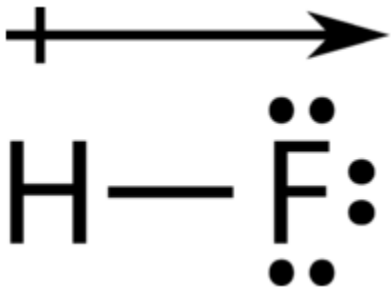


Figure 5.7.7: Use of crossed arrow to indicate polarity. (Credit: Joy Sheng Source: CK-12 Foundation; License: CC BY-NC 3.0)

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Simulation

If it wasn't for our understanding of bond polarity, we'd have a really tough time cleaning our clothing. Use this simulation to understand how the unique polarity of soap molecules helps to clean our clothing!

Summary

- The electronegativity of an atom determines how strongly it attracts electrons to itself.
- The polarity of a bond is affected by the electronegativity values of the two atoms involved in that bond.

Review

1. If two atoms bonded together have an electronegativity difference of 1.9, what is the bond type?
2. What would be the bond type for BH_2?
3. Your friend tells you that the LiF bond is covalent. Are they correct? Why or why not.

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5.8: Chemical Nomenclature

Nomenclature is the process of naming chemical compounds so that they can be easily identified as separate chemicals. The primary function of chemical nomenclature is to ensure that a spoken or written chemical name leaves no ambiguity concerning which chemical compound the name refers to—each chemical name should refer to a single substance. A less important aim is to ensure that each substance has a single name, although a limited number of alternative names is acceptable in some cases. Preferably, the name also conveys some information about the structure or chemistry of a compound. A common name will often suffice to identify a chemical compound in a particular set of circumstances. To be more generally applicable, the name should indicate at least the chemical formula. To be more specific still, the three-dimensional arrangement of the atoms may need to be specified.

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5.9: Writing Formulas for Ionic Compounds

Learning Objectives

- Write the correct formula for an ionic compound.
- Recognize polyatomic ions in chemical formulas.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like NaCl, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula $\mathrm{Na_2S}$. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charges to cancel each other out.

Example 5.9.1: Aluminum Nitride and Lithium Oxide

Write the formulas for aluminum nitride and lithium oxide.

Solution

Solution to Example 5.5.1

	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] [redacted]	$\mathrm{Li^+}$ [redacted]
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	$\mathrm{Al_1N_1}$	$\mathrm{Li_2O_1}$
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	$\mathrm{Li_2O}$

An alternative way to writing a correct formula for an ionic compound is to use the **crisscross method**. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped.

Example 5.9.2: The Crisscross Method for Lead (IV) Oxide

Write the formula for lead (IV) oxide.

Solution

Solution to Example 5.5.2

Crisscross Method	Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

Exercise 5.9.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a: available at the end of this section.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.

Example 5.9.3: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Solution to Example 5.5.3

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Na^+ [redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all subscripts that are 1.	Na_2S

Exercise 5.9.3

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a: available at the end of this section.

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1– charge. Table 5.9.1 lists the most common polyatomic ions.

Table 5.9.1: Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

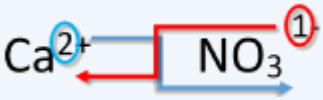
Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 5.9.4: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Solution to Example 5.5.4

Crisscross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 5.9.5

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion.

Solution

Solution to Example 5.5.5

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2–. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 5.9.5

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a: available at the end of this section.

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds.

Method 1

Compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 [or 2A] metal) and a nonmetallic element (bromine, a group 17 [or 7A] nonmetal). Therefore, it is most likely an ionic compound (in fact, it *is* ionic). In contrast, the compound NO_2 contains two elements that are both nonmetals

(nitrogen, from group 15 [or 5A], and oxygen, from group 16 [or 6A]. It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Method 2

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $\mathrm{Ba(NO_3)_2}$, you may recognize the "$\mathrm{NO_3}$" part as the nitrate ion, $\mathrm{NO_3^-}$. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the $\mathrm{Ba^{2+}}$ ion, with the 2+ charge balancing the overall 2– charge from the two nitrate ions. Thus, this compound is also ionic.

Example 5.9.6

Identify each compound as ionic or not ionic.

a. $\mathrm{Na_2O}$
b. $\mathrm{PCl_3}$
c. $\mathrm{NH_4Cl}$
d. $\mathrm{OF_2}$

Solution

Solution to Example 5.5.6

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, $\mathrm{Na_2O}$ is expected to be ionic via method 1.	$\mathrm{Na_2O}$, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, $\mathrm{PCl_3}$ is not ionic via method 1	$\mathrm{PCl_3}$, not ionic
c. The $\mathrm{NH_4}$ in the formula represents the ammonium ion, $\mathrm{NH_4^+}$, which indicates that this compound is ionic via method 2	$\mathrm{NH_4Cl}$, ionic
d. Both oxygen and fluorine are nonmetals. Therefore, $\mathrm{OF_2}$ is not ionic via method 1	$\mathrm{OF_2}$, not ionic

Exercise 5.9.6

Identify each compound as ionic or not ionic.

a. $\mathrm{N_2O}$
b. $\mathrm{FeCl_3}$
c. $\mathrm{(NH_4)_3PO_4}$
d. $\mathrm{SOCl_2}$

Answer a: available at the end of this section.

Summary

Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

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Answers

Exercise 5.9.2

CaO

CuS

Cu_2S

Exercise 5.9.3

NaBr

LiCl

MgO

Exercise 5.9.5

$MgCO_3$

$Al(CH_3COO)_3$

Exercise 5.9.6

not ionic

ionic

ionic

not ionic

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5.10: Naming Ionic Compounds

Learning Objectives

- To use the rules for naming ionic compounds.

After learning a few more details about the names of individual ions, you will be one step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system**, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 5.10.1*: The Modern and Common System of Cation Names*

Element	Stem	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 5.10.1 lists the elements that use the common system, along with their respective cation names.

Table 5.10.2: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 5.10.2 lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

Example 5.10.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. SO_3^{2-}
d. NH_4^+
e. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

Exercise 5.10.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. SO_4^{2-}
d. Ba^{2+}
e. HCO_3^-

Answer a: available at the end of this section.

Example 5.10.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Solution

a. Br^{-}
b. PO_4^{3-}
c. Cu^{2+}
d. Mg^{2+}

Exercise 5.10.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a: available at the end of this section.

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 5.10.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

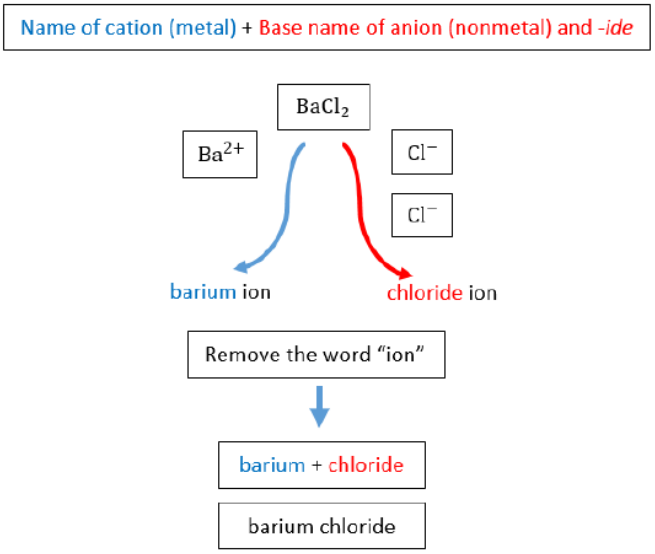


Figure 5.10.1: Naming $BaCl_2$
Naming formula: Name of metal cation + base
name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

Example 5.10.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

Exercise 5.10.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

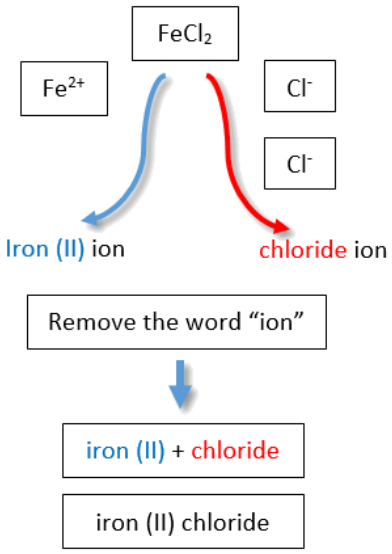
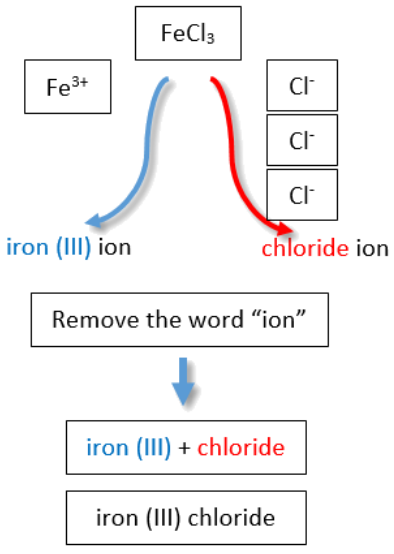
Answer a: available at the end of this section.

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the

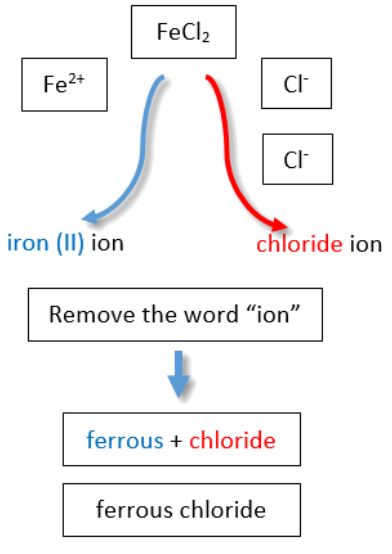
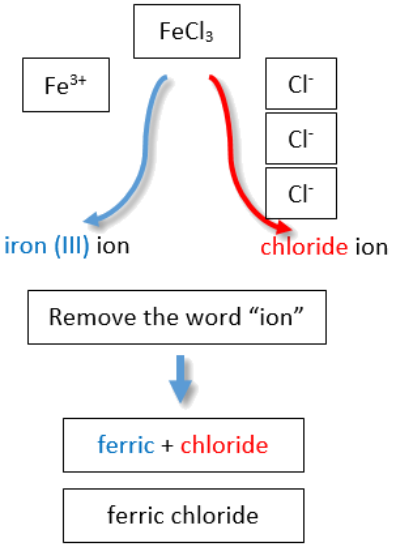
iron ion has a 2+ charge because there are two Cl^- ions in the formula (1– charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 5.10.2).

Table 5.10.3: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.

Name of cation (metal) + (Roman Numeral in parenthesis) + Base name of anion (nonmetal) and -ide	
	

If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 5.10.3) .

Table 5.10.4: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Old/Common System.

"Old" base name of cation (metal) and -ic or -ous + Base name of anion (nonmetal) and -ide	
-ous (for ions with lower charge)	-ic (for ions with higher charge)
	

Example 5.10.4:

Name each ionic compound.

a. Co_2O_3
b. $FeCl2$

Solution

Solutions to Example 5.7.4

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2– charge, so with three oxide ions, we have a total negative charge of 6–. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloride always has a 1– charge, so with two chloride ions, we have a total negative charge of 2–. • This means that the one iron ion must have a 2+ charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

Exercise 5.10.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a: available at the end of this section.

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium sulfate compound in Figure 5.10.6.

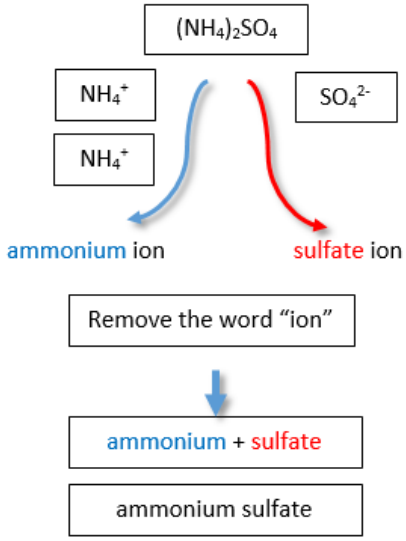


Figure 5.10.2: Naming Ionic Compounds with Polyatomic Ions

Example 5.10.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. (NH4)2S
b. AlPO4,
c. Fe3(PO4)2

Solution

Solutions to Example 5.7.5

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2– charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound’s name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6–, two phosphate ions are needed . The compound’s name is iron(II) phosphate.	iron(II) phosphate

Exercise 5.10.5A

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a: available at the end of this section.

Figure 5.10.1 is a synopsis of how to name simple ionic compounds.

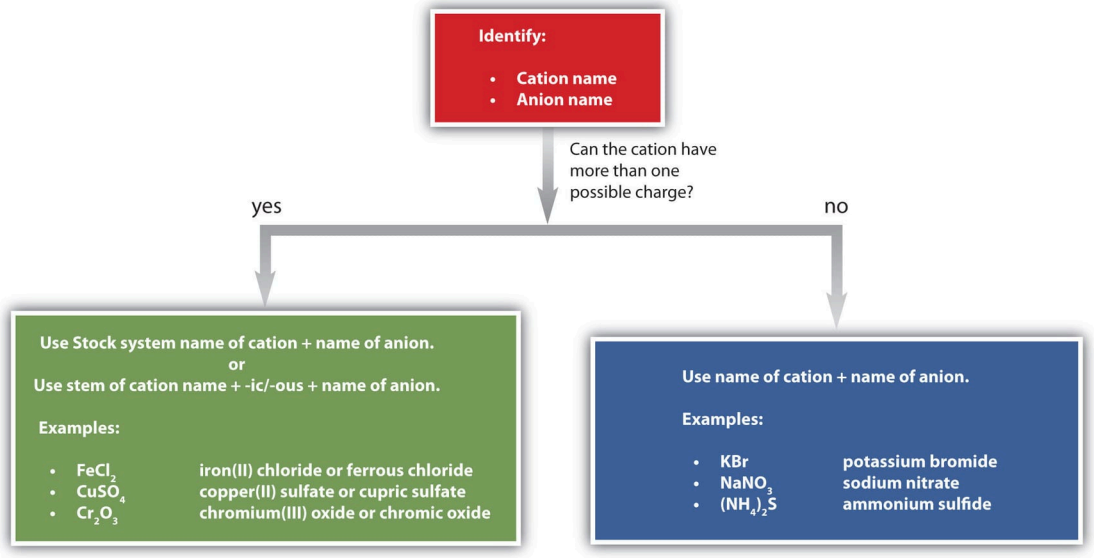


Figure 5.10.3: A Guide to Naming Simple Ionic Compounds.

Exercise 5.10.5B

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

Answer e: available at the end of this section.

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.

- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parentheses.
- Ternary compounds are composed of three or more elements.

Answers

Exercise 5.10.1

iron(II) ion
iron(III) ion
sulfate ion
barium ion
hydrogen carbonate ion or bicarbonate ion

Exercise 5.10.2

F^-

CO_3^{2-}

Sn^{2+}

K^+

Exercise 5.10.3

silver iodide
magnesium oxide
calcium phosphide

Exercise 5.10.4

gold(III) chloride
lead(IV) oxide
copper(II) oxide

Exercise 5.10.5A

ammonium phosphate
cobalt(III) nitrite

Exercise 5.10.5B

zinc bromide

Exercise 5.10.5B

aluminum oxide

Exercise 5.10.5B

ammonium phosphate

Exercise 5.10.5B

gold(III) fluoride or auric fluoride

Exercise 5.10.5B

silver fluoride

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5.11: Naming Covalent Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride $(NaCl)$. Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **electrons** in such a way that a **bond** forms between a pair of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

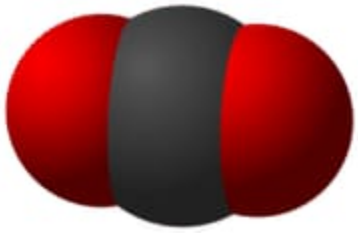


Figure 5.11.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.



Figure 5.11.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

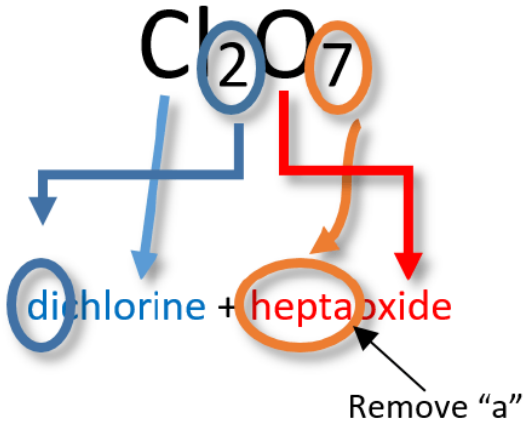
Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding

the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 5.11.1 lists these numerical prefixes.

Table 5.11.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

*This prefix is not used for the first element's name.



dichlorine heptoxide

Note

- Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetr**o**xide instead of tetr**ao**xide.
- The prefix is "mono" is not added to the first element's name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table 5.11.2.

Table 5.11.2

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the *mono-* prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono-* and the *a* of *hepta-* are dropped from the name when paired with oxide.

Exercise 5.11.1

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a: available at the end of this section.

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

Answers

Exercise 5.11.1

carbon tetrafluoride
selenium dichloride
sulfur trioxide

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5.12: Naming Acids

A spot test for gold has been in use for decades. The sample is first treated with nitric acid. Other metals may react or dissolve in this acid, but gold will not. Then the sample is added to a mixture of nitric acid and hydrochloric acid. Gold will only dissolve in this mixture. The term "acid test" arose from the California gold rush in the late 1840's when this combination was used to test for the presence of real gold. It has since come to mean, "tested and approved" in a number of fields.

Acids

An **acid** can be defined in several ways. The most straightforward definition is that an acid is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions (H^+) when dissolved in water.



Figure 5.12.1: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

This is a different type of compound than the others we have seen so far. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$\mathrm{HCl} \rightarrow \mathrm{H}^+ + \mathrm{Cl}^-$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here.

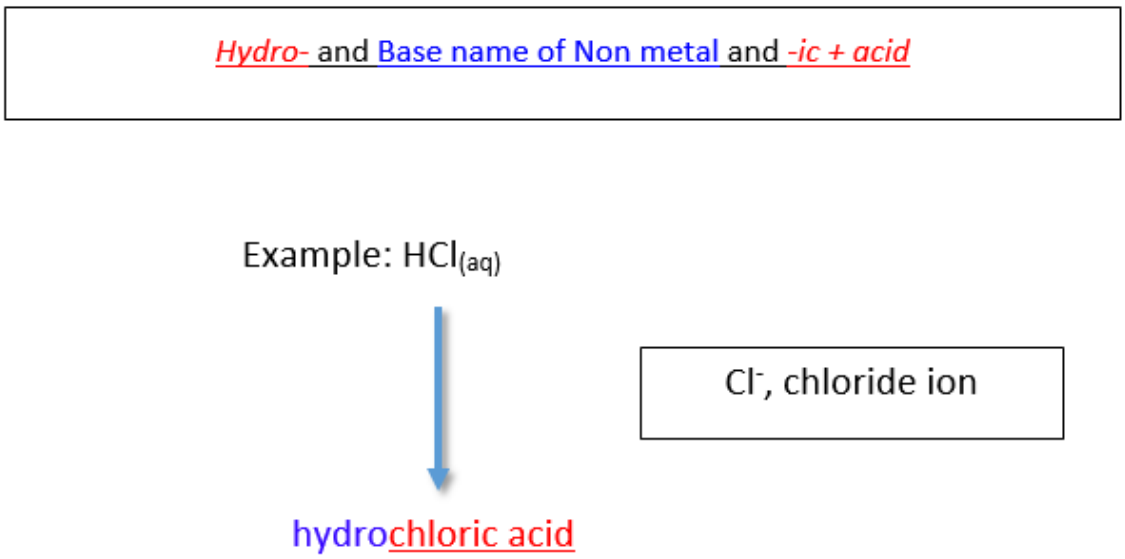
Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic.

Naming Binary acids (in aqueous form)

A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. The acid name begins with the prefix *hydro-*. followed by the base name of the anion, followed by the suffix *-ic*.

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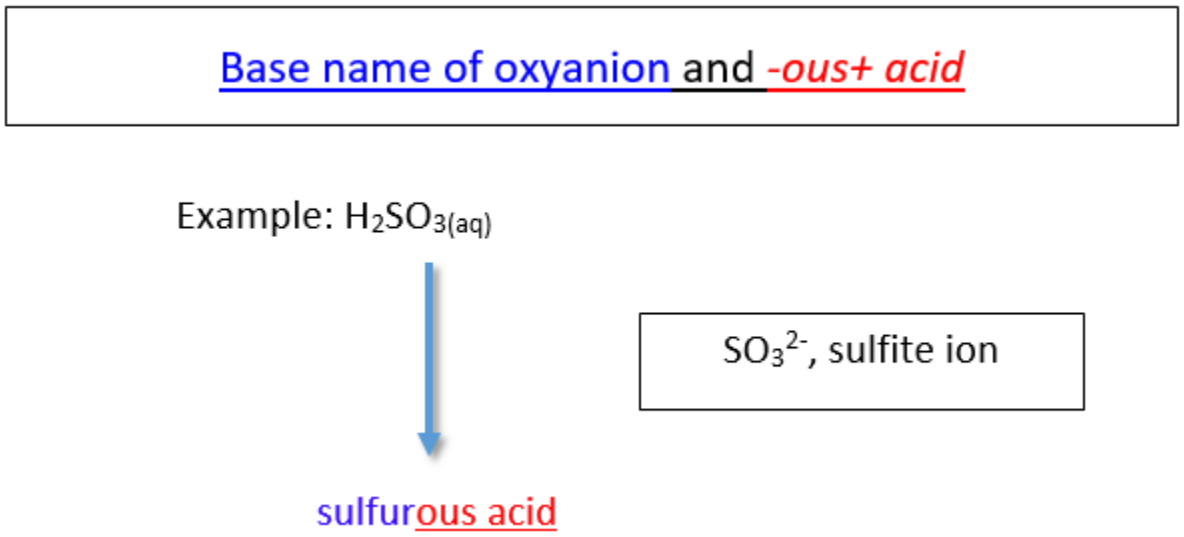
Formula for naming acids: Hydro- and Base name of nonmetal and -ic + acid. Example: HCl is hydrochloric acid.

Naming Oxyacids

An **oxyacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

a. Oxyanions with ***-ite*** ending.

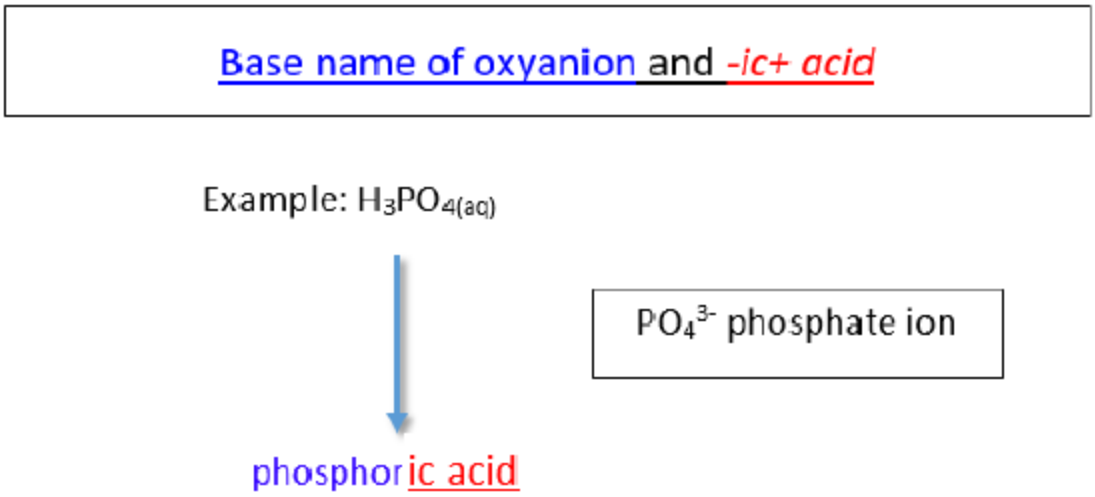
The name of the acid is the root of the anion followed by the suffix ***-ous***. There is no prefix.



Formula for naming oxyanions with -ite ending: Base name of oxyanion and -ous + acid. Example: H2SO3 is sulfurous acid.

b. Oxyanions with ***-ate*** ending.

The name of the acid is the root of the anion followed by the suffix ***-ic***. There is no prefix.



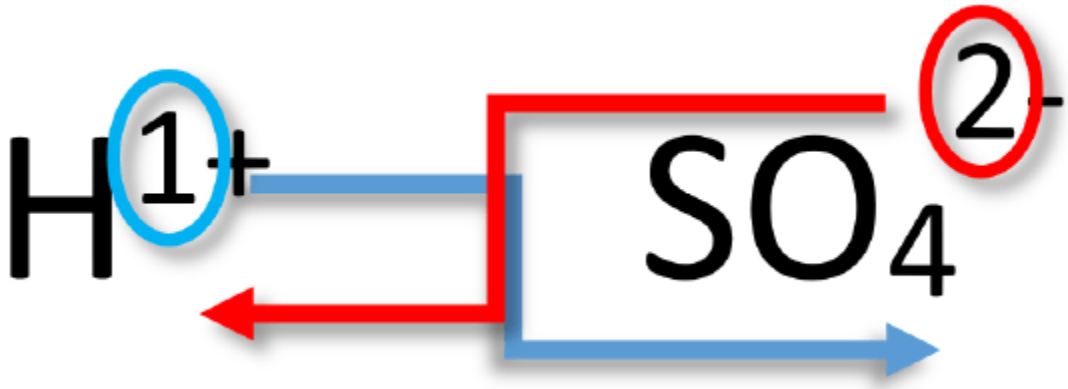
Formula for naming oxyanions with -ate ending: Base name of oxyanion and -ic + acid. Example: H3PO4 is phosphoric acid.

Note

The base name for sulfur containing oxyacid is ***sulfur-*** instead of just ***sulf-***. The same is true for a phosphorus containing oxyacid. The base name is ***phosphor-*** instead of simply ***phosph-***.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Another way to think about writing the correct formula is to utilize the crisscross method, shown below for sulfuric acid.



H2SO4: H has +1 charge and SO4 has -2 charge, so there must be 2 H+ and 1 SO4 for the charges to balance out.

Formula: H_2SO_4

Figure $5.12.2$***:*** *Crisscross approach to writing formula for sulfuric acid.*

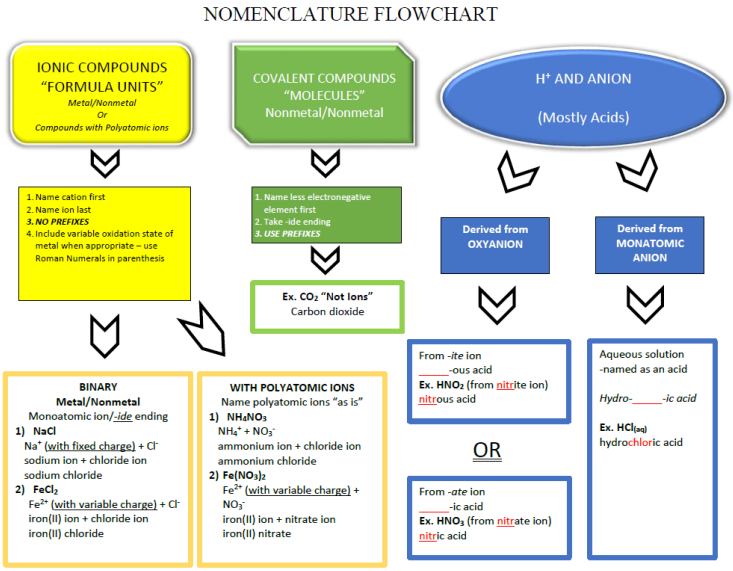
Summary

- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxyacids contain hydrogen, oxygen, and one other element.
- The name of the acid is based on the anion attached to the hydrogen.

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5.13: Nomenclature Summary



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6: THE SHAPE OF MOLECULES

6.1: VSEPR Theory
6.2: Molecule Polarity

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6.1: VSEPR Theory

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:



Figure 6.1.1: Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.

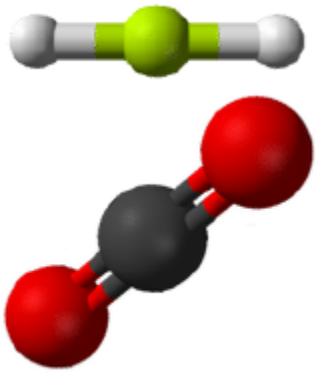


Figure 6.1.2: Beryllium hydride and carbon dioxide models. (CK12 Licence)

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

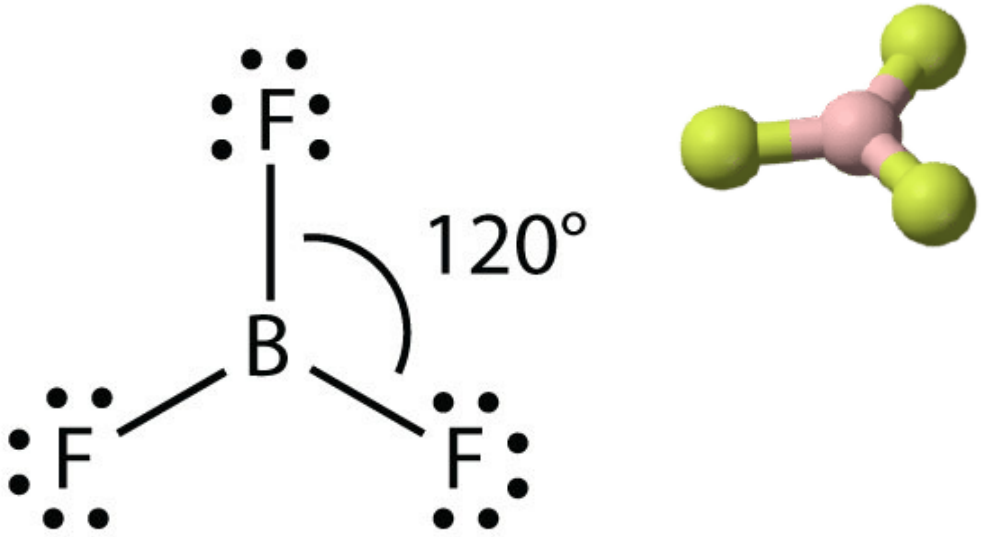


Figure 6.1.3: Boron trifluoride bonding. (CK12 Licence)

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

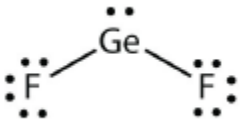


Figure 6.1.4: Germanium difluoride bonding.

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 6.1.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

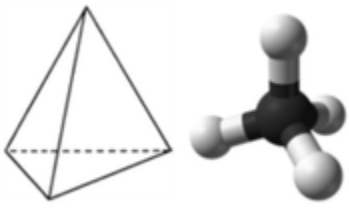


Figure 6.1.5: Tetrahedral structure of methane. (CK12 Licence)

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

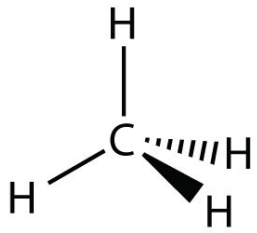


Figure 6.1.6: Methane bonding. (CK12 Licence)

NH_3 is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.

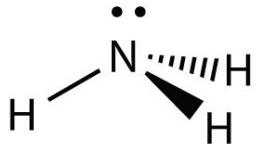


Figure 6.1.7: Ammonia bonding. (CK12 Licence)

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups, but only two of them are bonded to surrounding atoms.



Figure 6.1.8: Water bonding.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom, is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 6.1.9.



Figure 6.1.9: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

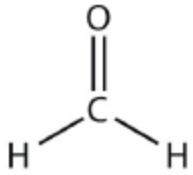


Figure 6.1.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table 6.1.1 summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table 6.1.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

Example 6.1.1

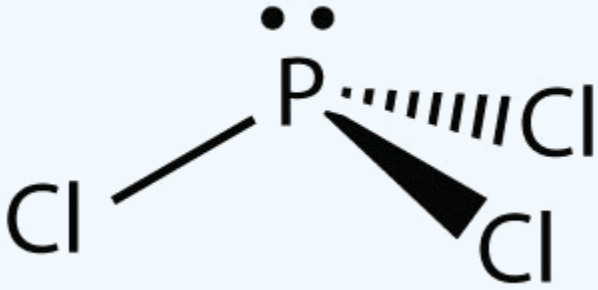
What is the approximate shape of each molecule?

a. PCl_3
b. NOF

Solution

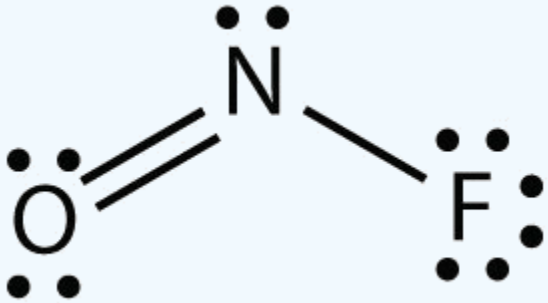
The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

Exercise 6.1.1

What is the approximate molecular shape of $\mathrm{CH_2Cl_2}$?

Answer available at the end of this section.

Exercise 6.1.2

Ethylene ($\mathrm{C_2H_4}$) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. (Hint: hydrogen is a terminal atom.)

Answer available at the end of this section.

Answer available at the end of this section.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

Answers

Exercise 6.1.1

Tetrahedral

Exercise 6.1.2

Exercise 6.1.2

Trigonal planar about both central C atoms.

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6.2: Molecule Polarity

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom, and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur), and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 6.2.1.

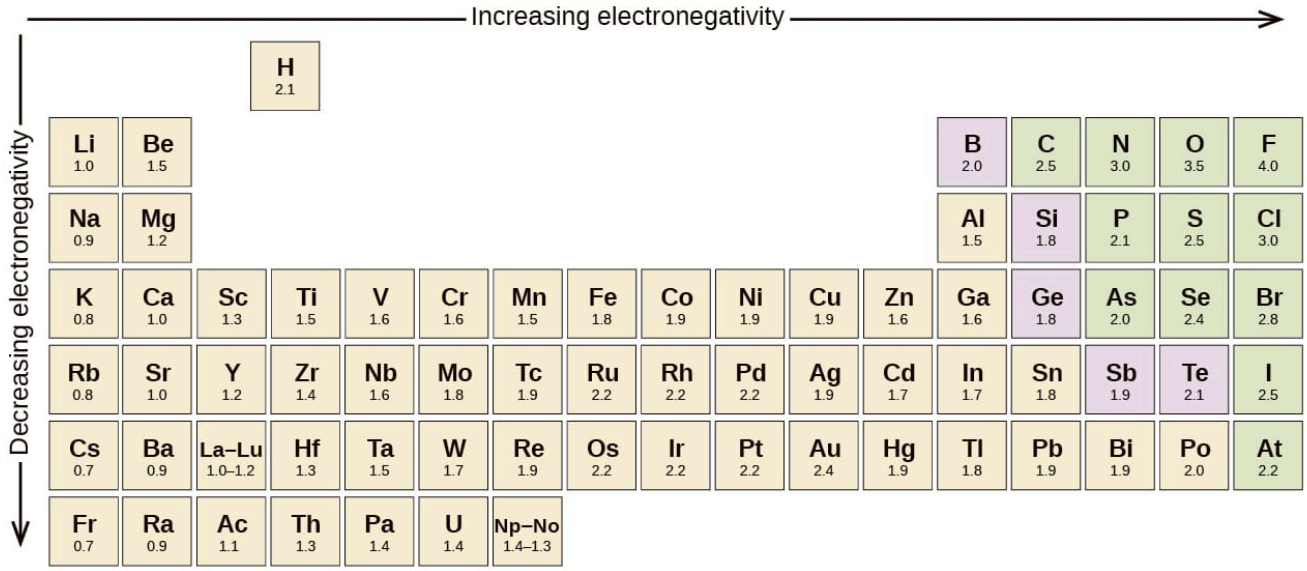


Figure 6.2.1: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

difference of the electronegativities of the two atoms involved in the covalent bond

Electronegativity Difference	Bond Type
0–0.4	pure covalent
0.5–2.0	polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

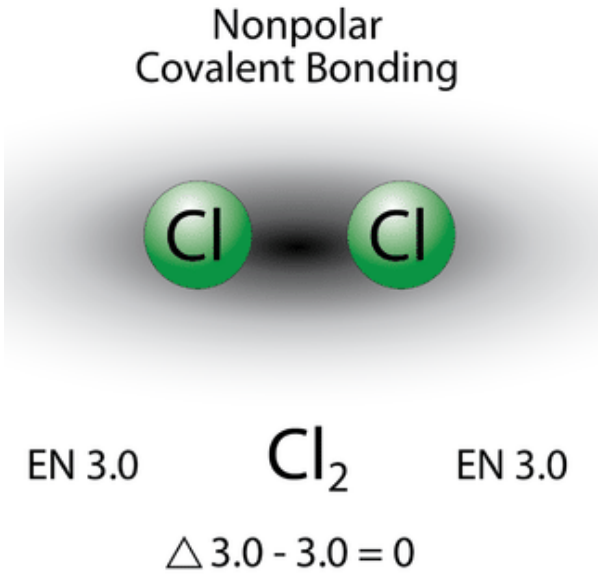


Figure 6.2.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the Cl_2 molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta EN = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

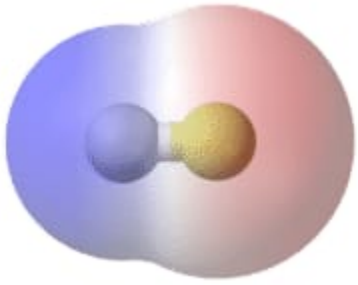


Figure 6.2.3: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

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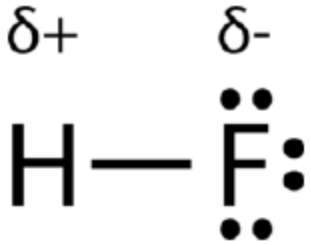


Figure 6.2.4: Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

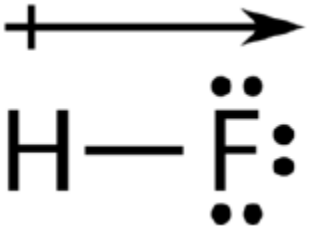


Figure 6.2.5: Use of crossed arrow to indicate polarity.

YouTube Video

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Electronegativity differences in bonding using Pauling scale. Differences in electronegativity classify bonds as covalent, polar covalent, or ionic.

Example 6.2.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 6.2.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 – 2.1 = 0.4. Thus we predict that this bond will be nonpolar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 – 2.1 = 1.4, so we predict that this bond will be polar covalent.

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Exercise 6.2.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a available at the end of this section.

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is generally useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical—bonded to the same element with no unshared pairs of electrons. **Polar molecules** are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well, as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming that you do, you can look at the structure of each one and decide if it is polar or not, *whether or not you know the individual atom's electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.



Figure 6.2.6: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $(\mathrm{CO_2})$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

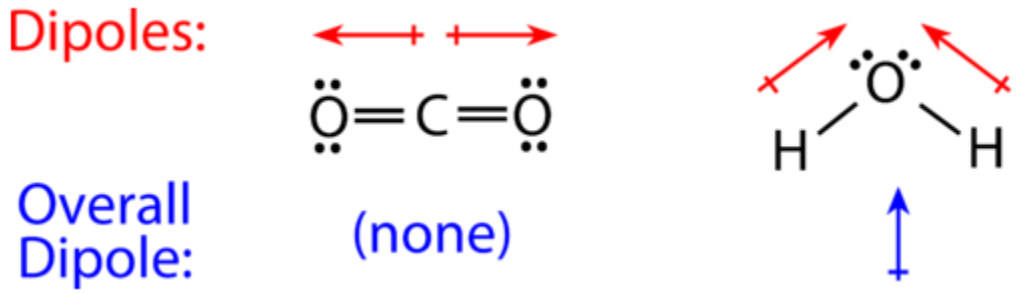


Figure 6.2.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if

one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $(\mathrm{BF_3})$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $(\mathrm{NH_3})$ is polar.

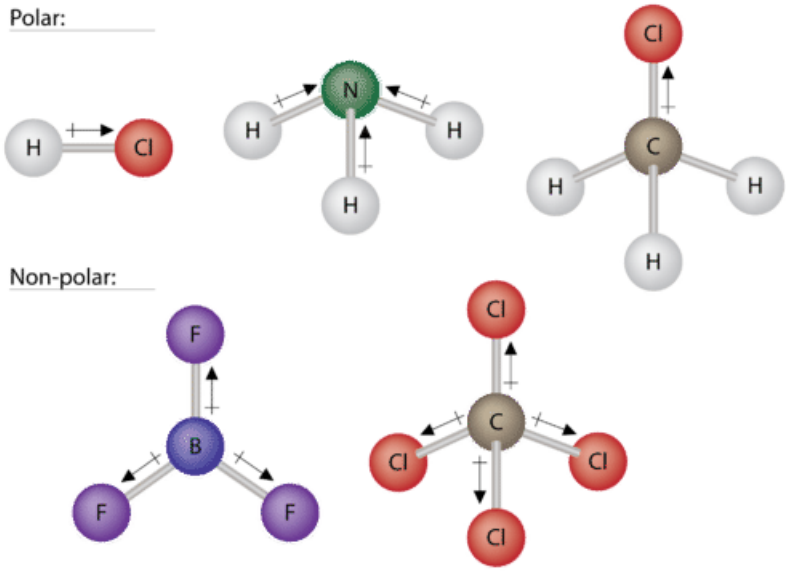


Figure $6.2.8$: Some examples of polar and nonpolar molecules based on molecular geometry.

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 do not cancel.

Steps to Identify Polar Molecules

1. Draw the Lewis structure.
2. Figure out the geometry (using VSEPR theory).
3. Visualize or draw the geometry.
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it).
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

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 $6.2.14$). 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.

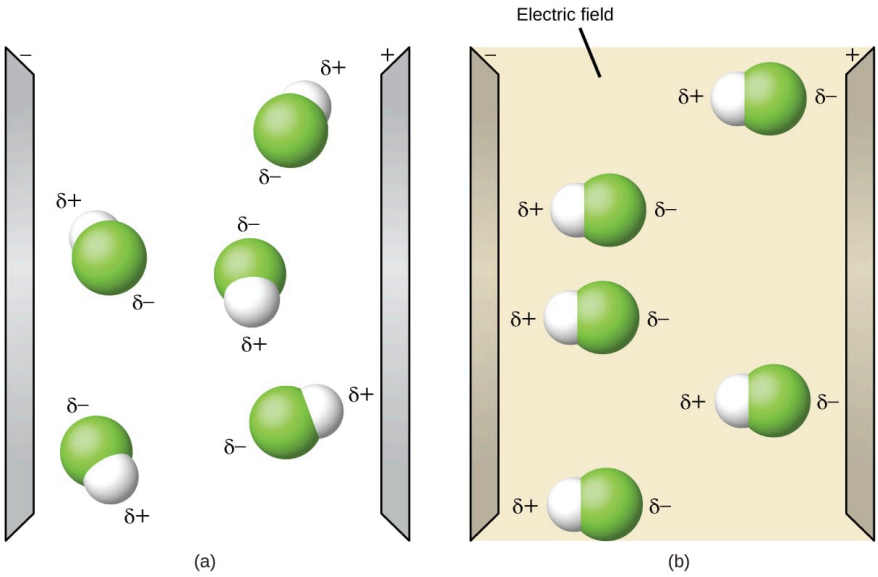


Figure 6.2.9: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Example 6.2.2:

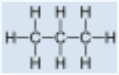
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

Exercise 6.2.2

Label each of the following as polar or nonpolar.

a. SO_3
b. NH_3

Answer a available at the end of this section.

Answer b available at the end of this section.

Contributions & Attributions

- StackExchange (thomij).

Answers

Exercise 6.2.1
likely ionic
polar covalent

Exercise 6.2.2
nonpolar

Exercise 6.2.2
polar

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7: INTERMOLECULAR FORCES

7.1: Interactions between Molecules
7.2: Properties of Liquids and Solids
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"
7.4: Types of Crystalline Solids

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7.1: Interactions between Molecules

In the winter, many people find the snow and ice beautiful; they enjoy getting out to ski or ice-skate. When the snow melts, however, the roads get very sloppy and messy. Some people look forward to spring, when the ice and snow are gone and the weather is warmer. All of these events and factors are dependent on the melting point of a solid and the freezing point of a liquid.

Melting Point

Solids are similar to liquids in that both are condensed states, with particles that are far closer together than those of a gas. However, while liquids are fluid, solids are not. The particles of most solids are packed tightly together in an orderly arrangement. The motion of individual atoms, ions, or molecules in a solid is restricted to **vibrational motion** about a fixed point. Solids are almost completely **incompressible** and are the densest of the three states of matter.

As a solid is heated, its particles vibrate more rapidly as it absorbs kinetic energy. Eventually, the organization of the particles within the solid structure begins to break down and the solid starts to melt. The **melting point** is the temperature at which a solid changes into a liquid. At its melting point, the disruptive vibrations of the particles of the solid overcome the attractive forces operating within the solid. As with boiling points, the melting point of a solid is dependent on the strength of those attractive forces. Sodium chloride (NaCl) is an ionic compound that consists of a multitude of strong ionic bonds. Sodium chloride melts at $801^{\circ}\mathrm{C}$. Ice (solid $\mathrm{H_2O}$) is a molecular compound of molecules that are held together by hydrogen bonds. Though hydrogen bonds are the strongest of the intermolecular forces, the strength of hydrogen bonds is much less than that of ionic bonds. The melting point of ice is 0 °C.



Figure 7.1.1: Melting ice cubes illustrate the process of this phase transition. (Public Domain; Moussa).

The melting point of a solid is the same as the **freezing point** of the liquid. At that temperature, the solid and liquid states of the substance are in equilibrium. For water, this equilibrium occurs at $0^{\circ}\mathrm{C}$.

$$\mathrm{H_2O}\,(s) \rightleftharpoons \mathrm{H_2O}\,(l)$$

We tend to think of solids as those materials that are solid at room temperature. However, all materials have melting points of some sort. Gases become solids at extremely low temperatures, and liquids will also become solid if the temperature is low enough. The table below gives the melting points of some common materials.

Table 7.1.1: Melting Points of Common Materials

Materials	Melting Point (°C)
Hydrogen	-259
Oxygen	-219
Diethyl ether	-116
Ethanol	-114
Water	0
Pure silver	961
Pure gold	1063
Iron	1538

Summary

The melting point is the temperature at which a solid changes into a liquid. Intermolecular forces have a strong influence on melting point.

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7.2: Properties of Liquids and Solids

Learning Objectives

- Describe the solid and liquid phases.

Solids and liquids are collectively called *condensed phases* because their particles are in virtual contact. The two states share little else, however.

Solids

In the solid state, the individual particles of a substance are in fixed positions with respect to each other because there is not enough thermal energy to overcome the intermolecular interactions between the particles. As a result, solids have a definite shape and volume. Most solids are hard, but some (like waxes) are relatively soft. Many solids composed of ions can also be quite brittle.



Figure 7.2.1: A crystalline arrangement of quartz crystal cluster. Some large crystals look the way they do because of the regular arrangement of atoms (ions) in their crystal structure. (Source: Wikipedia.)

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a crystal. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 7.2.1. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

Liquids

If the particles of a substance have enough energy to partially overcome intermolecular interactions, then the particles can move about each other while remaining in contact. This describes the liquid state. In a liquid, the particles are still in close contact, so liquids have a definite volume. However, because the particles can move about each other rather freely, a liquid has no definite shape and takes a shape dictated by its container.

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Figure 7.2.2: The formation of a spherical droplet of liquid water minimizes the surface area, which is the natural result of surface tension in liquids. (Source: Wikipedia.)

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either.

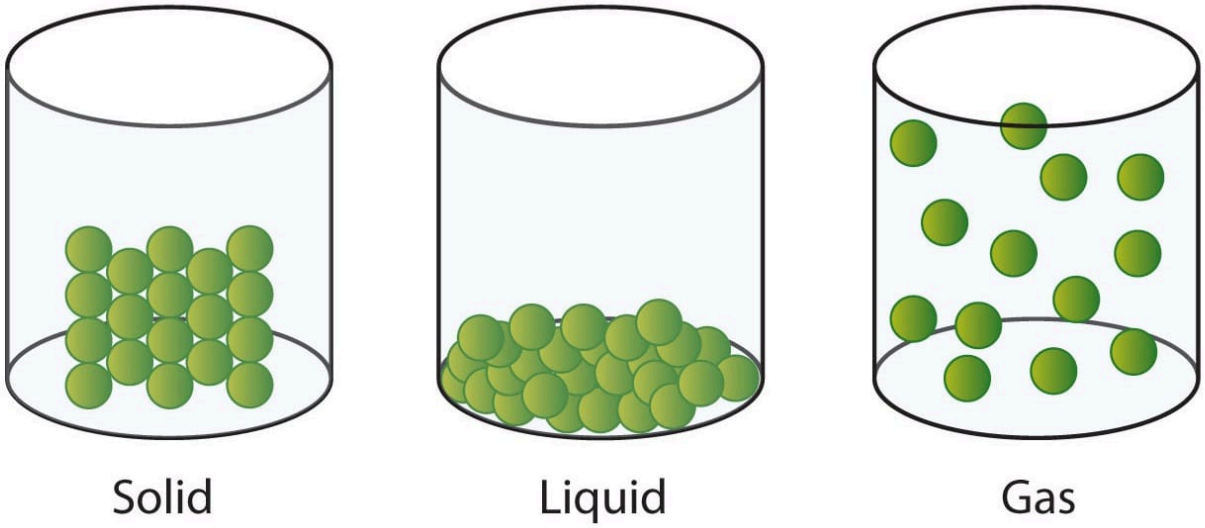


Figure 7.2.3: A Representation of the Solid, Liquid, and Gas States. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Figure 7.2.3 shows the differences among solids, liquids, and gases at the molecular level, while Table 7.2.1 lists the different characteristics of these states.

Table 7.2.1: Characteristics of the Three States of Matter

Characteristic	Solid	Liquid	Gas
shape	definite	indefinite	indefinite

Characteristic	Solid	Liquid	Gas
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

Example 7.2.1

What state or states of matter does each statement describe?

a. This state has a definite volume.
b. This state has no definite shape.
c. This state allows the individual particles to move about while remaining in contact.

Solution

a. This statement describes either the liquid state or the solid state.
b. This statement describes either the liquid state or the gas state.
c. This statement describes the liquid state.

Exercise 7.2.1

What state or states of matter does each statement describe?

a. This state has individual particles in a fixed position with regard to each other.
b. This state has individual particles far apart from each other in space.
c. This state has indefinite shape.

Answer a available at the end of this section.

Looking Closer: Water, the Most Important Liquid

Earth is the only known body in our solar system that has liquid water existing freely on its surface; life on Earth would not be possible without the presence of liquid water.

Water has several properties that make it a unique substance among substances. It is an excellent solvent; it dissolves many other substances and allows those substances to react when in solution. In fact, water is sometimes called the *universal solvent* because of this ability. Water has unusually high melting and boiling points (0°C and 100°C, respectively) for such a small molecule. The boiling points for similar-sized molecules, such as methane (BP = –162°C) and ammonia (BP = –33°C), are more than 100° lower. Though a liquid at normal temperatures, water molecules experience a relatively strong intermolecular interaction that allows them to maintain the liquid phase at higher temperatures than expected.

Unlike most substances, the solid form of water is less dense than its liquid form, which allows ice to float on water. In colder weather, lakes and rivers freeze from the top, allowing animals and plants to continue to live underneath. Water also requires an unusually large amount of energy to change temperature. While 100 J of energy will change the temperature of 1 g of Fe by 230°C, this same amount of energy will change the temperature of 1 g of H_2O by only 100°C. Thus, water changes its temperature slowly as heat is added or removed. This has a major impact on weather, as storm systems like hurricanes can be impacted by the amount of heat that ocean water can store.

Water's influence on the world around us is affected by these properties. Isn't it fascinating that such a small molecule can have such a big impact?

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Key Takeaway

- Solids and liquids are phases that have their own unique properties.

Answers

Exercise 7.2.1

solid
gas
liquid or gas

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7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"

Learning Objectives

- To describe the intermolecular forces in liquids.

The properties of liquids are intermediate between those of gases and solids, but are more similar to solids. In contrast to *intra*molecular forces, such as the covalent bonds that hold atoms together in molecules and polyatomic ions, *inter*molecular forces hold molecules together in a liquid or solid. Intermolecular forces are generally much weaker than covalent bonds. For example, it requires 927 kJ to overcome the intramolecular forces and break both O–H bonds in 1 mol of water, but it takes only about 41 kJ to overcome the intermolecular attractions and convert 1 mol of liquid water to water vapor at 100°C. (Despite this seemingly low value, the intermolecular forces in liquid water are among the strongest such forces known!) Given the large difference in the strengths of intra- and intermolecular forces, changes between the solid, liquid, and gaseous states almost invariably occur for molecular substances *without breaking covalent bonds*.

The properties of liquids are intermediate between those of gases and solids, but are more similar to solids.

Intermolecular forces determine bulk properties, such as the melting points of solids and the boiling points of liquids. Liquids boil when the molecules have enough thermal energy to overcome the intermolecular attractive forces that hold them together, thereby forming bubbles of vapor within the liquid. Similarly, solids melt when the molecules acquire enough thermal energy to overcome the intermolecular forces that lock them into place in the solid.

Intermolecular forces are electrostatic in nature; that is, they arise from the interaction between positively and negatively charged species. Like covalent and ionic bonds, intermolecular interactions are the sum of both attractive and repulsive components. Because electrostatic interactions fall off rapidly with increasing distance between molecules, intermolecular interactions are most important for solids and liquids, where the molecules are close together. These interactions become important for gases only at very high pressures, where they are responsible for the observed deviations from the ideal gas law at high pressures.

In this section, we explicitly consider three kinds of intermolecular interactions. There are two additional types of electrostatic interaction that you are already familiar with: the ion–ion interactions that are responsible for ionic bonding, and the ion–dipole interactions that occur when ionic substances dissolve in a polar substance such as water. The first two are often described collectively as van der Waals forces.

Dipole–Dipole Interactions

Polar covalent bonds behave as if the bonded atoms have localized fractional charges that are equal but opposite (i.e., the two bonded atoms generate a *dipole*). If the structure of a molecule is such that the individual bond dipoles do not cancel one another, then the molecule has a net dipole moment. Molecules with net dipole moments tend to align themselves so that the positive end of one dipole is near the negative end of another and vice versa, as shown in Figure $7.3.1a$.

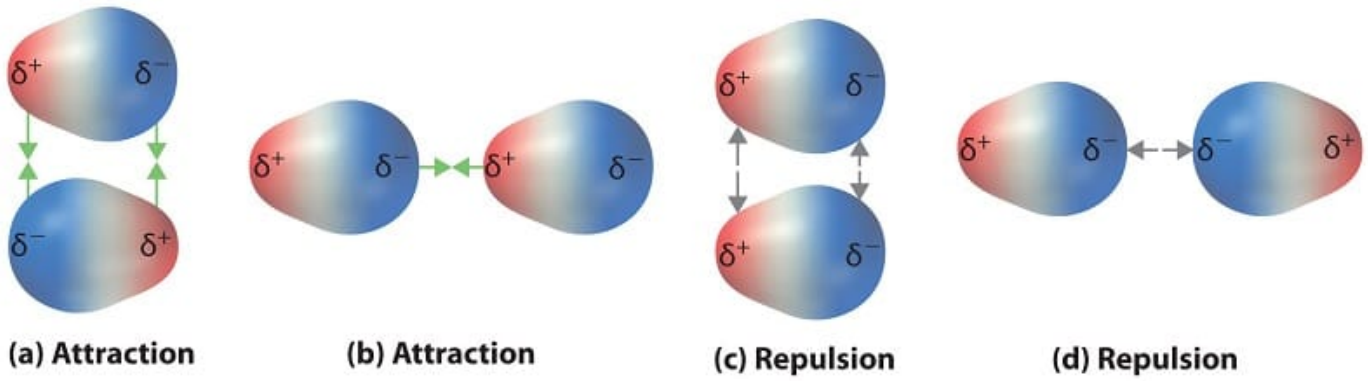


Figure 7.3.1: Attractive and Repulsive Dipole–Dipole Interactions. (a and b) Molecular orientations in which the positive end of one dipole (δ^+) is near the negative end of another (δ^-) (and vice versa) produce attractive interactions. (c and d) Molecular orientations that juxtapose the positive or negative ends of the dipoles on adjacent molecules produce repulsive interactions. (CC BY-SA-NC; anonymous)

These arrangements are more stable than arrangements in which two positive or two negative ends are adjacent (Figure 7.3.1*c*). Hence dipole–dipole interactions, such as those in Figure 7.3.1*b*, are *attractive intermolecular interactions*, whereas those in Figure 7.3.1*d* are *repulsive intermolecular interactions*. Because molecules in a liquid move freely and continuously, molecules always experience both attractive and repulsive dipole–dipole interactions simultaneously, as shown in Figure 7.3.2. On average, however, the attractive interactions dominate.

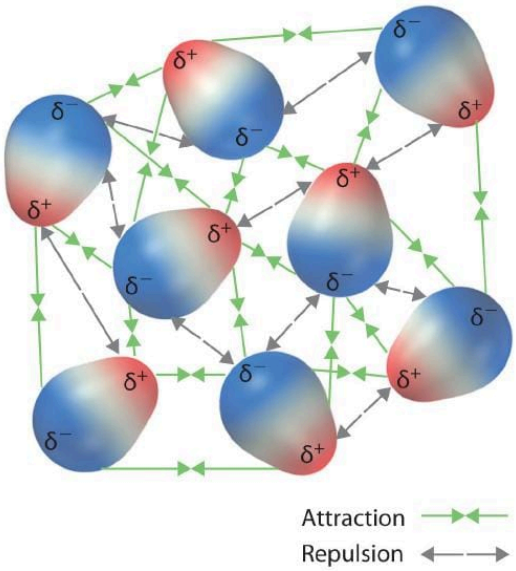


Figure 7.3.2: Both attractive and repulsive dipole–dipole interactions occur in a liquid sample with many molecules. (CC BY-SA-NC; anonymous) The green arrows pointing towards each other represent attraction. The gray arrows pointing away from each other represent repulsion

Because each end of a dipole possesses only a fraction of the charge of an electron, dipole–dipole interactions are substantially weaker than the interactions between two ions, each of which has a charge of at least ±1, or between a dipole and an ion, in which one of the species has at least a full positive or negative charge. In addition, the attractive interaction between dipoles falls off much more rapidly with increasing distance than do the ion–ion interactions. Recall that the attractive energy between two ions is proportional to $1/r$, where r is the distance between the ions. Doubling the distance ($r \rightarrow 2r$) decreases the attractive energy by one-half. In contrast, the energy of the interaction of two dipoles is proportional to $1/r^3$, so doubling the distance between the dipoles decreases the strength of the interaction by 2^3, or 8-fold. Thus a substance such as HCl, which is partially held together by dipole–dipole interactions, is a gas at room temperature and 1 atm pressure. Conversely, NaCl, which is held together by interionic interactions, is a high-melting-point solid. Within a series of compounds of similar molar mass, the strength of the intermolecular interactions increases as the dipole moment of the molecules increases, as shown in Table 7.3.1.

Table 7.3.1: Relationships Between the Dipole Moment and the Boiling Point for Organic Compounds of Similar Molar Mass

Compound	Molar Mass (g/mol)	Dipole Moment (D)	Boiling Point (K)
C_3H_6 (cyclopropane)	42	0	240
CH_3OCH_3 (dimethyl ether)	46	1.30	248
CH_3CN (acetonitrile)	41	3.9	355

The attractive energy between two ions is proportional to 1/r, whereas the attractive energy between two dipoles is proportional to 1/r6.

YouTube Video

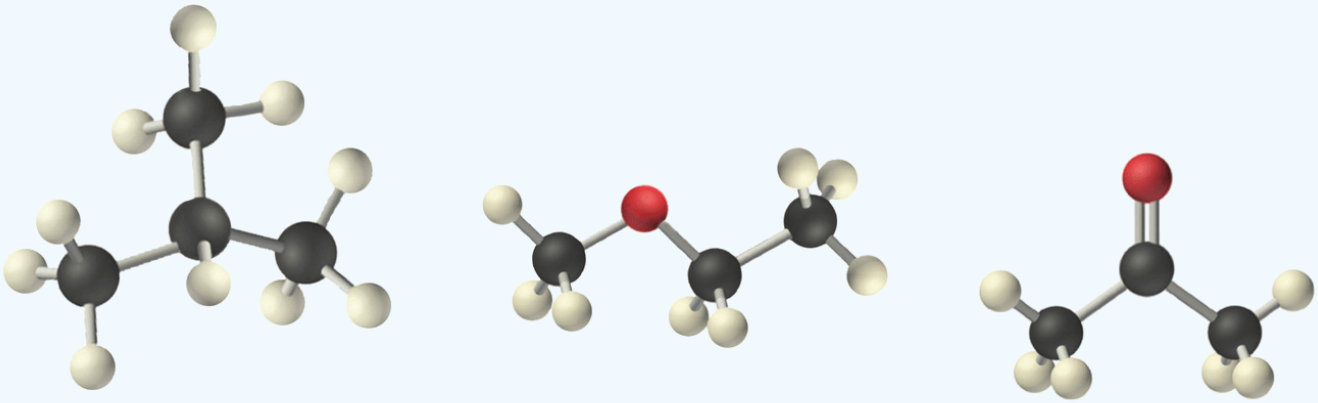
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Video Discussing Dipole Intermolecular Forces. Source: Dipole Intermolecular Force, YouTube(opens in new window) [youtu.be]

Example 7.3.1

Arrange ethyl methyl ether ($CH_3OCH_2CH_3$), 2-methylpropane [isobutane, $(CH_3)_2CHCH_3$], and acetone (CH_3COCH_3) in order of increasing boiling points. Their structures are as follows:



2-Methylpropane Ethyl methyl ether Acetone

Given: compounds.

Asked for: order of increasing boiling points.

Strategy:

Compare the molar masses and the polarities of the compounds. Compounds with higher molar masses and that are polar will have the highest boiling points.

Solution:

The three compounds have essentially the same molar mass (58–60 g/mol), so we must look at differences in polarity to predict the strength of the intermolecular dipole–dipole interactions and thus the boiling points of the compounds.

The first compound, 2-methylpropane, contains only C–H bonds, which are not very polar because C and H have similar electronegativities. It should therefore have a very small (but nonzero) dipole moment and a very low boiling point.

Ethyl methyl ether has a structure similar to H_2O; it contains two polar C–O single bonds oriented at about a 109° angle to each other, in addition to relatively nonpolar C–H bonds. As a result, the C–O bond dipoles partially reinforce one another and generate a significant dipole moment that should give a moderately high boiling point.

Acetone contains a polar C=O double bond oriented at about 120° to two methyl groups with nonpolar C–H bonds. The C–O bond dipole therefore corresponds to the molecular dipole, which should result in both a rather large dipole moment and a high boiling point.

Thus we predict the following order of boiling points:

2-methylpropane < ethyl methyl ether < acetone

This result is in good agreement with the actual data: 2-methylpropane, boiling point = –11.7°C, and the dipole moment (μ) = 0.13 D; methyl ethyl ether, boiling point = 7.4°C and μ = 1.17 D; acetone, boiling point = 56.1°C and μ = 2.88 D.

Exercise 7.3.1

Arrange carbon tetrafluoride (CF_4), ethyl methyl sulfide ($CH_3SC_2H_5$), dimethyl sulfoxide [$(CH_3)_2S{=}O$], and 2-methylbutane [isopentane, $(CH_3)_2CHCH_2CH_3$] in order of decreasing boiling points.

Answer available at the end of this section.

London Dispersion Forces

Thus far, we have considered only interactions between polar molecules. Other factors must be considered to explain why many nonpolar molecules, such as bromine, benzene, and hexane, are liquids at room temperature; why others, such as iodine and naphthalene, are solids. Even the noble gases can be liquefied or solidified at low temperatures, high pressures, or both (Table 7.3.2).

What kind of attractive forces can exist between nonpolar molecules or atoms? This question was answered by Fritz London (1900–1954), a German physicist who later worked in the United States. In 1930, London proposed that temporary fluctuations in the electron distributions within atoms and nonpolar molecules could result in the formation of short-lived instantaneous dipole moments, which produce attractive forces called London dispersion forces between otherwise nonpolar substances.

Table 7.3.2: Normal Melting and Boiling Points of Some Elements and Nonpolar Compounds

Substance	Molar Mass (g/mol)	Melting Point (°C)	Boiling Point (°C)
Ar	40	–189.4	–185.9
Xe	131	–111.8	–108.1
N_2	28	–210	–195.8
O_2	32	–218.8	–183.0
F_2	38	–219.7	–188.1
I_2	254	113.7	184.4
CH_4	16	–182.5	–161.5

Consider a pair of adjacent He atoms, for example. On average, the two electrons in each He atom are uniformly distributed

around the nucleus. Because the electrons are in constant motion, however, their distribution in one atom is likely to be asymmetrical at any given instant, resulting in an instantaneous dipole moment. As shown in part (a) in Figure 7.3.3, the instantaneous dipole moment on one atom can interact with the electrons in an adjacent atom, pulling them toward the positive end of the instantaneous dipole or repelling them from the negative end. The net effect is that the first atom causes the temporary formation of a dipole, called an induced dipole, in the second. Interactions between these temporary dipoles cause atoms to be attracted to one another. These attractive interactions are weak and fall off rapidly with increasing distance. London was able to show with quantum mechanics that the attractive energy between molecules due to temporary dipole–induced dipole interactions falls off as $1/r^6$. Doubling the distance therefore decreases the attractive energy by 2^6, or 64-fold.

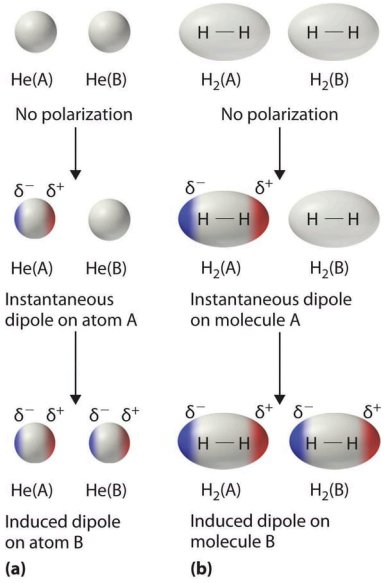


Figure 7.3.3: Instantaneous Dipole Moments. The formation of an instantaneous dipole moment on one He atom (a) or an H_2 molecule (b) results in the formation of an induced dipole on an adjacent atom or molecule.

Instantaneous dipole–induced dipole interactions between nonpolar molecules can produce intermolecular attractions just as they produce interatomic attractions in monatomic substances like Xe. This effect, illustrated for two H_2 molecules in part (b) in Figure 7.3.3, tends to become more pronounced as atomic and molecular masses increase (Table 7.3.2). For example, Xe boils at −108.1°C, whereas He boils at −269°C. The reason for this trend is that the strength of London dispersion forces is related to the ease with which the electron distribution in a given atom can be perturbed. In small atoms such as He, the two 1s electrons are held close to the nucleus in a very small volume, and electron–electron repulsions are strong enough to prevent significant asymmetry in their distribution. In larger atoms such as Xe, however, the outer electrons are much less strongly attracted to the nucleus because of filled intervening shells. As a result, it is relatively easy to temporarily deform the electron distribution to generate an instantaneous or induced dipole. The ease of deformation of the electron distribution in an atom or molecule is called its polarizability. Because the electron distribution is more easily perturbed in large, heavy species than in small, light species, we say that heavier substances tend to be much more *polarizable* than lighter ones.

For similar substances, London dispersion forces get stronger with increasing molecular size.

The polarizability of a substance also determines how it interacts with ions and species that possess permanent dipoles. Thus, London dispersion forces are responsible for the general trend toward higher boiling points with increased molecular mass and greater surface area in a homologous series of compounds, such as the alkanes (part (a) in Figure 7.3.4). The strengths of London dispersion forces also depend significantly on molecular shape because shape determines how much of one molecule can interact with its neighboring molecules at any given time. For example, part (b) in Figure 7.3.4 shows 2,2-dimethylpropane (neopentane) and *n*-pentane, both of which have the empirical formula C_5H_{12}. Neopentane is almost spherical, with a small surface area for intermolecular interactions, whereas *n*-pentane has an extended conformation that enables it to come into close contact with other *n*-pentane molecules. As a result, the boiling point of neopentane (9.5°C) is more than 25°C lower than the boiling point of *n*-pentane (36.1°C).

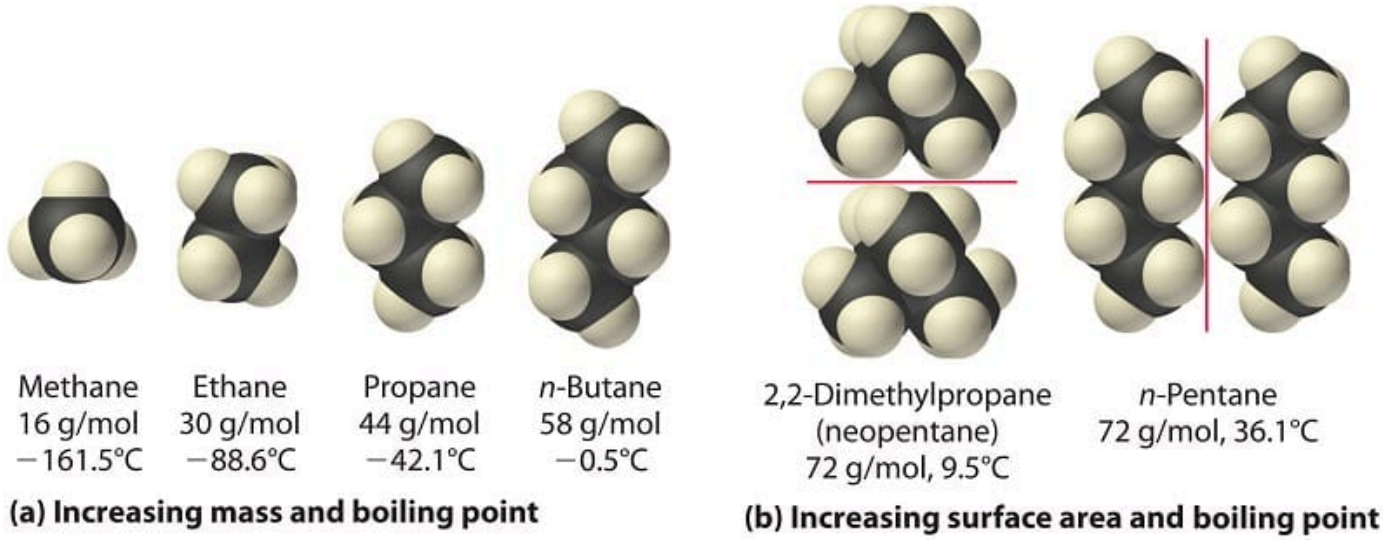


Figure 7.3.4: Mass and Surface Area Affect the Strength of London Dispersion Forces. (a) In this series of four simple alkanes, larger molecules have stronger London forces between them than smaller molecules do, and consequently have higher boiling points. (b) Linear *n*-pentane molecules have a larger surface area and stronger intermolecular forces than spherical neopentane molecules. As a result, neopentane is a gas at room temperature, whereas *n*-pentane is a volatile liquid.

All molecules, whether polar or nonpolar, are attracted to one another by London dispersion forces in addition to any other attractive forces that may be present. In general, however, dipole–dipole interactions in small polar molecules are significantly stronger than London dispersion forces, so the former predominate.

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Video Discussing London/Dispersion Intermolecular Forces. Source: Dispersion Intermolecular Force, YouTube(opens in new window) [youtu.be]

Example 7.3.2

Arrange *n*-butane, propane, 2-methylpropane [isobutene, $(CH_3)_2CHCH_3$], and *n*-pentane in order of increasing boiling points.

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Determine the intermolecular forces in the compounds, and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution:

The four compounds are alkanes and nonpolar, so London dispersion forces are the only important intermolecular forces. These forces are generally stronger with increasing molecular mass, so propane should have the lowest boiling point and *n*-pentane should have the highest, with the two butane isomers falling in between. Of the two butane isomers, 2-methylpropane is more compact, and *n*-butane has the more extended shape. Consequently, we expect intermolecular interactions for *n*-butane to be stronger due to its larger surface area, resulting in a higher boiling point. The overall order is thus as follows, with actual boiling points in parentheses: propane (−42.1°C) < 2-methylpropane (−11.7°C) < *n*-butane (−0.5°C) < *n*-pentane (36.1°C).

Exercise 7.3.2

Arrange GeH_4, $SiCl_4$, SiH_4, CH_4, and $GeCl_4$ in order of decreasing boiling points.

Answer available at the end of this section.

Hydrogen Bonds

Molecules with hydrogen atoms bonded to electronegative atoms such as O, N, and F (and to a much lesser extent, Cl and S) tend to exhibit unusually strong intermolecular interactions. These result in much higher boiling points than are observed for substances in which London dispersion forces dominate, as illustrated for the covalent hydrides of elements of groups 14–17 in Figure 7.3.5. Methane and its heavier congeners in group 14 form a series whose boiling points increase smoothly with increasing molar mass. This is the expected trend in nonpolar molecules, for which London dispersion forces are the exclusive intermolecular forces. In contrast, the hydrides of the lightest members of groups 15–17 have boiling points that are more than 100°C greater than predicted on the basis of their molar masses. The effect is most dramatic for water: if we extend the straight line connecting the points for H_2Te and H_2Se to the line for period 2, we obtain an estimated boiling point of −130°C for water! Imagine the implications for life on Earth if water boiled at −130°C rather than 100°C.

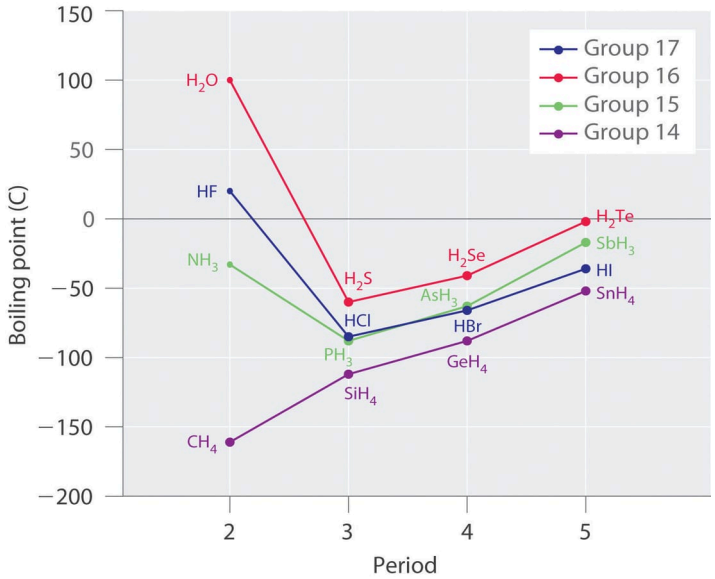


Figure 7.3.5: The Effects of Hydrogen Bonding on Boiling Points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (HF, NH_3, and H_2O) are anomalously high for compounds with such low molecular masses.

Group 14 is in purple, group 15 is in green, group 16 is red, and group 17 is blue. Graph of boiling point against period.

Why do strong intermolecular forces produce such anomalously high boiling points and other unusual properties, such as high enthalpies of vaporization and high melting points? The answer lies in the highly polar nature of the bonds between hydrogen and very electronegative elements such as O, N, and F. The large difference in electronegativity results in a large partial positive charge on hydrogen and a correspondingly large partial negative charge on the O, N, or F atom. Consequently, H–O, H–N, and H–F bonds have very large bond dipoles that can interact strongly with one another. Because a hydrogen atom is so small, these

dipoles can also approach one another more closely than most other dipoles. The combination of large bond dipoles and short dipole–dipole distances results in very strong dipole–dipole interactions called hydrogen bonds, as shown for ice in Figure 7.3.6 . A hydrogen bond is usually indicated by a dotted line between the hydrogen atom attached to O, N, or F (the *hydrogen bond donor*) and the atom that has the lone pair of electrons (the *hydrogen bond acceptor*). Because each water molecule contains two hydrogen atoms and two lone pairs, a tetrahedral arrangement maximizes the number of hydrogen bonds that can be formed. In the structure of ice, each oxygen atom is surrounded by a distorted tetrahedron of hydrogen atoms that form bridges to the oxygen atoms of adjacent water molecules. The bridging hydrogen atoms are *not* equidistant from the two oxygen atoms they connect, however. Instead, each hydrogen atom is 101 pm from one oxygen and 174 pm from the other. In contrast, each oxygen atom is bonded to two H atoms at the shorter distance and two at the longer distance, corresponding to two O–H covalent bonds and two O · · · H hydrogen bonds from adjacent water molecules, respectively. The resulting open, cagelike structure of ice means that the solid is actually slightly less dense than the liquid, which explains why ice floats on water, rather than sinks.

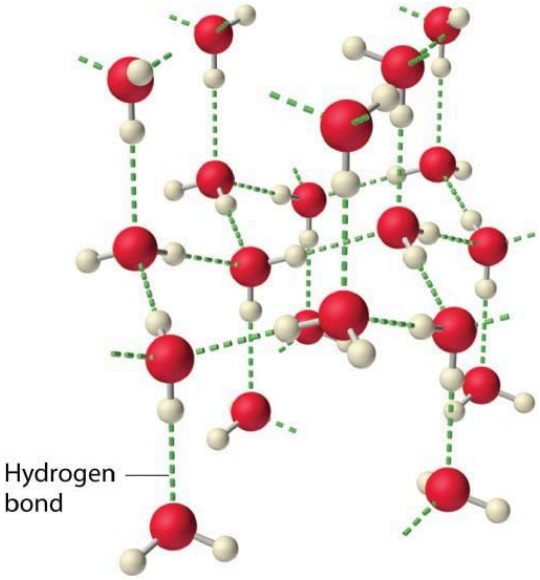


Figure 7.3.6: The Hydrogen-Bonded Structure of Ice

Each water molecule accepts two hydrogen bonds from two other water molecules and donates two hydrogen atoms to form hydrogen bonds with two more water molecules, producing an open, cage like structure. The structure of liquid water is very similar, but in the liquid, the hydrogen bonds are continually broken and formed because of rapid molecular motion.

Hydrogen bond formation requires both a hydrogen bond donor and a hydrogen bond acceptor.

Because ice is less dense than liquid water, rivers, lakes, and oceans freeze from the top down. In fact, the ice forms a protective surface layer that insulates the rest of the water, allowing fish and other organisms to survive in the lower levels of a frozen lake or sea. If ice were denser than the liquid, the ice formed at the surface in cold weather would sink as fast as it formed. Bodies of water would freeze from the bottom up, which would be lethal for most aquatic creatures. The expansion of water when freezing also explains why automobile or boat engines must be protected by “antifreeze” and why unprotected pipes in houses break if they are allowed to freeze.

YouTube Video

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Video Discussing Hydrogen Bonding Intermolecular Forces. Source: Hydrogen Bonding Intermolecular Force, YouTube(opens in

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Example 7.3.3

Considering CH_3OH, C_2H_6, Xe, and $(CH_3)_3N$, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Given: compounds

Asked for: formation of hydrogen bonds and structure

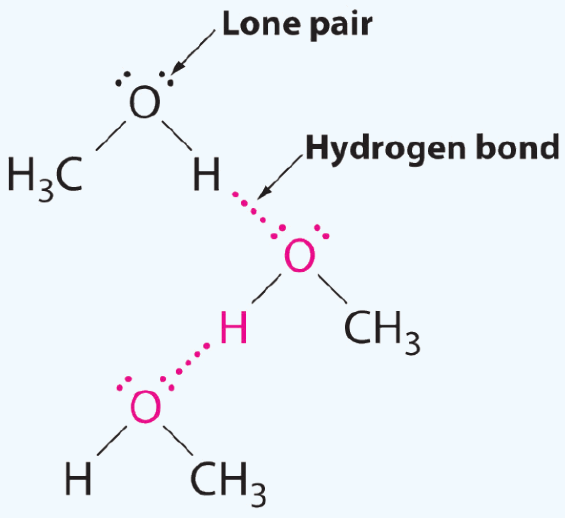
Strategy:

A. Identify the compounds with a hydrogen atom attached to O, N, or F. These are likely to be able to act as hydrogen bond donors.
B. Of the compounds that can act as hydrogen bond donors, identify those that also contain lone pairs of electrons, which allow them to be hydrogen bond acceptors. If a substance is both a hydrogen donor and a hydrogen bond acceptor, draw a structure showing the hydrogen bonding.

Solution:

A. Of the species listed, xenon (Xe), ethane (C_2H_6), and trimethylamine [$(CH_3)_3N$] do not contain a hydrogen atom attached to O, N, or F; hence they cannot act as hydrogen bond donors.

B. The one compound that can act as a hydrogen bond donor, methanol (CH_3OH), contains both a hydrogen atom attached to O (making it a hydrogen bond donor) and two lone pairs of electrons on O (making it a hydrogen bond acceptor); methanol can thus form hydrogen bonds by acting as either a hydrogen bond donor or a hydrogen bond acceptor. The hydrogen-bonded structure of methanol is as follows:



Hydrogen bonding in methanol

Exercise 7.3.3

Considering CH_3CO_2H, $(CH_3)_3N$, NH_3, and CH_3F, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Answer available at the end of this section.

Although hydrogen bonds are significantly weaker than covalent bonds, with typical dissociation energies of only 15–25 kJ/mol, they have a significant influence on the physical properties of a compound. Compounds such as HF can form only two hydrogen bonds at a time as can, on average, pure liquid NH_3. Consequently, even though their molecular masses are similar to that of water, their boiling points are significantly lower than the boiling point of water, which forms *four* hydrogen bonds at a time.

Example 7.3.4: Buckyballs

Arrange C_{60} (buckminsterfullerene, which has a cage structure), NaCl, He, Ar, and N_2O in order of increasing boiling points.

Given: compounds.

Asked for: order of increasing boiling points.

Strategy:

Identify the intermolecular forces in each compound and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution

Electrostatic interactions are strongest for an ionic compound, so we expect NaCl to have the highest boiling point. To predict the relative boiling points of the other compounds, we must consider their polarity (for dipole–dipole interactions), their ability to form hydrogen bonds, and their molar mass (for London dispersion forces). Helium is nonpolar and by far the lightest, so it should have the lowest boiling point. Argon and N_2O have very similar molar masses (40 and 44 g/mol, respectively), but N_2O is polar while Ar is not. Consequently, N_2O should have a higher boiling point. A C_{60} molecule is nonpolar, but its molar mass is 720 g/mol, much greater than that of Ar or N_2O. Because the boiling points of nonpolar substances increase rapidly with molecular mass, C_{60} should boil at a higher temperature than the other nonionic substances. The predicted order is thus as follows, with actual boiling points in parentheses:

$$\text{He } (-269^\circ\text{C}) < \text{Ar } (-185.7^\circ\text{C}) < N_2O\ (-88.5^\circ\text{C}) < C_{60}\ (>280^\circ\text{C}) < \text{NaCl } (1465^\circ\text{C}).$$

Exercise 7.3.4

Arrange 2,4-dimethylheptane, Ne, CS_2, Cl_2, and KBr in order of decreasing boiling points.

Answer available at the end of this section.

Example 7.3.5

Identify the most significant intermolecular force in each substance.

a. C_3H_8
b. CH_3OH
c. H_2S

Solution

a. Although C–H bonds are polar, they are only minimally polar. The most significant intermolecular force for this substance would be dispersion forces.
b. This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.

c. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram and VSEPR indicate that it is bent, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercise 7.3.6

Identify the most significant intermolecular force in each substance.

a. HF
b. HCl

Answer a available at the end of this section.

Summary

Intermolecular forces are electrostatic in nature and include van der Waals forces and hydrogen bonds. Molecules in liquids are held to other molecules by intermolecular interactions, which are weaker than the intramolecular interactions that hold the atoms together within molecules and polyatomic ions. Transitions between the solid and liquid, or the liquid and gas phases, are due to changes in intermolecular interactions, but do not affect intramolecular interactions. The three major types of intermolecular interactions are dipole–dipole interactions, London dispersion forces (these two are often referred to collectively as **van der Waals forces**), and hydrogen bonds. **Dipole–dipole interactions** arise from the electrostatic interactions of the positive and negative ends of molecules with permanent dipole moments; their strength is proportional to the magnitude of the dipole moment and to $1/r^3$, where r is the distance between dipoles. **London dispersion forces** are due to the formation of **instantaneous dipole moments** in polar or nonpolar molecules as a result of short-lived fluctuations of electron charge distribution, which in turn cause the temporary formation of an **induced dipole** in adjacent molecules; their energy falls off as $1/r^6$. Larger atoms tend to be more **polarizable** than smaller ones, because their outer electrons are less tightly bound and are therefore more easily perturbed. **Hydrogen bonds** are especially strong dipole–dipole interactions between molecules that have hydrogen bonded to a highly electronegative atom, such as O, N, or F. The resulting partially positively charged H atom on one molecule (the *hydrogen bond donor*) can interact strongly with a lone pair of electrons of a partially negatively charged O, N, or F atom on adjacent molecules (the *hydrogen bond acceptor*). Because of strong O · · · H hydrogen bonding between water molecules, water has an unusually high boiling point, and ice has an open, cage like structure that is less dense than liquid water.

Answers

Exercise 7.3.1

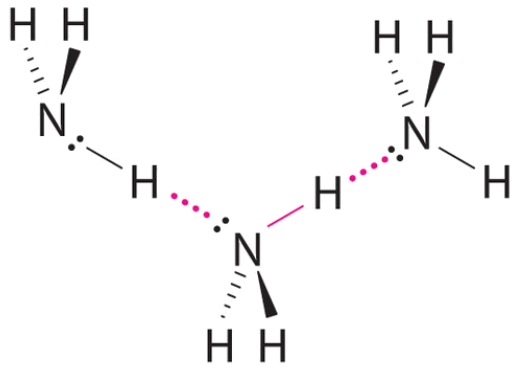
dimethyl sulfoxide (boiling point = 189.9°C) > ethyl methyl sulfide (boiling point = 67°C) > 2-methylbutane (boiling point = 27.8°C) > carbon tetrafluoride (boiling point = –128°C)

Exercise 7.3.2

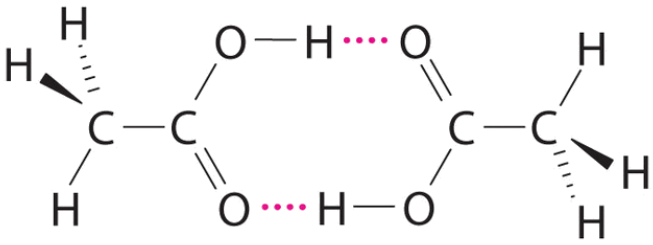
$GeCl_4$ (87°C) > $SiCl_4$ (57.6°C) > GeH_4 (–88.5°C) > SiH_4 (–111.8°C) > CH_4 (–161°C)

Exercise 7.3.3

CH_3CO_2H and NH_3;



Hydrogen bonding in ammonia



Hydrogen bonding in acetic acid

Hydrogen bonding in ammonia between nitrogen and hydrogen. hydrogen bonding in acetic acid is between oxygen and hydrogen.

Exercise 7.3.4

KBr (1435°C) > 2,4-dimethylheptane (132.9°C) > CS_2 (46.6°C) > Cl_2 (–34.6°C) > Ne (–246°C)

Exercise 7.3.6

hydrogen bonding

dipole-dipole interactions

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7.4: Types of Crystalline Solids

Learning Objectives

- Identify different types of solid substances.

As a society, we sometimes take things for granted. For example, it is often assumed that we will get electric power when we connect a plug to an electrical outlet. The wire that comprises that outlet is almost always copper, a material that conducts electricity well. The unique properties of the solid copper allow electrons to flow freely through the wire and into whatever device we connect it to. Then we can enjoy music, television, computer work, or whatever other activity we want to undertake. However, these activities—and the miracle of electricity itself—would not be possible without that copper wire!

Classes of Crystalline Solids

Crystalline substances can be described by the types of particles in them and the types of chemical bonding that take place between the particles. There are four types of crystals: (1) **ionic**, (2) **metallic**, (3) **covalent network**, and (4) **molecular**. Properties and several examples of each type are listed in the following table and are described in the table below.

Table 7.4.1: Crystalline Solids: Melting and Boiling Points

Type of Crystalline Solid	Examples (formulas)	Melting Point (°C)	Normal Boiling Point (°C)
Ionic	$NaCl$	801	1413
	CaF_2	1418	1533
Metallic	Hg	-39	630
	Na	371	883
	Au	1064	2856
	W	3410	5660
Covalent Network	B	2076	3927
	C (diamond)	3500	3930
	SiO_2	1600	2230
Molecular	H_2	-259	-253
	I_2	114	184
	NH_3	-78	-33
	H_2O	0	100

Ionic crystals - The ionic crystal structure consists of alternating positively-charged cations and negatively-charged anions (see figure below). The ions may either be monatomic or polyatomic. Generally, ionic crystals form from a combination of Group 1 or 2 metals and Group 16 or 17 nonmetals or nonmetallic polyatomic ions. Ionic crystals are hard and brittle and have high melting points. Ionic compounds do not conduct electricity as solids, but do conduct electricity when molten or in aqueous solution.

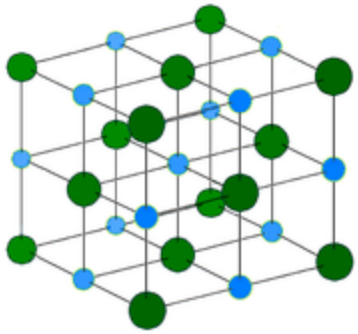


Figure 7.4.1: NaCl crystal.

Metallic crystal - Metallic crystals consist of metal cations surrounded by a "sea" of mobile valence electrons (see figure below). These electrons, also referred to as delocalized electrons, do not belong to any one atom, but are capable of moving through the entire crystal. As a result, metals are good conductors of electricity. As seen in the table above, the melting points of metallic crystals span a wide range.

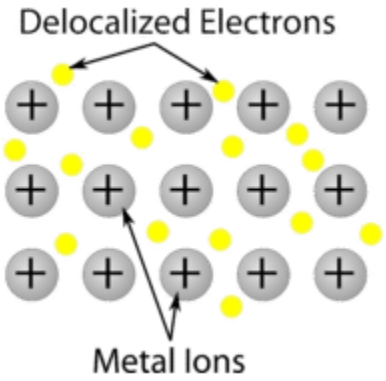


Figure 7.4.2: Metallic crystal lattice with free electrons able to move among positive metal atoms.

Covalent network crystals - A covalent network crystal consists of atoms at the lattice points of the crystal, with each atom being covalently bonded to its nearest neighbor atoms (see figure below). The covalently bonded network is three-dimensional and contains a very large number of atoms. Network solids include diamond, quartz, many metalloids, and oxides of transition metals and metalloids. Network solids are hard and brittle, with extremely high melting and boiling points. Being composed of atoms rather than ions, they do not conduct electricity in any state.

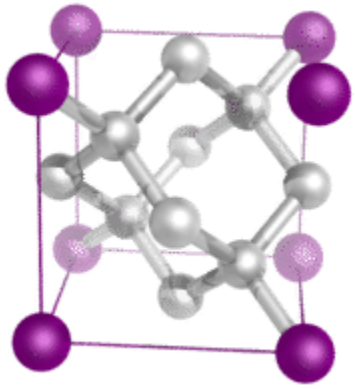


Figure 7.4.3: Diamond is a network solid and consists of carbon atoms covalently bonded to one another in a repeating three-dimensional pattern. Each carbon atom makes for single covalent bonds in a tetrahedral geometry.

Molecular crystals - Molecular crystals typically consist of molecules at the lattice points of the crystal, held together by relatively weak intermolecular forces (see figure below). The intermolecular forces may be dispersion forces in the case of nonpolar crystals, or dipole-dipole forces in the case of polar crystals. Some molecular crystals, such as ice, have molecules held together

by hydrogen bonds. When one of the noble gases is cooled and solidified, the lattice points are individual atoms rather than molecules. In all cases, the intermolecular forces holding the particles together are far weaker than either ionic or covalent bonds. As a result, the melting and boiling points of molecular crystals are much lower. Lacking ions or free electrons, molecular crystals are poor electrical conductors.

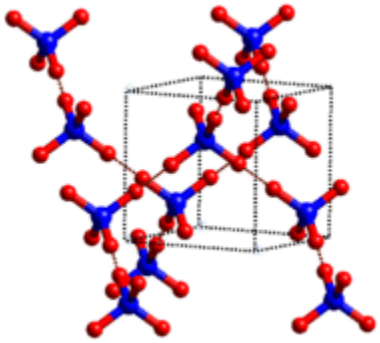


Figure 7.4.4: Ice crystal structure.

Some general properties of the four major classes of solids are summarized in Table 7.4.2.

Table 7.4.2: Properties of the Major Classes of Solids

Ionic Solids	Molecular Solids	Covalent Solids	Metallic Solids
poor conductors of heat and electricity	poor conductors of heat and electricity	poor conductors of heat and electricity*	good conductors of heat and electricity
relatively high melting point	low melting point	high melting point	melting points depend strongly on electron configuration
hard but brittle; shatter under stress	soft	very hard and brittle	easily deformed under stress; ductile and malleable
relatively dense	low density	low density	usually high density
dull surface	dull surface	dull surface	lustrous
***Many exceptions exist. For example, graphite has a relatively high electrical conductivity within the carbon planes, and diamond has the highest thermal conductivity of any known substance.**			

Example 7.4.1

Classify Ge, RbI, $C_6(CH_3)_6$, and Zn as ionic, molecular, covalent, or metallic solids and arrange them in order of increasing melting points.

Given: compounds

Asked for: classification and order of melting points

Strategy:

A. Locate the component element(s) in the periodic table. Based on their positions, predict whether each solid is ionic, molecular, covalent, or metallic.
B. Arrange the solids in order of increasing melting points based on your classification, beginning with molecular solids.

Solution:

A. **Germanium** lies in the p block just under Si, along the diagonal line of semi-metallic elements, which suggests that elemental Ge is likely to have the same structure as Si (the diamond structure). Thus Ge is probably a **covalent solid**.

RbI contains a metal from group 1 and a nonmetal from group 17, so it is **an ionic solid** containing Rb^+ and I^- ions.

The compound $C_6(CH_3)_6$ is a hydrocarbon (hexamethylbenzene), which consists of isolated molecules that stack to form **a molecular solid** with no covalent bonds between them.

Zn is a d-block element, so it is **a metallic solid**.

B. Arranging these substances in order of increasing melting points is straightforward, with one exception. We expect $C_6(CH_3)_6$ to have the lowest melting point and Ge to have the highest melting point, with RbI somewhere in between. The melting points of metals, however, are difficult to predict based on the models presented thus far. Because Zn has a filled valence shell, it should not have a particularly high melting point, so a reasonable guess is

$$C_6(CH_3)_6 < Zn < RbI < Ge\,.$$

The actual melting points are $C_6(CH_3)_6$, 166°C; Zn, 419°C; RbI, 642°C; and Ge, 938°C. This agrees with our prediction.

Exercise 7.4.1

Classify CO_2, $BaBr_2$, GaAs, and AgZn as ionic, covalent, molecular, or metallic solids and then arrange them in order of increasing melting points.

Answer available at the end of this section.

Summary

- Ionic crystals are composed of alternating positive and negative ions.
- Metallic crystals consist of metal cations surrounded by a "sea" of mobile valence electrons.
- Covalent crystals are composed of atoms which are covalently bonded to one another.
- Molecular crystals are held together by weak intermolecular forces.

Answers

Exercise 7.4.1

CO_2 (molecular) < AgZn (metallic) ~ $BaBr_2$ (ionic) < GaAs (covalent).

The actual melting points are: CO_2, about -15.6°C; AgZn, about 700°C; $BaBr_2$, 856°C; and GaAs, 1238°C.

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8: CHEMICAL REACTIONS

8.1: Evidence of a Chemical Reaction
8.2: The Chemical Equation
8.3: How to Write Balanced Chemical Equations
8.4: Classifying Chemical Reactions
8.5: Aqueous Solutions and Solubility
8.6: Precipitation Reactions
8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations
8.8: Acid–Base Reactions

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8.1: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

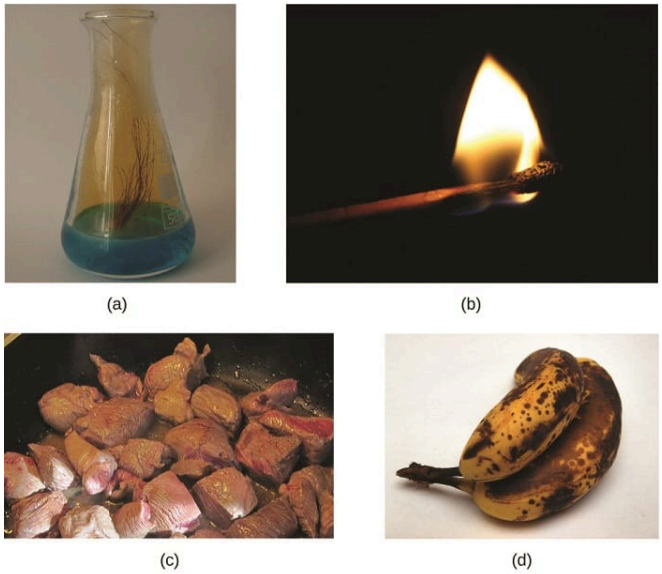


Figure 8.1.1: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes 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.)

To identify a chemical reaction, we look for a **chemical change**. 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 8.1.1).

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 8.1.1***:*** *Evidence of a Chemical Reaction*

Example 8.1.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

a. Freezing liquid mercury.
b. Adding yellow to blue to make green.
c. Cutting a piece of paper into two pieces.
d. Dropping a sliced orange into a vat of sodium dydroxide.
e. Filling a balloon with natural air.

Solution

A, B, C, & E involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is D.

Exercise 8.1.1

Which of the following is a chemical reaction?

a. Painting a wall blue.
b. A bicycle rusting.
c. Ice cream melting.
d. Scratching a key across a desk.
e. Making a sand castle.

Answer available at the end of this section.

Example 8.1.2: Evidence of a Chemical Reaction

Which of the following is not a chemical reaction?

a. Shattering glass with a baseball.
b. Corroding metal.
c. Fireworks exploding.
d. Lighting a match.
e. Baking a cake.

Solution

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is A.

Exercise 8.1.2

Which of the following is NOT a chemical reaction?

a. Frying an egg.
b. Slicing carrots.
c. A Macbook falling out of a window.
d. Creating ATP in the human body.
e. Dropping a fizzy tablet into a glass of water.

Answer available at the end of this section.

Summary

Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties.

Answers

Exercise 8.1.1

B

Exercise 8.1.2

B and C

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8.2: The Chemical Equation

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that—what is produced—or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\text{Flour} + \text{Butter} + \text{Sugar} + \text{Vanilla} + \text{Baking Soda} + \text{Eggs} + \text{Chocolate Chips}}_{\text{Ingredients} = \text{Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\mathrm{SO_2} + \mathrm{O_2}$, are reactants and sulfur trioxide, $\mathrm{SO_3}$, is the product.

$$\underbrace{2\,\mathrm{SO_2\,(g)} + \mathrm{O_2\,(g)}}_{\text{Reactants}} \rightarrow \underbrace{2\,\mathrm{SO_3\,(g)}}_{\text{Products}}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants} \rightarrow \text{Products}$$

There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table $8.2.1$ shows a listing of symbols used in chemical equations.

Table $8.2.1$: Symbols Used in Chemical Equations

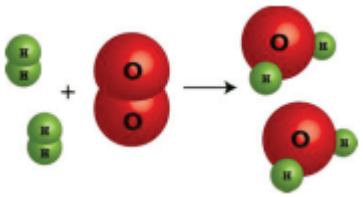
Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state

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Symbol	Description	Symbol	Description
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:
"*Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor.*"

3. They could write the equation in chemical shorthand.

$$2\text{H}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{H}_2\text{O}\,(g)$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$\text{Ca}(\text{NO}_3)_2\,(aq) + 2\text{NaOH}\,(aq) \rightarrow \text{Ca}(\text{OH})_2\,(s) + 2\text{NaNO}_3\,(aq)$$

How much easier is that to read? Let's try it in reverse. Look at the following reaction in shorthand and write the word equation for the reaction:

$$\text{Cu}\,(s) + \text{AgNO}_3\,(aq) \rightarrow \text{Cu}(\text{NO}_3)_2\,(aq) + \text{Ag}\,(s)$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).
3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table 8.2.1*: Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine

Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

Example 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
b. Gaseous propane, $\mathrm{C_3H_8}$, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane ($\mathrm{C_3H_8}$) and oxygen ($\mathrm{O_2}$)

Product: carbon dioxide ($\mathrm{CO_2}$) and water ($\mathrm{H_2O}$)

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g)$$

Exercise 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.
b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{LiCl}\,(aq) + \mathrm{H_2O}\,(l)$
c. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a available at the end of this section.

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

Answers

Exercise 8.2.1

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

$Cu(s) + O_2(g) \rightarrow CuO(s)$

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8.3: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products; and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

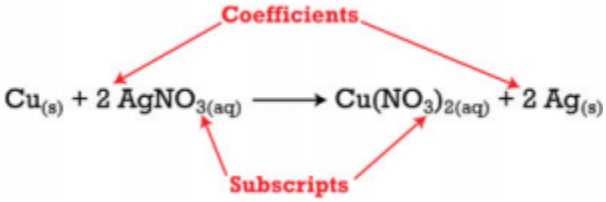
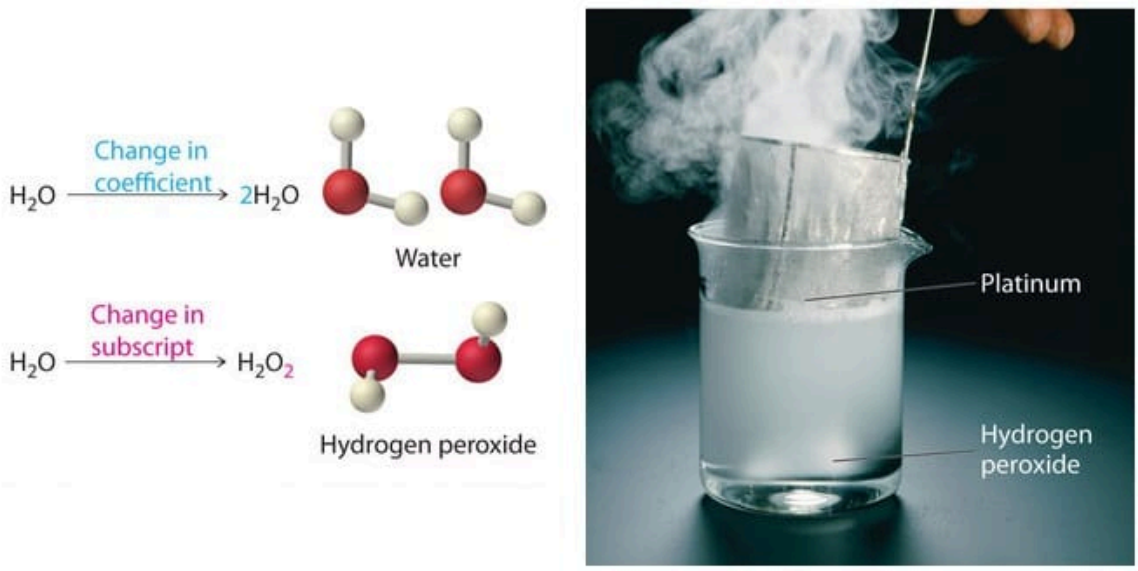


Figure 8.3.1: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 8.3.1.



Original molecule $\mathrm{H_2O}$: if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make $\mathrm{H_2O_2}$, that's hydrogen peroxide.

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

Example 8.3.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane ($\mathrm{C_7H_{16}}$).

$$\mathrm{C_7H_{16}\,(l) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)}$$

Solution

Solutions to Example 8.3.1

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assum that the final balanced chemical equation contains 1 molecule or formula unit of this substance.

Steps	Example
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which cont carbon atom, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow \underline{7}CO_2(g) + H_2O(g)$$ • 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which hydrogen atoms, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow 7\,CO_2(g) + \underline{8}H_2O(g)$$ • 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex s O_2, on the reactant side: $$C_7H_{16}(l) + \underline{11}O_2(g) \longrightarrow 7\,CO_2(g) + 8\,H_2O(g)$$ • 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen ato oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

Example 8.3.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Solutions to Example 7.4.2

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assum that the final balanced chemical equation contains 1 molecule or formula unit of this substance.

Steps	Example
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means tha must be 8 CO_2 molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $$C_8H_{18}(l) + \underline{\frac{25}{2}}O_2\ (g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(g)$$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminat fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $$\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$$
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each sid Balancing equations requires some practice on your part as well as some common sense. If you find yourself u large coefficients or if you have spent several minutes without success, go back and make sure that you have w formulas of the reactants and products correctly.

Example 8.3.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Solutions to Example 7.4.3

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $\mathrm{Pb(NO_3)_2\,(aq) + NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the Na to balance the chloride ions. $\mathrm{Pb(NO_3)_2\,(aq) + \underline{2}NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on reactant side, 1 Na atom on product side • 2 Cl atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow \underline{2}NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides
4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow 2\,NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides

Exercise 8.3.1

Is each chemical equation balanced?

a. $\mathrm{2\,Hg(\ell)^+\ O_2\,(g)\ \rightarrow Hg_2O_2\,(s)}$
b. $\mathrm{C_2H_4\,(g) + 2\,O_2(g) \longrightarrow 2\,CO_2\,(g) + 2\,H_2O\,(g)}$
c. $\mathrm{Mg(NO_3)_2\,(s) + 2\,Li\,(s)\ \rightarrow Mg(s)^+\ 2\,LiNO_3\,(s)}$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 8.3.2

Balance the following chemical equations.

a. $\mathrm{N_2\,(g) + O_2\,(g) \longrightarrow NO_2\,(g)}$
b. $\mathrm{Pb(NO_3)_2\,(aq) + FeCl_3\,(aq) \longrightarrow Fe(NO_3)_3\,(aq) + PbCl_2\,(s)}$
c. $\mathrm{C_6H_{14}\,(l) + O_2(g) \longrightarrow CO_2\,(g) + H_2O\,(g)}$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- **Chemical reaction -** The process in which one or more substances are changed into one or more new substances.
- **Reactants -** The starting materials in a reaction.
- **Products -** Materials present at the end of a reaction.
- **Balanced chemical equation -** A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts -** Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient -** A small whole number that appears in front of a formula in a balanced chemical equation.

Answers

Exercise 8.3.1
yes

Exercise 8.3.1
no

Exercise 8.3.1
yes

Exercise 8.3.2
N_2 (g) + $2O_2$ (g) → $2NO_2$ (g)

Exercise 8.3.2
$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

Exercise 8.3.2
$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

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8.4: Classifying Chemical Reactions

Learning Objectives

- Classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- Predict the products of simple reactions.

The chemical reactions we have described are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Most chemical reactions can be classified into one or more of five basic types: acid–base reactions, exchange reactions, condensation reactions (and the reverse, cleavage reactions), and oxidation–reduction reactions. The general forms of these five kinds of reactions are summarized in Table 8.4.1, along with examples of each. It is important to note, however, that many reactions can be assigned to more than one classification, as you will see in our discussion.

Table 8.4.1: Basic Types of Chemical Reactions

Name of Reaction	General Form	Examples
Oxidation–Reduction (redox)	oxidant + reductant → reduced oxidant + oxidized reductant	$C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$
Acid–Base	acid + base → salt	$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
Exchange: Single Replacement	AB + C → AC + B	$ZnCl_2(aq) + Mg(s) \rightarrow MgCl_2(aq) + Zn(s)$
Exchange: Double Replacement	AB + CD → AD + CB	$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
Combination (Synthesis)	A + B → AB	$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$
Decomposition	AB → A + B	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

The classification scheme is only for convenience; the same reaction can be classified in different ways, depending on which of its characteristics is most important. Oxidation–reduction reactions, in which there is a net transfer of electrons from one atom to another, and condensation reactions are discussed in this section. Acid–base reactions are one kind of exchange reaction—the formation of an insoluble salt, such as barium sulfate, when solutions of two soluble salts are mixed together.

Combination Reactions

A **combination reaction** is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$$A + B \rightarrow AB$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

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$$2\text{Na}\,(s) + \text{Cl}_2\,(g) \rightarrow 2\text{NaCl}\,(s)$$

Notice that in order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2).

One type of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide:

$$2\text{Mg}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{MgO}\,(s)$$

Sulfur reacts with oxygen to form sulfur dioxide:

$$\text{S}\,(s) + \text{O}_2\,(g) \rightarrow \text{SO}_2\,(g)$$

When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide:

$$2\text{S}\,(s) + 3\text{O}_2\,(g) \rightarrow 2\text{SO}_3\,(g)$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide:

$$2\text{Fe}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{FeO}\,(s)$$

$$4\text{Fe}\,(s) + 3\text{O}_2\,(g) \rightarrow 2\text{Fe}_2\text{O}_3\,(s)$$

Example 8.4.1: Combustion of Solid Potassium

Potassium is a very reactive alkali metal that must be stored under oil in order to prevent it from reacting with air. Write the balanced chemical equation for the combination reaction of potassium with oxygen.

Solution

Solutions to Example 7.10.1

Steps	Example Solution
Plan the problem.	Make sure formulas of all reactants and products are correct before balancing the equation. Oxygen gas is a diatomic molecule. Potassium oxide is an ionic compound and so its formula is constructed by the crisscross method. Potassiu an ion becomes K^+, while the oxide ion is O^{2-}.
Solve.	The skeleton (unbalanced) equation: $$\text{K}\,(s) + \text{O}_2\,(g) \rightarrow \text{K}_2\text{O}\,(s)$$ The equation is then easily balanced with coefficients. $$4\text{K}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{K}_2\text{O}\,(s)$$
Think about your result.	Formulas are correct and the resulting combination reaction is balanced.

Combination reactions can also take place when an element reacts with a compound to form a new compound composed of a larger number of atoms. Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation:

$$2\text{CO}\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g)$$

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Two compounds may also react to form a more complex compound. A very common example is the reactions of oxides with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide:

$$\mathrm{CaO}\,(s) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{Ca(OH)_2}\,(aq)$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$\mathrm{SO_3}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2SO_4}\,(aq)$$

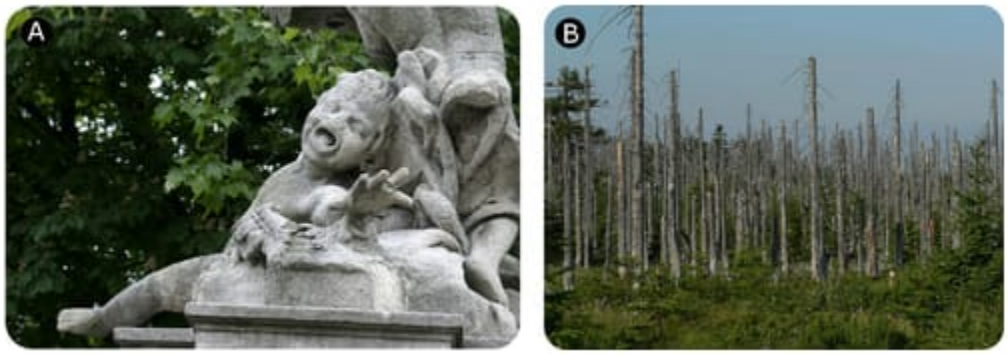


Figure 8.4.1: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain.

Exercise 8.4.1

a. Write the chemical equation for the synthesis of silver bromide, AgBr.
b. Predict the products for the following reaction: $\mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$

Answer a: available at the end of this section.

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$\mathrm{AB} \rightarrow \mathrm{A} + \mathrm{B}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas:

$$2\mathrm{HgO}\,(s) \rightarrow 2\mathrm{Hg}\,(l) + \mathrm{O_2}\,(g)$$

YouTube Video

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Video $8.4.2$***:*** *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.*

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide:

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g)$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water:

$$2\mathrm{NaOH}\,(s) \rightarrow \mathrm{Na_2O}\,(s) + \mathrm{H_2O}\,(g)$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water:

$$\mathrm{H_2CO_3}\,(aq) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

Example $8.4.2$: Electrolysis of Water

When an electric current is passed through pure water, it decomposes into its elements. Write a balanced equation for the decomposition of water.

Solution

Solutions to Example 7.10.2

Steps	Example Solution
Plan the problem.	Water is a binary compound composed of hydrogen and oxygen. The hydrogen and oxygen gases produced in the rea are both diatomic molecules.
Solve.	The skeleton (unbalanced) equation: $$\mathrm{H_2O}\,(l) \overset{\mathrm{elec}}{\rightarrow} \mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$ Note the abbreviation "elec" above the arrow to indicate the passage of an electric current to initiate the reaction. E the equation. $$2\mathrm{H_2O}\,(l) \overset{\mathrm{elec}}{\rightarrow} 2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$
Think about your result.	The products are elements and the equation is balanced.

Exercise $8.4.2$

Write the chemical equation for the decomposition of:

a. $\mathrm{Al_2O_3}$
b. $\mathrm{Ag_2S}$

Answer a available at the end of this section.

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Single Replacement Reactions

A third type of reaction is the single replacement reaction, in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$\mathrm{A} + \mathrm{BC} \rightarrow \mathrm{AC} + \mathrm{B}$$

In this general reaction, element A is a metal and replaces element B, also a metal, in the compound. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$\mathrm{Y} + \mathrm{XZ} \rightarrow \mathrm{XY} + \mathrm{Z}$$

where Y is a nonmetal and replaces the nonmetal Z in the compound with X.

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\mathrm{Mg}\,(s) + \mathrm{Cu(NO_3)_2}\,(aq) \rightarrow \mathrm{Mg(NO_3)_2}\,(aq) + \mathrm{Cu}\,(s)$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

Hydrogen Replacement

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$\mathrm{Zn}\,(s) + 2\mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\,(aq) + \mathrm{H_2}\,(g)$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal. Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All Group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2\mathrm{Na}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{NaOH}\,(aq) + \mathrm{H_2}\,(g)$$

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Figure 8.4.2: (First image) Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Second image) Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite.

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine:

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l)$$

The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

Example 8.4.3

What are the products of the reaction between solid aluminum (Al) and iron (III) oxide $(\mathrm{Fe_2O_3})$?

Solution

Solutions to Example 7.10.3

Steps	Example Solution
Plan the problem.	To predict the products, we need to know that aluminum will replace iron and form aluminum oxide (the metal will rep the metal ion in the compound). Aluminum has a charge of $+3$ and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be $\mathrm{Al_2O_3}$. Since iron is replaced in the compound by aluminum, the iror now be the single element in the products.
Solve.	The unbalanced equation will be: $$\mathrm{Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe}$$ and the balanced equation will be: $$\mathrm{2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe}$$

Steps	Example Solution
Think about your result.	This is a single replacement reaction, and when balanced the coefficients accurately reflect that the iron and aluminu the same charge in this reaction.

Exercise 8.4.3

a. Write the chemical equation for the single replacement reaction between zinc solid and lead (II) nitrate solution to produce zinc nitrate solution and solid lead. (Note that zinc forms ions with a $+2$ charge.)
b. Predict the products for the following reaction: $\mathrm{Fe} + \mathrm{CuSO_4}$. (In this reaction, assume iron forms ions with a $+2$ charge.)

Answer a available at the end of this section.

Double Replacement Reactions

A **double-replacement reaction** is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement (also called double-displacement) reaction is:

$$\mathrm{AB} + \mathrm{CD} \rightarrow \mathrm{AD} + \mathrm{BC}$$

In this reaction, A and C are positively-charged cations, while B and D are negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

Formation of a Precipitate

A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs:

$$2\mathrm{KI}\,(aq) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow 2\mathrm{KNO_3}\,(aq) + \mathrm{PbI_2}\,(s) \tag{8.4.1}$$

There are very strong attractive forces that occur between $\mathrm{Pb^{2+}}$ and $\mathrm{I^-}$ ions and the result is a brilliant yellow precipitate (Figure 8.4.3). The other product of the reaction, potassium nitrate, remains soluble.



Figure 8.4.3: Lead (II) iodide precipitates when potassium iodide is mixed with lead (II) nitrate (Equation 8.4.1). (CC BY-SA 3.0; PRHaney).

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas:

$$\mathrm{Na_2S}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{H_2S}\,(g)$$

Formation of a Molecular Compound

Another kind of double-replacement reaction is one that produces a molecular compound as one of its products. Many examples in this category are reactions that produce water. When aqueous hydrochloric acid is reacted with aqueous sodium hydroxide, the products are aqueous sodium chloride and water:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Example 8.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow$ (hydrogen cyanide gas is formed).

Solution

Solutions to Example 7.10.4

Steps	Example Solution
Plan the problem.	The production of a gas drives the reaction.
Solve.	The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partner balanced equation is: $$\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow \mathrm{NaBr}\,(aq) + \mathrm{HCN}\,(g)$$

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Steps	Example Solution
Think about your result.	This is a double replacement reaction. All formulas are correct and the equation is balanced.

Exercise 8.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $(NH_4)_2SO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow$ (a precipitate of barium sulfate forms).

Answer a: available at the end of this section.

Occasionally, a reaction will produce both a gas and a molecular compound. The reaction of a sodium carbonate solution with hydrochloric acid produces aqueous sodium chloride, carbon dioxide gas, and water:

$$Na_2CO_3\,(aq) + 2HCl\,(aq) \rightarrow 2\ \text{DEL}\ NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Answers

Exercise 8.4.1

$2Ag + Br_2 \rightarrow 2AgBr$

$CO_2\,(g) + H_2O\,(l) \rightarrow H_2CO_3$

Exercise 8.4.2

$2Al_2O_3 \rightarrow 4Al + 3O_2$

$Ag_2S \rightarrow 2Ag + S$

Exercise 8.4.3

$Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$

$Fe + CuSO_4 \rightarrow Cu + FeSO_4$

Exercise 8.4.4

$$(NH_4)_2SO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow 2NH_4NO_3\,(aq) + BaSO_4\,(s)$$

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8.5: Aqueous Solutions and Solubility

Learning Objectives

- Define and give examples of electrolytes.

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 8.5.1).

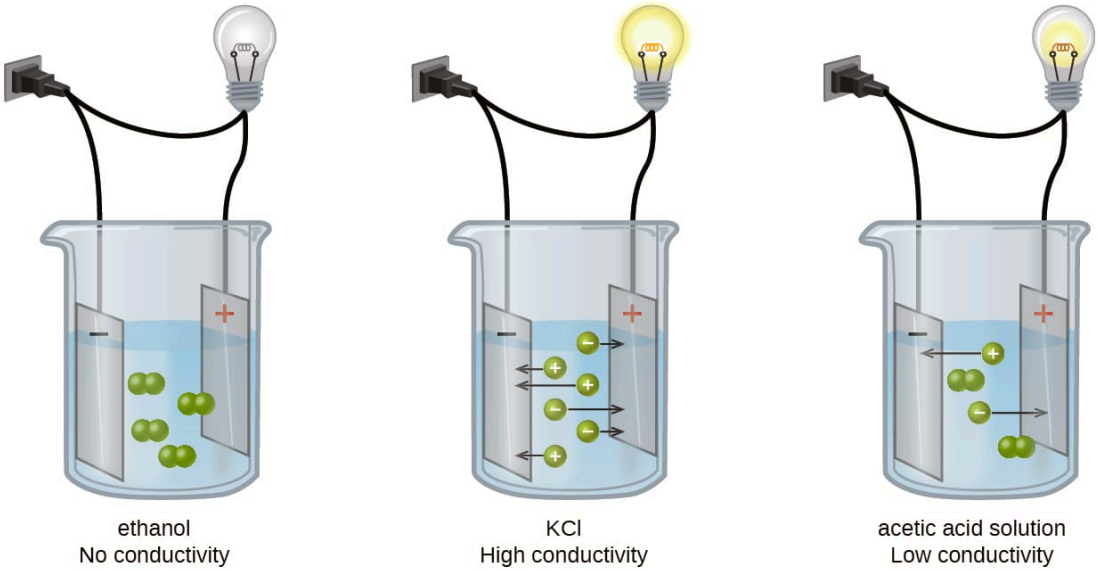


Figure 8.5.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte. This diagram shows three separate beakers. Each has a wire plugged into a wall outlet. In each case, the wire leads from the wall to the beaker and is split resulting in two ends. One end leads to a light bulb and continues on to a rectangle labeled with a plus sign. The other end leads to a rectangle labeled with a minus sign. The rectangles are in a solution. In the first beaker, labeled “Ethanol No Conductivity,” four pairs of linked small green spheres suspended in the solution between the rectangles. In the second beaker, labeled “K C l Strong Conductivity,” six individual green spheres, three labeled plus and three labeled minus are suspended in the solution. Each of the six spheres has an arrow extending from it pointing to the rectangle labeled with the opposite sign. In the third beaker, labeled “Acetic acid solution Weak conductivity,” two pairs of joined green spheres and two individual spheres, one labeled plus and one labeled minus are shown suspended between the two rectangles. The plus labeled sphere has an arrow pointing to the rectangle labeled minus and the minus labeled sphere has an arrow pointing to the rectangle labeled plus.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 8.5.2. 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.

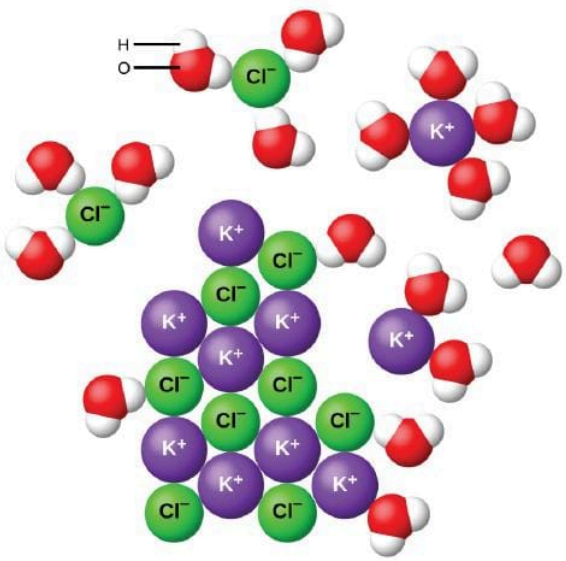


Figure 8.5.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown. The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled C l superscript minus mixed and touching near the center of the diagram. Outside of this cluster of spheres are seventeen clusters of three spheres, which include one red and two white spheres. A red sphere in one of these clusters is labeled O. A white sphere is labeled H. Two of the green C l superscript minus spheres are surrounded by three of the red and white clusters, with the red spheres closer to the green spheres than the white spheres. One of the K superscript plus purple spheres is surrounded by four of the red and white clusters. The white spheres of these clusters are closest to the purple spheres.

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.

Let us consider what happens at the microscopic level when we add solid KCl 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 penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 8.5.2 shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system, as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (Tables 8.5.1 and 8.5.2).

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Table 8.5.1: Solubility Rules for Soluble Substances

Soluble in Water	Important Exceptions (Insoluble)
All Group IA and NH_4^+ salts	none
All nitrates, chlorates, perchlorates and acetates	none
All sulfates	$CaSO_4$, $BaSO_4$, $SrSO_4$, $PbSO_4$
All chlorides, bromides, and iodides	AgX, Hg_2X_2, PbX_2 (X= Cl, Br, or I)

Table 8.5.2: Solubility Rules for Sparingly Soluble Substances

Sparingly Soluble in Water	Important Exceptions (Soluble)
All carbonates and phosphates	Group IA and NH_4^+ salts
All hydroxides	Group IA and NH_4^+ salts; Ba^{2+}, Sr^{2+}, Ca^{2+} sparingly soluble
All sulfides	Group IA, IIA and NH_4^+ salts; MgS, CaS, BaS sparingly soluble
All oxalates	Group IA and NH_4^+ salts
Special note: The following electrolytes are of only moderate solubility in water: CH_3COOAg, Ag_2SO_4, $KClO_4$. They will precipitate only if rather concentrated solutions are used.	

As an example on how to use the solubility rules, predict if a precipitate will form when solutions of cesium bromide and lead (II) nitrate are mixed.

$$\mathrm{Cs^+}\,(aq) + \mathrm{Br^-}\,(aq) + \blacksquare\,(aq) + 2\mathrm{NO_3^-}\,(aq) \rightarrow ?$$

The potential precipitates from a double-replacement reaction are cesium nitrate and lead (II) bromide. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, the cesium and nitrate ions are spectator ions and the lead (II) bromide is a precipitate. The balanced net ionic reaction is:

$$\blacksquare\,(aq) + 2\mathrm{Br^-}\,(aq) \rightarrow \mathrm{PbBr_2}\,(s)$$

Example 8.5.1: Solubility

Classify each compound as soluble or insoluble

a. Zn(NO3)2
b. PbBr2
c. Sr3(PO4)2

Solution

a. All nitrates are soluble in water, so Zn(NO3)2 is soluble.
b. All bromides are soluble in water, except those combined with Pb^{2+}, so PbBr2 is insoluble.
c. All phosphates are insoluble, so Sr3(PO4)2 is insoluble.

Exercise 8.5.1: Solubility

Classify each compound as soluble or insoluble.

a. $Mg(OH)_2$
b. KBr
c. $Pb(NO_3)_2$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Summary

Substances that dissolve in water to yield ions are called electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water. Solubility rules allow prediction of what products will be insoluble in water.

Contributions & Attributions

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-

Answers

Exercise 8.5.1: Solubility
insoluble

Exercise 8.5.1: Solubility
soluble

Exercise 8.5.1: Solubility
soluble

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8.6: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate, a reddish precipitate of silver dichromate is produced.

$$\mathrm{AgNO_3\,(aq) + K_2Cr_2O_7\,(aq)\ \rightarrow Ag_2Cr_2O_7\,(s) + KNO_3\,(aq)} \tag{8.6.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$\mathrm{AC + BD\ \rightarrow \underset{insoluble}{AD} + BC} \tag{8.6.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called **double-displacement reactions**. Precipitation reactions are used to isolate metals that have been extracted from their ores, and to recover precious metals for recycling.

YouTube Video

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Video: Mixing potassium dichromate and silver nitrate together to initiate a precipitation reaction (Equation 8.6.1).

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $NaCl$ solution is mixed with 500 mL of aqueous KBr solution, the final solution has a volume of 1.00 L and contains $\mathrm{Na^+\,(aq)}$, $\mathrm{Cl^-\,(aq)}$, $\mathrm{K^+\,(aq)}$, and $\mathrm{Br^-\,(aq)}$. As you will see in (Figure 8.6.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

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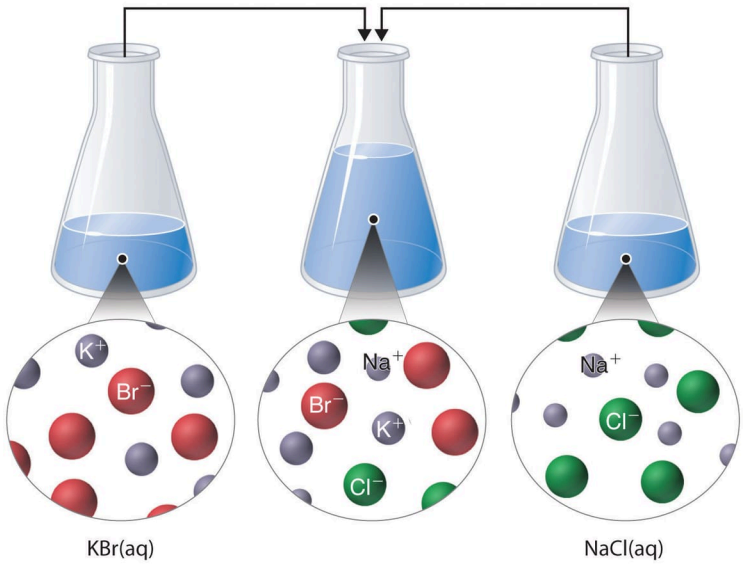


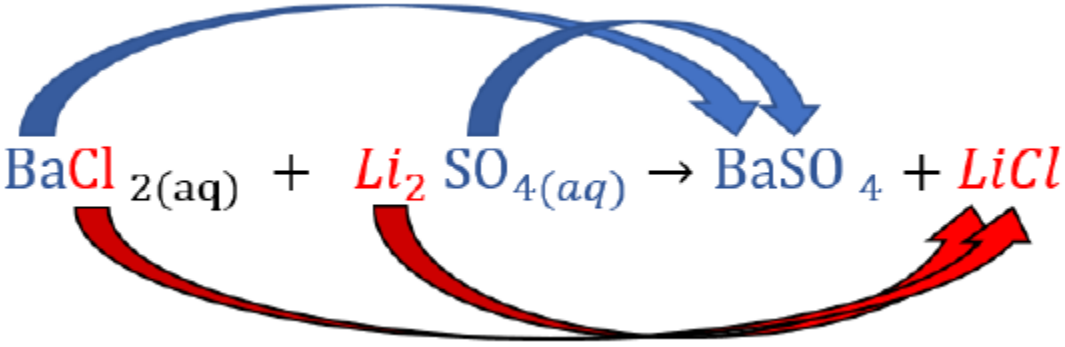
Figure 8.6.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

Change the partners of the anions and cations on the reactant side to form new compounds (products):



Chemical equation of the reactants barium chloride and lithium sulfate forming the products barium sulfate and lithium chloride.

Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains $\mathrm{Ba^{2+}}$, $\mathrm{Cl^-}$, $\mathrm{Li^+}$, and $\mathrm{SO_4{}^{2-}}$ ions. The only possible exchange reaction is to form LiCl and $\mathrm{BaSO_4}$.

Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4 + LiCl}$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq)\ \rightarrow BaSO_4\,(s) + LiCl\,(aq)}$$

Table 7.5.1 from the previous section shows that LiCl is soluble in water, but $\mathrm{BaSO_4}$ is not soluble in water.

Balance the equation:

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq)\ \rightarrow BaSO_4\,(s) + 2\,LiCl\,(aq)}$$

Although soluble barium salts are toxic, $\mathrm{BaSO_4}$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a "barium milkshake" or a "barium enema"—a suspension of very fine $\mathrm{BaSO_4}$ particles in water.



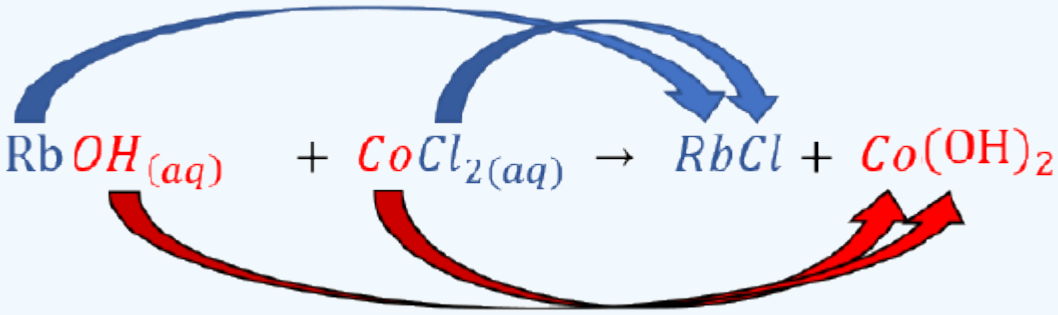
Figure 8.6.2: An x-ray of the digestive organs of a patient who has swallowed a "barium milkshake." A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. (Public Domain; Glitzy queen00 via Wikipedia).

Example 8.6.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

Solution

Solutions to Example 7.6.1

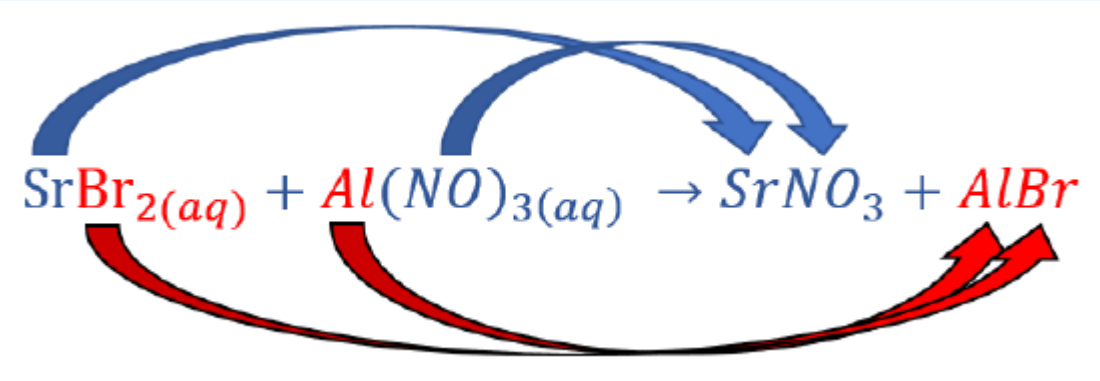
Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants rubidium hydroxide and cobalt(II) chloride forming the products rubidium chloride and cobalt hydroxide.
Correct the formulas of the products based on the charges of the ions.	$\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl + Co(OH)_2}$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$
Balance the equation.	Coefficients already balanced. $\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$

Example 8.6.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution

Solutions for Example 7.6.2

Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants strontium bromide and aluminum nitrate forming the products strontium nitrate and aluminum bromide.
Correct the formulas of the products based on the charges of the ions.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow Sr(NO_3)_2 + AlBr_3}$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow Sr(NO_3)_2\,(aq) + AlBr_3\,(aq)}$ According to Table 7.5.1 from the previous section, both $\mathrm{AlBr_3}$ (rule 4) and $\mathrm{Sr(NO_3)_2}$ (rule 2) are soluble.
If all possible products are soluble, then no net reaction will occur.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow}$ **NO REACTION**

Exercise 8.6.2

Using the information in Table 7.5.1 from the previous section, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

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Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributions & Attributions

- Modified by Joshua Halpern (Howard University)

Answers

Exercise 8.6.2

$Fe(OH)_2$ precipitate is formed.

Exercise 8.6.2

$Hg_3(PO_4)_2$ precipitate is formed.

Exercise 8.6.2

No Reaction.

Exercise 8.6.2

$CaCO_3$ is precipitate formed.

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8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations.

$$\underbrace{\mathrm{BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2\,NaCl\,(aq)}}_{\text{Complete Chemical Equation}} \tag{8.7.1}$$

However, Equation 8.7.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Below is the complete ionic equation:

$$\underbrace{\mathrm{Ba^{2+}\,(aq)} + \overbrace{\mathrm{2Cl^-(aq)}}^{\text{spectator}} + \overbrace{\mathrm{2Na^+(aq)}}^{\text{spectator}} + \mathrm{SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s)} + \overbrace{\mathrm{2Na^+(aq)}}^{\text{spectator}} + \overbrace{\mathrm{Cl^-(aq)}}^{\text{spectator}}}_{\text{Complete Ionic Equation}} \tag{8.7.2}$$

Equation 8.7.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $\mathrm{BaSO_4\,(s)}$ will form when *any* solution containing $\mathrm{Ba^{2+}\,(aq)}$ is mixed with *any* solution containing $\mathrm{SO_4{}^{2-}\,(aq)}$ (provided concentrations are not extremely small). This happens independently of the $\mathrm{Cl^-\,(aq)}$ and $\mathrm{Na^+\,(aq)}$ ions in Equation 8.7.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $\mathrm{BaSO_4}$ the net ionic equation is

$$\underbrace{\mathrm{Ba^{2+}\,(aq) + SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s)}}_{\text{Net Ionic Equation}} \tag{8.7.3}$$

Example 8.7.1

a. When a solution of $\mathrm{AgNO_3}$ is added to a solution of $\mathrm{CaCl_2}$, insoluble AgCl precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.
b. Write the balanced net ionic equation to describe any reaction that occurs when the solutions of $\mathrm{Na_2SO_4}$ and $\mathrm{NH_4I}$ are mixed.

Solution

Solutions to Example 7.7.1

Equation Type	Example 8.7.1a	Example 8.7.1b
Complete Chemical Equation	$\mathrm{2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow\ 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)}$ The proper states and formulas of all products are written and the chemical equation is balanced.	$\mathrm{Na_2SO_4\,(aq) + 2\,NH_4I\,(aq) \longrightarrow 2\,NaI\,(aq) + (NH_4)_2SO_4\,(aq)}$ Both products are aqueous so there is no net ionic equation that can be written.
Complete Ionic Equation	$\mathrm{2\,Ag^+\,(aq) + 2\,NO_3{}^-\,(aq) + Ca^{2+}\,(aq) + 2\,Cl^-\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca^{2+}\,(aq) + 2\,NO_3{}^-\,(aq)}$ AgCl is a solid so it does not break up into ions in solution.	

Equation Type	Example 8.7.1a	Example 8.7.1b
Net Ionic Equation	$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$ All spectator ions are removed. The chemical equation is written using the lowest common coefficients.	NaI and $(NH_4)_2SO_4$ are both soluble. No net ionic equation. There is no reaction.

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $BaCl_2$ solution, for instance, is often used as a test for the presence of $SO_4{}^{2-}(aq)$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $BaSO_4$. Thus, if $BaCl_2$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the $SO_4{}^{2-}$ ion.



Figure 8.7.1: The three common silver halide precipitates: AgI, $AgBr$ and $AgCl$ (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$AgNO_3$ solutions are often used in a similar way to test for halide ions. If $AgNO_3$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of Cl^- ions, a cream-colored precipitate indicates the presence of Br^- ions, and a yellow precipitate indicates the presence of I^- ions (Figure 8.7.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $AgNO_3$ is added to tap water, a white precipitate is almost always formed. The Cl^- ions in tap water usually come from the Cl_2 which is added to municipal water supplies to kill microorganisms.

Exercise 8.7.1

Write balanced net ionic equations to describe any reaction that occurs when the following solutions are mixed.

a. $K_2CO_3 + SrCl_2$
b. $FeSO_4 + Ba(NO_3)_2$

Answer a available at the end of this section.

Answer b available at the end of this section.

Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then

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the rules of stoichiometry may be applied.

Contributions & Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

Answers

Exercise 8.7.1

$$\mathrm{Sr^{2+}(aq) + CO_3{}^{2-}(aq) \longrightarrow SrCO_3(s)}$$

Exercise 8.7.1

$$\mathrm{Ba^{2+}(aq) + SO_4{}^{2-}(aq) \longrightarrow Ba(SO_4)(s)}$$

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8.8: Acid–Base Reactions

Learning Objectives

- Identify when a reaction will evolve a gas.

Neutralization Reactions

Acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen-ion-containing compound and the base is a hydroxide-ion-containing compound, water is also a product. The general reaction is as follows:

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{salt}$$

The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:

$$\mathrm{NaOH\,(aq) + HCl\,(aq)\ \rightarrow NaCl\,(aq) + H_2O\ (\ell)} \tag{8.8.1}$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$\mathrm{2\,NaOH\,(aq) + H_2SO_4\,(aq)\ \rightarrow Na_2SO_4\,(aq) + 2\,H_2O\ (\ell)} \tag{8.8.2}$$

Example 8.8.1: Neutralizing Nitric Acid

Nitric acid ($HNO_3(aq)$) can be neutralized by calcium hydroxide ($Ca(OH)_2(aq)$). Write a balanced chemical equation for the reaction between these two compounds and identify the salt that it produces.

Solution

Solutions to Example 7.8.1

Steps	Explanation	Equation
Write the unbalanced equation.	This is a double displacement reaction, so the cations and anions swap to create new products.	$Ca(OH)_2(aq) + HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(\ell)$
Balance the equation.	Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of HNO_3 to provide H^+ ions	$Ca(OH)_2(aq) + \underline{2}HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + \underline{2}H_2O(\ell)$
Additional step: identify the salt.		The salt formed is calcium nitrate.

Exercise 8.8.1

Hydrocyanic acid ($\mathrm{HCN\,(aq)}$) can be neutralized by potassium hydroxide ($\mathrm{KOH\,(aq)}$). Write a balanced chemical equation for the reaction between these two compounds and identify the salt that it produces.

Answer

\[\ce{KOH (aq) + HCN(aq) → KCN (aq) + H2O(ℓ)} \nonumber \]

Gas Evolving Reactions

A gas evolution reaction is a chemical process that produces a gas, such as oxygen or carbon dioxide. In the following examples, an acid reacts with a carbonate, producing salt, carbon dioxide, and water, respectively. For example, nitric acid reacts with sodium carbonate to form sodium nitrate, carbon dioxide, and water (Table 8.8.1):

$$2\,HNO_3(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaNO_3(aq) + CO_2(g) + H_2O(l)$$

Sulfuric acid reacts with calcium carbonate to form calcium sulfate, carbon dioxide, and water:

$$H_2SO_4\,(aq) + CaCO_3\,(aq) \longrightarrow CaSO_4\,(aq) + CO_2(g) + H_2O\,(l)$$

Hydrochloric acid reacts with calcium carbonate to form calcium chloride, carbon dioxide, and water:

$$2\,HCl\,(aq) + CaCO_3\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Figure 8.8.1 demonstrates this type of reaction:

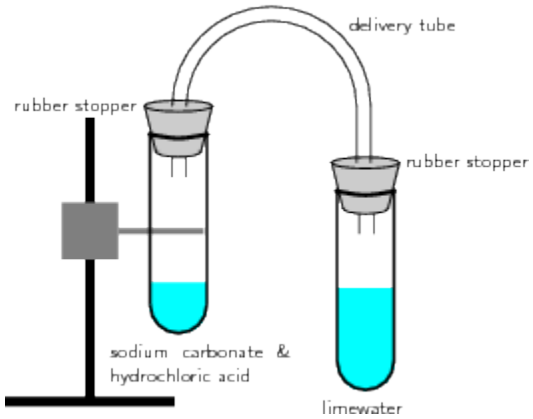


Figure 8.8.1: Reaction of acids with carbonates. In this reaction setup, lime water (water + calcium hydroxide) is poured into one of the test tubes and sealed with a stopper. A small amount of hydrochloric acid is carefully poured into the remaining test tube. A small amount of sodium carbonate is added to the acid, and the tube is sealed with a rubber stopper. The two tubes are connected. As a result of the acid-carbonate reaction, carbon dioxide is produced and the lime water turns milky.

In this reaction setup, lime water, a dilute calcium hydroxide ($Ca(OH)_2$) solution, is poured into one of the test tubes and sealed with a stopper. A small amount of hydrochloric acid is carefully poured into the remaining test tube. A small amount of sodium carbonate is added to the acid, and the tube is sealed with a rubber stopper. The two tubes are connected. As a result of the acid-carbonate reaction, carbon dioxide is produced and the lime water turns milky.

Table 8.8.1: Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
sulfide	none	H_2S	$2\,HCl\,(aq) + K_2S \rightarrow H_2S\,(g) + 2\,KCl\,(aq)$
carbonates and bicarbonates	H_2CO_3	CO_2	$2\,HCl\,(aq) + K_2CO_2 \rightarrow H_2O\,(l) + CO_2\,(g) + 2\,KCl\,(aq)$
sulfites and bisulfites	H_2SO_3	SO_2	$2\,HCl\,(aq) + K_2SO_2 \rightarrow H_2O\,(l) + SO_2\,(g) + 2\,KCl\,(aq)$
ammonia	NH_4OH	NH_3	\(\ce{NH4Cl(aq) + KOH \rightarrow H2O (l) + NH3(g) + 2KCl (aq)}\)

The gas-evolving experiment lime water is illustrated in the following video:

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YouTube Video

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Video $8.8.1$***:*** *Carbon Dioxide (CO_2) & Limewater (Chemical Reaction). As the reaction proceeds, the limewater on the turns from clear to milky; this is due to the $CO_2(g)$ reacting with the aqueous calcium hydroxide to form calcium carbonate, which is only slightly soluble in water.*

When this experiment is repeated with nitric or sulfuric acid instead of HCl, it yields the same results: the clear limewater turns milky, indicating the production of carbon dioxide. Another method to chemically generate gas is the oxidation of metals in acidic solutions. This reaction will yield a metal salt and hydrogen gas.

$$2\,\mathrm{HCl\,(aq)} + \mathrm{Zn\,(s)} \rightarrow \mathrm{ZnCl_2\,(aq)} + \mathrm{H_2\,(g)}$$

Here, hydrochloric acid oxidizes zinc to produce an aqueous metal salt and hydrogen gas bubbles. Recall that oxidation refers to a loss of electrons, and reduction refers to the gain of electrons. In the above redox reaction, neutral zinc is oxidized to Zn^{2+}, and the acid, H^+, is reduced to $H_2(g)$. The oxidation of metals by strong acids is another common example of a gas evolution reaction.

Contributors & Affiliations

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9: STOICHIOMETRY AND THE MOLE

9.1: Avogadro's Number

9.2: Conversions Between Moles and Particles

9.3: Molar Mass

9.4: Conversions Between Moles and Mass

9.5: Conversions Between Mass and Number of Particles

9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"

9.7: Stoichiometry

9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!

9.9: Mole-Mass and Mass-Mass Calculations

9.10: Yields

9.11: Limiting Reagents

9.12: Mass Percent Composition from a Chemical Formula

9.13: Calculating Empirical Formulas

9.14: Calculating Molecular Formulas

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9.1: Avogadro's Number



Figure $9.1.1$ (Public Domain; User:OSX/Wikimedia Commons via Wikipedia)

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

It certainly is easy to count objects such as bananas, or something as large as elephants (as long as you stay out of their way). However, counting grains of sugar from a sugar canister would take a long, long time. Atoms and molecules are extremely small—far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years, without stopping, you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists of the past needed a name that could stand for a very large number of items. Amadeo Avogadro (1776-1856), an Italian scientist, provided such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.



Figure 9.1.2: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry. (Public Domain; C. Sentier via Wikipedia)

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as $NaCl$ and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.



Figure 9.1.3: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year. ((left; Left: Michael David Hill, 2005 (Mikiwikipikidikipedia); Right: chrisbb@prodigy.net;(left) Michael David Hill, 2005 (Mikiwikipikidikipedia); (right) chrisbb@prodigy.net via Left: http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; Right: http://www.flickr.com/photos/chrisbrenschmidt/436990097;(left) http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; (right) http://www.flickr.com/photos/chrisbrenschmidt/436990097) CC-BY 2.5; (right) CC-BY 2.0)

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Summary

- A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

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9.2: Conversions Between Moles and Particles



Figure 9.2.1 (Public Domain; Diane A. Reid/National Cancer Institute via Wikipedia)

Big numbers or little numbers?

Do you hate to type subscripts and superscripts? Even with a good word-processing program, having to click on an icon to get a superscript and then remembering to click off after you type the number can be a real hassle. If we did not know about moles and just knew about numbers of atoms or molecules (those big numbers that require lots of superscripts), life would be much more complicated and we would make many more typing errors.

Conversions Between Moles and Atoms

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Example 9.2.1: Converting Number of Particles to Moles

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Step 1: List the known quantities and plan the problem.

Known

- number of C atoms $= 4.72 \times 10^{24}$
- 1 mole $= 6.02 \times 10^{23}$ atoms

Unknown

- 4.72 x 10^{24} = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures.

Suppose that you want to know how many hydrogen atoms are in a mole of water molecules. First, you need to know the chemical formula for water, which is H_2O. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen are in two water molecules? There are $2 \times 2 = 4$ hydrogen atoms. How about in a dozen? In that case, a dozen is 12; so $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers (4 and 24), you multiply the given number of molecules by two atoms of hydrogen per molecule. So, to find the number of hydrogen atoms in a mole of water molecules, the problem can be solved using conversion factors:

$$1 \text{ mol } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2O} = 1.20 \times 10^{24} \text{ atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.

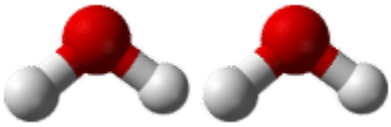


Figure 9.2.2: Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Example 9.2.2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is the sample?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $4.89 \times 10^{25} = \text{O}$ atoms
- 1 mole $= 6.02 \times 10^{23}$ molecules H_2SO_4

Unknown

- mol of H_2SO_4 molecules

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ molecule } H_2SO_4}{4 \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{6.02 \times 10^{23} \text{ molecules } H_2SO_4} = 20.3 \text{ mol } H_2SO_4$$

Step 3: Think about your result.

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The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.

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Summary

- Methods are described for conversions between moles, atoms, and molecules.

Review

1. What conversion factor would we need to convert moles of helium to atoms of helium?
2. I want to convert atoms to moles. My friend tells me to multiply the number of atoms by 6.02×10^{23} atoms/mole. Is my friend correct?
3. Why do you need to know the formula for a molecule in order to calculate the number of moles of one of the atoms?
4. How many atoms of fluorine are in 5.6×10^{22} molecules of MgF_2?

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9.3: Molar Mass



Figure 9.3.1 (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

When creating a solution, how do I know how much of each substance to put in?

I want to make a solution that contains 1.8 moles of potassium dichromate. You do not have a balance calibrated in moles, but you do have one calibrated in grams. If you know the relationship between moles and the number of grams in a mole, you can use your balance to measure out the needed amount of material.

Molar Mass

Molar mass is defined as the mass of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of lithium is $6.94\ \text{g}$, the molar mass of zinc is $65.38\ \text{g}$, and the molar mass of gold is $196.97\ \text{g}$. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole, or g/mol.

Molar Masses of Compounds

The molecular formula of the compound carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen:

$$12.01\ \text{amu} + 2\,(16.00\ \text{amu}) = 44.01\ \text{amu}$$

The **molecular mass** of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is $44.01\ \text{amu}$.

The molar mass of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of $44.01\ \text{g}$, while one mole of sodium sulfide formula units has a mass of $78.04\ \text{g}$. The molar masses are $44.01\ \text{g/mol}$ and $78.04\ \text{g/mol}$ respectively. In both cases, that is the mass of 6.02×10^{23} representative particles. The representative particle of CO_2 is the molecule, while for Na_2S it is the formula unit.

Example 9.3.1: Molar Mass of a Compound

Calcium nitrate, $\text{Ca}(\text{NO}_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Solution

Step 1: List the known and unknown quantities and plan the problem.

Known

- Formula $= \mathrm{Ca(NO_3)_2}$
- Molar mass $\mathrm{Ca} = 40.08\ \mathrm{g/mol}$
- Molar mass $\mathrm{N} = 14.01\ \mathrm{g/mol}$
- Molar mass $\mathrm{O} = 16.00\ \mathrm{g/mol}$

Unknown

- molar mass $Ca(NO_3)_2$

First we need to analyze the formula. Since the Ca lacks a subscript, there is one Ca atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit and each nitrate ion consists of one nitrogen atom and three oxygen atoms per formula unit. Thus, $1\ \mathrm{mol}$ of calcium nitrate contains $1\ \mathrm{mol}$ of Ca atoms, $2\ \mathrm{mol}$ of N atoms, and $6\ \mathrm{mol}$ of O atoms.

Step 2: Calculate

Use the molar masses of each atom together with the number of atoms in the formula and add together.

$$1\ \mathrm{mol\ Ca} \times \frac{40.08\ \mathrm{g\ Ca}}{1\ \mathrm{mol\ Ca}} = 40.08\ \mathrm{g\ Ca}$$

$$2\ \mathrm{mol\ N} \times \frac{14.01\ \mathrm{g\ N}}{1\ \mathrm{mol\ N}} = 28.02\ \mathrm{g\ N}$$

$$6\ \mathrm{mol\ O} \times \frac{16.00\ \mathrm{g\ O}}{1\ \mathrm{mol\ O}} = 96.00\ \mathrm{g\ O}$$

$$\text{Molar mass of } \mathrm{Ca(NO_3)_2} = 40.08\ \mathrm{g} + 28.02\ \mathrm{g} + 96.00\ \mathrm{g} = 164.10\ \mathrm{g/mol}$$

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Summary

- Calculations are described for the determination of molar mass of an atom or a compound.

Review

1. What is the molar mass of Pb?
2. Where do you find the molar mass of an element?
3. How many moles of Cl are in one mole of the $CaCl_2$?
4. How many moles of H are in one mole of the compound $(NH_4)_3PO_4$?
5. Calculate the molar mass of $CaCl_2$.

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9.4: Conversions Between Moles and Mass



Figure 9.4.1 (Public Domain; User:Estormiz/Wikimedia Commons via Wikipedia)

How can we get more product?

Chemical manufacturing plants are always seeking to improve their processes. One way that improvement comes about is through measuring the amount of material produced in a reaction. By knowing how much is made, the scientists and engineers can try different ways of getting more product at less cost.

Conversions Between Moles and Mass

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment, you need 3.00 moles of calcium chloride $(\mathrm{CaCl_2})$. Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. The molar mass of $\mathrm{CaCl_2}$ is $110.98\ \mathrm{g/mol}$. The conversion factor that can be used is then based on the equality that $1\ \mathrm{mol} = 110.98\ \mathrm{g}\ \mathrm{CaCl_2}$. Dimensional analysis will allow you to calculate the mass of $\mathrm{CaCl_2}$ that you should measure.

$$3.00\ \mathrm{mol\ CaCl_2} \times \frac{110.98\ \mathrm{g\ CaCl_2}}{1\ \mathrm{mol\ CaCl_2}} = 333\ \mathrm{g\ CaCl_2}$$

When you measure the mass of $333\ \mathrm{g}$ of $\mathrm{CaCl_2}$, you are measuring 3.00 moles of $\mathrm{CaCl_2}$.



Figure 9.4.2: Calcium chloride is used as a drying agent and as a road deicer. (Public Domain; Martin Walker (User:Walkerma/Wikimedia Commons) via Wikipedia)

Example 9.4.1: Converting Moles to Mass

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molar mass of $\mathrm{Cr} = 52.00\ \mathrm{g/mol}$
- $0.560\ \mathrm{mol\ Cr}$

Unknown

- $0.560\ \mathrm{mol\ Cr}$ = ? g

One conversion factor will allow us to convert from the moles of Cr to mass.

Step 2: Calculate.

$$0.560\ \mathrm{mol\ Cr} \times \frac{52.00\ \mathrm{g\ Cr}}{1\ \mathrm{mol\ Cr}} = 29.1\ \mathrm{g\ Cr}$$

Step 3: Think about your result.

Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because of the $0.560\ \mathrm{mol}$.

A similar conversion factor utilizing molar mass can be used to convert from the mass of a substance to moles. In a laboratory situation, you may perform a reaction and produce a certain amount of a product which can be massed. It will often then be necessary to determine the number of moles of the product that was formed. The next problem illustrates this situation.

Example 9.4.2: Converting Mass to Moles

A certain reaction produces $2.81\ \mathrm{g}$ of copper (II) hydroxide, $\mathrm{Cu(OH)_2}$. Determine the number of moles produced in the reaction.

Step 1: List the known quantities and plan the problem.

Known

- mass = 2.81g

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Unknown

- mol $Cu(OH)_2$

One conversion factor will allow us to convert from mass to moles.

Step 2: Calculate.

First, it is necessary to calculate the molar mass of $\mathrm{Cu(OH)_2}$ from the molar masses of Cu, O, and H. The molar mass is $97.57\ \mathrm{g/mol}$.

$$2.81\ \mathrm{g\ Cu(OH)_2} \times \frac{1\ \mathrm{mol\ Cu(OH)_2}}{97.57\ \mathrm{g\ Cu(OH)_2}} = 0.0288\ \mathrm{mol\ Cu(OH)_2}$$

Step 3: Think about your result.

The relatively small mass of product formed results in a small number of moles.

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Simulation

Practice converting mass to moles in everyday objects in this simulation:

How many atoms are in a gold ring, a grain of sand, or the human body?

Summary

- Calculations involving conversions between moles of a material and the mass of that material are described.

Review

1. You have 19.7 grams of a material and wonder how many moles were formed. Your friend tells you to multiply the mass by grams/mole. Is your friend correct?
2. How many grams of MgO are in 3.500 moles?
3. How many moles of H_2O are in 15.2 grams of pure ice?

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9.5: Conversions Between Mass and Number of Particles

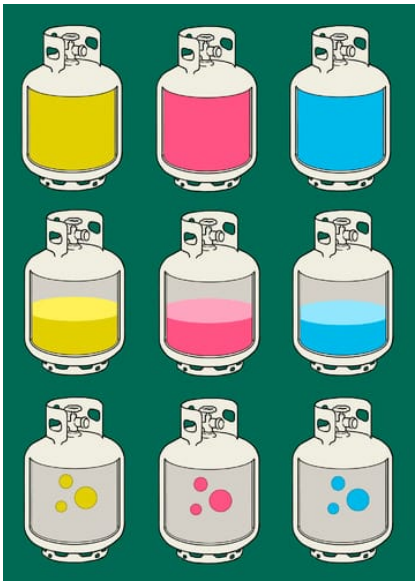


Figure 9.5.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

How much gas is there?

Avogadro was interested in studying gases. He theorized that equal volumes of gases under the same conditions contained the same number of particles. Other researchers studied how many gas particles were in specific volumes of gases. Eventually, scientists were able to develop the relationship between number of particles and mass, using the idea of moles.

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and the mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to grams. In order to convert from mass to number of particles or vice-versa, a conversion to moles is required.



Figure 9.5.2: Conversion from number of particles to mass, or from mass to number of particles requires two steps. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 9.5.1: Converting Mass to Particles

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molar mass $\mathrm{Cl_2} = 70.90\ \mathrm{g/mol}$
- $20.0\ \mathrm{g}\ \mathrm{Cl_2}$

Unknown

- number of molecules of Cl_2

Use two conversion factors. The first converts grams of $\mathrm{Cl_2}$ to moles. The second converts moles of $\mathrm{Cl_2}$ to the number of molecules.

Step 2: Calculate.

$$20.0\ \mathrm{g\ Cl_2} \times \frac{1\ \mathrm{mol\ Cl_2}}{70.90\ \mathrm{g\ Cl_2}} \times \frac{6.02 \times 10^{23}\ \mathrm{molecules\ Cl_2}}{1\ \mathrm{mol\ Cl_2}} = 1.70 \times 10^{23}\ \mathrm{molecules\ Cl_2}$$

The problem is done using two consecutive conversion factors. There is no need to explicitly calculate the moles of $\mathrm{Cl_2}$.

Step 3: Think about your result.

Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

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Summary

- Calculations are illustrated for conversions between mass and number of particles.

Review

1. Why can't we convert directly from number of particles to grams?
2. How many atoms of chlorine are present in 1.70×10^{23} molecules Cl_2?
3. How many molecules of BH_3 are in 14.32 grams BH_3?

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9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"

Learning Objectives

- Use chemical formulas as conversion factors.

Figure 9.6.1 shows that we need 2 hydrogen atoms and 1 oxygen atom to make one water molecule. If we want to make two water molecules, we will need 4 hydrogen atoms and 2 oxygen atoms. If we want to make five molecules of water, we need 10 hydrogen atoms and 5 oxygen atoms. The ratio of atoms we will need to make any number of water molecules is the same: 2 hydrogen atoms to 1 oxygen atom.

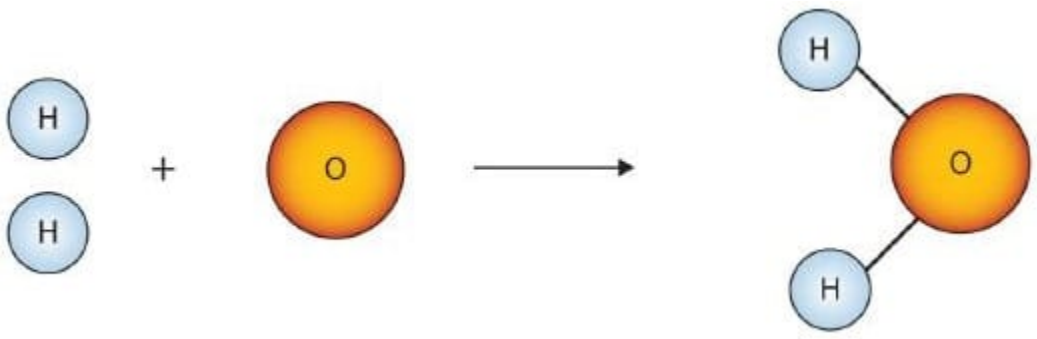


Figure 9.6.1 Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

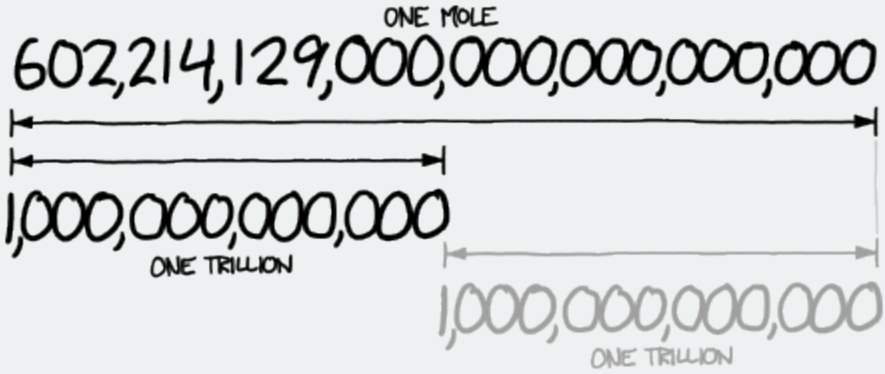
Using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of moles of molecules to the number of moles of atoms. For example, in 1 mol of water (H_2O) we can construct the relationships given in (Table 9.6.1).

Table 9.6.1: Molecular Relationships for Water

1 Molecule of H_2O Has	1 Mol of H_2O Has	Molecular Relationships
2 H atoms	2 mol of H atoms	$\frac{2\text{ mol H atoms}}{1\text{ mol }H_2O\text{ molecules}}$ or $\frac{1\text{ mol }H_2O\text{ molecules}}{2\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }H_2O\text{ molecules}}$ or $\frac{1\text{ mol }H_2O\text{ molecules}}{1\text{ mol O atoms}}$

The Mole is big

A mole represents a very large number! The number 602,214,129,000,000,000,000,000 looks about twice as long as a trillion, which means it's about a trillion trillion.



(CC BY-SA NC; what if? [what-if.xkcd.com]).

A trillion trillion kilograms is how much a planet weighs. If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Table 9.6.2: Molecular and Mass Relationships for Ethanol

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular and Mass Relationships
2 C atoms	2 mol of C atoms	$\frac{2\text{ mol C atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{2\text{ mol C atoms}}$
6 H atoms	6 mol of H atoms	$\frac{6\text{ mol H atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{6\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{1\text{ mol O atoms}}$
2 (12.01 amu) C 24.02 amu C	2 (12.01 g) C 24.02 g C	$\frac{24.02\text{ g C}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{24.02\text{ g C}}$
6 (1.008 amu) H 6.048 amu H	6 (1.008 g) H 6.048 g H	$\frac{6.048\text{ g H}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{6.048\text{ g H}}$
1 (16.00 amu) O 16.00 amu O	1 (16.00 g) O 16.00 g O	$\frac{16.00\text{ g O}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{16.00\text{ g O}}$

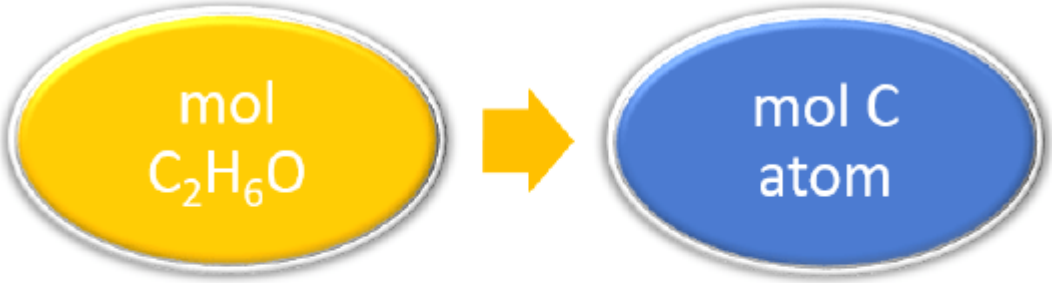
The following example illustrates how we can use the relationships in Table 9.6.2 as conversion factors.

Example 9.6.1: Ethanol

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?

Solution

Solutions to Example 6.5.1

Steps for Problem Solving	If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?
Identify the "given" information and what the problem is asking you to "find."	Given: 2.5 mol C_2H_6O Find: mol C atoms
List other known quantities.	1 mol C_2H_6O = 2 mol C
Prepare a concept map and use the proper conversion factor.	 $\frac{2\ mol\ C}{1\ mol\ C_2H_6O}$
Cancel units and calculate.	Note how the unit *mol C_2H_6O molecules* cancels algebraically. $2.5\ \cancel{\text{mol } C_2H_6O \text{ molecules}} \times \frac{2\text{ mol C atoms}}{1\ \cancel{\text{mol } C_2H_6O \text{ molecules}}} = 5.0\text{ mol C atoms}$
Think about your result.	There are twice as many C atoms in one C_2H_6O molecule, so the final amount should be double.

Exercise 9.6.1

If a sample contains 6.75 mol of Na_2SO_4, how many moles of sodium atoms, sulfur atoms, and oxygen atoms does it have?

Answer available at the end of this section.

Answer available at the end of this section.

The fact that 1 mol equals 6.022×10^{23} items can also be used as a conversion factor.

Example 9.6.2: Oxygen Mass

Determine the mass of Oxygen in 75.0g of C_2H_6O.

Solution

Solutions to Example 6.5.2

Steps for Problem Solving	Determine the mass of Oxygen in 75.0g of C_2H_6O
Identify the "given" information and what the problem is asking you to "find."	Given: 75.0g C_2H_6O Find: g O
List other known quantities.	1 mol O = 16.0g O 1 mol C_2H_6O = 1 mol O 1 mol C_2H_6O = 46.07g C_2H_6O
Prepare a concept map and use the proper conversion factor.	 The conversion factors are 1 mol C2H6O over 46.07 g C2H6O, 1 mol O over 1 mol C2H6O, and 16.00 g O over 1 mole O.
Cancel units and calculate.	$75.0\ \cancel{\text{g C}_2\text{H}_6\text{O}} \times \frac{1\ \cancel{\text{mol C}_2\text{H}_6\text{O}}}{46.07\ \cancel{\text{g C}_2\text{H}_6\text{O}}} \times \frac{1\ \cancel{\text{mol O}}}{1\ \cancel{\text{mol C}_2\text{H}_6\text{O}}} \times \frac{16.00\ \text{g O}}{1\ \cancel{\text{mol O}}} = 26.0\ \text{g O}$
Think about your result.	

Exercise 9.6.2

a. How many molecules are present in 16.02 mol of C_4H_{10}? How many C atoms are in 16.02 mol?
b. How many moles of each type of atom are in 2.58 mol of Na_2SO_4?

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Summary

In any given formula, the ratio of the number of moles of molecules (or formula units) to the number of moles of atoms can be

used as a conversion factor.

Answers

Exercise 9.6.1

13.5 mol Na atoms, 6.75 mol S atoms, and 27.0 mol O atoms

Exercise 9.6.1

Exercise 9.6.2

9.647 x 10^{24} C_4H_{10} molecules and 3.859 x 10^{25} C atoms

Exercise 9.6.2

5.16 mol Na atoms, 2.58 mol S atoms, and 10.3 mol O atoms

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9.7: Stoichiometry

Learning Objectives

- Define *stoichiometry*.
- Relate quantities in a balanced chemical reaction on a molecular basis.

Consider a classic recipe for pound cake: 1 pound of eggs, 1 pound of butter, 1 pound of flour, and 1 pound of sugar. (That's why it's called "pound cake.") If you have 4 pounds of butter, how many pounds of sugar, flour, and eggs do you need? You would need 4 pounds each of sugar, flour, and eggs.

Now suppose you have 1.00 g H_2. If the chemical reaction follows the balanced chemical equation

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

then what mass of oxygen do you need to make water?

Curiously, this chemical reaction question is very similar to the pound cake question. Both of them involve relating a quantity of one substance to a quantity of another substance or substances. The relating of one chemical substance to another using a balanced chemical reaction is called **stoichiometry**. Using stoichiometry is a fundamental skill in chemistry; it greatly broadens your ability to predict what will occur and, more importantly, how much is produced.

Let us consider a more complicated example. A recipe for pancakes calls for 2 cups (c) of pancake mix, 1 egg, and 1/2 c of milk. We can write this in the form of a chemical equation:

2 c mix + 1 egg + 1/2 c milk → 1 batch of pancakes

If you have 9 c of pancake mix, how many eggs and how much milk do you need? It might take a little bit of work, but eventually you will find you need 4½ eggs and 2¼ c milk.

How can we formalize this? We can make a conversion factor using our original recipe and use that conversion factor to convert from a quantity of one substance to a quantity of another substance. This is similar to the way we constructed a conversion factor between feet and yards in Chapter 2. Because one recipe's worth of pancakes requires 2 c of pancake mix, 1 egg, and 1/2 c of milk, we actually have the following mathematical relationships that relate these quantities:

2 c pancake mix ⇔ 1 egg ⇔ 1/2 c milk

where ⇔ is the mathematical symbol for "is equivalent to." This does not mean that 2 c of pancake mix equal 1 egg. However, *as far as this recipe is concerned*, these are the equivalent quantities needed for a single recipe of pancakes. So, any possible quantities of two or more ingredients must have the same numerical ratio as the ratios in the equivalence.

We can deal with these equivalences in the same way we deal with equalities in unit conversions: we can make conversion factors that essentially equal 1. For example, to determine how many eggs we need for 9 c of pancake mix, we construct the conversion factor:

$$\frac{1\,egg}{2\,c\,pancake\,mix}$$

This conversion factor is, in a strange way, equivalent to 1 because the recipe relates the two quantities. Starting with our initial quantity and multiplying by our conversion factor,

$$9\,\cancel{c\,pancake\,mix} \times \frac{1\,egg}{\cancel{2\,c\,pancake\,mix}} = 4.5\,eggs$$

Note how the units *cups pancake mix* canceled, leaving us with units of *eggs*. This is the formal, mathematical way of getting our amounts to mix with 9 c of pancake mix. We can use a similar conversion factor for the amount of milk:

$$9\,c\,\cancel{pancake\ mix} \times \frac{\frac{1}{2}c\,milk}{2\,c\,\cancel{pancake\ mix}} = 2.25\,c\,milk$$

Again, units cancel, and new units are introduced.

A balanced chemical equation is nothing more than *a recipe for a chemical reaction*. The difference is that a balanced chemical equation is written in terms of atoms and molecules, not cups, pounds, and eggs.

For example, consider the following chemical equation:

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

We can interpret this as, literally, "two hydrogen molecules react with one oxygen molecule to make two water molecules." That interpretation leads us directly to some equivalencies, just as our pancake recipe did:

2H2 molecules ⇔ 1O2 molecule ⇔ 2H2O molecules

These equivalences allow us to construct conversion factors:

$$\frac{2\,molecules\,H_2}{1\,molecule\,O_2}\ \frac{2\,molecules\,H_2}{2\,molecules\,H_2O}\ \frac{1\,molecule\,H_2}{2\,molecules\,H_2O}$$

and so forth. These conversions can be used to relate quantities of one substance to quantities of another. For example, suppose we need to know how many molecules of oxygen are needed to react with 16 molecules of H2. As we did with converting units, we start with our given quantity and use the appropriate conversion factor:

$$\cancel{16\,molecules\,H_2} \times \frac{1\,molecules\,O_2}{\cancel{2\,molecules\,H_2}} = 8\,molecules\,O_2$$

Note how the unit *molecules H2* cancels algebraically, just as any unit does in a conversion like this. The conversion factor came directly from the coefficients in the balanced chemical equation. This is another reason why a properly balanced chemical equation is important.

Example 9.7.1

How many molecules of SO3 are needed to react with 144 molecules of Fe2O3 given this balanced chemical equation?

$$Fe_2O_3 + 3\,SO_3 \longrightarrow Fe_2(SO_4)_3$$

Solution

We use the balanced chemical equation to construct a conversion factor between Fe2O3 and SO3. The number of molecules of Fe2O3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO3 go on the top. Thus, the appropriate conversion factor is

$$\frac{3\,molecules\,SO_3}{1\,molecule\,Fe_2O_3}$$

Starting with our given amount and applying the conversion factor, the result is

$$144\ \cancel{molecules}\ Fe_2O_3 \times \frac{3\,molecules\,SO_3}{1\ \cancel{molecule}\ Fe_2O_3} = 432\,molecules\,SO_3$$

We need 432 molecules of SO3 to react with 144 molecules of Fe2O3.

Exercise 9.7.1

How many molecules of H_2 are needed to react with 29 molecules of N_2 to make ammonia if the balanced chemical equation is:

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3?$$

Answer available at the end of this section.

Chemical equations also allow us to make conversions regarding the number of atoms in a chemical reaction, because a chemical formula lists the number of atoms of each element in a compound. The formula H_2O indicates that there are two hydrogen atoms and one oxygen atom in each molecule, and these relationships can be used to make conversion factors:

$$\frac{2\,atoms\,H}{1\,molecule\,H_2O} \quad \frac{1\,molecule\,H_2O}{1\,atom\,O}$$

Conversion factors like this can also be used in stoichiometry calculations.

Example 9.7.2

How many molecules of NH_3 can you make if you have 228 atoms of H_2?

Solution

From the formula, we know that one molecule of NH_3 has three H atoms. Use that fact as a conversion factor:

$$228\ \cancel{atoms\,H} \times \frac{1\,molecule\,NH_3}{3\ \cancel{atoms\,H}} = 76\,NH_3$$

Exercise 9.7.2

How many molecules of $Fe_2(SO_4)_3$ can you make from 777 atoms of S?

Answer available at the end of this section.

Summary

Quantities of substances can be related to each other using balanced chemical equations.

Answers

Exercise 9.7.1

87 molecules

Exercise 9.7.2

259 molecules

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9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!

Learning Objectives

- Balance a chemical equation in terms of moles.
- Use the balanced equation to construct conversion factors in terms of moles.
- Calculate moles of one substance from moles of another substance using a balanced chemical equation.

Consider this balanced chemical equation:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

We interpret this as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:

$$100\,H_2 + 50\,O_2 \longrightarrow 100\,H_2O$$

This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:

$$5{,}000\,H_2 + 2{,}500\,O_2 \longrightarrow 5{,}000\,H_2O$$

Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:

$$12.044 \times 10^{23} H_2 + 6.022 \times 10^{23} O_2 \rightarrow 12.044 \times 10^{23} H_2O$$

These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as "two moles of hydrogen react with one mole of oxygen to make two moles of water." All balanced chemical reactions are balanced in terms of moles.

Example 9.8.1

Interpret this balanced chemical equation in terms of moles.

$$P_4 + 5\,O_2 \longrightarrow P_4O_{10}$$

Solution

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Exercise 9.8.1

Interpret this balanced chemical equation in terms of moles.

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

Answer available at the end of this section.

In Section 4.1, we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalents—equivalents between the reactants and the products. However, now we understand that *these equivalents are expressed in terms of moles*. Consider the chemical equation

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

This chemical reaction gives us the following equivalents:

$$2 \text{ mol } H_2 \Leftrightarrow 1 \text{ mol } O_2 \Leftrightarrow 2 \text{ mol } H_2O$$

Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H_2 and 1 mol of O_2 and use it to convert from moles of one substance to moles of another:

$$17.6\ \cancel{mol\ H_2} \times \frac{1\ mol\ O_2}{2\ \cancel{mol\ H_2}} = 8.80\ mol\ O_2$$

Note how the mol H_2 unit cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation,** when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won't always be so simple!

Example 9.8.2

For the balanced chemical equation

$$2\,C_4H_{10}\,(g) + 13\,O_2 \longrightarrow 8\,CO_2\,(g) + 10\,H_2O(\ell)$$

if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is

$$13 \text{ mol } O_2 \Leftrightarrow 8 \text{ mol } CO_2$$

We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154\ \cancel{mol\ O_2} \times \frac{8\ mol\ CO_2}{13\ \cancel{mol\ O_2}} = 94.8\ mol\ CO_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they do not contribute to the number of significant figures in the final answer.

Exercise 9.8.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer available at the end of this section.

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but

moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

Answers

Exercise 9.8.1

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

Exercise 9.8.2

118 mol

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9.9: Mole-Mass and Mass-Mass Calculations

Learning Objectives

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Suppose we know we have 123.2 g of Cl_2. How can we determine how many moles of $AlCl_3$ we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} = 1.738\ mol\,Cl_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\,mol\,AlCl_3}{3\,mol\,Cl_2}$$

Using this conversion factor with the molar quantity we calculated above, we get

$$1.738\ \cancel{mol\,Cl_2} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

So, we will get 1.159 mol of $AlCl_3$ if we react 123.2 g of Cl_2.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2 \ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 9.9.1

How many moles of HCl will be produced when 249 g of AlCl3 are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of AlCl3 to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of AlCl3 is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$249\ \cancel{g\ AlCl_3} \times \frac{1\,mol\ AlCl_3}{133.33\ \cancel{g\ AlCl_3}} = 1.87\,mol\ AlCl_3$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl3 and the number of moles of HCl:

$$frac6\,mol\ HCl2\ \cancel{mol\ AlCl_3}$$

Applying this conversion factor to the quantity of AlCl3, we get

$$1.87\ \cancel{mol\ AlCl_3} \times \frac{6\,mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.61\,mol\ HCl$$

Alternatively, we could have done this in one line:

$$249\ \cancel{g\ AlCl_3} \times \frac{1\ \cancel{mol\ AlCl_3}}{133.33\ \cancel{g\ AlCl_3}} \times \frac{6\,mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.60\,mol\ HCl$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Exercise 9.9.1

How many moles of Al2O3 will be produced when 23.9 g of H2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer available at the end of this section.

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

Example 9.9.2

How many grams of NH3 will be produced when 33.9 mol of H2 are reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have

$$33.9\ \cancel{mol\ H_2} \times \frac{2\ mol\ NH_3}{3\ \cancel{mol\ H_2}} = 22.6\ mol\ NH_3$$

Now, using the molar mass of NH_3, which is 17.03 g/mol, we get

$$22.6\ \cancel{mol\ NH_3} \times \frac{17.03\ g\ NH_3}{1\ \cancel{mol\ NH_3}} = 385\ g\ NH_3$$

Exercise 9.9.2

How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

$$\mathrm{N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)}$$

Answer available at the end of this section.

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO_3 that can be produced by the reaction of 45.3 g of SO_2 and O_2:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

First, we convert the given amount, 45.3 g of SO_2, to moles of SO_2 using its molar mass (64.06 g/mol):

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} = 0.707\ mol\ SO_2$$

Second, we use the balanced chemical reaction to convert from moles of SO_2 to moles of SO_3:

$$0.707\ \cancel{mol\ SO_2} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} = 0.707\ mol\ SO_3$$

Finally, we use the molar mass of SO_3 (80.06 g/mol) to convert to the mass of SO_3:

$$0.707\ \cancel{mol\ SO_3} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We can also perform all three steps sequentially, writing them on one line as

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO_3, which is what we are looking for, as our final answer.

Example 9.9.3

What mass of Mg will be produced when 86.4 g of K are reacted?

$$\mathrm{MgCl_2\,(s) + 2\,K\,(s) \longrightarrow Mg\,(s) + 2\,KCl\,(s)}$$

Solution

We will simply follow the steps

$$\text{mass K} \rightarrow \text{mol K} \rightarrow \text{mol Mg} \rightarrow \text{mass Mg}$$

In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4\ \cancel{g\,K} \times \frac{1\,mol\,K}{39.09\ \cancel{g\,K}} \times \frac{1\ \cancel{mol\,Mg}}{2\ \cancel{mol\,K}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 26.87\,g\,Mg$$

Exercise 9.9.3

What mass of H_2 will be produced when 122 g of Zn are reacted?

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Answer available at the end of this section.

Summary

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

Answers

Exercise 9.9.1

0.442 mol

Exercise 9.9.2

30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

Exercise 9.9.3

3.77 g

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9.10: Yields

Learning Objective

- Define and determine theoretical yields, actual yields, and percent yields.

In all the previous calculations we have performed involving balanced chemical equations, we made two assumptions:

1. The reaction goes exactly as written.
2. The reaction proceeds completely.

In reality, such things as side reactions occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as methane, some CO is produced as well as CO_2. However, we will continue to ignore side reactions, unless otherwise noted. The second assumption, that the reaction proceeds completely, is more troublesome. Many chemical reactions do not proceed to completion as written, for a variety of reasons (some of which we will consider in Chapter 13). When we calculate an amount of product assuming that all the reactant reacts, we calculate the **theoretical yield**, an amount that is theoretically produced as calculated using the balanced chemical reaction.

In many cases, however, this is not what really happens. In many cases, less—sometimes, much less—of a product is made during the course of a chemical reaction. The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made.

Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{9.10.1}$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%. Again, if percent yield is greater than 100%, an error has been made.

Example 9.10.1

A worker reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $Zn(NO_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?

$$\mathrm{Zn\,(s) + 2\,HNO_3\,(aq) \longrightarrow Zn(NO_3)_2\,(aq) + H_2\,(g)}$$

Solution

A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of Zn (65.39 g/mol) and $Zn(NO_3)_2$ (189.41 g/mol). In three steps, the mass-mass calculation is:

$$30.5\ \cancel{g\ Zn} \times \frac{1\ \cancel{mol\ Zn}}{65.39\ \cancel{g\ Zn}} \times \frac{1\ \cancel{mol\ Zn(NO_3)_2}}{1\ \cancel{mol\ Zn}} \times \frac{189.41\ g\ Zn(NO_3)_2}{1\ \cancel{mol\ Zn(NO_3)_2}} = 88.3\ g\ Zn(NO_3)_2$$

Thus, the theoretical yield is 88.3 g of $Zn(NO_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $Zn(NO_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100 (Equation 9.10.1):

$$\frac{65.2\ g\ Zn(NO_3)_2}{88.3\ g\ Zn(NO_3)_2} \times 100\% = 73.8\%$$

The worker achieved almost three-fourths of the possible yield.

Exercise 9.10.1

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2. What is the theoretical yield and the percent yield?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

*Technically, this is a reversible reaction (with double arrows), but for this exercise consider it irreversible (single arrow).

Answer

theoretical yield = 20.1 g; percent yield = 10.2%

Chemistry is Everywhere: Actual Yields in Drug Synthesis and Purification

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield might be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Even purifications of complex molecules into drug-quality purity are subject to percent yields. Consider the purification of impure albuterol. Albuterol ($C_{13}H_{21}NO_2$; accompanying figure) is an inhaled drug used to treat asthma, bronchitis, and other obstructive pulmonary diseases. It is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. The details of the steps do not concern us; only the percent yields do:

A child using an albuterol inhaler, a container of albuterol medication, and a molecular model of albuterol are shown in three combined images.

impure albuterol → intermediate A	percent yield = 70%
intermediate A → intermediate B	percent yield = 100%
intermediate B → intermediate C	percent yield = 40%
intermediate C → intermediate D	percent yield = 72%
intermediate D → purified albuterol	percent yield = 35%
overall percent yield = 70% × 100% × 40% × 72% × 35% = 7.5%	

That is, only about *one-fourteenth* of the original material was turned into the purified drug. This demonstrates one reason why some drugs are so expensive—a lot of material is lost in making a high-purity pharmaceutical.



Figure 9.10.1 A child using an albuterol inhaler, the container of albuterol medication, and a molecular model of the albuterol molecule. Source: Photo on far left © Thinkstock. Photo in center courtesy of Intropin, http://commons.wikimedia.org/wiki/Fi...te_%281%29.JPG.

Summary

Theoretical yield is the calculated yield using the balanced chemical reaction. Actual yield is what is actually obtained in a chemical reaction. Percent yield is a comparison of the actual yield with the theoretical yield.

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9.11: Limiting Reagents

Learning Objectives

- Identify a limiting reagent from a set of reactants.
- Calculate how much product will be produced from the limiting reagent.
- Calculate how much reactant(s) remains when the reaction is complete.

In addition to the assumption that reactions proceed all the way to completion, one additional assumption we have made about chemical reactions is that all the reactants are present in the proper quantities to react to products; this is not always the case. In Figure 9.11.2 we are taking hydrogen atoms and oxygen atoms (left) to make water molecules (right). However, there are not enough oxygen atoms to use up all the hydrogen atoms. We run out of oxygen atoms and cannot make any more water molecules, so the process stops when we run out of oxygen atoms.

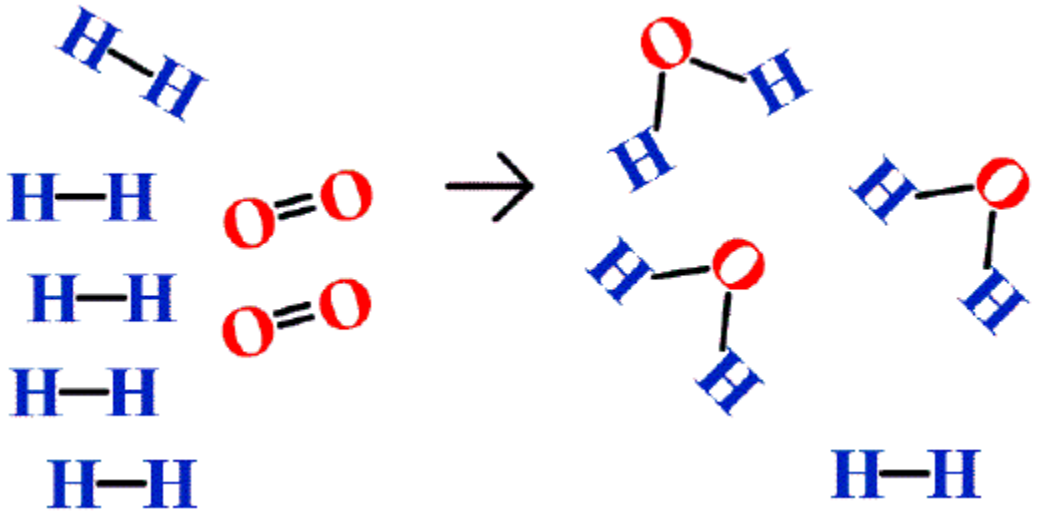


Figure 9.11.2: Making Water. In this scenario for making water molecules, we run out of O atoms before we use up all the H atoms. Similar situations exist for many chemical reactions when one reactant runs out before the other.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reagent; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass), assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

For example, consider this reaction:

$$4As(s) + 3O_2(g) \rightarrow 2As_2O_3(s)$$

Suppose we start a reaction with 50.0 g of As and 50.0 g of O2. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that each reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The calculations are as follows:

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{2\,mol\,As_2O_3}{4\ \cancel{mol\,As}} = 0.334\,mol\,As_2O_3$$

$$50.0\ \cancel{g\,O_2} \times \frac{1\ \cancel{mol\,O_2}}{32.00\ \cancel{g\,O_2}} \times \frac{2\,mol\,As_2O_3}{3\ \cancel{mol\,O_2}} = 1.04\,mol\,As_2O_3$$

Comparing these two answers, it is clear that 0.334 mol of As2O3 is less than 1.04 mol of As2O3, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is "in excess."

Identifying the limiting reagent, then, is straightforward. However, there are usually two associated questions: (1) what mass of product (or products) is then actually formed? and (2) what mass of what reactant is left over? The first question is straightforward to answer: simply perform a conversion from the number of moles of product formed to its mass, using its molar mass. For As2O3, the molar mass is 197.84 g/mol; knowing that we will form 0.334 mol of As2O3 under the given conditions, we will get

$$0.334\ \cancel{mol\,As_2O_3} \times \frac{197.84\,g\,As_2}{1\ \cancel{mol\,As_2O_3}} = 66.1\,g\,As_2O_3$$

The second question is somewhat more convoluted to answer. First, we must do a mass-mass calculation relating the limiting reagent (here, As) to the other reagent (O2). Once we determine the mass of O2 that reacted, we subtract that from the original amount to determine the amount left over. According to the mass-mass calculation,

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{3\ \cancel{mol\,O_2}}{4\ \cancel{mol\,As}} \times \frac{32.00\,g\,O_2}{1\ \cancel{mol\,O_2}} = 16.0\,g\,O_2\ reacted$$

Because we reacted 16.0 g of our original O2, we subtract that from the original amount, 50.0 g, to get the mass of O2 remaining:

50.0 g O2 – 16.0 g O2 reacted = 34.0 g O2 left over

You must remember to perform this final subtraction to determine the amount remaining; a common error is to report the 16.0 g as the amount remaining.

Example 9.11.1

A 5.00 g quantity of Rb is combined with 3.44 g of MgCl2 according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of Mg is formed, and what mass of what reactant is left over?

Solution

Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.

$$5.00\ \cancel{g\,Rb} \times \frac{1\ \cancel{mol\,Rb}}{85.47\ \cancel{g\,Rb}} \times \frac{1\ \cancel{mol\,Mg}}{2\ \cancel{mol\,Rb}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.711\,g\,Mg$$

$$3.44\ \cancel{g\,MgCl_2} \times \frac{1\ \cancel{mol\,MgCl_2}}{95.21\ \cancel{g\,MgCl_2}} \times \frac{1\ \cancel{mol\,Mg}}{1\ \cancel{mol\,MgCl_2}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.878\,g\,Mg$$

The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgCl2 reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount.

$$5.00\ \cancel{g\ Rb} \times \frac{1\ \cancel{mol\ Rb}}{85.47\ \cancel{g\ Rb}} \times \frac{1\ \cancel{mol\ MgCl_2}}{2\ \cancel{mol\ Rb}} \times \frac{95.21\ g\ Mg}{1\ \cancel{mol\ MgCl_2}} = 2.78\ g\ MgCl_2\ \ reacted$$

Because we started with 3.44 g of MgCl2, we have

3.44 g MgCl2 – 2.78 g MgCl2 reacted = 0.66 g MgCl2 left

Exercise 9.11.1

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?

$$\underbrace{22.7\ g}_{MgO(s)} + \underbrace{17.9\ g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer available at the end of this section.

Summary

The limiting reagent is the reactant that produces the least amount of product. Mass-mass calculations can determine how much product is produced and how much of the other reactants remain.

Answers

Exercise 9.11.1

H2S is the limiting reagent; 1.5 g of MgO are left over.

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9.12: Mass Percent Composition from a Chemical Formula

Learning Objectives

- Determine the percent composition of each element in a compound from the chemical formula.

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. This is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same, given that the compound is pure.

Example 9.12.1

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Solution

Solutions to Example 6.7.1

Steps for Problem Solving	**Calculate the percent composition of dichlorine heptoxide (Cl_2O_7).**
Identify the "given" information and what the problem is asking you to "find."	Given : Cl_2O_7 Find: % Composition (% Cl and %O)
List other known quantities.	Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g Molar mass of Cl_2O_7 = 182.90 g/mol
Cancel units and calculate.	$\%\text{Cl} = \frac{70.90\text{ g Cl}}{182.90\text{ g}} \times 100\% = 38.76\%\text{ Cl}$ $\%\text{ O} = \frac{112.00\text{ g O}}{182.90\text{ g}} \times 100\% = 61.24\%\text{ O}$ Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compour molar mass of the compound and multiplying by 100%.

Steps for Problem Solving	Calculate the percent composition of dichlorine heptoxide (Cl_2O_7).
Think about your result.	The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \text{Cl}$ and $61.24\%\ \text{O}$. Suppose that you needed to know the masses of chlorine and oxygen present in a $12.50\ \text{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50\ \text{g}\ Cl_2O_7 \times \frac{38.76\ \text{g Cl}}{100\ \text{g}\ Cl_2O_7} = 4.845\ \text{g Cl}$$

$$12.50\ \text{g}\ Cl_2O_7 \times \frac{61.24\ \text{g O}}{100\ \text{g}\ Cl_2O_7} = 7.655\ \text{g O}$$

The sum of the two masses is $12.50\ \text{g}$, the mass of the sample size.

Exercise 9.12.1

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a: available at the end of this section.

Summary

- Processes are described for calculating the percent composition of a compound based on the chemical formula.

Answers

Exercise 9.12.1

78.32% Ba and 21.67% F

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9.13: Calculating Empirical Formulas

Learning Objectives

- Define empirical formula.
- Determine empirical formula from percent composition of a compound.

In the early days of chemistry, there were few tools for the detailed study of compounds. Much of the information regarding the composition of compounds came from the elemental analysis of inorganic materials. The "new" field of organic chemistry (the study of carbon compounds) faced the challenge of not being able to characterize a compound completely. The relative amounts of elements could be determined, but so many of these materials had carbon, hydrogen, oxygen, and possibly nitrogen in simple ratios. We did not know exactly how many of these atoms were actually in a specific molecule.

Determining Empirical Formulas

An empirical formula tells us the relative ratios of different atoms in a compound. The ratios hold true on the *molar* level as well. Thus, H_2O is composed of two atoms of hydrogen and 1 atom of oxygen. Likewise, **1.0 mole of H_2O** is composed **of 2.0 moles of hydrogen** and **1.0 mole of oxygen**. We can also work backwards from molar ratios because *if we know the molar amounts of each element in a compound, we can determine the empirical formula*.

> In a procedure called elemental analysis, an unknown compound can be analyzed in the laboratory in order to determine the percentages of each element contained within it. These percentages can be transformed into the mole ratio of the elements, which leads to the empirical formula.

Empirical Formula: In Steps

Steps to determine empirical formula:

1. Assume a $100\ \text{g}$ sample of the compound so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.
3. In order to find a whole-number ratio, divide the moles of each element by whichever of the moles from step 2 is the smallest.
4. If all the moles at this point are whole numbers (or very close), the empirical formula can be written with the moles as the subscript of each element.
5. In some cases, one or more of the moles calculated in step 3 will not be whole numbers. Multiply each of the moles by the smallest whole number that will convert each into a whole number. Write the empirical formula.

Example 9.13.1

A compound of iron and oxygen is analyzed and found to contain 69.94% iron and 30.06% oxygen. Find the empirical formula of the compound.

Solution

Solutions to Example 6.8.1

Steps for Problem Solving	**Find the empirical formula of a compound of 69.94% iron and 30.06% oxygen.**
Identify the "given" information and what the problem is asking you to "find."	Given: $\%$ of $\mathrm{Fe} = 69.94\%$ $\%$ of $\mathrm{O} = 30.06\%$ Find: Empirical formula $= \mathrm{Fe}_?\mathrm{O}_?$
Calculate	
a. Assume a $100\ \mathrm{g}$ sample, convert the same % values to grams.	$69.94\ \mathrm{g\ Fe}$ $30.06\ \mathrm{g\ O}$
b. Convert to moles.	$69.94\ \mathrm{g\ Fe} \times \dfrac{1\ \mathrm{mol\ Fe}}{55.85\ \mathrm{g\ Fe}} = 1.252\ \mathrm{mol\ Fe}$ $30.06\ \mathrm{g\ O} \times \dfrac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g\ O}} = 1.879\ \mathrm{mol\ O}$
c. Divide both moles by the smallest of the results.	$\mathrm{Fe}: \dfrac{1.252\ \mathrm{mol}}{1.252}$ $\mathrm{O}: \dfrac{1.879\ \mathrm{mol}}{1.252}$ The "non-whole number" empirical formula of the compound is $\mathrm{Fe}_1\mathrm{O}_{1.5}$
Multiply each of the moles by the smallest whole number that will convert each into a whole number.	Fe:O = 2 (1:1.5) = 2:3 Since the moles of O is still not a whole number, both moles can be multiplied by 2, while rounding to a whole
Write the empirical formula.	The empirical formula of the compound is $\mathrm{Fe}_2\mathrm{O}_3$.

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Steps for Problem Solving	Find the empirical formula of a compound of 69.94% iron and 30.06% oxygen.
Think about your result.	The subscripts are whole numbers and represent the mole ratio of the elements in the compound. The compound i compound iron (III) oxide.

Exercise $9.13.1$

Mercury forms a compound with chlorine that is 73.9% mercury and 26.1% chlorine by mass. What is the empirical formula?

Answer available at the end of this section.

Summary

- A process is described for the calculation of the empirical formula of a compound, based on the percent composition of that compound.

Answers

Exercise $9.13.1$

$HgCl_2$

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9.14: Calculating Molecular Formulas

Learning Objectives

- Understand the difference between empirical formulas and molecular formulas.
- Determine molecular formula from percent composition and molar mass of a compound.

Below, we see two carbohydrates: glucose and sucrose. Sucrose is almost exactly twice the size of glucose, although their empirical formulas are very similar. Some people can distinguish them on the basis of taste, but it's not a good idea to go around tasting chemicals. The best way to tell glucose and sucrose apart is to determine the molar masses—this approach allows you to easily tell which compound is which.

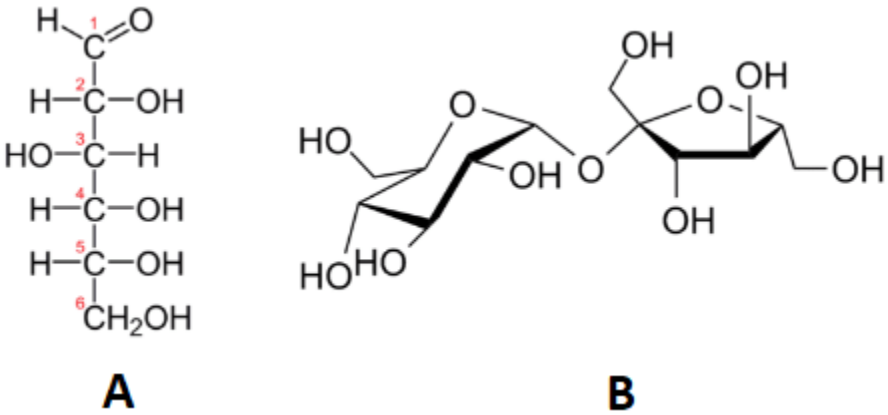


Figure 9.14.1: (A) the molecular structure of glucose and (B) the molecular structure of sucrose.

Molecular Formulas

Molecular formulas give the kind and number of atoms of each element present in the molecular compound. In many cases, the molecular formula is the same as the empirical formula. The chemical formula will always be some *integer multiple* (n) of the empirical formula (i.e. integer multiples of the subscripts of the empirical formula).

$$\text{Molecular Formula} = n(\text{Empirical formula})$$

therefore

$$n = \frac{\text{Molecular Formula}}{\text{Empirical Formula}}$$

The integer multiple, n, can also be obtained by dividing the molar mass, MM, of the compound by the empirical formula mass, EFM (the molar mass represented by the empirical formula).

$$n = \frac{MM(molarmass)}{EFM(empiricalformulamolarmass)}$$

Table 9.14.1 shows the comparison between the empirical and molecular formula of methane, acetic acid, and glucose, and the different values of n. The molecular formula of methane is $\mathrm{CH_4}$ and because it contains only one carbon atom, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole number multiple of the empirical formula. Acetic acid is an organic acid that is the main component of vinegar. Its molecular formula is $\mathrm{C_2H_4O_2}$. Glucose is a simple sugar that cells use as a primary source of energy. Its molecular formula is $\mathrm{C_6H_{12}O_6}$. The structures of both molecules are shown in Figure 9.14.2. They are very different compounds, yet both have the same empirical formula of $\mathrm{CH_2O}$.

Table 9.14.1: Molecular Formula and Empirical Formula of Various Compounds.

Name of Compound	Molecular Formula	Empirical Formula	n
Methane	CH_4	CH_4	1
Acetic acid	$C_2H_4O_2$	CH_2O	2
Glucose	$C_6H_{12}O_6$	CH_2O	6

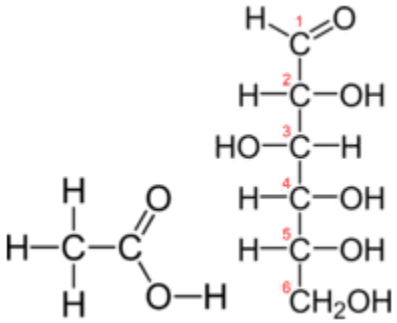


Figure 9.14.2: Acetic acid (left) has a molecular formula of $C_2H_4O_2$, while glucose (right) has a molecular formula of $C_6H_{12}O_6$. Both have the empirical formula CH_2O.

Empirical formulas can be determined from the percent composition of a compound as discussed in section 6.8. In order to determine its molecular formula, it is necessary to know the molar mass of the compound. Chemists use an instrument called a mass spectrometer to determine the molar mass of compounds. In order to go from the empirical formula to the molecular formula, follow these steps:

1. Calculate the **empirical formula molar mass (EFM)**.
2. Divide the molar mass of the compound by the empirical formula molar mass. The result should be a whole number or very close to a whole number.
3. Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.

Example 9.14.1

The empirical formula of a compound of boron and hydrogen is BH_3. Its molar mass is $27.7\ g/mol$. Determine the molecular formula of the compound.

Solution

Solutions to Example 6.9.1

Steps for Problem Solving	Determine the molecular formula of BH_3.
Identify the "given" information and what the problem is asking you to "find."	Given: Empirical formula $= BH_3$ Molar mass $= 27.7\ g/mol$ Find: Molecular formula $=?$

Steps for Problem Solving	Determine the molecular formula of BH_3.
Calculate the empirical formula mass (EFM).	$\text{Empirical formula molar mass (EFM)} = 13.84\ \text{g/mol}$
Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.	$\frac{\text{molar mass}}{\text{EFM}} = \frac{27.7 g/mol}{13.84 g/mol} = 2$
Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.	$BH_3 \times 2 = B_2H_6$
Write the molecular formula.	The molecular formula of the compound is B_2H_6.
Think about your result.	The molar mass of the molecular formula matches the molar mass of the compound.

Exercise 9.14.1

Vitamin C (ascorbic acid) contains 40.92 % C, 4.58 % H, and 54.50 % O, by mass. The experimentally determined molecular mass is 176 amu. What are the empirical and chemical formulas for ascorbic acid?

Answer Empirical Formula available at the end of this section.

Summary

- A procedure is described that allows the calculation of the exact molecular formula for a compound.

Answers

Exercise 9.14.1

$C_3H_4O_3$

$C_6H_8O_6$

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10: GASES

10.1: Kinetic Molecular Theory- A Model for Gases
10.2: What is Gas Pressure?
10.3: Boyle’s Law - Pressure and Volume
10.4: Charles’s Law- Volume and Temperature
10.5: Gay-Lussac's Law- Temperature and Pressure
10.6: Avogadro’s Law- Volume and Moles
10.7: The Combined Gas Law
10.8: The Ideal Gas Law
10.9: Molar Volume of a Gas at STP
10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!
10.11: Gas Stoichiometry
10.12: Dalton's Law of Partial Pressures

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10.1: Kinetic Molecular Theory- A Model for Gases

Learning Objectives

- State the major concepts behind the kinetic theory of gases.
- Relate the general properties of gases to the kinetic theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas.

Figure 10.1.1 shows a representation of how we mentally picture the gas phase.

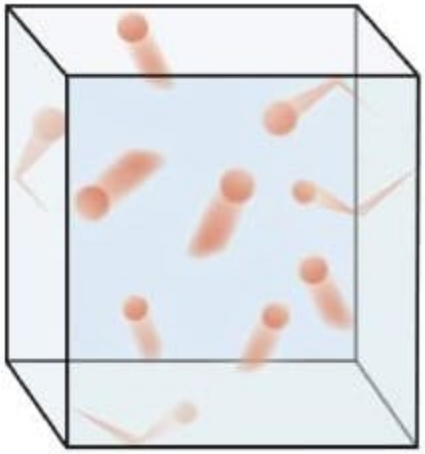


Figure 10.1.1: The Kinetic Theory of Gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

- The physical behavior of gases is explained by the kinetic theory of gases.
- An ideal gas adheres exactly to the kinetic theory of gases.

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10.2: What is Gas Pressure?

Learning Objectives

- Define *pressure*.
- Learn the units of pressure and how to convert between them.

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being "held in" by the earth's gravity, rather than the gas being in a container. The pressure of the atmosphere is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

However, "average atmospheric pressure at sea level" is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mmHg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalents:

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr}$$

We can use these equivalents as with any equivalence—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, 1 atm = 101,325 Pa.

Example 10.2.1: Pressure Conversion

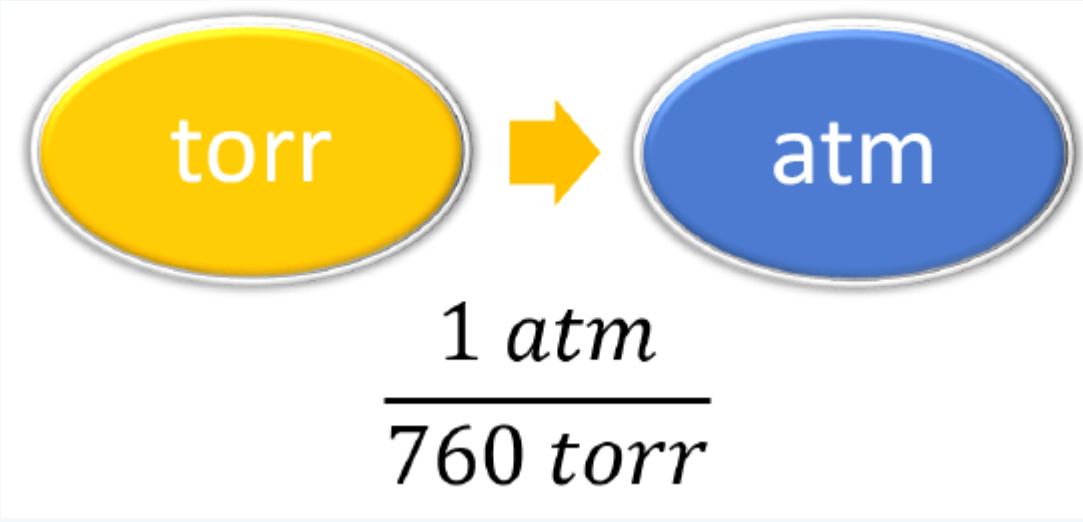
How many atmospheres are there in 595 torr?

Solution

Solutions to Example 11.3.1

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 595 torr Find: ? atm

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Steps for Problem Solving	Unit Conversion
List other known quantities.	1 atm = 760 torr
Prepare a concept map.	
Cancel units and calculate.	$595 \ \cancel{torr} \times \frac{1 \, atm}{760 \ \cancel{torr}} = 0.783 \ atm$
Think about your result.	595 torr is less than 760 torr so the final answer should be less than 1 atm.

Exercise 10.2.1

How many atmospheres are there in 1,022 torr?

Answer available at the end of this section.

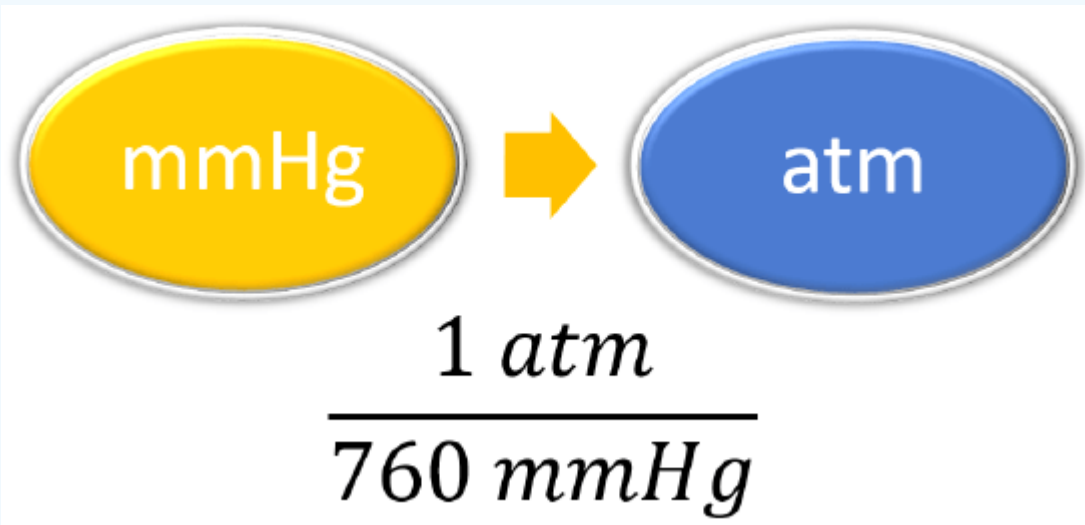
Example 10.2.2: Mars

The atmosphere on Mars is largely CO_2 at a pressure of 6.01 mmHg. What is this pressure in atmospheres?

Solution

Solutions to Example 11.3.2

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 6.01mmHg Find: ? atm
List other known quantities.	1 atm = 760 mmHg

Steps for Problem Solving	Unit Conversion
Prepare a concept map.	 $\frac{1\ atm}{760\ mmHg}$
Cancel units and calculate.	$6.01\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}} = 0.00791\ atm = 7.91 \times 10^{-3} atm$
Think about your result.	6.01 is a very small number relative to 760 mmHg, just like the value in atmospheres.

Exercise 10.2.2

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer available at the end of this section.

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

Answers

Exercise 10.2.1

1.345 atm

Exercise 10.2.2

652 torr

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10.3: Boyle's Law - Pressure and Volume

Learning Objectives

- Learn what is meant by the term *gas laws*.
- Learn and apply Boyle's Law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Take pressure (*P*) and volume (*V*), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [*n*]), if the temperature (*T*) of the gas was kept constant, pressure and volume were related: as one increases, the other decreases. As one decreases, the other increases. This means that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{ constant at constant n and T}$$

If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled P_1 and V_1 and the new conditions are labeled P_2 and V_2, we have

$$P_1V_1 = \text{constant} = P_2V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1V_1 = P_2V_2 \text{ at constant n and T}$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law is called **Boyle's Law**, after the English scientist Robert Boyle, who first announced it in 1662. Figure $10.3.1$ shows two representations of how Boyle's Law works.

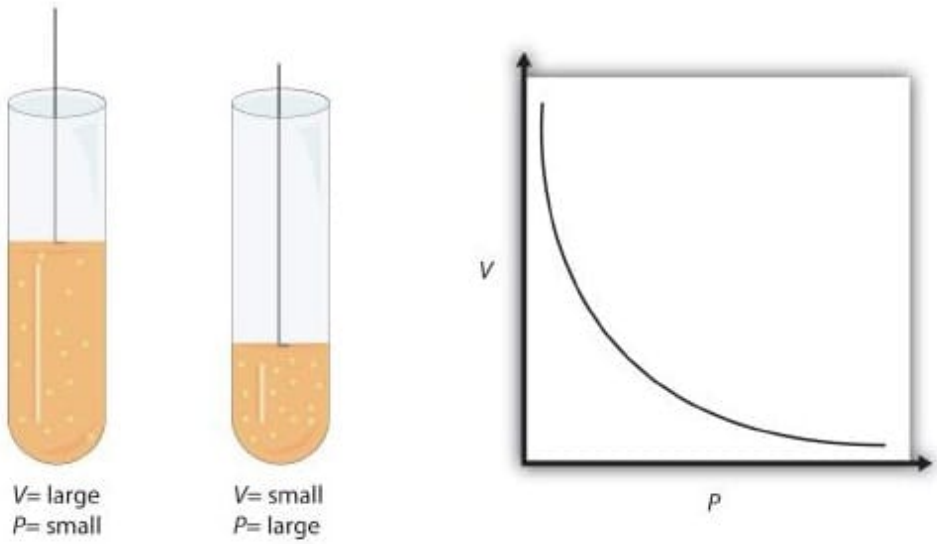


Figure $10.3.1$: Boyle's Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot *P* versus *V* for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

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Boyle's Law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle's Law there are two pressure variables; they must have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won't matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

Example 10.3.1

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

Solution

Solutions to Example 11.8.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 2.44 atm and V_1 = 4.01 L P_2 = 1.93 atm Find: V_2 = ? L
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{P_1 \times V_1}{P_2}$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.44 \cancel{\text{ atm}} \times 4.01 \text{ L}}{1.93 \cancel{\text{ atm}}} = 5.07 \text{ L}$$
Think about your result.	We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreas 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answ sense based on Boyle's Law.

Exercise 10.3.1

If P_1 = 334 torr, V_1 = 37.8 mL, and P_2 = 102 torr, what is V_2?

Answer available at the end of this section.

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As mentioned, you can use any units for pressure and volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

Example 10.3.2:

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

Solution

Solutions to Example 11.8.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 722 torr and V_1 = 88.8 mL V_2 = 0.633 L Find: P_2 = ? torr
List other known quantities.	1 L = 1000 mL to have the same units for volume.
Plan the problem.	1. Perform the conversion of the second volume unit from L to mL. 2. Rearrange the equation algebraically to solve for P_2. $$P_2 = \frac{P_1 \times V_1}{V_2}$$
Cancel units and calculate.	1. $$0.663 \cancel{L} \times \frac{1000\,ml}{1\,\cancel{L}} = 663\,ml$$ 2. Substitute the known quantities into the equation and solve. $$P_2 = \frac{722\text{ torr} \times 88.8\,\cancel{\text{mL}}}{663\,\cancel{\text{mL}}} = 96.7\text{ torr}$$
Think about your result.	When the volume increased, the pressure decreased, which is as expected for Boyle's Law.

Exercise 10.3.2

If V_1 = 456 mL, P_1 = 308 torr, and P_2 = 1.55 atm, what is V_2?

Answer available at the end of this section.

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Summary

- The behavior of gases can be modeled with gas laws.
- Boyle's Law relates the pressure and volume of a gas at constant temperature and amount.

Answers

Exercise 10.3.1

124 mL

Exercise 10.3.2

119 mL

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10.4: Charles's Law- Volume and Temperature

Learning Objectives

- Learn and apply Charles's Law.

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stop of molecular motion.

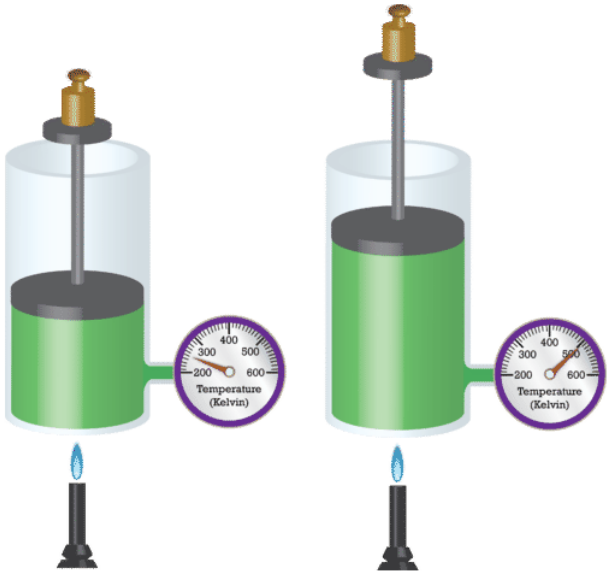


Figure 10.4.1: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 10.4.1: Temperature-Volume Data

Temperature (K)	**Volume** (mL)	$\frac{V}{T} = k \left(\frac{\mathrm{mL}}{\mathrm{K}}\right)$
50	20	0.40
100	40	0.40

Temperature (K)	Volume (mL)	$\frac{V}{T} = k\left(\frac{\mathrm{mL}}{\mathrm{K}}\right)$
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

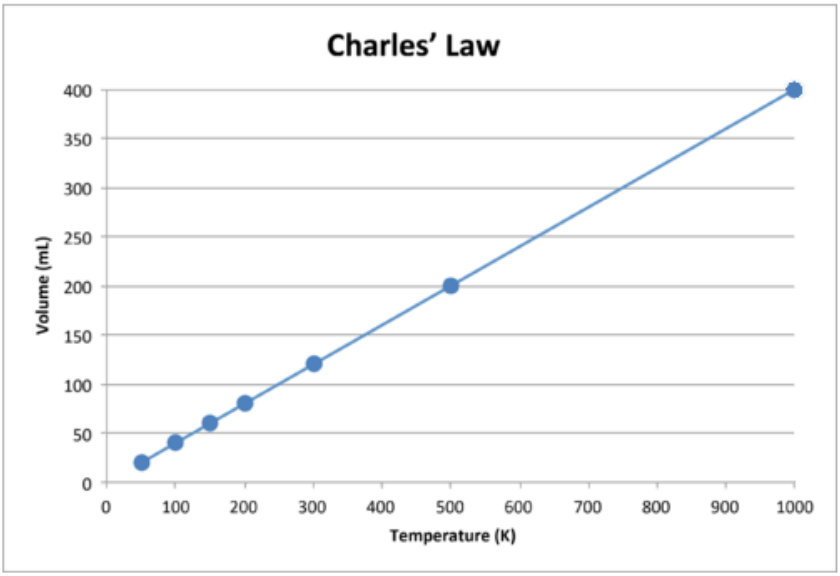


Figure $10.4.2$:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\mathrm{K} = {}^{\circ}\mathrm{C} + 273$.

Example $10.4.1$:

A balloon is filled to a volume of $2.20\ \mathrm{L}$ at a temperature of $22^{\circ}\mathrm{C}$. The balloon is then heated to a temperature of $71^{\circ}\mathrm{C}$. Find the new volume of the balloon.

Solution

Solutions to Example 11.5.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 2.20\ \text{L}$ and $T_1 = 22^\circ\text{C} = 295\ \text{K}$ $T_2 = 71^\circ\text{C} = 344\ \text{K}$ Find: V_2 = ? L
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times T_2}{T_1}$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.20\ \text{L} \times 344\ \cancel{\text{K}}}{295\ \cancel{\text{K}}} = 2.57\ \text{L}$$
Think about your result.	The volume increases as the temperature increases. The result has three significant figures.

Exercise 10.4.1

If V_1 = 3.77 L and T_1 = 255 K, what is V_2 if T_2 = 123 K?

Answer available at the end of this section.

Example 10.4.2:

A sample of a gas has an initial volume of 34.8 L and an initial temperature of −67°C. What must be the temperature of the gas for its volume to be 25.0 L?

Solution

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Solutions to Example 11.5.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Given:T_1 = -27°C and V_1 = 34.8 L V_2 = 25.0 L Find: T_2 = ? K
List other known quantities.	K = -27°C + 273
Plan the problem.	1. Convert the initial temperature to Kelvin 2. Rearrange the equation algebraically to solve for T_2. $$T_2 = \frac{V_2 \times T_1}{V_1}$$
Cancel units and calculate.	1. –67°C + 273 = 206 K 2. Substitute the known quantities into the equation and solve. $$T_2 = \frac{25.0 \cancel{\text{L}} \times 206 \text{ K}}{34.8 \cancel{\text{L}}} = 148 \text{ K}$$
Think about your result.	This is also equal to –125°C. As temperature decreases, volume decreases—which it does in this example.

Exercise 10.4.2

If V_1 = 623 mL, T_1 = 255°C, and V_2 = 277 mL, what is T_2?

Answer available at the end of this section.

Summary

- Charles's Law relates the volume and temperature of a gas at constant pressure and amount.

Answers

Exercise 10.4.1

1.82 L

Exercise 10.4.2

235 K, or –38°C

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10.5: Gay-Lussac's Law- Temperature and Pressure

Learning Objectives

- Explain Gay-Lussac's Law.

Propane tanks are widely used with barbeque grills. But it's not fun to find out halfway through grilling that you have run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778-1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.



Figure $10.5.1$: Joseph Gay-Lussac.

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \quad \text{and} \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume, its pressure continually decreases until the gas condenses to a liquid.

Example $10.5.1$

The gas in an aerosol can is under a pressure of 3.00 atm at a temperature of 25°C. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of 845°C?

Solution

Solutions to Example 11.10.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $P_1 = 3.00\ \text{atm}$ $T_1 = 25^\circ\text{C} = 298\ \text{K}$ $T_2 = 845^\circ\text{C} = 1118\ \text{K}$ Find: $P_2 = ?\ \text{atm}$
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for P_2. $$P_2 = \frac{P_1 \times T_2}{T_1}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$P_2 = \frac{3.00\ \text{atm} \times 1118\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.3\ \text{atm}$$
Think about your result.	The pressure increases dramatically due to a large increase in temperature.

Summary

- Pressure and temperature at constant volume are directly proportional.

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10.6: Avogadro’s Law- Volume and Moles

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. How much air should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles (or number of particles) of gas when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

where n is the number of moles of gas and k is a constant. Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

Example 10.6.1

A balloon has been filled to a volume of $1.90\ \text{L}$ with $0.0920\ \text{mol}$ of helium gas. If $0.0210\ \text{mol}$ of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

Solution

Solutions to Example 11.11.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 1.90\ \text{L}$ $n_1 = 0.0920\ \text{mol}$ Find: $V_2 = ?\ \text{L}$
List other known quantities.	Note that the final number of moles has to be calculated by adding the original number of moles to the moles o helium. $n_2 = 0.0920 + 0.0210 = 0.1130\ \text{mol}$

Steps for Problem Solving	
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times n_2}{n_1}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{1.90\ \text{L} \times 0.1130\ \cancel{\text{mol}}}{0.0920\ \cancel{\text{mol}}} = 2.33\ \text{L}$$
Think about your result.	Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly.

Exercise 10.6.1

A 12.8 L volume of gas contains .000498 moles of oxygen gas. At constant temperature and pressure, what volume does .0000136 moles of the gas fill?

Answer available at the end of this section.

Summary

- Calculations for relationships between volume and number of moles of a gas can be performed using Avogadro's Law.

Answers

Exercise 10.6.1

0.350 L

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10.7: The Combined Gas Law

Learning Objectives

- Learn and apply the Combined Gas Law.

One thing we notice about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that we can propose a gas law that combines pressure, volume, and temperature. This gas law is known as the **Combined Gas Law**, and its mathematical form is

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \; at \; constant \; n$$

This allows us to follow changes in all three major properties of a gas. Again, the usual warnings apply about how to solve for an unknown algebraically (isolate it on one side of the equation in the numerator), units (they must be the same for the two similar variables of each type), and units of temperature must be in Kelvin.

Example 10.7.1:

A sample of gas at an initial volume of 8.33 L, an initial pressure of 1.82 atm, and an initial temperature of 286 K simultaneously changes its temperature to 355 K and its volume to 5.72 L. What is the final pressure of the gas?

Solution

Solutions to Example 11.4.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: V_1 = 8.33 L, P_1 = 1.82 atm, and T_1 = 286 K V2 = 5.72 L and T_2 = 355 K Find: P_2 = ? atm
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

Steps for Problem Solving	
Calculate.	Now substitute the known quantities into the equation and solve. $$P_2 = \frac{(1.82\,atm)(8.33\,\cancel{L})(355\,\cancel{K})}{(286\,\cancel{K})(5.72\,\cancel{L})} = 3.22atm$$
Think about your result.	Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas we changing.

Exercise 10.7.1

If P_1 = 662 torr, V_1 = 46.7 mL, T_1 = 266 K, P_2 = 409 torr, and T_2 = 371 K, what is V_2?

Answer available at the end of this section.

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be in the numerator and all by itself on one side of the equation.

Summary

- The Combined Gas Law relates pressure, volume, and temperature of a gas.

Answers

Exercise 10.7.1

105 mL

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10.8: The Ideal Gas Law

Learning Objectives

- Explain the Ideal Gas Law.

There are a number of chemical reactions that require ammonia. In order to carry out the reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

Ideal Gas Law

The Combined Gas Law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. Putting these together leaves us with the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **Ideal Gas Law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable R for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R$$

The Ideal Gas Law is conveniently rearranged to look this way, with the multiplication signs omitted:

$$PV = nRT$$

The variable R in the equation is called the **ideal gas constant**.

Evaluating the Ideal Gas Constant

The value of R, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: kPa, atm, or $\mathrm{mm\ Hg}$. Therefore, R can have three different values.

We will demonstrate how R is calculated when the pressure is measured in kPa. The volume of $1.00\ \mathrm{mol}$ of any gas at STP (Standard temperature, 273.15 K and pressure, 1 atm) is measured to be $22.414\ \mathrm{L}$. We can substitute $101.325\ \mathrm{kPa}$ for pressure, $22.414\ \mathrm{L}$ for volume, and $273.15\ \mathrm{K}$ for temperature into the ideal gas equation and solve for R.

$$\begin{aligned} R &= \frac{PV}{nT} \\ &= \frac{101.325\ \mathrm{kPa} \times 22.414\ \mathrm{L}}{1.000\ \mathrm{mol} \times 273.15\ \mathrm{K}} \\ &= 8.314\ \mathrm{kPa \cdot L/K \cdot mol} \end{aligned}$$

This is the value of R that is to be used in the ideal gas equation when the pressure is given in kPa. The table below shows a summary of this and the other possible values of R. It is important to choose the correct value of R to use for a given problem.

Table 10.8.1*: Values of the Ideal Gas Constant*

Unit of P	Unit of V	Unit of n	Unit of T	Value and Unit of R
kPa	L	mol	K	$8.314\ \text{J/K} \cdot \text{mol}$
atm	L	mol	K	$0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol}$
mm Hg	L	mol	K	$62.36\ \text{L} \cdot \text{mm Hg/K} \cdot \text{mol}$

Notice that the unit for R when the pressure is in kPa has been changed to $\text{J/K} \cdot \text{mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

Example 10.8.1 Oxygen Gas

What volume is occupied by $3.76\ \text{g}$ of oxygen gas at a pressure of $88.4\ \text{kPa}$ and a temperature of 19°C? Assume the oxygen is ideal.

Solution

Solutions to Example 11.5.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: • $P = 88.4\ \text{kPa}$ • $T = 19^\circ\text{C} = 292\ \text{K}$ Mass $\text{O}_2 = 3.76\ \text{g}$ Find: *V* = ? L
List other known quantities.	$\text{O}_2 = 32.00\ \text{g/mol}$ $R = 8.314\ \text{J/K} \cdot \text{mol}$
Plan the problem.	1. First, determine the number of moles of O_2 from the given mass and the molar mass. 2. Then, rearrange the equation algebraically to solve for V $$V = \frac{nRT}{P}$$
Calculate.	1. $$3.76\ \cancel{\text{g}} \times \frac{1\ \text{mol O}_2}{32.00\ \cancel{\text{g}}\ \text{O}_2} = 0.1175\ \text{mol O}_2$$ 2. Now substitute the known quantities into the equation and solve. $$V = \frac{nRT}{P} = \frac{0.1175\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{J/K}} \cdot \cancel{\text{mol}} \times 292\ \cancel{\text{K}}}{88.4\ \cancel{\text{kPa}}} = 3.23\ \text{L O}_2$$

Steps for Problem Solving	
Think about your result.	The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar v $(22.4\,\mathrm{L/mol})$ since the pressure and temperature are reasonably close to standard. The result has three significant because of the values for T and P. Since a joule $(\mathrm{J}) = \mathrm{kPa} \cdot \mathrm{L}$, the units cancel out correctly, leaving a volume in lite

Example 10.8.2: Argon Gas

A 4.22 mol sample of Ar has a pressure of 1.21 atm and a temperature of 34°C. What is its volume?

Solution

Solutions to Example 11.5.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: n = 4.22 mol P = 1.21 atm T = 34°C Find: *V* = ? L
List other known quantities.	none
Plan the problem.	1. The first step is to convert temperature to Kelvin. 2. Then, rearrange the equation algebraically to solve for V $$V = \frac{nRT}{P}$$
Calculate.	1. 34 + 273 = 307 K 2. Now substitute the known quantities into the equation and solve. $$V = \frac{(4.22\ \cancel{mol})(0.08205 \frac{L.\cancel{atm}}{\cancel{mol.K}})(307\ \cancel{K})}{1.21\ \cancel{atm}}$$ $$= 87.9\ L$$
Think about your result.	The number of moles of Ar is large so the expected volume should also be large.

Exercise 10.8.1

A 0.0997 mol sample of O_2 has a pressure of 0.692 atm and a temperature of 333 K. What is its volume?

Answer available at the end of this section.

Exercise 10.8.2

For a 0.00554 mol sample of H_2, P = 23.44 torr and T = 557 K. What is its volume?

Answer available at the end of this section.

Summary

- The Ideal Gas Law is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas.

Answers

Exercise 10.8.1

3.94 L

Exercise 10.8.2

8.21 L

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10.9: Molar Volume of a Gas at STP



Figure 10.9.1 (Public Domain; User:Mark.murphy/Wikimedia Commons via Wikipedia)

How do scuba divers know if they will run out of gas?

Knowing how much gas is available for a dive is crucial to a diver's survival. The tank on the diver's back is equipped with gauges to indicate how much gas is present and what the pressure is. A basic knowledge of gas behavior allows the diver to assess how long they can stay underwater without developing problems.

Avogadro's Hypothesis and Molar Volume

Volume is a third way to measure the amount of matter, after item count and mass. With liquids and solids, volume varies greatly depending on the density of the substance. This is because solid and liquid particles are packed close together with very little space in between the particles. However, gases are largely composed of empty spaces between the actual gas particles (see figure below).

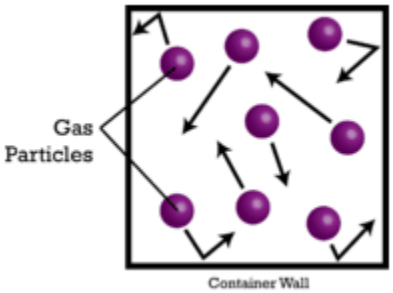


Figure 10.9.2: Gas particles are very small compared to the large amounts of empty space between them. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In 1811, Amadeo Avogadro explained that the volumes of all gases can be easily determined. **Avogadro's hypothesis** states that equal volumes of all gases at the same temperature and pressure contain equal numbers of particles. Since the total volume that a gas occupies is made up primarily of the empty space between the particles, the actual size of the particles themselves is nearly negligible. A given volume of a gas with small light particles, such as hydrogen $(\mathrm{H_2})$, contains the same number of particles as the same volume of a heavy gas with large particles, such as sulfur hexafluoride, $\mathrm{SF_6}$.

Gases are compressible, meaning that when put under high pressure, the particles are forced closer to one another. This decreases the amount of empty space and reduces the volume of the gas. Gas volume is also affected by temperature. When a gas is heated, its molecules move faster and the gas expands. Because of the variation in gas volume due to pressure and temperature changes, the comparison of gas volumes must be done at standard temperature and pressure. **Standard temperature and pressure (STP)** is defined as $0^{\circ}\mathrm{C}\ (273.15\ \mathrm{K})$ and $1\ \mathrm{atm}$ pressure. The **molar volume** of a gas is the volume of one mole of a gas at STP. At STP, one mole (6.02×10^{23} representative particles) of any gas occupies a volume of $22.4\ \mathrm{L}$ (figure below).

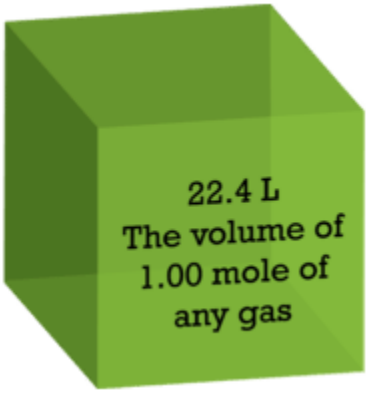
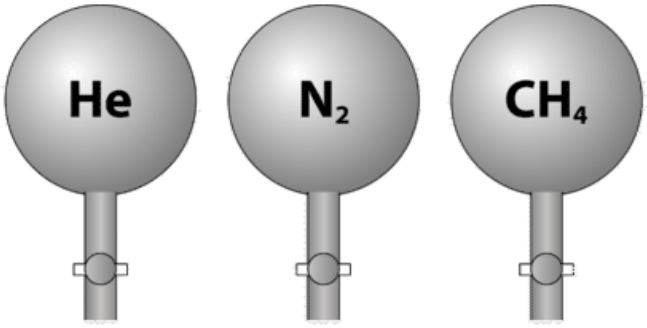


Figure $10.9.3$: A mole of any gas occupies $22.4\ \mathrm{L}$ at standard temperature and pressure ($0^{\circ}\mathrm{C}$ and $1\ \mathrm{atm}$). (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The figure below illustrates how molar volume can be seen when comparing different gases. Samples of helium (He), nitrogen $(\mathrm{N_2})$, and methane $(\mathrm{CH_4})$ are at STP. Each contains 1 mole or 6.02×10^{23} particles. However, the mass of each gas is different and corresponds to the molar mass of that gas: $4.00\ \mathrm{g/mol}$ for He, $28.0\ \mathrm{g/mol}$ for $\mathrm{N_2}$, and $16.0\ \mathrm{g/mol}$ for $\mathrm{CH_4}$.



Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Figure $10.9.4$: Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of particles. At standard temperature and pressure, 1 mole of any gas occupies $22.4\ \mathrm{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Summary

- Equal volumes of gases at the same conditions contain the same number of particles.
- Standard temperature and pressure is abbreviated (STP).
- Standard temperature is 0°C (273.15 K) and standard pressure is 1 atm.
- At STP, one mole of any gas occupies a volume of 22.4 L

Review

- A container is filled with gas, what do we know about the space actually taken up by a gas?
- Why do we need to do all our comparisons at the same temperature and pressure?
- At standard temperature and pressure, 1 mole of gas is always equal to how many liters?

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10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!

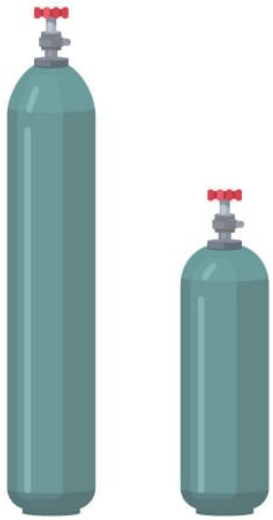


Figure 10.10.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How can you tell how much gas is in these containers?

Small gas tanks are often used to supply gases for chemistry reactions. A gas gauge will give some information about how much is in the tank, but quantitative estimates are needed so that the reaction will be able to proceed to completion. Knowing how to calculate the necessary parameters for gases is very helpful to avoid running out earlier than desired.

Conversions Between Moles and Gas Volume

Molar volume at STP can be used to convert from moles to gas volume and from gas volume to moles. The equality of $1\ \text{mol} = 22.4\ \text{L}$ is the basis for the conversion factor.

Example 10.10.1: Converting Gas Volume to Moles

Many metals react with acids to produce hydrogen gas. A certain reaction produces $86.5\ \text{L}$ of hydrogen gas at STP. How many moles of hydrogen were produced?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $86.5\ \text{L}\ \text{H}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- moles of H_2

Apply a conversion factor to convert from liters to moles.

Step 2: Calculate.

$$86.5\ \text{L}\ \text{H}_2 \times \frac{1\ \text{mol}\ \text{H}_2}{22.4\ \text{L}\ \text{H}_2} = 3.86\ \text{mol}\ \text{H}_2$$

Step 3: Think about your result.

The volume of gas produced is nearly four times larger than the molar volume. The fact that the gas is hydrogen plays no role in the calculation.

Example $10.10.2$**: Convertig Moles to Gas Volume**

What volume does $4.96\ \text{mol}$ of O_2 occupy at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $4.96\ \text{mol}\ \text{O}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- volume of O_2

Step 2: Calculate.

$$4.96\ \text{mol} \times 22.4\ \text{L/mol} = 111.1\ \text{L}$$

Step 3: Think about your result.

The volume seems correct given the number of moles.

Example $10.10.3$**: Converting Volume to Mass**

If we know the volume of a gas sample at STP, we can determine how much mass is present. Assume we have $867\ \text{L}$ of N_2 at STP. What is the mass of the nitrogen gas?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $867\ \text{L}\ \text{N}_2$
- $1\ \text{mol} = 22.4\ \text{L}$
- Molar mass $\text{N}_2 = 28.02\ \text{g/mol}$

Unknown

- mass of N_2

Step 2: Calculate.

We start by determining the number of moles of gas present. We know that 22.4 liters of a gas at STP equals one mole, so:

$$867\ \text{L} \times \frac{1\ \text{mol}}{22.4\ \text{L}} = 38.7\ \text{mol}$$

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We also know the molecular weight of N_2 $(28.0\ g/mol)$, so we can then calculate the weight of nitrogen gas in 867 liters:

$$38.7\ mol \times \frac{28\ g}{1\ mol} = 1083.6\ g\ N_2$$

Step 3: Think about your result.

In a multi-step problem, be sure that the units check out.

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Summary

- Conversions between moles and volume of a gas are shown.

Review

1. In the problems above, why was the gas always at standard temperature and pressure?
2. A container contains 45.2 L of N_2 gas at STP. How many moles of N_2 gas are in the container?
3. If the gas in the previous problem was CH_4 gas at STP instead of N_2 gas, then how many moles of CH_4 gas would there be?

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10.11: Gas Stoichiometry

The Haber cycle reaction of gaseous nitrogen and hydrogen to form ammonia is a critical step in the production of fertilizer from ammonia. It is important to have an excess of the starting materials so that a maximum yield of ammonia can be achieved. By knowing how much ammonia is needed for the manufacture of a batch of fertilizer, the proper amounts of nitrogen and hydrogen gases can be incorporated into the process.

Gas Stoichiometry

You have learned how to use molar volume to solve stoichiometry problems for chemical reactions involving one or more gases at STP. Now, we can use the ideal gas law to expand our treatment of chemical reactions to solve stoichiometry problems for reactions that occur at any temperature and pressure.

Example 10.11.1

What volume of carbon dioxide is produced by the combustion of $25.21\ \text{g}$ of ethanol $(\text{C}_2\text{H}_5\text{OH})$ at 54°C and $728\ \text{mm Hg}$?

Solution

Before using the ideal gas law, it is necessary to write and balance the chemical equation. Recall that in most combustion reactions, the given substance reacts with O_2 to form CO_2 and H_2O. Here is the balanced equation for the combustion of ethanol:

$$\text{C}_2\text{H}_5\text{OH}\,(l) + 3\text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g) + 3\text{H}_2\text{O}\,(l)$$

Step 1: List the known quantities and plan the problem.

Known

- Mass $\text{C}_2\text{H}_5\text{OH} = 25.21\ \text{g}$
- Molar mass $\text{C}_2\text{H}_5\text{OH} = 46.08\ \text{g/mol}$
- $P = 728\ \text{mm Hg}$
- $T = 54^\circ\text{C} = 327\ \text{K}$

Unknown

The number of moles of carbon dioxide gas is first calculated by stoichiometry. Then, the ideal gas law is used to calculate the volume of CO_2 produced.

Step 2: Solve.

$$25.21\ \text{g}\ \text{C}_2\text{H}_5\text{OH} \times \frac{1\ \text{mol}\ \text{C}_2\text{H}_5\text{OH}}{46.08\ \text{g}\ \text{C}_2\text{H}_5\text{OH}} \times \frac{2\ \text{mol}\ \text{CO}_2}{1\ \text{mol}\ \text{C}_2\text{H}_5\text{OH}} = 1.094\ \text{mol}\ \text{CO}_2$$

The moles of carbon dioxide (n) is now substituted into $PV = nRT$ to solve for the volume.

$$V = \frac{nRT}{P} = \frac{1.094\ \text{mol} \times 62.36\ \text{L} \cdot \text{mm Hg/K} \cdot \text{mol} \times 327\ \text{K}}{728\ \text{mm Hg}} = 30.6\ \text{L}$$

Step 3: Think about your result.

The mass of ethanol is slightly more than one half mole, meaning that the mole ratio results in slightly more than one mole of carbon dioxide being produced. Because of the elevated temperature and reduced pressure compared to STP, the resulting volume is larger than $22.4\ \text{L}$.

Summary

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10.12: Dalton's Law of Partial Pressures

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are 96.5% carbon dioxide and 3% nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus contributes pressure well over $2700\ \text{mm Hg}$. And there is no oxygen present, so humans could not breathe there. Not that anyone would want to go to Venus—the surface temperature is usually over 460°C.

Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about 78% nitrogen and 21% oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up 78% of the gas particles in a given sample of air, it exerts 78% of the pressure. If the overall atmospheric pressure is $1.00\ \text{atm}$, then the pressure of just the nitrogen in the air is $0.78\ \text{atm}$. The pressure of the oxygen in the air is $0.21\ \text{atm}$.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of nitrogen is represented by P_{N_2}. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's law can be expressed with the following equation:

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, P_1 and P_2, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300\ \text{mm Hg}$ and $P_2 = 500\ \text{mm Hg}$, then $P_{\text{total}} = 800\ \text{mm Hg}$.

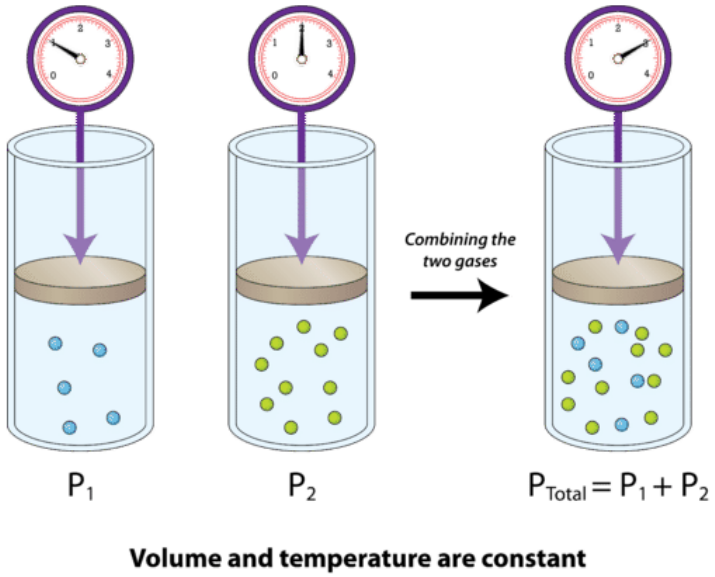


Figure $10.12.1$: Dalton's law says that the pressure of a gas mixture is equal to the partial pressures of the combining gases. (CC BY-NC; CK-12)

Summary

- The partial pressure of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture.
- The total pressure in a system is equal to the sums of the partial pressures of the gases present, according to Dalton's Law

of Partial Pressures.

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11: SOLUTIONS

11.1: Solutions - Homogeneous Mixtures
11.2: Solutions of Solids Dissolved in Water
11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
11.4: Solution Dilution (M1V1 = M2V2)
11.5: Calculating Solution Concentration- Mass Percent
11.6: Solution Stoichiometry

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11.1: Solutions - Homogeneous Mixtures

Learning Objectives

- Learn terminology involving solutions.
- Explain the significance of the statement "like dissolves like."
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater or lesser presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious.



Figure 11.1.1: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water, while air is a solution of a gaseous solute ($\mathrm{O_2}$) in a gaseous solvent ($\mathrm{N_2}$). In all cases, however, the overall phase of the solution is the same phase as the solvent. Table 11.1.1 lists some common types of solutions, with examples of each.

Table 11.1.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol ($\mathrm{C_2H_5OH}$) in $\mathrm{H_2O}$ (alcoholic beverages)
liquid	solid	salt water
solid	gas	$\mathrm{H_2}$ gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

Example 11.1.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 11.1.1

A solution is made by dissolving 3.33 g of $HCl(g)$ in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer available at the end of this section.

Like Dissolves Like

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, does dissolve in water—it is this oxygen that the fish take in through their gills. The reason we can enjoy carbonated sodas is also due to a nonpolar compound that dissolves in water. Pepsi-cola and all the other sodas have carbon dioxide gas, CO_2, a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Table 11.1.2: Summary of Solubilities

Solute (Polarity of Compound)	Solvent (Polarity of Compound)	Dominant Intermolecular Force	Is Solution Formed?
Polar	Polar	Dipole-Dipole Force and/or Hydrogen Bond	yes
Non-polar	Non-polar	Dispersion Force	yes
Polar	Non-polar		no
Non-polar	Polar		no
Ionic	Polar	Ion-Dipole	yes
Ionic	Non-polar		no

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges, a solution does not form. When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole

interactions, and hydrogen bonding (as described in Chapter 10). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water.



Figure 11.1.2: Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

Example 11.1.2: Polar and Nonpolar Solvents

Would $\mathrm{I_2}$ be more soluble in $\mathrm{CCl_4}$ or $\mathrm{H_2O}$? Explain your answer.

Solution

$\mathrm{I_2}$ is nonpolar. Of the two solvents, $\mathrm{CCl_4}$ is nonpolar and $\mathrm{H_2O}$ is polar, so $\mathrm{I_2}$ would be expected to be more soluble in $\mathrm{CCl_4}$.

Exercise 11.1.2

Would $\mathrm{C_3H_7OH}$ be more soluble in $\mathrm{CCl_4}$ or $\mathrm{H_2O}$? Explain your answer.

Answer available at the end of this section.

Example 11.1.3

Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol ($\mathrm{CH_3OH}$)
b. sodium sulfate (\ce{Na2SO4}\))
c. octane ($\mathrm{C_8H_{18}}$)

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

a. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
b. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.

c. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

Exercise 11.1.3: Toluene

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer available at the end of this section.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.

Answers

Exercise 11.1.1

solute: $HCl(g)$

solvent: CH_3OH

Exercise 11.1.2

H_2O, because both experience hydrogen bonding.

Exercise 11.1.3**: Toluene**

Octane (C_8H_{18}) will dissolve. It is also non-polar.

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11.2: Solutions of Solids Dissolved in Water

Learning Objectives

- Define electrolytes and non electrolytes
- Explain why solutions form.
- Discuss the idea of water as the "universal solvent".
- Explain how water molecules attract ionic solids when they dissolve in water.

We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. We also know that solutions have constant composition, and that this composition can be varied up to a point to maintain the homogeneous nature of the solution. But how exactly do solutions form? Why is it that oil and water will not form a solution, and yet vinegar and water will? Why could we dissolve table salt in water, but not in vegetable oil? The reasons why solutions will form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Solubility and Saturation

Table salt (NaCl) readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is specified as the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table $11.2.1$ lists the solubilities of some simple ionic compounds. These solubilities vary widely. NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table $11.2.1$*: Solubilities of Some Ionic Compounds*

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated, but rather concentrated. In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

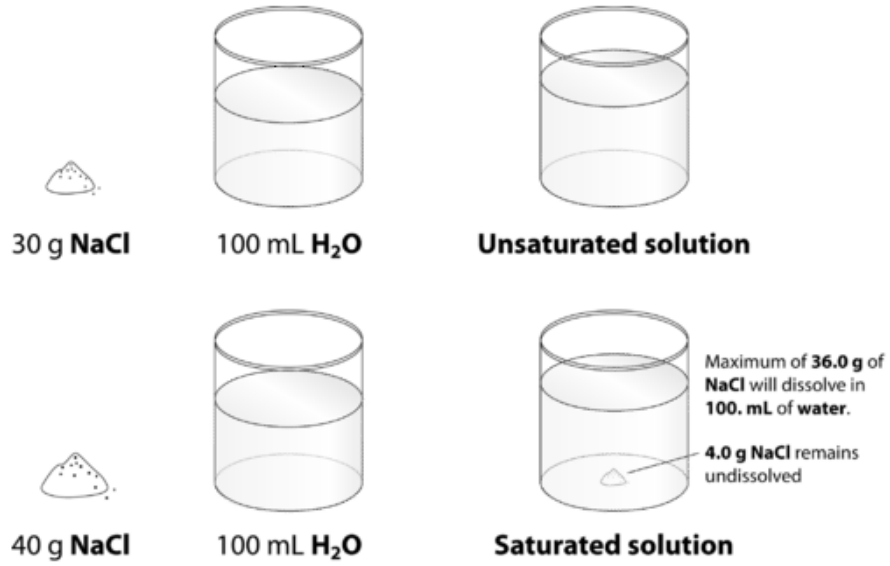


Figure 11.2.1: When $30.0\ \text{g}$ of NaCl is added to $100\ \text{mL}$, it all dissolves, forming an unsaturated solution. When $40.0\ \text{g}$ is added, $36.0\ \text{g}$ dissolves and $4.0\ \text{g}$ remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

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Electrolyte Solutions: Dissolved Ionic Solids

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 (good conductor). If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, the substance is a weak electrolyte (does not conduct electricity as well).

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 11.2.1).

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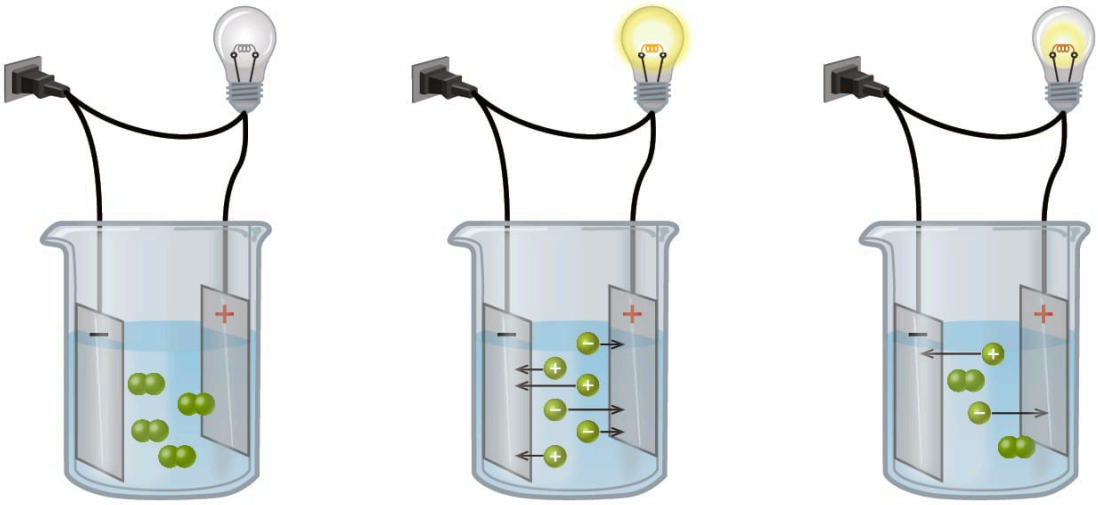


Figure 11.2.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte. This diagram shows three separate beakers. Each has a wire plugged into a wall outlet. In each case, the wire leads from the wall to the beaker and is split resulting in two ends. One end leads to a light bulb and continues on to a rectangle labeled with a plus sign. The other end leads to a rectangle labeled with a minus sign. The rectangles are in a solution. In the first beaker, labeled “Ethanol No Conductivity,” four pairs of linked small green spheres suspended in the solution between the rectangles. In the second beaker, labeled “K C l Strong Conductivity,” six individual green spheres, three labeled plus and three labeled minus are suspended in the solution. Each of the six spheres has an arrow extending from it pointing to the rectangle labeled with the opposite sign. In the third beaker, labeled “Acetic acid solution Weak conductivity,” two pairs of joined green spheres and two individual spheres, one labeled plus and one labeled minus are shown suspended between the two rectangles. The plus labeled sphere has an arrow pointing to the rectangle labeled minus and the minus labeled sphere has an arrow pointing to the rectangle labeled plus.

Water and other polar molecules are attracted to ions, as shown in Figure 11.2.2. 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.

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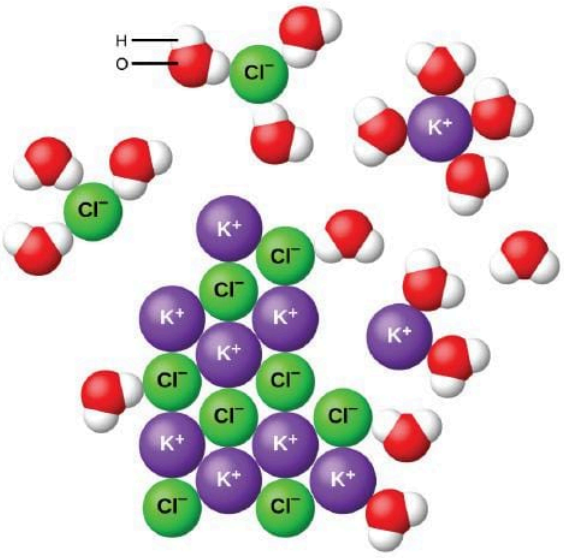


Figure 11.2.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown. The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled C l superscript minus mixed and touching near the center of the diagram. Outside of this cluster of spheres are seventeen clusters of three spheres, which include one red and two white spheres. A red sphere in one of these clusters is labeled O. A white sphere is labeled H. Two of the green C l superscript minus spheres are surrounded by three of the red and white clusters, with the red spheres closer to the green spheres than the white spheres. One of the K superscript plus purple spheres is surrounded by four of the red and white clusters. The white spheres of these clusters are closest to the purple spheres.

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.

Example 11.2.1: Identifying Ionic Compounds

Which compound(s) will dissolve in solution to separate into ions?

a. LiF
b. $\mathrm{P_2F_5}$
c. $\mathrm{C_2H_5OH}$

Solution

LiF will separate into ions when dissolved in solution, because it is an ionic compound. $\mathrm{P_2F_5}$ and $\mathrm{C_2H_5OH}$ are both covalent and will stay as molecules in a solution.

Exercise 11.2.1

Which compounds will dissolve in solution to separate into ions?

a. $C_6H_{12}O_{11}$, glucose
b. CCl_4
c. $CaCl_2$
d. $AgNO_3$

Answer available at the end of this section.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per $100\ \text{g}$ of solvent. The solubility of sodium chloride in water is $36.0\ \text{g}$ per $100\ \text{g}$ water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (Figure $11.2.4$).

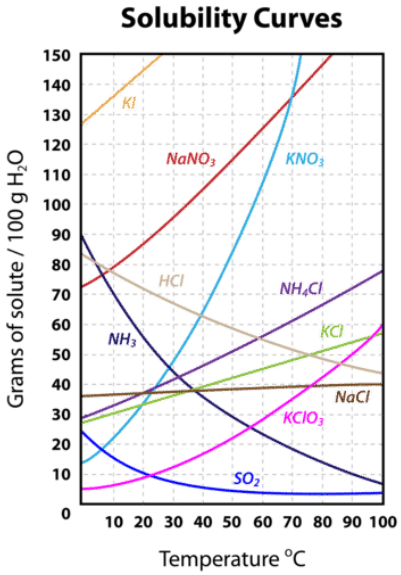


Figure $11.2.4$: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, NH_3, and SO_2—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ \text{g}$ of KNO_3 is added to $100\ \text{g}$ of water at 30°C. According to the solubility curve, approximately $48\ \text{g}$ of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since $48\ \text{g}$ is less than $80\ \text{g}$. We can also determine that there will be $80 - 48 = 32\ \text{g}$ of undissolved KNO_3 remaining at the bottom of the container. Now suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about $107\ \text{g}$. Now the solution is unsaturated since it contains only the original $80\ \text{g}$ of dissolved solute. Now suppose the solution is cooled all the way down to 0°C. The solubility at 0°C is about $14\ \text{g}$, meaning that $80 - 14 = 66\ \text{g}$ of the KNO_3 will re-crystallize.

Summary

- Solubility is the specific amount of solute that can dissolve in a given amount of solvent.
- Saturated and unsaturated solutions are defined.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).

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- Water is considered the universal solvent since it can dissolve both ionic and polar solutes, as well as some nonpolar solutes (in very limited amounts).
- The solubility of a solid in water increases with an increase in temperature.

Vocabulary

- **Miscible** - Liquids that have the ability to dissolve in each other.
- **Immiscible** - Liquids that do not have the ability to dissolve in each other.
- **Electrostatic attraction** - The attraction of oppositely charged particles.

Answers

Exercise 11.2.1

c & d

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11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)

Learning Objectives

- Use molarity to determine quantities in chemical reactions.
- Use molarity as a conversion factor in calculations.

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{11.3.1}$$

The symbol for molarity is M or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[\text{Ag}^+]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to perform calculations with, but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” This is read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> Be sure to note that molarity is calculated as the total volume of the **entire** solution, not just volume of solvent! The solute contributes to total volume.

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

Step 1: First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4 \, \cancel{g HCl} \times \frac{1 \, mol \, HCl}{36.5 \, \cancel{g HCl}} = 0.614 \, mol \, HCl$$

Step 2: Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614 \, mol \, HCl}{1.56 L \, solution} = 0.394 \, MHCl$$

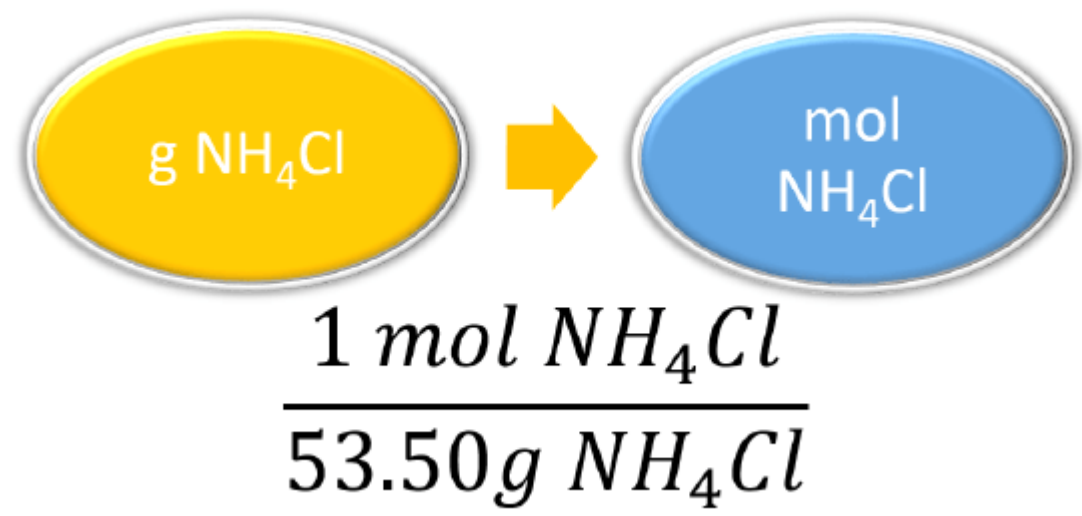
Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

Example 11.3.1

A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

Solution

Solutions to Example 13.6.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$ Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$ Find: Molarity = ? M
List other known quantities.	Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters. $\text{M} = \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \cancel{\text{g NH}_4\text{Cl}} \times \frac{1\ \text{mol NH}_4\text{Cl}}{53.50\ \cancel{\text{g NH}_4\text{Cl}}} = 0.7893\ \text{mol NH}_4\text{Cl}$ $\frac{0.7893\ \text{mol NH}_4\text{Cl}}{0.5000\ \text{L solution}} = 1.579\ \text{M}$
Think about your result.	The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol}\ \text{NH}_4\text{Cl}$. Four significant figures are appropriate.

Exercise 11.3.1A

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

Exercise 11.3.1B

What is the concentration, in $\mathrm{mol/L}$, where $137\ \mathrm{g}$ of NaCl has been dissolved in enough water to make $500\ \mathrm{mL}$ of solution?

Answer available at the end of this section.

Using Molarity in Calculations

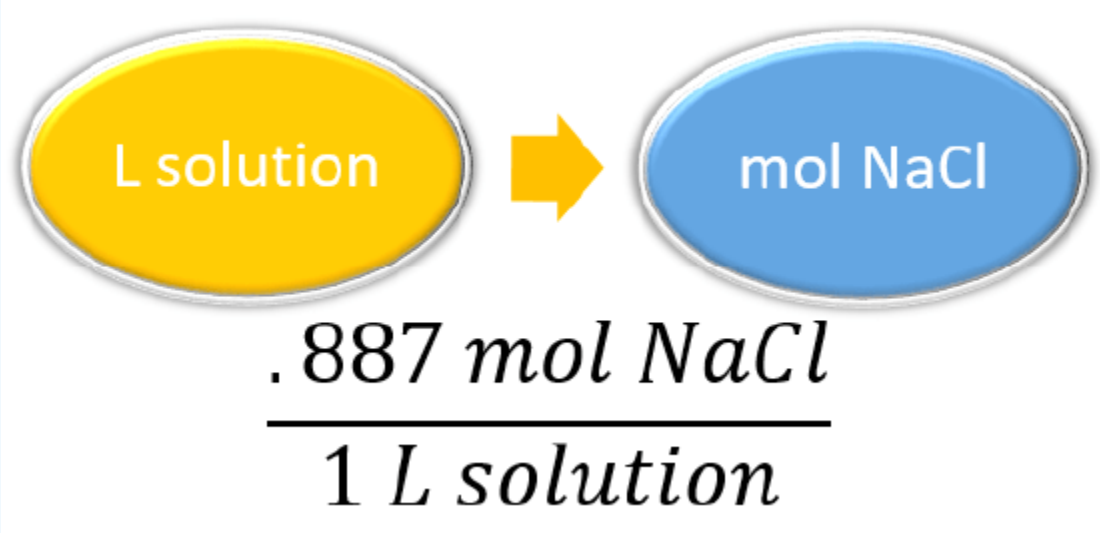
Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition.

Example 11.3.2: Determining Moles of Solute, Given the Concentration and Volume of a Solution

For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

Solution



$$0.108\ \cancel{L\ solution} \times \frac{0.887\ mol\ NaCl}{1\ \cancel{L\ solution}} = 0.0958\ mol\ NaCl$$

If we used the definition approach, we get the same answer, but now we are using conversion factor skills.

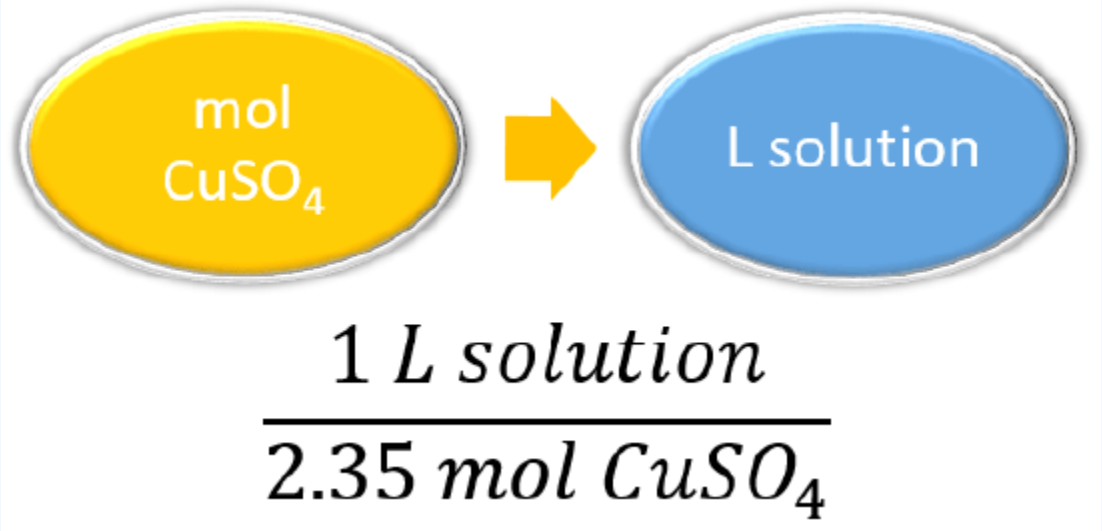
Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 11.3.3: Determining Volume of a Solution, Given the Concentration and Moles of Solute

Using concentration as a conversion factor, how many liters of 2.35 M CuSO4 are needed to obtain 4.88 mol of CuSO4?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:



$$4.88\ \cancel{mol\ CuSO_4} \times \frac{1\ L\ solution}{2.35\ \cancel{mol\ CuSO_4}} = 2.08\ L\ of\ solution$$

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams. The following example illustrates this.

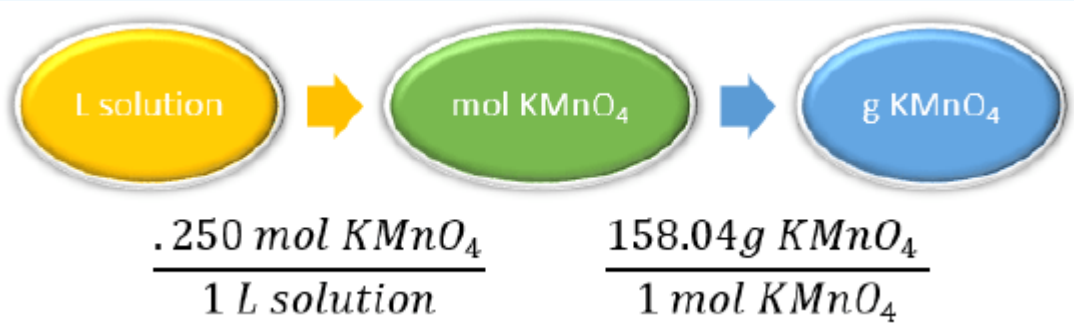
Example 11.3.4

A chemist needs to prepare $3.00\ \text{L}$ of a $0.250\ \text{M}$ solution of potassium permanganate (KMnO_4). What mass of KMnO_4 does she need to make the solution?

Solution

Solutions to Example 13.6.4

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Molarity $= 0.250\ \text{M}$ Volume $= 3.00\ \text{L}$ Find: Mass $\text{KMnO}_4 = ?\ \text{g}$

Steps for Problem Solving	
List other known quantities.	Molar mass $\mathrm{KMnO_4} = 158.04\ \mathrm{g/mol}$ 0.250 mol $KMnO_4$ to 1 L of $KMnO_4$ solution
Plan the problem.	
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $\mathrm{mol\ KMnO_4} = 0.250\ \mathrm{M\ KMnO_4} \times 3.00\ \mathrm{L} = 0.750\ \mathrm{mol\ KMnO_4}$ $3.00\ \cancel{L\ solution} \times \frac{0.250\ \cancel{\mathrm{mol\ KMnO_4}}}{1\ \cancel{L\ solution}} \times \frac{158.04\ \mathrm{g\ KMnO_4}}{1\ \cancel{\mathrm{mol\ KMnO_4}}} = 119\ \mathrm{g\ KMnO_4}$
Think about your result.	When $119\ \mathrm{g}$ of potassium permanganate is dissolved into water to make $3.00\ \mathrm{L}$ of solution, the molarity is $0.250\ \mathrm{M}$.

Exercise 11.3.4A

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer available at the end of this section.

Exercise 11.3.4B

Answer the problems below using concentration as a conversion factor.

a. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
b. What volume of 1.50 M HCl solution contains 10.0 g of hydrogen chloride?

Answer a available at the end of this section.

How to Indicate Concentration

- Square brackets are often used to represent concentration, e.g., [NaOH] = 0.50 M.
- Use the capital letter M for molarity, not a lower case m (this is a different concentration unit called *molality*).

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Watch as the Flinn Scientific Tech Staff demonstrates "How To Prepare Solutions."

It is important to note that there are many different ways you can set up and solve your chemistry equations. Some students prefer to answer multi-step calculations in one long step, while others prefer to work out each step individually. Neither method is necessarily better or worse than the other method—whichever makes the most sense to ***you*** is the one that you should use. In this text, we will typically use unit analysis (also called dimension analysis or factor analysis).

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

Answers

Exercise $11.3.1A$

1.57 M $C_6H_{12}O_6$

Exercise $11.3.1B$

4.69 M NaCl

Exercise $11.3.4A$

1.74 L

Exercise $11.3.4B$

6.12 g

183 mL or 0.183L

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11.4: Solution Dilution (M1V1 = M2V2)

Learning Objectives

- Explain how concentrations can be changed in the lab.
- Understand how stock solutions are used in the laboratory.

We are often concerned with how much solute is dissolved in a given amount of solution. We will begin our discussion of solution concentration with two related and relative terms: **dilute** and **concentrated**.

- A **dilute** solution is one in which there is a relatively small amount of solute dissolved in the solution.
- A **concentrated** solution contains a relatively large amount of solute.

These two terms do not provide any quantitative information (actual numbers), but they are often useful in comparing solutions in a more general sense. These terms also do not tell us whether or not the solution is saturated or unsaturated, or whether the solution is "strong" or "weak". These last two terms will have special meanings when we discuss acids and bases, so be careful not to confuse them.

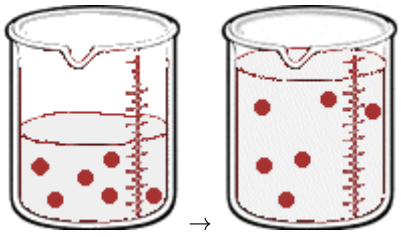
Stock Solutions

It is often necessary to have a solution with a concentration that is very precisely known. Solutions containing a precise mass of solute in a precise volume of solution are called **stock (or standard) solutions.** To prepare a standard solution, a piece of lab equipment called a volumetric flask should be used. These flasks range in size from 10 mL to 2000 mL and are carefully calibrated to a single volume. On the narrow stem is a **calibration mark**. The precise mass of solute is dissolved in a bit of the solvent, and this is added to the flask. Then, enough solvent is added to the flask until the level reaches the calibration mark.

Often, it is convenient to prepare a series of solutions of known concentrations by first preparing a single **stock solution**, as described in the previous section. **Aliquots** (carefully measured volumes) of the stock solution can then be diluted to any desired volume. In other cases, it may be inconvenient to weigh a small mass of sample accurately enough to prepare a small volume of a dilute solution. Each of these situations requires that a solution be diluted to obtain the desired concentration.

Dilutions of Stock (or Standard) Solutions

Imagine we have a salt water solution with a certain concentration. That means we have a certain amount of salt (a certain mass or a certain number of moles) dissolved in a certain volume of solution. Next, we will dilute this solution. This is done by adding more water, not more salt:



Before Dilution and After Dilution

The molarity of solution 1 is

$$M_1 = \frac{\text{moles}_1}{\text{liter}_1}$$

and the molarity of solution 2 is

$$M_2 = \frac{\text{moles}_2}{\text{liter}_2}$$

rearrange the equations to find moles:

$$\text{moles}_1 = M_1 \text{liter}_1$$

and

$$\text{moles}_2 = M_2 \text{liter}_2$$

What stayed the same and what changed between the two solutions? By adding more water, we changed the volume of the solution. Doing so also changed its concentration. **However, the number of moles of solute did not change.** So,

$$moles_1 = moles_2$$

Therefore

$$\boxed{M_1 V_1 = M_2 V_2} \tag{11.4.1}$$

where

- M_1 and M_2 are the concentrations of the original and diluted solutions
- V_1 and V_2 are the volumes of the two solutions

Preparing dilutions is a common activity in the chemistry lab and elsewhere. Once you understand the above relationship, the calculations are simple.

Suppose that you have $100.\ \text{mL}$ of a $2.0\ \text{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \text{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2.

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \text{M} \times 100.\ \text{mL}}{500.\ \text{mL}} = 0.40\ \text{M HCl}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves calculating what amount of a highly concentrated solution is required to make a desired quantity of solution of lesser concentration. The highly concentrated solution is typically referred to as the stock solution.

Example 11.4.1: Diluting Nitric Acid

Nitric acid (HNO_3) is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \text{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \text{L}$ of a $0.50\ \text{M}$ solution?

Solution

Solutions to Example13.7.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: M_1, Stock $\mathrm{HNO_3} = 16\ \mathrm{M}$ $V_2 = 8.00\ \mathrm{L}$ $M_2 = 0.50\ \mathrm{M}$ Find: Volume stock $\mathrm{HNO_3}\ (V_1) = ?\ \mathrm{L}$
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_1. $$V_1 = \frac{M_2 \times V_2}{M_1}$$
Calculate and cancel units.	Now substitute the known quantities into the equation and solve. $$V_1 = \frac{0.50\ \mathrm{M} \times 8.00\ \mathrm{L}}{16\ \mathrm{M}} = 0.25\ \mathrm{L}$$
Think about your result.	$0.25\ \mathrm{L}\ (250\ \mathrm{mL})$ of the stock $\mathrm{HNO_3}$ needs to be diluted with water to a final volume of $8.00\ \mathrm{L}$. The dilution is by a f 32 to go from $16\ \mathrm{M}$ to $0.5\ \mathrm{M}$.

Exercise 11.4.1

A 0.885 M solution of KBr with an initial volume of 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

Answer

135.4 mL

Note that the calculated volume will have the same dimensions as the input volume, and dimensional analysis tells us that in this case we don't need to convert to liters, since L cancels when we divide M (mol/L) by M (mol/L).

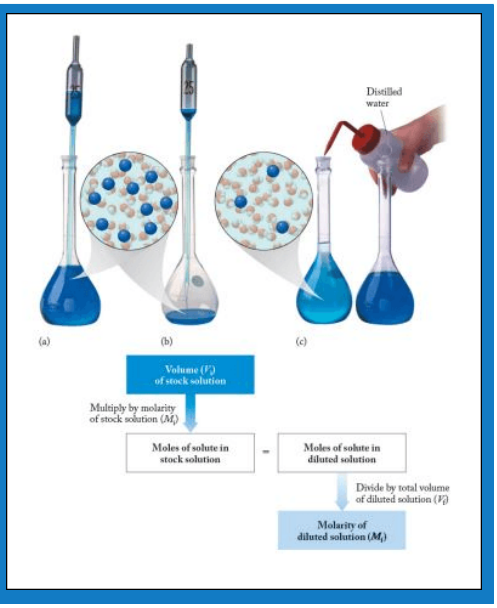
Diluting and Mixing Solutions

Diluting and Mixing Solutions

YouTube Video

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How to Dilute a Solution by CarolinaBiological



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11.5: Calculating Solution Concentration- Mass Percent

Learning Objectives

- Express the amount of solute in a solution in various concentration units.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms with meanings that depend on various factors.

Introduction

Concentration is the measure of how much of a given substance is mixed with another substance. Solutions are said to be either dilute or concentrated. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution.

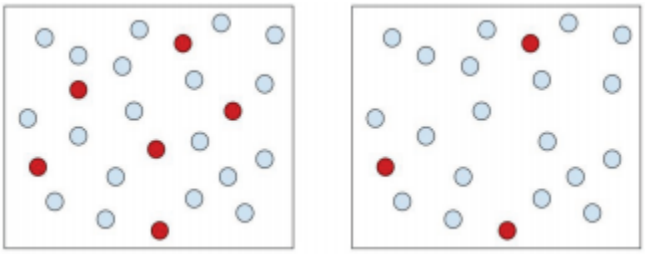


Figure $11.5.1$: The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. To make juice, you have to mix the frozen juice concentrate from inside these containers with three or four times the container size full of water. Therefore, you are diluting the concentrated juice. In terms of solute and solvent, the concentrated solution has a lot of solute versus the dilute solution that would have a smaller amount of solute.

The terms "concentrated" and "dilute" provide qualitative methods of describing concentration. Although qualitative observations are necessary and have their place in every part of science, including chemistry, we have seen throughout our study of science that there is a definite need for quantitative measurements in science. This is particularly true in solution chemistry. In this section, we will explore some quantitative methods of expressing solution concentration.

Mass Percent

There are several ways of expressing the concentration of a solution by using a percentage. The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \mathrm{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving $25.0\ \mathrm{g}$ of sugar into $100.0\ \mathrm{g}$ of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\text{ g sugar}}{125.0\text{ g solution}} \times 100\% = 20.0\%\text{ sugar}$$

Example 11.5.1

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\%\text{ m/m} = \frac{36.5\text{ g}}{355\text{ g}} \times 100\% = 10.3\%$$

Exercise 11.5.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer available at the end of this section.

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g\text{ solute}}{100\text{ gsolution}}$ or $\frac{100\text{ gsolution}}{g\text{ solute}}$

For example, if you need to make $3000.0\ \text{g}$ of a 5.00% solution of sodium chloride, the mass of solute needs to be determined.

Solution

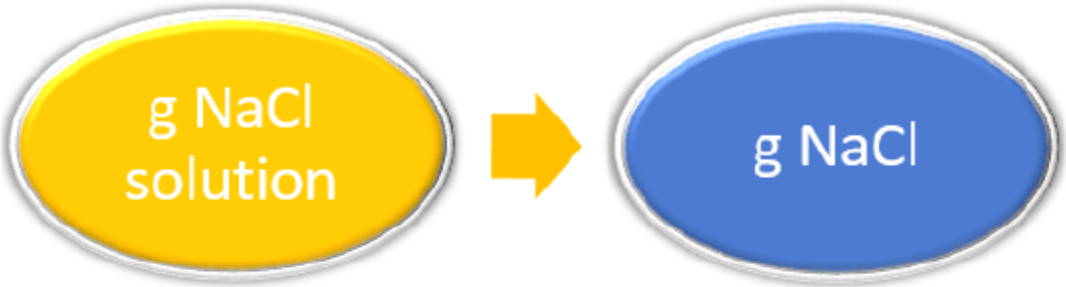
Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:



$$\frac{5.00\ g\ NaCl}{100\ g\ NaCl\ solution}$$

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0\ \cancel{g\ NaCl\ solution} \times \frac{5.00\ g\ NaCl}{100\ \cancel{g\ NaCl\ solution}} = 150.0g\ NaCl$$

You would need to weigh out $150\ \text{g}$ of NaCl and add it to $2850\ \text{g}$ of water. Notice that it was necessary to subtract the mass of the $\text{NaCl}\ (150\ \text{g})$ from the mass of solution $(3000\ \text{g})$ to calculate the mass of the water that would need to be added.

Exercise 11.5.1

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg, 3.00 % (by mass) H_2O_2 solution?

Answer

180 g H_2O_2

Contributors and Attributions

Answers

Exercise 11.5.1

7.90 %

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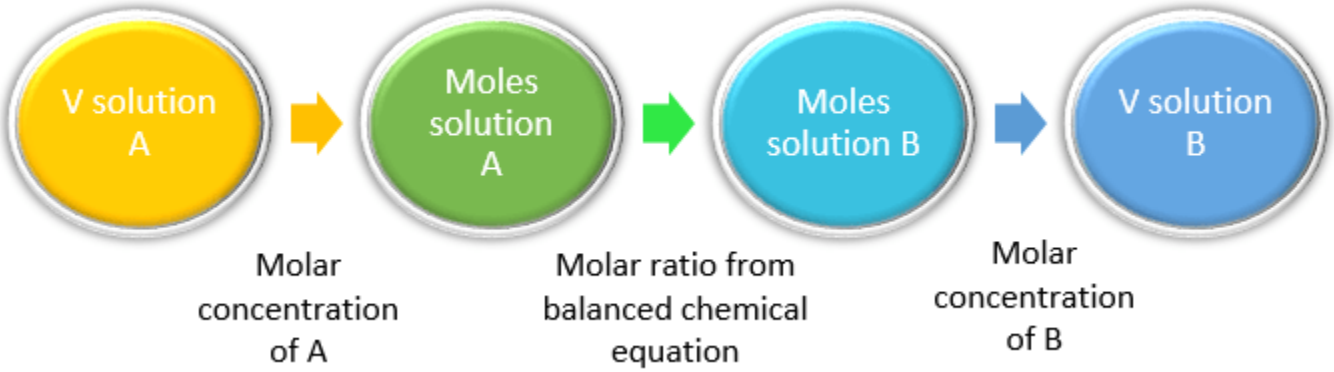
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11.6: Solution Stoichiometry

Learning Objectives

- Determine amounts of reactants or products in aqueous solutions.

As we learned previously, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are "switched" (they *replace* each other). Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and therefore their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead (II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead (II) chloride.

$$\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow PbCl_2\,(s) + 2\,NaNO_3\,(aq)} \tag{11.6.1}$$

In the reaction shown above, if we mixed 0.123 L of a 1.00 M solution of NaCl with 1.50 M solution of $\mathrm{Pb(NO_3)_2}$, we could calculate the volume of $\mathrm{Pb(NO_3)_2}$ solution needed to completely precipitate the $\mathrm{Pb^{2+}}$ ions.

The molar concentration can also be expressed as the following:

$$1.00\,M\,\mathrm{NaCl} = \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}}$$

and

$$1.50\,M\,\mathrm{Pb(NO_3)_2} = \frac{1.50\ mol\ \mathrm{Pb(NO_3)_2}}{1\ L\ \mathrm{Pb(NO_3)_2 solution}}$$

First, we must examine the reaction stoichiometry in the balanced reaction (Equation 11.6.1). In this reaction, one mole of $\mathrm{Pb(NO_3)_2}$ reacts with two moles of NaCl to give one mole of $\mathrm{PbCl_2}$ precipitate. Thus, the concept map utilizing the stoichiometric ratios is:



so the volume of lead (II) nitrate that reacted is calculated as:

$$0.123\ L\ \text{NaCl solution} \times \frac{1.00\ mol\ \text{NaCl}}{1\ L\ \text{NaCl solution}} \times \frac{1\ mol\ \text{Pb(NO}_3)_2}{2\ mol\ \text{NaCl}} \times \frac{1\ L\ \text{Pb(NO}_3)_2\ \text{solution}}{1.5\ mol\ \text{Pb(NO}_3)_2} = 0.041\ \text{Pb(NO}_3)_2\ L\ \text{solution}$$

This volume makes intuitive sense for two reasons: (1) the number of moles of $\text{Pb(NO}_3)_2$ required is half of the number of moles of NaCl, based off of the stoichiometry in the balanced reaction (Equation 11.6.1); (2) the concentration of $\text{Pb(NO}_3)_2$ solution is 50% greater than the NaCl solution, so less volume is needed.

Example 11.6.1

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate all Ba^{2+} in the solution?

Solution

Solutions to Example 13.8.1

Steps for Problem Solving	Example 11.6.1
Identify the "given" information and what the problem is asking you to "find."	Given: 275 mL $BaCl_2$ 0.250 M BaCl_2 or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M Na_2SO_4 or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume Na_2SO_4 solution.
Set up and balance the chemical equation.	$\text{Na}_2\text{SO}_4\,(\text{aq}) + \text{BaCl}_2\,(\text{aq}) \longrightarrow \text{BaSO}_4\,(\text{s}) + 2\text{NaCl}\,(\text{aq})$ An insoluble product is formed after the reaction.
List other known quantities.	1 mol of Na_2SO_4 to 1 mol $BaCl_2$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	mL $BaCl_2$ solution → L $BaCl_2$ solution → mol $BaCl_2$ solution → mol Na_2SO_4 solution $\frac{1\ L}{1000\ mL}$ $\frac{0.250\ mol\ BaCl_2}{1\ L\ BaCl_2 solution}$ $\frac{1\ mol\ Na_2SO_4}{1\ mol\ BaCl_2}$ $\frac{1\ L\ Na_2SO_4\ solution}{0.500\ mol}$

Steps for Problem Solving	Example 11.6.1
Cancel units and calculate.	$275\ \cancel{mL\ BaCl_2\ solution} \times \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \times \frac{0.250\ \cancel{mol\ BaCl_2}}{1\ \cancel{L\ BaCl_2\ solution}} \times \frac{1\ \cancel{mol\ Na_2SO}}{1\ \cancel{mol\ BaCl}$ $\times \frac{1\ L\ Na_2SO_4\ solution}{0.500\ \cancel{mol Na_2SO_4}}$ = 0.1375 L sodium sulfate
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than barium chloride. Also, the units are correct.

Exercise 11.6.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

Answer available at the end of this section.

Answers

Exercise 11.6.1

0.250 L LiOH solution

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12: CHEMICAL EQUILIBRIUM

12.1: The Idea of Dynamic Chemical Equilibrium
12.2: The Equilibrium Constant, Keq
12.3: Heterogeneous Equilibria
12.4: Calculations with Equilibrium Constants
12.5: Le Chatelier's Principle

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12.1: The Idea of Dynamic Chemical Equilibrium

Chapter 1

1. Chapter 1: The Chemical World
2. 1.1: The Scope of Chemistry
3. 1.2: Chemicals Compose Ordinary Things
4. 1.3: Hypothesis, Theories, and Laws
5. 1.4: The Scientific Method: How Chemists Think
6. 1.5: A Beginning Chemist: How to Succeed
7. 1.E: Exercises

- Chapter 2

1. Chapter 2: Measurement and Problem Solving
2. 2.1: Taking Measurements
3. 2.2: Scientific Notation: Writing Large and Small Numbers
4. 2.3: Significant Figures: Writing Numbers to Reflect Precision
5. 2.4: Significant Figures in Calculations
6. 2.5: The Basic Units of Measurement
7. 2.6: Problem Solving and Unit Conversions
8. 2.7: Solving Multistep Conversion Problems
9. 2.8: Units Raised to a Power
10. 2.9: Density
11. 2.10: Numerical Problem-Solving Strategies and the Solution Map
12. 2.E: Measurement and Problem Solving (Exercises)

- Chapter 3

1. Chapter 3: Matter and Energy
2. 3.1: In Your Room
3. 3.2: What is Matter?
4. 3.3: Classifying Matter According to Its State: Solid, Liquid, and Gas
5. 3.4: Classifying Matter According to Its Composition
6. 3.5: Differences in Matter: Physical and Chemical Properties
7. 3.6: Changes in Matter: Physical and Chemical Changes
8. 3.7: Conservation of Mass: There is No New Matter
9. 3.8: Energy
10. 3.9: Energy and Chemical and Physical Change
11. 3.10: Temperature: Random Motion of Molecules and Atoms
12. 3.11: Temperature Changes: Heat Capacity
13. 3.12: Energy and Heat Capacity Calculations
14. 3.E: Exercises

- Chapter 4

1. Chapter 4: Atoms and Elements
2. 4.1: Experiencing Atoms at Tiburon
3. 4.2: Indivisible: The Atomic Theory
4. 4.3: The Nuclear Atom
5. 4.4: The Properties of Protons, Neutrons, and Electrons
6. 4.5: Elements: Defined by Their Number of Protons
7. 4.6: Looking for Patterns: The Periodic Table
8. 4.7: Ions: Losing and Gaining Electrons
9. 4.8: Isotopes: When the Number of Neutrons Varies
10. 4.9: Atomic Mass: The Average Mass of an Element's Atoms

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- Chapter 5

1. Chapter 5: Molecules and Compounds
2. 5.1: Sugar and Salt
3. 5.2: Compounds Display Constant Composition
4. 5.3: Chemical Formulas: How to Represent Compounds
5. 5.4: A Molecular View of Elements and Compounds
6. 5.5: Writing Formulas for Ionic Compounds
7. 5.6: Nomenclature: Naming Compounds
8. 5.7: Naming Ionic Compounds
9. 5.8: Naming Molecular Compounds
10. 5.9: Naming Acids
11. 5.10: Nomenclature Summary
12. 5.11: Formula Mass: The Mass of a Molecule or Formula Unit

- Chapter 6

1. 6: Chemical Composition
2. 6.1: How Much Sodium?
3. 6.2: Counting Nails by the Pound
4. 6.3: Counting Atoms by the Gram
5. 6.4: Counting Molecules by the Gram
6. 6.5: Chemical Formulas as Conversion Factors
7. 6.6: Mass Percent Composition of Compounds
8. 6.7: Mass Percent Composition from a Chemical Formula
9. 6.8: Calculating Empirical Formulas for Compounds
10. 6.9: Calculating Molecular Formulas for Compounds

- Chapter 7

1. Chapter 7: Chemical Reactions
2. 7.1: Grade School Volcanoes, Automobiles, and Laundry Detergents
3. 7.2: Evidence of a Chemical Reaction
4. 7.3: Chemical Equations
5. 7.4: How to Write Balanced Chemical Equations
6. 7.5: Aqueous Solutions and Solubility: Compounds Dissolved in Water
7. 7.6: Precipitation Reactions
8. 7.7: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations
9. 7.8: Acid–Base and Gas Evolution Reactions
10. 7.9: Oxidation–Reduction Reactions
11. 7.10: Classifying Chemical Reactions
12. 7.11: The Activity Series: Predicting Spontaneous Redox Reactions

- Chapter 8

1. Chapter 8: Quantities in Chemical Reactions
2. 8.1: Climate Change: Too Much Carbon Dioxide
3. 8.2: Stoichiometry
4. 8.3: Mole-to-Mole Conversions
5. 8.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions
6. 8.5: Limiting Reactant and Theoretical Yield
7. 8.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants
8. 8.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

- Chapter 9

1. Chapter 9: Electrons in Atoms and the Periodic Table
2. 9.1: Blimps, Balloons, and Models of the Atom
3. 9.2: Light is Visible Electromagnetic Radiation

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4. 9.3: The Electromagnetic Spectrum
5. 9.4: The Bohr Model: Atoms with Orbits
6. 9.5: The Quantum-Mechanical Model: Atoms with Orbitals
7. 9.6: Quantum-Mechanical Orbitals and Electron Configurations
8. 9.7: Electron Configurations and the Periodic Table
9. 9.8: The Explanatory Power of the Quantum-Mechanical Model
10. 9.9: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character
11. 9.E: Electrons in Atoms and the Periodic Table (Exercises)

Chapter 10

1. Chapter 10: Chemical Bonding
2. 10.1: Bonding Models and AIDS Drugs
3. 10.2: Representing Valence Electrons with Dots
4. 10.3: Lewis Structures of Ionic Compounds: Electrons Transferred
5. 10.4: Covalent Lewis Structures: Electrons Shared
6. 10.5: Writing Lewis Structures for Covalent Compounds
7. 10.6: Resonance: Equivalent Lewis Structures for the Same Molecule
8. 10.7: Predicting the Shapes of Molecules
9. 10.8: Electronegativity and Polarity: Why Oil and Water Don't Mix

- Chapter 11

1. Chapter 11: Gases
2. 11.1: Extra-Long Straws
3. 11.2: Kinetic Molecular Theory: A Model for Gases
4. 11.3: Pressure: The Result of Constant Molecular Collisions
5. 11.4: Boyle's Law: Pressure and Volume
6. 11.5: Charles's Law: Volume and Temperature
7. 11.5A: Gay-Lussac's Law: Temperature and Pressure
8. 11.6: The Combined Gas Law: Pressure, Volume, and Temperature
9. 11.7: Avogadro's Law: Volume and Moles
10. 11.8: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles
11. 11.9: Mixtures of Gases: Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen
12. 11.10: Gases in Chemical Reactions

- Chapter 12

1. Chapter 12: Liquids, Solids, and Intermolecular Forces
2. 12.1: Interactions between Molecules
3. 12.2: Properties of Liquids and Solids
4. 12.3: Surface Tension and Viscosity
5. 12.4: Evaporation and Condensation
6. 12.5: Melting, Freezing, and Sublimation
7. 12.6: Types of Intermolecular Forces: Dispersion, Dipole–Dipole, Hydrogen Bonding, and Ion-Dipole
8. 12.7: Types of Crystalline Solids
9. 12.8: Water: A Remarkable Molecule

- Chapter 13

1. Chapter 13: Solutions
2. 13.1: Prelude - Tragedy in Cameroon
3. 13.2: Solutions: Homogeneous Mixtures
4. 13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy
5. 13.4: Solutions of Gases in Water
6. 13.5: Solution Concentration: Mass Percent
7. 13.6: Solution Concentration: Molarity
8. 13.7: Solution Dilution
9. 13.8: Solution Stoichiometry

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10. 13.9: Freezing Point Depression and Boiling Point Elevation
11. 13.10: Osmosis

- Chapter 14

1. Chapter 14: Acids and Bases
2. 14.1: Sour Patch Kids and International Spy Movies
3. 14.2: Acids: Properties and Examples
4. 14.3: Bases: Properties and Examples
5. 14.4: Molecular Definitions of Acids and Bases
6. 14.5: Reactions of Acids and Bases
7. 14.6: Acid–Base Titration
8. 14.7: Strong and Weak Acids and Bases
9. 14.8: Water: Acid and Base in One
10. 14.9: The pH and pOH Scales: Ways to Express Acidity and Basicity
11. 14.10: Buffers: Solutions That Resist pH Change

- Chapter 15

1. Chapter 15: Chemical Equilibrium
2. 15.1: Life is Controlled Disequilibrium
3. 15.2: The Rate of a Chemical Reaction
4. 15.3: The Idea of Dynamic Chemical Equilibrium
5. 15.4: The Equilibrium Constant: A Measure of How Far a Reaction Goes
6. 15.5: Heterogeneous Equilibria: The Equilibrium Expression for Reactions Involving a Solid or a Liquid
7. 15.6: Calculating and Using Equilibrium Constants
8. 15.7: Disturbing a Reaction at Equilibrium: Le Châtelier's Principle
9. 15.8: The Effect of a Concentration Change on Equilibrium
10. 15.9: The Effect of a Volume Change on Equilibrium
11. 15.10: The Effect of Temperature Changes on Equilibrium
12. 15.11: The Solubility-Product Constant
13. 15.12: The Path of a Reaction and the Effect of a Catalyst

- Chapter 16

1. Chapter 16: Oxidation and Reduction
2. 16.1: The End of the Internal Combustion Engine?
3. 16.2: Oxidation and Reduction: Some Definitions
4. 16.3: Oxidation States: Electron Bookkeeping
5. 16.4: Balancing Redox Equations
6. 16.5: The Activity Series: Predicting Spontaneous Redox Reactions
7. 16.6: Batteries: Using Chemistry to Generate Electricity
8. 16.7: Electrolysis: Using Electricity to Do Chemistry
9. 16.8: Corrosion: Undesirable Redox Reactions

- Chapter 17

1. Chapter 17: Radioactivity and Nuclear Chemistry
2. 17.1: Diagnosing Appendicitis
3. 17.2: The Discovery of Radioactivity
4. 17.3: Types of Radioactivity: Alpha, Beta, and Gamma Decay
5. 17.4: Detecting Radioactivity
6. 17.5: Natural Radioactivity and Half-Life
7. 17.6: Radiocarbon Dating: Using Radioactivity to Measure the Age of Fossils and Other Artifacts
8. 17.7: The Discovery of Fission and the Atomic Bomb
9. 17.8: Nuclear Power: Using Fission to Generate Electricity
10. 17.9: Nuclear Fusion: The Power of the Sun
11. 17.10: The Effects of Radiation on Life
12. 17.11: Radioactivity in Medicine

Learning Objectives

- Describe the three possibilities that exist when reactants come together.
- Describe what is occurring in a system at equilibrium.

Think for a minute about sitting down to a table to eat dinner. There are three possibilities that could happen when you eat dinner. You could (1) finish your entire dinner, (2) you could not want any of it and leave it all on your plate, or (3) you could eat some of it and leave some of it. Reactions have the same possibilities. Reactions do not always proceed all the way from start to finish. You may have reactions that (1) go to completion so that at the end, the reaction vessel contains all products and only products. Some reactions (2) may not start at all, so at the end the reaction vessel contains all reactants and only reactants. And some reactions (3) may start but not go to completion, that is, the reaction might start but not go completely to products. In this last case, the reaction vessel would contain some reactants and some products. In this section, we are going to take a closer look at the third type of reaction.

Reversible Reactions and Equilibrium

Consider the hypothetical reaction:

$$\mathrm{A} + \mathrm{B} \rightarrow \mathrm{C} + \mathrm{D}.$$

If we looked at this reaction using what we have learned, this reaction will keep going, forming C and D until A and B run out. This is what we call an "irreversible reaction" or a "reaction that goes to completion".

Some reactions, however, are **reversible**, meaning the reaction can go backwards in which products react to form reactants, so that: $\mathrm{A} + \mathrm{B} \leftarrow \mathrm{C} + \mathrm{D}$. The direction of the arrow shows that C and D are reacting to form A and B. What if the two reactions, the forward reaction and the reverse reaction, were occurring at the same time? What would this look like? If you could peer into the reaction, you would be able to find A, B, C, and D particles. A and B would react to form C and D at the same time that C and D are reacting to form A and B.

If the forward and reverse reactions are happening at the same rate, the reaction is said to be at **equilibrium** or **dynamic equilibrium**. At this point, the concentrations of A, B, C, and D are not changing (or, are constant) and we would see no difference in our reaction container, but reactions are still occurring in both directions. It is important to point out that having constant amounts of reactants and products does NOT mean that the concentration of the reactants is equal to the concentration of the products. It means they are not changing. These reactions appear to have stopped before one of the reactants has run out.

Chemists use a double-headed arrow, $\rightleftharpoons$, to show that a reaction is at equilibrium. We would write the example reaction as:

$$\mathrm{A} + \mathrm{B} \rightleftharpoons \mathrm{C} + \mathrm{D}.$$

The arrow indicates that both directions of the reaction are happening.

Another way to think about reversible and irreversible reactions is to compare them to two types of games of tag. Reversible reactions are in many ways like a traditional game of tag: the "it" person can become "not it" and somebody who is "not it" is tagged and becomes "it". In this way, it is a reversible change. It is also like a reaction at equilibrium, because overall no change is occurring. There is always a constant number of "it" people and "not it" people in the game. Also, having constant numbers of "it" and "not it" people in our game does not mean that the number of "it" people (reactants) is equal to the number of "not it" people (products). Furthermore, this is similar to equilibrium in that this game never truly ends (unless everybody gets tired of playing). The game could go on forever. We could write this as the following reversible reaction:

$$\text{"It"} \rightleftharpoons \text{"Not it"}$$

Irreversible reactions (those that only go in one direction from reactants to products and cannot reach a state of equilibrium) are more like a game of sharks and minnows. In sharks and minnows, almost everybody starts out as a minnow. Once tagged, they become a shark. However, the difference here is that once you are a shark you are always a shark; there is no way to go back to becoming a minnow. The game continues until everybody has been tagged and becomes a shark. This is similar to irreversible

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reactions in that the reactants turn into products, but can't change back. Furthermore, the reaction will proceed until the reactants have been used up and there are not any more left. We could write the reaction as:

$$\text{Minnow} \rightarrow \text{Shark}$$

Here's another example of a reversible reaction—dissolving salt in a beaker of water, described by the following reaction:

$$\text{NaCl}\,(\text{s}) \leftrightarrow \text{NaCl}\,(\text{aq})$$

If you keep adding more and more solid salt, eventually you'll reach the point where no more salt dissolves, and the excess sits at the bottom of the beaker. At this point we have a saturated solution. Has the dissolving reaction stopped? It would appear so, but that's not the case (wouldn't that be too easy?).

What happens in our saturated solution, which has reached the point of **equilibrium**, is that both the forward

$$\text{NaCl}\,(\text{s}) \rightarrow \text{NaCl}\,(\text{aq})$$

and reverse

$$\text{NaCl}\,(\text{aq}) \rightarrow \text{NaCl}\,(\text{s})$$

reactions are still going on, but at the same ***rate***. This in effect cancels out any observable, or measurable, changes in our system. At the same rate that solid NaCl produces aqueous NaCl (dissolved salt), the dissolved salt is re-crystallizing to form more solid NaCl.

- **Equilibrium** is the state at which the rate of the forward reaction equals the rate of the reverse reaction.
- At the point of equilibrium, no more measurable or observable changes in the system can be noted.

It is important for you to understand that equilibrium means the ***rates*** of the forward and reverse reactions are equal; it does ***not*** mean that there are equal amounts of reactants and products present at equilibrium.

For example, the following reaction was allowed to come to the point of equilibrium, and concentrations of all reaction participants were measured at that time:

$$\text{H}_2\,(\text{g}) + \text{I}_2\,(\text{g}) \leftrightarrow 2\,\text{HI}\,(\text{g})$$

At equilibrium:

$$[\text{H}_2] = 0.022\ \text{M}$$

$$[\text{I}_2] = 0.022\ \text{M}$$

$$[\text{HI}] = 0.156\ \text{M}$$

For this particular reversible reaction, there is more HI at equilibrium (0.156 M) than there is of H_2 and I_2 (both at 0.022M). We say that the product side of the reaction is favored.

> Equilibrium does not mean equal amounts at equilibrium!

Here is an example to help you understand how equilibrium works: imagine yourself on a escalator that is going down. You start at the top (reactants) and end up at the bottom (products). But when you are partway down, you start walking up the escalator as it continues going down. If you match your rate of walking up to the same rate that the escalator is going down, you make no progress and appear to be at a standstill. To an observer it would look as if you and the escalator had come to a stop, when actually both upward and downward movements continue.

Equilibrium is **dynamic**—both forward and reverse reactions continue, even though the reaction appears to have stopped. And this equilibrium does not need to occur right in the middle of two floors—you could be near the bottom, near the top, or anywhere in between when you carry out your reverse process.

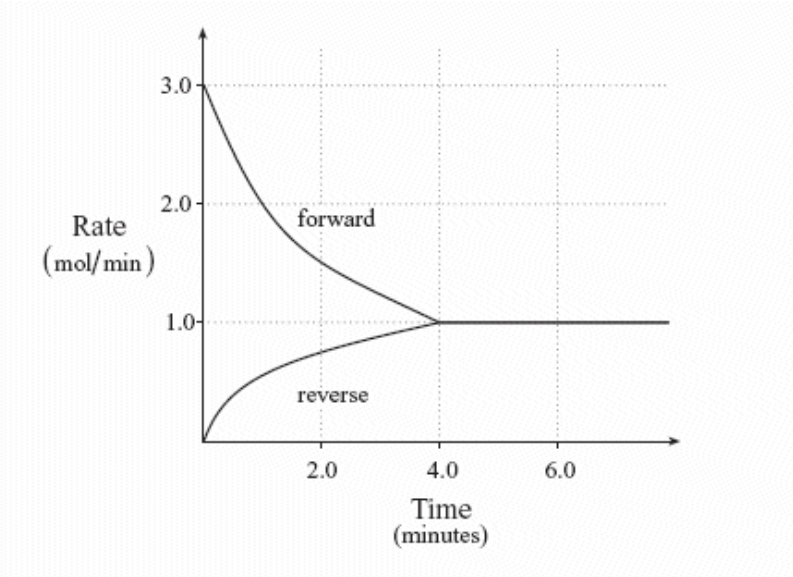
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In order for a reversible reaction to reach the point of equilibrium, the reaction must be carried out in a closed system—no additional reactants can be added or products removed. If, in our last example, the product HI was removed as it formed, the reaction would never reach the point of equilibrium; instead, H_2 and I_2 would continue to react to produce HI until one or both of the reactants was used up.

If reactants are constantly being added, and products removed as they form, the system would appear to be at equilibrium because to an outside observer it would appear that the reaction has stopped—but that would not be the case. This situation—with new material constantly being added as products are removed—is called a **steady state system**. A factory with an assembly line is a steady state system—new raw materials are constantly being added; finished products are removed. A campfire with wood being added to the fire is another steady state system. Be careful not to confuse steady state with equilibrium.

How do the rates of the forward and reverse reactions change as the reaction heads towards equilibrium (before it reaches equilibrium)?

If we start our above reaction with H_2 and I_2, and with no HI, the two gases will react at a certain rate. But remember that the rate of a reaction slows down over time, as the reactants get used up (and lower their concentrations). Eventually, however, the amount of the product HI increases, and it will begin producing H_2 and I_2. Thus the rate of the reverse reaction starts out slowly (there is no HI present), but will speed up as the concentration of HI increases. Eventually both rates will level off (not always to the same level as shown by this example, however):



Chemists have found that there is a mathematical relationship that exists between the concentration of the reactants and products, once equilibrium has been reached, that is independent of the initial concentration of the participants. For any general reaction,

$$a\,\mathrm{A} + b\,\mathrm{B} \rightleftharpoons c\,\mathrm{D} + d\,\mathrm{D} \tag{12.1.1}$$

an equilibrium constant expression can be written as:

$$\mathrm{K_{eq}} = \frac{[\mathrm{C}]^{\mathrm{c}} \times [\mathrm{D}]^{\mathrm{d}}}{[\mathrm{A}]^{\mathrm{a}} \times [\mathrm{B}]^{\mathrm{b}}} \tag{12.1.2}$$

This mathematical relationship exists for all equilibrium systems, and produces a constant ratio called the ***equilibrium constant, K_{eq}*** .

Law of Mass-Action

Equation 12.1.2 is sometimes called the Law of Mass-Action.

This relationship will be very important to us for the next few units, so it is important that you understand how to set this relationship up and what it tells us about an equilibrium system.

The products of the reaction (C and D) are placed in the numerator, and their concentrations are raised to the power of the coefficients from the balanced equation. The reactants (A and B) are placed in the denominator, with their concentrations raised to the power of *their* coefficients.

Example 12.1.1

For the reaction between hydrogen and iodine gas to produce hydrogen iodide:

$$\mathrm{H_2\,(g) + I_2\,(g) \;\rightleftharpoons\; 2\,HI\,(g)} \tag{12.1.3}$$

the equilibrium constant expression will be:

$$\mathrm{K_{eq} = \frac{[HI]^2}{[H_2] \times [I_2]}} \tag{12.1.4}$$

Using the example we examined in our last section, equilibrium concentrations for each substance were measured at equilibrium and found to be:

At equilibrium:
$[\mathrm{H_2}] = 0.022\ \mathrm{M}$
$[\mathrm{I_2}] = 0.022\ \mathrm{M}$
$[\mathrm{HI}] = 0.156\ \mathrm{M}$

We substitute these values into our equilibrium expression and solve for K_{eq}:

$$\mathrm{K_{eq} = \frac{[HI]^2}{[H_2] \times [I_2]}} = \frac{(0.156)^2}{(0.022)(0.022)} = 50.3$$

The value of K_{eq}, which has no units, is a constant for any particular reaction, and its value does not change unless the temperature of the system is changed. It does not depend on the initial concentrations used to reach the point of equilibrium.

For example, the following data were obtained for equilibrium concentrations of H_2, I_2 and HI, and the value of K_{eq} was calculated for each trial:

Trial
$[\mathrm{HI}]$
$[\mathrm{H_2}]$
$[\mathrm{I_2}]$
K_{eq}
1
0.156
0.0220
0.0220
50.3
2
0.750
0.106
0.106
50.1
3

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1.00
0.820
0.0242
50.4
4
1.00
0.0242
0.820
50.4
5
1.56
0.220
0.220
50.3

Aside from accounting for slight experimental variation between trials, the value for K_{eq} is the same despite differences in equilibrium concentrations for the individual participants.

There is one other important point to make at this time.

K_{eq} relates the concentrations of products to reactants at equilibrium.

For aqueous solutions, concentration is often measured as mol · L^{-1}. For gases, concentration is often measured as partial pressure.

The concentrations of both aqueous solutions and gases change during the progress of a reaction. For reactions involving a solid or a liquid, while the ***amounts*** of the solid or liquid will change during a reaction, their ***concentrations*** (much like their densities) will ***not*** change during the reaction. Instead, their values will remain constant. Because they are constant, their values **are not included in the equilibrium constant expression**.

Example 12.1.2

For example, consider the reaction showing the formation of solid calcium carbonate from solid calcium oxide and carbon dioxide gas:

$$\mathrm{CaO\,(s) + CO_2\,(g) \;\leftrightarrow CaCO_3\,(s)}$$

The equilibrium constant for this reaction is (before modification):

$$\mathrm{K_{eq} = \frac{[CaCO_3]}{[CaO] \times [CO_2]}}$$

But we remove those participants whose state is either a solid or a liquid, which leaves us with the following equilibrium constant expression:

$$\mathrm{K_{eq} = \frac{1}{[CO_2]}}$$

Summary

- There are a few possible ways a reaction can go: it can go to completion ($\mathrm{reactants \rightarrow products}$), and it can occur but not go to completion. Instead, it will reach chemical equilibrium ($\mathrm{reactants \rightleftharpoons products}$).
- Chemical equilibrium occurs when the number of particles becoming products is equal to the number of particles becoming reactants.
- A dynamic equilibrium is a state where the rate of the forward reaction is equal to the rate of the reverse reaction.

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Vocabulary

- **Equilibrium** - A state that occurs when the rate of the forward reaction is equal to the rate of the reverse reaction.

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12.2: The Equilibrium Constant, Keq

Learning Objectives

- Write equilibrium constant expressions.
- Use equilibrium constant expressions to solve for unknown concentrations.
- Use known concentrations to solve for the equilibrium constants.
- Explain what the value of K means in terms of relative concentrations of reactants and products.

In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

The Equilibrium Constant

Equilibrium reactions are those that do not go to completion, but are in a state where the reactants are reacting to yield products and the products are reacting to produce reactants. In a reaction at equilibrium, the equilibrium concentrations of all reactants and products can be measured. The **equilibrium constant (K)** is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants. Sometimes, subscripts are added to the equilibrium constant symbol K, such as K_{eq}, K_{c}, K_{p}, K_{a}, K_{b}, and K_{sp}. These are all equilibrium constants and are subscripted to indicate special types of equilibrium reactions.

There are some rules about writing equilibrium constant expressions that need to be learned:

1. Concentrations of products are multiplied on the top of the expression. Concentrations of reactants are multiplied together on the bottom.
2. Coefficients in the equation become exponents in the equilibrium constant expression.
3. Solids, liquids, and solvents are assigned a value of 1, so their concentrations do not affect the value of K.

Example 12.2.1

Write the equilibrium constant expression for:

$$\text{CO}\,(g) + 3\text{H}_2\,(g) \rightleftharpoons \text{CH}_4\,(g) + \text{H}_2\text{O}\,(g)$$

Solution

$$K = \frac{[\text{CH}_4]\,[\text{H}_2\text{O}]}{[\text{CO}]\,[\text{H}_2]^3}$$

*Note that the coefficients become exponents. Also, note that the concentrations of products in the numerator are *multiplied*. The same is true of the reactants in the denominator.

Example 12.2.2

Write the equilibrium constant expression for:

$$2\text{TiCl}_3\,(s) + 2\text{HCl}\,(g) \rightleftharpoons 2\text{TiCl}_4\,(s) + \text{H}_2\,(g)$$

Solution

$$K = \frac{[\mathrm{H_2}]}{[\mathrm{HCl}]^2}$$

*Note that the solids have a value of 1, and multiplying or dividing by 1 does not change the value of K.

Example 12.2.2

Write the equilibrium constant expression for:

$$\mathrm{P_4}\,(s) + 6\mathrm{Cl_2}\,(g) \rightleftharpoons 4\mathrm{PCl_3}\,(s)$$

Solution

$$K = \frac{1}{[\mathrm{Cl_2}]^6}$$

*Note that the only product is a solid, which is defined to have a value of 1. That leaves just 1 on top in the numerator.

Example 12.2.3

Write the equilibrium constant expression for:

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

Solution

$$K = [\mathrm{H^+}]\,[\mathrm{OH^-}]$$

*Note that the water is the solvent, and thus has a value of 1. Dividing by 1 does not change the value of K.

Equilibrium Constant Expressions

The equilibrium constant value is the ratio of the concentrations of the products over the reactants. This means that we can use the value of K to predict whether there are more products or reactants at equilibrium for a given reaction. What can the value of K_{eq} tell us about a reaction?

- **If K_{eq} is very large**, the concentration of the products is much greater than the concentration of the reactants. The reaction essentially "goes to completion"; all, or most of, the reactants are used up to form the products.
- **If K_{eq} is very small**, the concentration of the reactants is much greater than the concentration of the products. The reaction does not occur to any great extent—most of the reactants remain unchanged, and there are few products produced.
- **When K_{eq} is not very large or very small** (close to a value of 1) then roughly equal amounts of reactants and products are present at equilibrium.

Here are some examples to consider:

Reaction	Chemical Equations	Equilibrium Constant
the decomposition of ozone, $\mathrm{O_3}$	$2\,\mathrm{O_3}\,(\mathrm{g}) \rightleftharpoons 3\,\mathrm{O_2}\,(\mathrm{g})$	$\mathrm{K_{eq}} = 2.0 \times 10^{57}$
	K_{eq} is very large, indicating that mostly $\mathrm{O_2}$ is present in an equilibrium system, with very little $\mathrm{O_3}$.	
production of nitrogen monoxide	$\mathrm{N_2}\,(\mathrm{g}) + \mathrm{O_2}\,(\mathrm{g}) \rightleftharpoons 2\,\mathrm{NO}\,(\mathrm{g})$	$\mathrm{K_{eq}} = 1.0 \times 10^{-25}$
	Very little NO is produced by this reaction; $\mathrm{N_2}$ and $\mathrm{O_2}$ do not react readily to produce NO (this is lucky for us—otherwise we would have little oxygen to breath in our atmosphere!).	

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Reaction	Chemical Equations	Equilibrium Constant
reaction of carbon monoxide and water	$\mathrm{CO\,(g) + H_2O\,(g) \rightleftharpoons CO_2\,(g) + H_2\,(g)}$	$\mathrm{K_{eq} = 5.09\ (at\ 700\ K)}$
	The concentrations of the reactants are very close to the concentrations of the products at equilibrium.	

If the equilibrium constant is 1 or nearly 1, it indicates that the molarities of the reactants and products are about the same. If the equilibrium constant value is a large number, like 100, or a very large number, like 1×10^{15}, it indicates that the products (numerator) are a great deal larger than the reactants. This means that at equilibrium, the great majority of the material is in the form of products and it is said that the "products are strongly favored". If the equilibrium constant is small, like 0.10, or very small, like 1×10^{-12}, it indicates that the reactants are much larger than the products and the reactants are strongly favored. With large K values, most of the material at equilibrium is in the form of products and with small K values, most of the material at equilibrium is in the form of the reactants.

The equilibrium constant expression is an equation that we can use to solve for K or for the concentration of a reactant or product.

Example 12.2.4

Determine the value of K for the reaction

$$\mathrm{SO_2\,(g) + NO_2\,(g) \rightleftharpoons SO_3\,(g) + NO\,(g)}$$

when the equilibrium concentrations are: $\mathrm{[SO_2] = 1.20\ M}$, $\mathrm{[NO_2] = 0.60\ M}$, $\mathrm{[NO] = 1.6\ M}$, and $\mathrm{[SO_3] = 2.2\ M}$.

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{\mathrm{[SO_3]\,[NO]}}{\mathrm{[SO_2]\,[NO_2]}}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)\,(1.6)}{(1.20)\,(0.60)} = 4.9$$

Example 12.2.5

Consider the following reaction:

$$\mathrm{CO\,(g) + H_2O\,(g) \rightleftharpoons H_2\,(g) + CO_2\,(g)}$$

with $K = 1.34$. If the $\mathrm{[H_2O] = 0.100\ M}$, $\mathrm{[H_2] = 0.100\ M}$, and $\mathrm{[CO_2] = 0.100\ M}$ at equilibrium, what is the equilibrium concentration of CO?

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{\mathrm{[H_2]\,[CO_2]}}{\mathrm{[CO]\,[H_2O]}}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)\,(0.100)}{\mathrm{[CO]}\,(0.100)}$$

Solving for $[\mathrm{CO}]$, we get: $[\mathrm{CO}] = 0.0746\ \mathrm{M}$

Summary

- The equilibrium constant expression is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants.
- If the value of K is greater than 1, the products in the reaction are favored. If the value of K is less than 1, the reactants in the reaction are favored. If K is equal to 1, neither reactants nor products are favored.

Vocabulary

- **Equilibrium constant** (K) - A mathematical ratio that shows the concentrations of the products divided by the concentrations of the reactants.

Contributions & Attributions

-

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12.3: Heterogeneous Equilibria

Learning Objectives

- To understand how different phases affect equilibria.

When the products and reactants of an equilibrium reaction form a single phase, whether gas or liquid, the system is a homogeneous equilibrium. In such situations, the concentrations of the reactants and products can vary over a wide range. In contrast, a system whose reactants, products, or both are in more than one phase is a heterogeneous equilibrium, such as the reaction of a gas with a solid or liquid.

As noted in the previous section, the equilibrium constant expression is actually a ratio of activities. To simplify the calculations in general chemistry courses, the activity of each substance in the reaction is often approximated using a ratio of the molarity of a substance compared to the standard state of that substance. For substances that are liquids or solids, the standard state is just the concentration of the substance within the liquid or solid. Because the molar concentrations of pure liquids and solids normally do not vary greatly with temperature, the ratio of the molarity to the standard state for substances that are liquids or solids always has a value of 1. For example, for a compound such as CaF_2(s), the term going into the equilibrium expression is [CaF_2]/[CaF_2] which cancels to unity. Thus, when the activities of the solids and liquids (including solvents) are incorporated into the equilibrium expression, they do not change the value.

Consider the following reaction, which is used in the final firing of some types of pottery to produce brilliant metallic glazes:

$$\mathrm{CO_2\,(g) + C\,(s)} \rightleftharpoons \mathrm{2\,CO\,(g)} \tag{12.3.1}$$

The glaze is created when metal oxides are reduced to metals by the product, carbon monoxide. The equilibrium constant expression for this reaction is as follows:

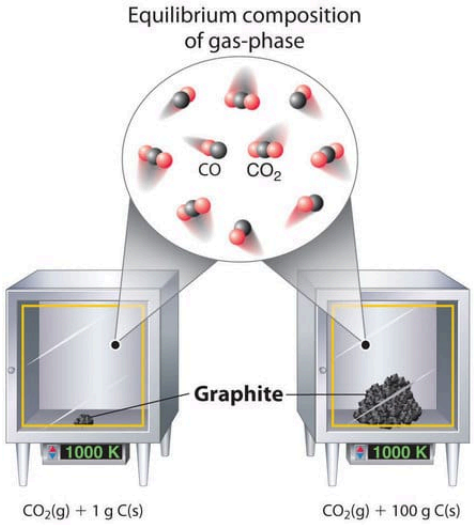
$$K = \frac{a_{\mathrm{CO}}^2}{a_{\mathrm{CO_2}} a_C} = \frac{[\mathrm{CO}]^2}{[\mathrm{CO_2}][1]} = \frac{[\mathrm{CO}]^2}{[\mathrm{CO_2}]} \tag{12.3.2}$$

The equilibrium constant for this reaction can also be written in terms of the partial pressures of the gases:

$$K_p = \frac{(P_{CO})^2}{P_{CO_2}} \tag{12.3.3}$$

Incorporating all the constant values into K' or K_p allows us to focus on the substances whose concentrations change during the reaction.

Although the activities of pure liquids or solids are not written explicitly in the equilibrium constant expression, these substances must be present in the reaction mixture for chemical equilibrium to occur. Whatever the concentrations of CO and $\mathrm{CO_2}$, the system described in Equation 12.3.1 will reach chemical equilibrium only if a stoichiometric amount of solid carbon or excess solid carbon has been added so that some is still present once the system has reached equilibrium. As shown in Figure 12.3.1, it does not matter whether 1 g or 100 g of solid carbon is present; in either case, the composition of the gaseous components of the system will be the same at equilibrium.



$$CO_2(g) + C(s) \rightleftharpoons 2CO(g)$$

Figure 12.3.2: Effect of the Amount of Solid Present on Equilibrium in a Heterogeneous Solid–Gas System. In the system, the equilibrium composition of the gas phase at a given temperature, 1000 K in this case, is the same whether a small amount of solid carbon (left) or a large amount (right) is present.

Example 12.3.1

Write each expression for K, incorporating all constants, and K_p for the following equilibrium reactions.

a. $\mathrm{PCl_3\,(l) + Cl_2\,(g) \rightleftharpoons PCl_5\,(s)}$
b. $\mathrm{Fe_3O_4\,(s) + 4\,H_2\,(g) \rightleftharpoons 3\,Fe\,(s) + 4\,H_2O\,(g)}$

Given: balanced equilibrium equations.

Asked for: expressions for K and K_p.

Strategy:

Find K by writing each equilibrium constant expression as the ratio of the concentrations of the products and reactants, each raised to its coefficient in the chemical equation. Then express K_p as the ratio of the partial pressures of the products and reactants, each also raised to its coefficient in the chemical equation.

Solution

This reaction contains a pure solid (PCl_5) and a pure liquid (PCl_3). Their activities are equal to 1, so when incorporated into the equilibrium constant expression, they do not change the value. So

$$K = \frac{1}{(1)[Cl_2]}$$

and

$$K_p = \frac{1}{(1)P_{Cl_2}}$$

This reaction contains two pure solids (Fe_3O_4 and Fe), which are each assigned a value of 1 in the equilibrium constant expressions:

$$K = \frac{(1)[H_2O]^4}{(1)[H_2]^4}$$

and

$$K_p = \frac{(1)(P_{H_2O})^4}{(1)(P_{H_2})^4}$$

Exercise 12.3.1

Write the expressions for K and K_p for the following reactions.

a. $\mathrm{CaCO_3\,(s) \rightleftharpoons CaO\,(s) + CO_2\,(g)}$
b. $\underset{glucose}{\mathrm{C_6H_{12}O_6\,(s)}} + \mathrm{6\,O_2\,(g) \rightleftharpoons 6\,CO_2\,(g) + 6\,H_2O\,(g)}$

Answer a available at the end of this section.

For reactions carried out in solution, the solvent is assumed to be pure, and therefore is assigned an activity equal to 1 in the equilibrium constant expression. The activities of the solutes are approximated by their molarities. The result is that the equilibrium constant expressions appear to only depend upon the concentrations of the solutes.

> *The activities of pure solids, pure liquids, and solvents are defined as having a value of '1'. Often, it is said that these activities are "left out" of equilibrium constant expressions. This is an unfortunate use of words. The activities are not "left out" of equilibrium constant expressions. Rather, because they have a value of '1', they do not change the value of the equilibrium constant when they are multiplied together with the other terms.* The activities of the solutes are approximated by their molarities.

Summary

An equilibrated system that contains products and reactants in a single phase is a homogeneous equilibrium; a system whose reactants, products, or both are in more than one phase is a heterogeneous equilibrium.

Contributors and Attributions

- Anonymous
- Modified by Tom Neils (Grand Rapids Community College)

Answers

Exercise 12.3.1

$K = [\mathrm{CO_2}]$ and $K_p = P_{\mathrm{CO_2}}$

$K = \frac{[CO_2]^6[H_2O]^6}{[O_2]^6}$ and $K_p = \frac{(P_{CO_2})^6(P_{H_2O})^6}{(P_{O_2})^6}$

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12.4: Calculations with Equilibrium Constants

Iron is an important component of red blood cells. Patients who have low iron will usually be anemic and have a lower than normal number of red blood cells. One way to assess serum iron concentration is with the use of Ferrozine, a complex organic molecule. Ferrozine forms a product with Fe^{3+}, producing a pink color. In order to determine factors affecting the reaction, we need to measure the equilibrium constant. If the equilibrium does not lie far in the direction of products, precautions need to be taken when using this material to measure iron in serum.

Calculations with Equilibrium Constants

The general value of the equilibrium constant gives us information about whether the reactants or the products are favored at equilibrium. Since the product concentrations are in the numerator of the equilibrium expression, a $K_{\mathrm{eq}} > 1$ means that the products are favored over the reactants. A $K_{\mathrm{eq}} < 1$ means that the reactants are favored over the products.

Though it would often seem that the K_{eq} value would have various units depending on the values of the exponents in the expression, the general rule is that any units are dropped. All K_{eq} values will be reported as having no units.

Example 12.4.1

Equilibrium occurs when nitrogen monoxide gas reacts with oxygen gas to form nitrogen dioxide gas:

$$2\mathrm{NO}\,(g) + \mathrm{O_2}\,(g) \rightleftharpoons 2\mathrm{NO_2}\,(g)$$

At equilibrium at $230^{\circ}\mathrm{C}$, the concentrations are measured to be $[\mathrm{NO}] = 0.0542\ \mathrm{M}$, $[\mathrm{O_2}] = 0.127\ \mathrm{M}$, and $[\mathrm{NO_2}] = 15.5\ \mathrm{M}$. Calculate the equilibrium constant at this temperature.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $[\mathrm{NO}] = 0.0542\ \mathrm{M}$
- $[\mathrm{O_2}] = 0.127\ \mathrm{M}$
- $[\mathrm{NO_2}] = 15.5\ \mathrm{M}$

Unknown

The equilibrium expression is first written according to the general form in the text. The equilibrium values are substituted into the expression and the value calculated.

Step 2: Solve.

$$K_{\mathrm{eq}} = \frac{[\mathrm{NO_2}]^2}{[\mathrm{NO}]^2\,[\mathrm{O_2}]}$$

Substituting in the concentrations at equilibrium:

$$K_{\mathrm{eq}} = \frac{(15.5)^2}{(0.0542)^2\,(0.127)} = 6.44 \times 10^5$$

Step 3: Think about your result.

The equilibrium concentration of the product $\mathrm{NO_2}$ is significantly higher than the concentrations of the reactants NO and $\mathrm{O_2}$. As a result, the K_{eq} value is much larger than 1, an indication that the product is favored at equilibrium.

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The equilibrium expression only shows those substances whose concentrations are variable during the reaction. A pure solid or a pure liquid does not have a concentration that will vary during a reaction. Therefore, an equilibrium expression omits pure solids and liquids, and only shows the concentrations of gases and aqueous solutions. The decomposition of mercury (II) oxide can be shown by the following equation, followed by its equilibrium expression.

$$2\text{HgO}\,(s) \rightleftharpoons 2\text{Hg}\,(l) + \text{O}_2\,(g) \qquad K_{\text{eq}} = [\text{O}_2]$$

The stoichiometry of an equation can also be used in a calculation of an equilibrium constant. At 40°C, solid ammonium carbamate decomposes to ammonia and carbon dioxide gases.

$$\text{NH}_4\text{CO}_2\text{NH}_2\,(s) \rightleftharpoons 2\text{NH}_3\,(g) + \text{CO}_2\,(g)$$

At equilibrium, the $[\text{CO}_2]$ is found to be $4.71 \times 10^{-3}\ \text{M}$. Can the K_{eq} value be calculated from that information alone? Because the ammonium carbamate is a solid, it is not present in the equilibrium expression.

$$K_{\text{eq}} = [\text{NH}_3]^2\,[\text{CO}_2]$$

The stoichiometry of the chemical equation indicates that as the ammonium carbamate decomposes, $2\ \text{mol}$ of ammonia gas is produced for every $1\ \text{mol}$ of carbon dioxide. Therefore, at equilibrium, the concentration of the ammonia will be twice the concentration of carbon dioxide. So,

$$[\text{NH}_3] = 2 \times \left(4.71 \times 10^{-3}\right) = 9.42 \times 10^{-3}\ \text{M}$$

Substituting these values into the K_{eq} expression:

$$K_{\text{eq}} = \left(9.42 \times 10^{-3}\right)^2 \left(4.71 \times 10^{-3}\right) = 4.18 \times 10^{-7}$$

Using Equilibrium Constants

The equilibrium constants are known for a great many reactions. Hydrogen and bromine gases combine to form hydrogen bromide gas. At 730°C, the equation and K_{eq} are given below.

$$\text{H}_2\,(g) + \text{Br}_2\,(g) \rightleftharpoons 2\text{HBr}\,(g) \qquad K_{\text{eq}} = 2.18 \times 10^6$$

A certain reaction is begun with only HBr. When the reaction mixture reaches equilibrium at 730°C, the concentration of bromine gas is measured to be $0.00243\ \text{M}$. What is the concentration of the H_2 and the HBr at equilibrium?

Since the reaction begins with only HBr and the mole ratio of H_2 to Br_2 is 1:1, the concentration of H_2 at equilibrium is also $0.00243\ \text{M}$. The equilibrium expression can be rearranged to solve for the concentration of HBr at equilibrium:

$$\begin{aligned} K_{\text{eq}} &= \frac{[\text{HBr}]^2}{[\text{H}_2]\,[\text{Br}_2]} \\ [\text{HBr}] &= \sqrt{K_{\text{eq}}\,[\text{H}_2]\,[\text{Br}_2]} \\ &= \sqrt{2.18 \times 10^6\,(0.00243)\,(0.00243)} = 3.59\ \text{M} \end{aligned}$$

Since the value of the equilibrium constant is very high, the concentration of HBr is much greater than that of H_2 and Br_2 at equilibrium.

Summary

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12.5: Le Chatelier's Principle

Learning Outcomes

- Define Le Chatelier's principle.
- Predict how the change in amounts of substances, temperature, or pressure will affect amounts of reactants and products present at equilibrium.

Le Chatelier's Principle

Chemical equilibrium was studied by the French chemist Henri Le Chatelier (1850 - 1936) and his description of how a system responds to a stress to equilibrium has become known as **Le Chatelier's principle**: When a chemical system that is at equilibrium is disturbed by a stress, the system will respond in order to relieve the stress. Stresses to a chemical system involve changes in the concentrations of reactants or products, changes in the temperature of the system, or changes in the pressure of the system. We will discuss each of these stresses separately. The change to the equilibrium position in every case is either a favoring of the forward reaction or a favoring of the reverse reaction. When the forward reaction is favored, the concentrations of products increase, while the concentrations of reactants decrease. When the reverse reaction is favored, the concentrations of the products decrease, while the concentrations of reactants increase.

Original Equilibrium	Favored Reaction	Result
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Forward: $\mathrm{A} \rightarrow \mathrm{B}$	$[\mathrm{A}]$ decreases; $[\mathrm{B}]$ increases
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Reverse: $\mathrm{A} \leftarrow \mathrm{B}$	$[\mathrm{A}]$ increases; $[\mathrm{B}]$ decreases



Figure 9.6.1: *Henri Le Chatelier.*

Effect of Concentration

A change in concentration of one of the substances in an equilibrium system typically involves either the addition or the removal of one of the reactants or products. Consider the Haber-Bosch process for the industrial production of ammonia from nitrogen and hydrogen gases.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g)$$

If the concentration of one substance in a system is increased, the system will respond by favoring the reaction that removes that substance. When more $\mathrm{N_2}$ is added, the forward reaction will be favored because the forward reaction uses up $\mathrm{N_2}$ and converts it to $\mathrm{NH_3}$. The forward reaction speeds up temporarily as a result of the addition of a reactant. The position of equilibrium shifts as more $\mathrm{NH_3}$ is produced. The concentration of $\mathrm{NH_3}$ increases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ decrease. After some time passes, equilibrium is reestablished with new concentrations of all three substance. As can be seen in the figure below, if more $\mathrm{N_2}$ is added, a new equilibrium is achieved by the system. The new concentration of $\mathrm{NH_3}$ is higher because of the favoring of the forward reaction. The new concentration of the $\mathrm{H_2}$ is lower .The concentration of $\mathrm{N_2}$ is higher than in the original equilibrium, but

went down slightly following the addition of the N_2 that disturbed the original equilibrium. By responding in this way, the value of the equilibrium constant for the reaction, K_{eq}, does not change as a result of the stress to the system.

In other words, the amount of each substance is different but the ratio of the amount of each remains the same.

If more NH_3 were added, the reverse reaction would be favored. This "favoring" of a reaction means temporarily speeding up the reaction in that direction until equilibrium is reestablished. Recall that once equilibrium is reestablished, the rates of the forward and reverse reactions are again equal. The addition of NH_3 would result in increased formation of the reactants, N_2 and H_2.

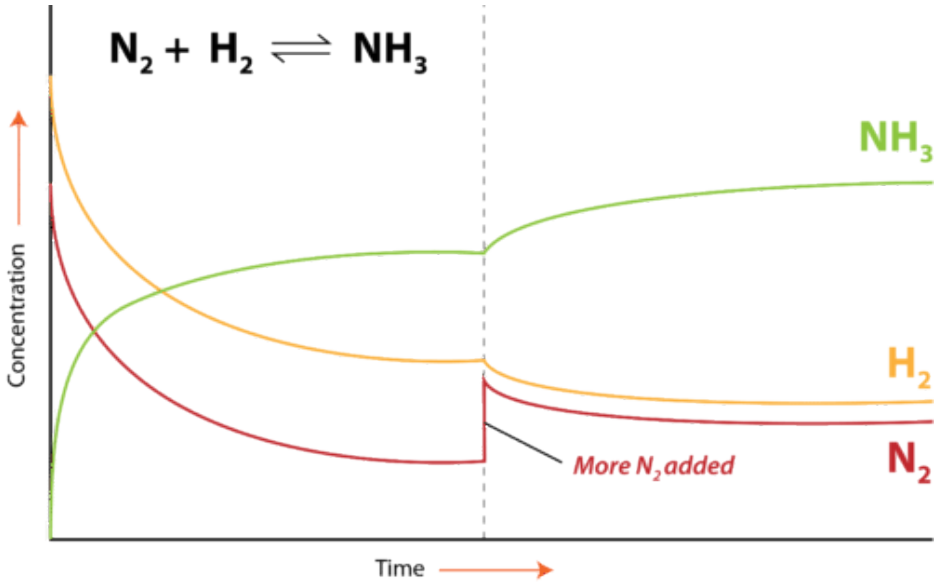


Figure 9.6.2: *The Haber-Bosch process is an equilibrium between reactants N_2 and H_2 and product NH_3.*

An equilibrium can also be disrupted by the removal of one of the substances. If the concentration of a substance is decreased, the system will respond by favoring the reaction that replaces that substance. In the industrial Haber-Bosch process, NH_3 is removed from the equilibrium system as the reaction proceeds. As a result, the forward reaction is favored so that more NH_3 is produced. The concentrations of N_2 and H_2 decrease. Continued removal of NH_3 will eventually force the reaction to go to completion until all of the reactants are used up. If either N_2 or H_2 were removed from the equilibrium system, the reverse reaction would be favored and the concentration of NH_3 would decrease.

The effect of changes in concentration on an equilibrium system according to Le Chatelier's principle is summarized in the table below.

Table 9.6.1

Stress	Response
addition of reactant	forward reaction favored
addition of product	reverse reaction favored
removal of reactant	reverse reaction favored
removal of product	forward reaction favored

Example 12.5.1

Given this reaction at equilibrium:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

How will it affect the reaction if the equilibrium is stressed by each change?

1. H2 is added.
2. NH3 is added.
3. NH3 is removed.

Solution

1. If H2 is added, there is now more reactant, so the reaction will shift to the right (toward products) to reduce the added H2.
2. If NH3 is added, there is now more product, so the reaction will shift to the left (toward reactants) to reduce the added NH3.
3. If NH3 is removed, there is now less product, so the reaction will shift to the right (toward products) to replace the product removed.

12.5.1Exe

Given this reaction at equilibrium:

$$CO(g) + Br_2(g) \rightleftharpoons COBr_2(g)$$

How will it affect the reaction if the equilibrium is stressed by each change?

1. Br2 is removed.
2. COBr2 is added.

Answer available at the end of this section.

Effect of Temperature

Increasing or decreasing the temperature of a system at equilibrium is also a stress to the system. The equation for the Haber-Bosch process is written again below, as a thermochemical equation (i.e. it contains information about the energy gained or lost when the reaction occurs).

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g) + 91\ \mathrm{kJ}$$

The forward reaction is the exothermic direction: the formation of $\mathrm{NH_3}$ releases heat which is why that is shown as a product. The reverse reaction is the endothermic direction: as $\mathrm{NH_3}$ decomposes to $\mathrm{N_2}$ and $\mathrm{H_2}$, heat is absorbed. An increase in the temperature for this is like adding a product because heat is being released by the reaction. If we add a product then the reaction proceeds towards the formation of more reactants. Reducing the temperature for this system would be similar to removing a product which would favor the formation of more products. The amount of $\mathrm{NH_3}$ will increase and the amount of $\mathrm{N_2}$ and $\mathrm{H_2}$ will decrease.

For changes in concentration, the system responds in such a way that the value of the equilibrium constant, K_{eq} is unchanged. However, a change in temperature shifts the equilibrium and the K_{eq} value either increases or decreases. As discussed in the previous section, values of K_{eq} are dependent on the temperature. When the temperature of the system for the Haber-Bosch process is increased, the resultant shift in equilibrium towards the reactants means that the K_{eq} value decreases. When the temperature is decreased, the shift in equilibrium towards the products means that the K_{eq} value increases.

Le Chatelier's principle as related to temperature changes can be illustrated easily be the reaction in which dinitrogen tetroxide is in equilibrium with nitrogen dioxide.

$$\mathrm{N_2O_4}\,(g) + \mathrm{heat} \rightleftharpoons 2\mathrm{NO_2}\,(g)$$

Dinitrogen tetroxide $(\mathrm{N_2O_4})$ is colorless, while nitrogen dioxide $(\mathrm{NO_2})$ is dark brown in color. When $\mathrm{N_2O_4}$ breaks down into $\mathrm{NO_2}$, heat is absorbed (endothermic) according to the forward reaction above. Therefore, an increase in temperature (adding heat) of the system will favor the forward reaction. Conversely, a decrease in temperature (removing heat) will favor the reverse reaction.

Example 12.5.2

Predict the effect of increasing the temperature on this equilibrium.

$$PCl_3 + Cl_2 \rightleftharpoons PCl_5 + 60kJ$$

Solution

Because energy is listed as a product, it is being produced, so the reaction is exothermic. If the temperature is increasing, a product is being added to the equilibrium, so the equilibrium shifts to minimize the addition of extra product: it shifts to the left (back toward reactants).

12.5.2Exe

Predict the effect of decreasing the temperature on this equilibrium.

$$N_2O_4 + 57kJ \rightleftharpoons 2NO_2$$

Answer available at the end of this section.

Effect of Pressure

Changing the pressure of an equilibrium system in which gases are involved is also a stress to the system. A change in the pressure on a liquid or a solid has a negligible effect. We will return again the equilibrium for the Haber-Bosch process. Imagine the gases are contained in a closed system in which the volume of the system is controlled by an adjustable piston as shown in the figure below.

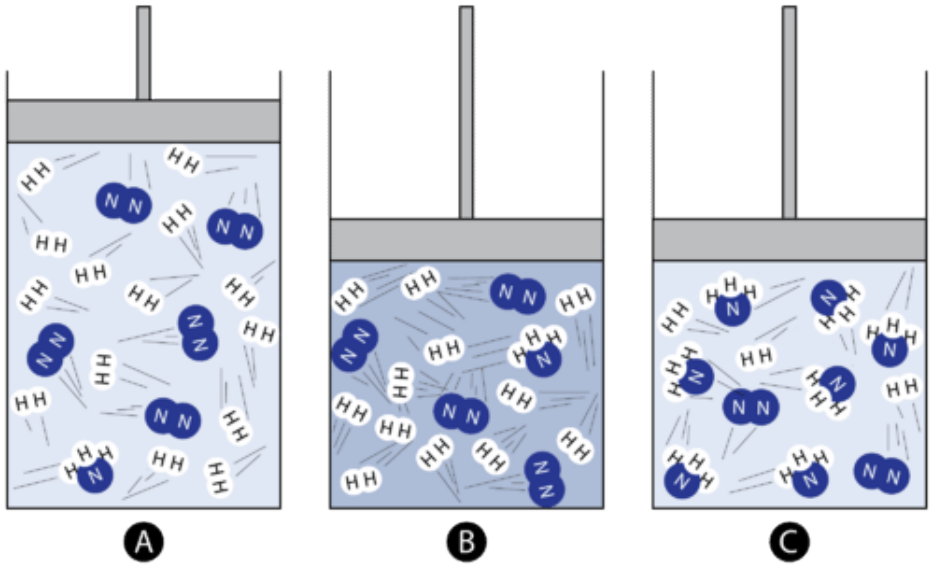


Figure 9.6.3: *Effect of pressure on equilibrium. (A) is at equilibrium, (B) increase of pressure through decreased volume, and (C) equilibrium is reestablished.*

On the far left, the reaction system contains primarily $\mathrm{N_2}$ and $\mathrm{H_2}$, with only one molecule of $\mathrm{NH_3}$ present. As the piston is pushed inwards, the pressure of the system increases according to Boyle's law. This is a stress to the equilibrium. In the middle image, the same number of molecules is now confined in a smaller space and so the pressure has increased. According to Le Chatelier's principle, the system responds in order to relieve the stress. In the image on the right, the forward reaction has been favored and more $\mathrm{NH_3}$ is produced. The overall result is a decrease in the number of gas molecules in the entire system. This in turn decreases the pressure and provides a relief to the original stress of a pressure increase. An increase in pressure on an equilibrium system

favors the reaction which products fewer total moles of gas. In this case, it is the forward reaction that is favored.

A decrease in pressure on the above system could be achieved by pulling the piston outward, increasing the container volume. The equilibrium would respond by favoring the reverse reaction in which $\mathrm{NH_3}$ decomposes to $\mathrm{N_2}$ and $\mathrm{H_2}$. This is because the overall number of gas molecules would increase and so would the pressure. A decrease in pressure on an equilibrium system favors the reaction which produces more total moles of gas. This is summarized in the table below.

Table 9.6.2

Stress	Response
pressure increase	reaction produces fewer gas molecules
pressure decrease	reaction produces more gas molecules

Like changes in concentration, the K_{eq} value for a given reaction is unchanged by a change in pressure. The amounts of each substance will change but the ratio will not. It is important to remember when analyzing the effect of a pressure change on equilibrium that only gases are affected. If a certain reaction involves liquids or solids, they should be ignored. For example, calcium carbonate decomposes according to the equilibrium reaction:

$$\mathrm{CaCO_3\,(s) \rightleftharpoons CaO\,(s) + O_2\,(g)}$$

Oxygen is the only gas in the system. An increase in the pressure of the system slows the rate of decomposition of $\mathrm{CaCO_3}$ because the reverse reaction is favored. When a system contains equal moles of gas on both sides of the equation, pressure has no effect on the equilibrium position, as in the formation of HCl from $\mathrm{H_2}$ and $\mathrm{Cl_2}$.

$$\mathrm{H_2\,(g) + Cl_2\,(g) \rightleftharpoons 2HCl\,(g)}$$

Example 12.5.3

What is the effect on this equilibrium if pressure is increased?

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Solution

According to Le Chatelier's principle, if pressure is increased, then the equilibrium shifts to the side with the fewer number of moles of gas. This particular reaction shows a total of 4 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts to the right (toward the products side).

12.5.3 Exe

What is the effect on this equilibrium if pressure is decreased?

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

Answer available at the end of this section.

Application of Le Chatelier's Principle

Oxygen transport by the blood

In aerobic respiration, oxygen is transported to the cells where it is combined with glucose and metabolized to carbon dioxide, which then moves back to the lungs from which it is expelled.

hemoglobin + O_2 $\rightleftharpoons$ oxyhemoglobin

The partial pressure of O_2 in the air is 0.2 atm, sufficient to allow these molecules to be taken up by hemoglobin (the red pigment of blood) in which it becomes loosely bound in a complex known as oxyhemoglobin. At the ends of the capillaries which deliver the blood to the tissues, the O_2 concentration is reduced by about 50% owing to its consumption by the cells. This shifts the equilibrium to the left, releasing the oxygen so it can diffuse into the cells.

Key Takeaways

- In a reaction at equilibrium, the introduction of more products will shift the mass balance towards more reactants, and the introduction of more reactants will lead to the formation of more products, but the *ratio* of *Products/Reactants* (*equilibrium constant*), *K* is unchanged.
- If temperature is changed, the numeric value *K* will change. If a reaction is *exothermic* (releases heat), an increase in the **temperature** will force the equilibrium to the left, causing the system to absorb heat and thus partially offsetting the rise in temperature. The opposite effect occurs for *endothermic* reactions, which are shifted to the right by rising temperature.
- The effect of **pressure** on an equilibrium is significant only for reactions which involve different numbers of moles of gases on the two sides of the equation. An increase in the total pressure will shift to the side with fewer moles of gas. A decrease in pressure will shift to the side with more moles of gas.

Contributors and Attributions

-

- Allison Soult, Ph.D. (Department of Chemistry, University of Kentucky)

Answers

1. shift to the left (toward reactants)
2. shift to the left (toward reactants)

Equilibrium shifts to the left (toward reactants).

Reaction shifts to the left (toward reactants).

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13: ELECTROLYTES, ACIDS AND BASES

13.1: Electrolytes
13.2: Acids- Properties and Examples
13.3: Bases- Properties and Examples
13.4: Strong and Weak Acids and Bases
13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory
13.6: Autodissociation of Water and Kw
13.7: The pH and pOH Scales
13.8: Acid-Base Reactions
13.9: Acid–Base Titrations
13.10: Buffers- Solutions that Resist pH Change

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13.1: 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 13.1.1).

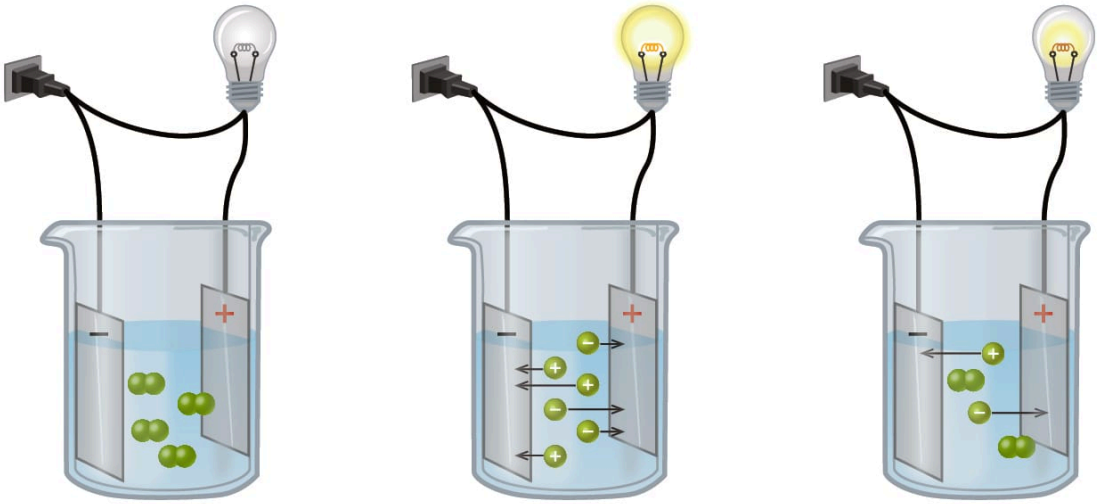


Figure 13.1.1: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 13.1.2. 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.

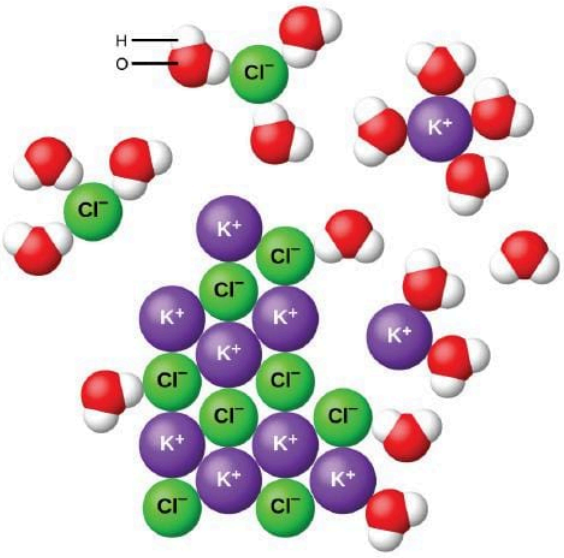


Figure 13.1.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

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 13.1.2 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.

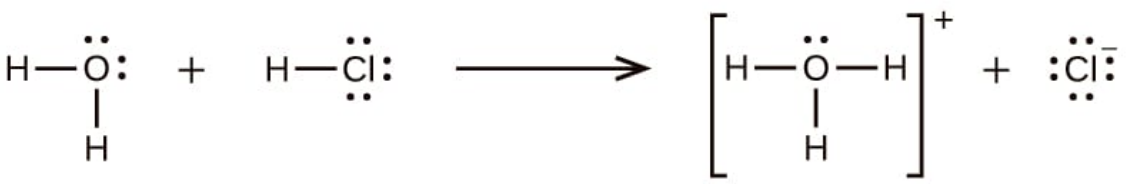
Covalent Electrolytes

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.

$$\mathrm{H_2O\,(l) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq)}$$

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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13.2: Acids- Properties and Examples

Learning Objectives

- Examine properties of acids.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{13.2.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

It should not be hard for you to name several common acids (but you might find that listing bases is a little more difficult). Below is a partial list of some common acids, along with some chemical formulas:

Table 13.2.1: Common Acids and Their Uses

Chemist Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) and stomach acid is HCl	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.

Chemist Name	Common Name	Uses
sulfuric acid, H_2SO_4		Used in car batteries, and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fizz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $C_6H_8O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$	aspirin	The active ingredient in aspirin.

What exactly makes an acid an acid, and what makes a base act as a base? Take a look at the formulas given in the above table and take a guess.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. A colorless, highly pungent solution of hydrogen chloride (HCl) in water. Hydrochloric acid is usually prepared by treating HCl with water.

$$\mathrm{HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)}$$

Hydrochloric acid can therefore be used to prepare chloride salts. Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases.

Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula $\mathrm{H_2SO_4}$. Sulfuric acid is a diprotic acid and has a wide range of applications including use in domestic acidic drain cleaners, as an electrolyte in lead-acid batteries, and in various cleaning agents. It is also a central substance in the chemical industry.



Figure 13.2.1: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating property, removing water (H_2O) from other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 13.2.2).

$$M + H_2SO_4 \longrightarrow M(SO_4) + H_2 \tag{13.2.2}$$

Nitric Acid

Nitric acid (HNO_3) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide ($NO_2\,(g)$) with water.

$$3\,NO_2\,(g) + H_2O\ (l) \longrightarrow 2\,HNO_3\,(ag) + NO\,(g)$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals (e.g., nitric acid with magnesium, manganese or zinc will liberate H_2 gas):

$$Mg + 2\,HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$$

$$Mn + 2\,HNO_3 \longrightarrow Mn(NO_3)_2 + H_2$$

$$Zn + 2\,HNO_3 \longrightarrow Zn(NO_3)_2 + H_2$$

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard it poses is chemical burn, as it carries out acid hydrolysis with proteins (amide) and fats (ester) which consequently decomposes living tissue (Figure 13.2.2). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin

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Figure 13.2.2: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula $\mathrm{H_2CO_3}$ and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of $\mathrm{H_2CO_3\,(aq)}$. Carbonic acid, which is a weak acid, forms two kinds of salts: the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve, producing calcium bicarbonate—which leads to many limestone features such as stalactites and stalagmites. Carbonic acid is a polyprotic acid, specifically it is diprotic, meaning that it has two protons which may dissociate from the parent molecule.

When carbon dioxide dissolves in water, it exists in chemical equilibrium (discussed in Chapter 15), producing carbonic acid:

$$\mathrm{CO_2 + H_2O \rightleftharpoons H_2CO_3}$$

The reaction can be pushed to favor the reactants to generate $\mathrm{CO_2\,(g)}$ from solution, which is key to the bubbles observed in carbonated beverages (Figure 13.2.3).



Figure 13.2.3: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

Formic Acid

Formic acid ($\mathrm{HCO_2H}$) is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally,

most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

Citric Acid

Citric acid ($C_6H_8O_7$) is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the TCA cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is used as a flavoring and preservative in food and beverages, especially soft drinks.

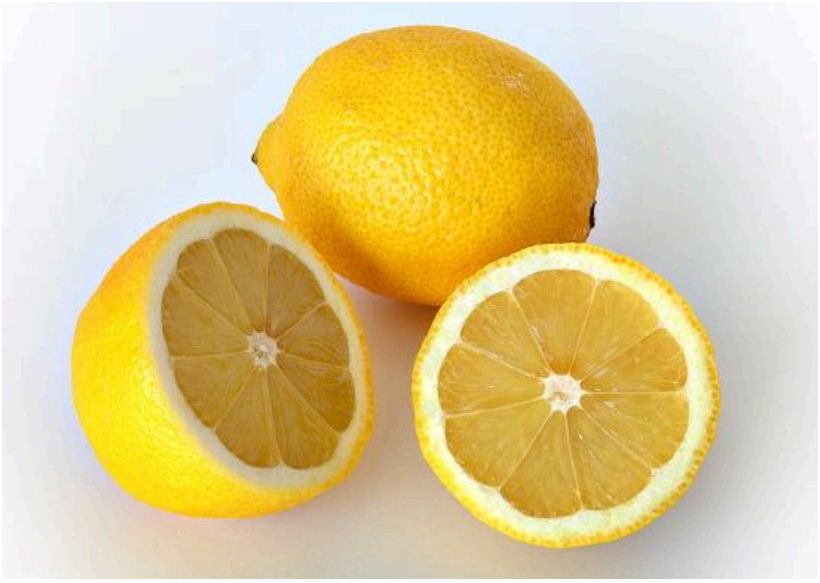


Figure 13.2.4: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years.

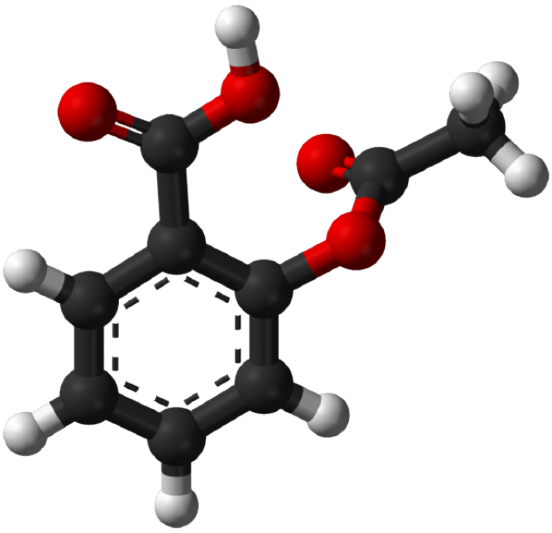


Figure 13.2.5: Ball-and-stick model of the aspirin molecule. (Public Domain; Ben Mills).

Aspirin is a white, crystalline, weakly acidic substance.

Summary

A brief summary of key aspects of several acids commonly encountered by students was given. Acids are a distinct class of compounds because of the properties of their aqueous solutions.

Contributions & Attributions

-

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

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13.3: Bases- Properties and Examples

Learning Objectives

- Examine properties of bases.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base, while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.



Figure 13.3.1: Phenolphthalein indicator in presence of base.

Warning!

Tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words—don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate. Various common bases and corresponding uses are given in Table 13.3.2 .

Table 13.3.1: Common Bases and Corresponding Uses

Some Common Bases	Uses
sodium hydroxide, NaOH (lye or caustic soda)	Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners.

Some Common Bases	Uses
potassium hydroxide, KOH (lye or caustic potash)	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$ (slaked lime)	Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil.
aluminum hydroxide	Used in water purification and as an ingredient in antacids.
ammonia, NH_3	Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula NaOH. It is a white solid ionic compound consisting of sodium cations $\mathrm{Na^+}$ and hydroxide anions $\mathrm{OH^-}$.

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction:

$$\mathrm{NaOH\,(s)\ \rightarrow Na^+\,(aq) + OH^-\,(aq)}$$

The resulting solution is usually colorless and odorless and feels slippery when it comes in contact with skin.



Figure 13.3.1: Sample of sodium hydroxide as pellets in a watch glass. (Public Domain; Walkerma.)

Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula KOH, and is commonly called caustic potash. Along with sodium hydroxide (NaOH), this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its corrosive nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathrm{KOH\,(s)\ \rightarrow K^+\,(aq) + OH^-\,(aq)}$$

Concentrated aqueous solutions are sometimes called *potassium lyes*.

Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $\mathrm{Mg(OH)_2}$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.



Figure 13.3.1: Bottle of Antacid tablets. (CC BY 2.,5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $\mathrm{Mg(OH)_2}$. However, a weak concentration of dissociated ions can be found in solution:

$$\mathrm{Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}\,(aq) + 2\,OH^-\,(aq)}$$

Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathrm{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$\mathrm{Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,OH^-\,(aq)}$$

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathrm{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

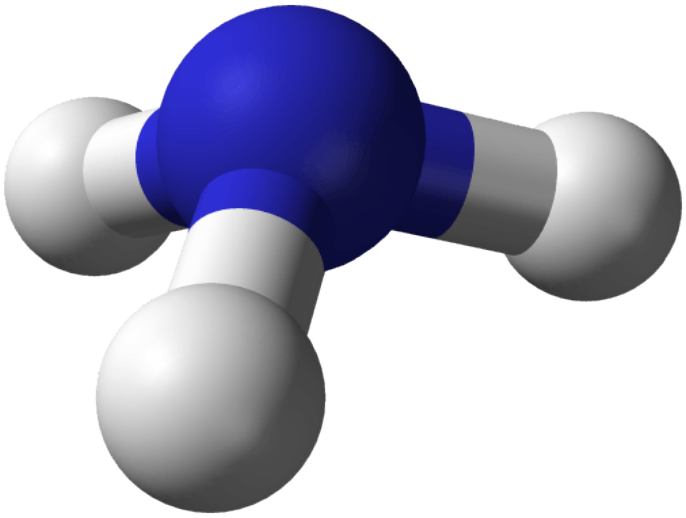


Figure 13.3.1: Ball-and-stick model of the ammonia molecule. (Public Domain; Ben Mills).

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from $\mathrm{H_2O}$ to yield ammonium and hydroxide ions:

$$\mathrm{NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4{}^+\,(aq) + OH^-\,(aq)}$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

- A brief summary of properties of bases was given.
- The properties of bases mostly contrast those of acids.
- Bases have many, varied uses.

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13.4: Strong and Weak Acids and Bases

Learning Objectives

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.
- Determine if a salt produces an acidic or a basic solution.

Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider HCl(aq). When HCl is dissolved in H_2O, it completely dissociates into H^+(aq) and Cl^-(aq) ions; all the HCl molecules become ions:

$$HCl \overset{100\%}{\rightarrow} H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2 \overset{\sim 5\%}{\longrightarrow} H^+(aq) + C_2H_3O_2^-(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as **a reversible reaction**:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

As it turns out, there are very few strong acids, which are given in Table 13.4.1. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

> Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.

Table 13.4.1: Strong Acids and Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
$HClO_3$	$Mg(OH)_2$
$HClO_4$	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in

solution, it is a **weak base**. There are very few strong bases (Table 13.4.1); any base not listed is a weak base. All strong bases are OH^- compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

Example 13.4.1: Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$
c. C_5H_5N

Solution

a. Because HCl is listed in Table 13.4.1, it is a strong acid.
b. Because $Mg(OH)_2$ is listed in Table 13.4.1, it is a strong base.
c. The nitrogen in C_5H_5N would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

Exercise 13.4.1

Identify each acid or base as strong or weak.

a. RbOH
b. $\mathrm{HNO_2}$

Answer a available at the end of this section.

Example 13.4.2: Characterizing Base Ionization

Write the balanced chemical equation for the dissociation of $Ca(OH)_2$ and indicate whether it proceeds 100% to products or not.

Solution

This is an ionic compound of Ca^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:

$$\mathrm{Ca(OH)_2 \longrightarrow Ca^{2+}(aq) + 2\,OH^-(aq)}$$

Because $Ca(OH)_2$ is listed in Table 13.4.1, this reaction proceeds 100% to products.

Exercise 13.4.2

Write the balanced chemical equation for the dissociation of hydrazoic acid (HN_3) and indicate whether it proceeds 100% to products or not.

Answer a available at the end of this section.

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Key Takeaways

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.
- Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions.

Contributions & Attributions

Answers

Exercise 13.4.1

strong base

weak acid

Exercise 13.4.2

The reaction is as follows:

$$\mathrm{HN_3} \longrightarrow \mathrm{H^+(aq) + N_3^-(aq)}$$

It does not proceed 100% to products because hydrazoic acid is not a strong acid.

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13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major classifications of substances known as acids or bases. The theory developed by Svante Arrhenius in 1883, the Arrhenius definition, states that an acid produces H^+ in solution and a base produces OH^-. Later, two more sophisticated and general theories were proposed. These theories are the Brønsted-Lowry and Lewis definitions of acids and bases. This section will cover the Arrhenius and Brønsted-Lowry theories; the Lewis theory is discussed elsewhere.

The Arrhenius Theory of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. An **Arrhenius acid** is a compound that **increases the concentration of H^+ ions** that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$\mathrm{HCl\,(aq)} \rightarrow \mathrm{H^+\,(aq) + Cl^-\,(aq)}$$

In this reaction, hydrochloric acid (HCl) dissociates completely into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$\mathrm{HCl\,(aq) + H_2O\,(l)} \rightarrow \mathrm{H_3O^+\,(aq) + Cl^-\,(aq)}$$

An **Arrhenius base** is a compound that **increases the concentration of OH^- ions** that are present when added to water. The dissociation is represented by the following equation:

$$\mathrm{NaOH\ \ (aq)} \rightarrow \mathrm{Na^+\ \ (aq) + OH^-\ \ (aq)}$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium (Na^+) and hydroxide (OH^-) ions when dissolved in water, thereby releasing OH^- ions into solution.

> Arrhenius acids are substances which produce hydrogen ions in solution and Arrhenius bases are substances which produce hydroxide ions in solution.

Limitations to the Arrhenius Theory

The Arrhenius theory has many more limitations than the other two theories. The theory does not explain the weak base ammonia (NH_3), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself. The Arrhenius definition of acid and base is also limited to aqueous (i.e., water) solutions.

The Brønsted-Lowry Theory of Acids and Bases

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, **a Brønsted-Lowry**

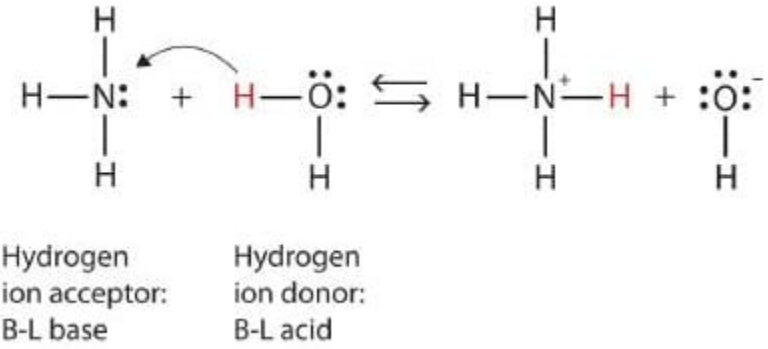
acid is a **proton donor (PD)**, while **a Brønsted-Lowry base** is a **proton acceptor (PA)**.

A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$\mathrm{NH_3\,(aq) + H_2O\ (\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)} \tag{13.5.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:

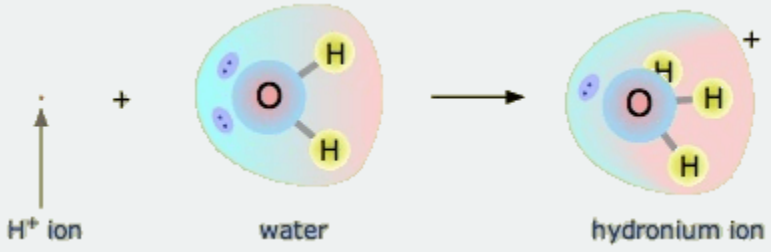


Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

The Hydronium Ion

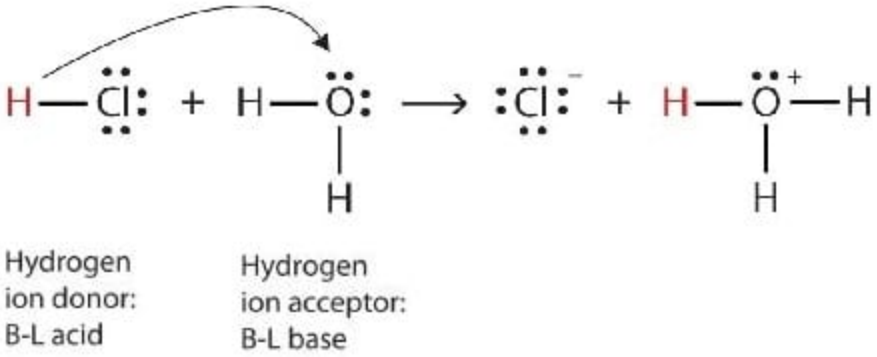
A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $\mathrm{H_5O_2^+}$ or $\mathrm{H_9O_4^+}$ rather than $\mathrm{H_3O^+}$. It is simpler, however, to use $\mathrm{H_3O^+}$ to represent the hydronium ion.



With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$\mathrm{HCl\,(g) + H_2O\ (\ell)\ \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{13.5.2}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense, but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.
- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.
- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Example 13.5.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule, just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$C_6H_5NH_2\,(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3{}^+\,(aq) + OH^-\,(aq)$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

Exercise 13.5.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$H_2PO_4{}^- + H_2O \rightleftharpoons HPO_4{}^{2-} + H_3O^+$$

Answer available at the end of this section.

Exercise 13.5.2

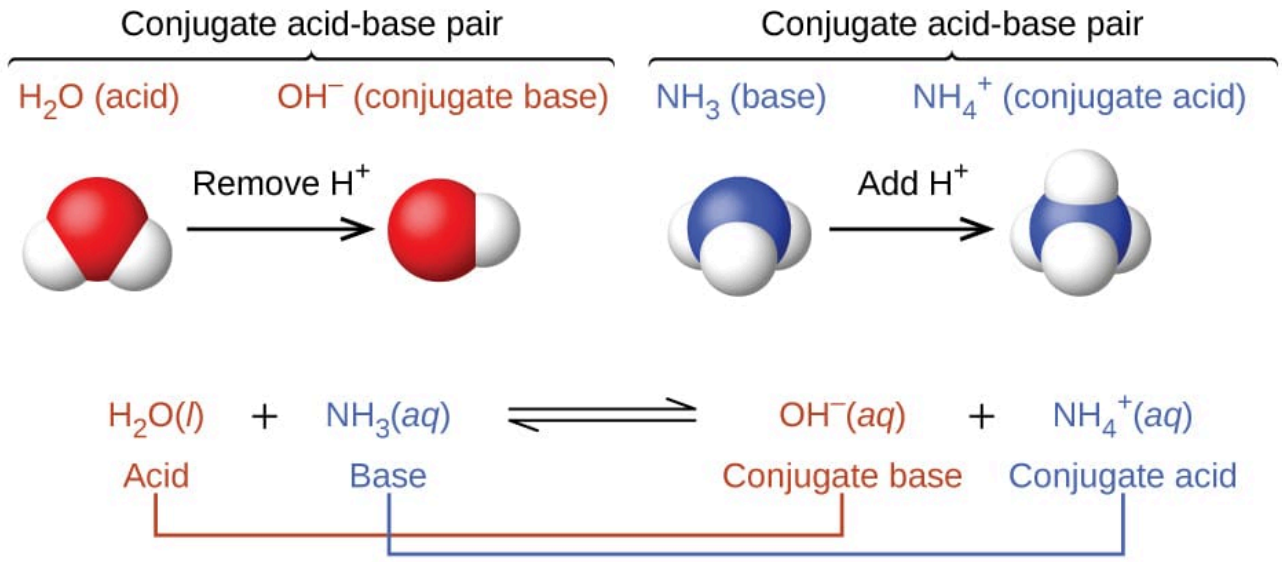
Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. HPO_4^{2-}
c. H_3PO_4
d. NH_4^+
e. $CH_3NH_3^+$

Answer available at the end of this section.

Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, OH^-, and the conjugate acid of ammonia, NH_4^+ :



This figure has three parts in two rows. In the first row, two diagrams of acid-base pairs are shown. On the left, a space filling model of H subscript 2 O is shown with a red O atom at the center and two smaller white H atoms attached in a bent shape. Above this model is the label “H subscript 2 O (acid)” in purple. An arrow points right, which is labeled “Remove H superscript plus.” To the right is another space filling model with a single red O atom to which a single smaller white H atom is attached. The label in purple above this model reads, “O H superscript negative (conjugate base).” Above both of these red and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” To the right is a space filling model with a central blue N atom to which three smaller white H atoms are attached in a triangular pyramid arrangement. A label in green above reads “N H subscript 3 (base).” An arrow labeled “Add H superscript plus” points right. To the right of the arrow is another space filling model with a blue central N atom and four smaller white H atoms in a tetrahedral arrangement. The green label above reads “N H subscript 3 superscript plus (conjugate acid).” Above both of these blue and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” The second row of the figure shows the chemical reaction, H subscript 2 O (l) is shown in purple, and is labeled below in purple as “acid,” plus N H subscript 3 (a q) in green, labeled below in green as “base,” followed by a double sided arrow arrow and O H superscript negative (a q) in purple, labeled in purple as “conjugate base,” plus N H subscript 4 superscript plus (a q)” in green, which is labeled in green as “conjugate acid.” The acid on the left side of the equation is connected to the conjugate base on the right with a purple line. Similarly, the base on the left is connected to the conjugate acid on the right side.

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-.

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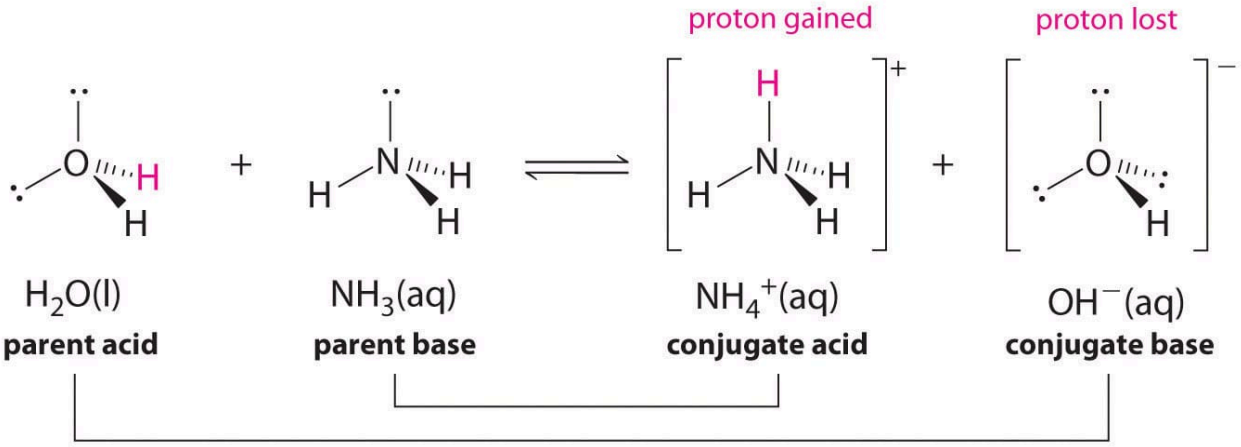


Figure 13.5.1. The pairing of parent acids and bases with conjugate acids and bases.

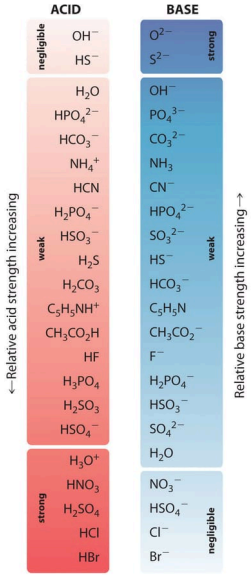


Figure 13.5.1: The Relative Strengths of Some Common Conjugate Acid-Base Pairs. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

Example 13.5.2

Identify the conjugate acid-base pairs in this equilibrium.

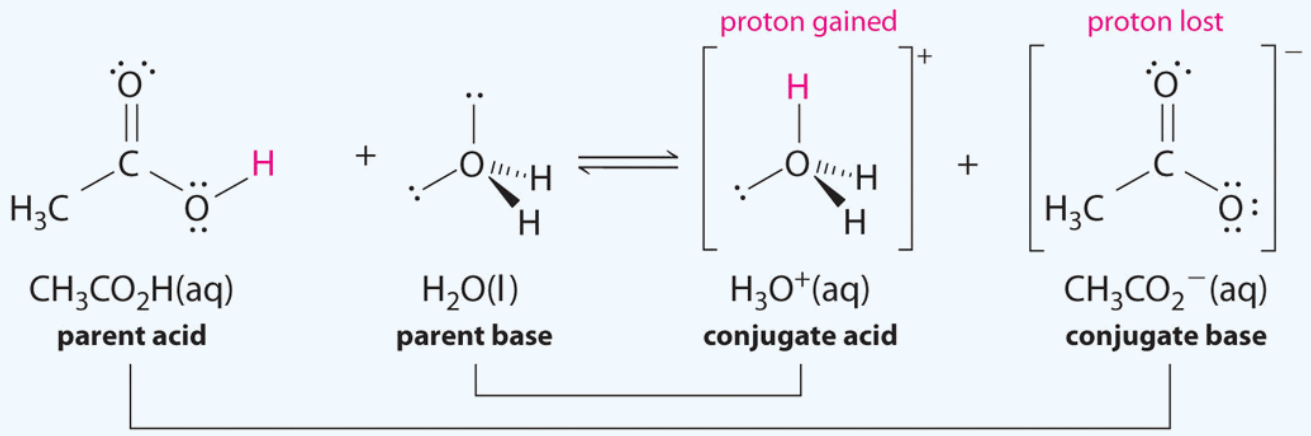
$$\mathrm{CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-}$$

Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid-base pairs:

- the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and
- the parent base and its conjugate acid (H_3O^+/H_2O).



Example 13.5.3

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H2O and OH^-, where H2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base.

The other pair consists of (CH3)3N and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and (CH3)3N is the conjugate base.

Exercise 13.5.3

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer available at the end of this section.

Answers

Exercise 13.5.1

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: H2O

Exercise 13.5.2

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H^+. This eliminates HCl, H_3PO_4, NH_4^+ and $CH_3NH_3^+$ because they are Bronsted-Lowry acids. They all give away protons. In the case of HPO_4^{2-}, consider the following equation:

$$HPO_4^{2-}(aq) + H_2O(l) \rightarrow PO_4^{3-}(aq) + H_3O^+(aq)$$

Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-}. Now consider the following equation:

$$HPO_4^{2-}(aq) + H_2O(l) \rightarrow H_2PO_4^- + OH^-(aq)$$

In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $H_2PO_4^-$ and OH^-. Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) HPO_4^{2-}.**

Exercise 13.5.3

H2O (acid) and OH^- (base); NH_2^- (base) and NH3 (acid)

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13.6: Autodissociation of Water and Kw

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H_3O^+ and OH^- in aqueous solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4{}^+ + OH^-$$

where H_2O acts as an acid (in red).

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^-$$

where H_2O acts as an base (in blue).

It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{13.6.1}$$

This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**.

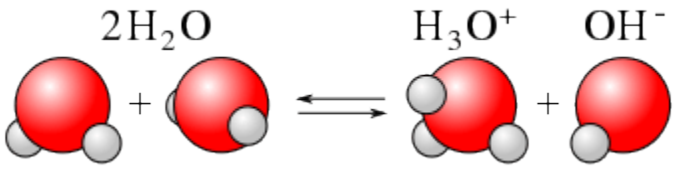


Figure 13.6.1: Autoionization of water, resulting in hydroxide and hydronium ions.

At this level, the concentration of both $H_3O^+(aq)$ and $OH^-(aq)$ in a sample of pure H_2O is about $1.0 \times 10^{-7}\ M$ (at room temperature). If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H_3O^+]=[OH^-] = 1.0 \times 10^{-7} \tag{13.6.2}$$

for *any* sample of pure water because H2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14}:

$$[H_3O^+]\times[OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

- For acids, the concentration of $H_3O^+(aq)$ (i.e., $[H_3O^+]$) is greater than $1.0 \times 10^{-7}\ M$.
- For bases the concentration of $OH^-(aq)$ (i.e., $[OH^-]$) is greater than $1.0 \times 10^{-7}\ M$.

However, the *product* of the two concentrations—$[H_3O^+][OH^-]$—is *always* equal to 1.0×10^{-14}, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{13.6.3}$$

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$) has to be for the product to equal 1.0×10^{-14}; or if you know $[OH^-]$), you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

Warning: Temperature Matters

The degree of autoionization of water (Equation 13.6.1)—and hence the value of K_w—changes with temperature, so Equations 13.6.2 - 13.6.3 are accurate only at room temperature.

Example 13.6.1: Hydroxide Concentration

What is $[OH^-]$) of an aqueous solution if $[H_3O^+]$ is $1.0 \times 10^{-4} M$?

Solution

Solutions to Example 14.7.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[H_3O^+] = 1.0 \times 10^{-4}\ M$ Find: [OH¯] = ? M
List other known quantities.	none
Plan the problem.	Using the expression for K_w, (Equation 13.6.3), rearrange the equation algebraically to solve for [OH¯]. $$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} M$$ It is assumed that the concentration unit is molarity, so $[OH^-]$ is 1.0×10^{-10} M.
Think about your result.	The concentration of the acid is high (> 1×10^{-7} M), so $[OH^-]$ should be low.

Exercise 13.6.1

What is $[OH^-]$ in a 0.00032 M solution of H_2SO_4?

Hint available at the end of this section.

Answer available at the end of this section.

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H_3O^+ or OH^- ions in the formula unit because $[H_3O^+]$ or $[OH^-]$) may not be the same as the concentration of the acid or base itself.

Example 13.6.2: Hydronium Concentration

What is $[H_3O^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

Solution

Solutions to Example 14.7.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[Ca(OH)_2] = 0.0044\ M$ Find: $[H_3O^+]$ = ? M
List other known quantities.	We begin by determining $[OH^-]$. The concentration of the solute is 0.0044 M, but because $Ca(OH)_2$ is a strong there are two OH^- ions in solution for every formula unit dissolved, so the actual $[OH^-]$ is two times this: $$[OH^-] = 2\ \times 0.0044\ \text{M} = 0.0088\,\text{M}\cdot$$
Plan the problem.	Use the expression for K_w (Equation 13.6.3) and rearrange the equation algebraically to solve for $[H_3O^+]$. $$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$[H_3O^+] = \frac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$$ $[H_3O^+]$ has decreased significantly in this basic solution.
Think about your result.	The concentration of the base is high (> 1 x 10^{-7} M) so Extra close brace or missing open brace should b

Exercise 13.6.2

What is $[H_3O^+]$ of an aqueous solution if $[OH^-]$ is $1.0 \times 10^{-9}\ M$?

Answer available at the end of this section.

In any aqueous solution, the product of $[H_3O^+]$ and $[OH^-]$ equals 1.0×10^{-14} (at room temperature).

Contributions & Attributions

Answers

Exercise 13.6.1

Assume **both** protons ionize from the molecule...although this is not the case.

Exercise 13.6.1

$3.1 \times 10^{-11}\ M$

Exercise 13.6.2

1.0×10^{-5} M

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13.7: The pH and pOH Scales

Learning Objectives

- Define pH and pOH.
- Determine the pH of acidic and basic solutions.
- Determine the hydronium ion concentration and pOH from pH.

As we have seen, $[H_3O^+]$ and $[OH^-]$ values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of $[H_3O^+]$:

$$pH = -\log[H_3O^+] \tag{13.7.1}$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H_3O^+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the pH scale. The pH scale is the range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. Figure $13.7.1$ illustrates this relationship, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H_3O^+]$, which will give a positive value for pH.

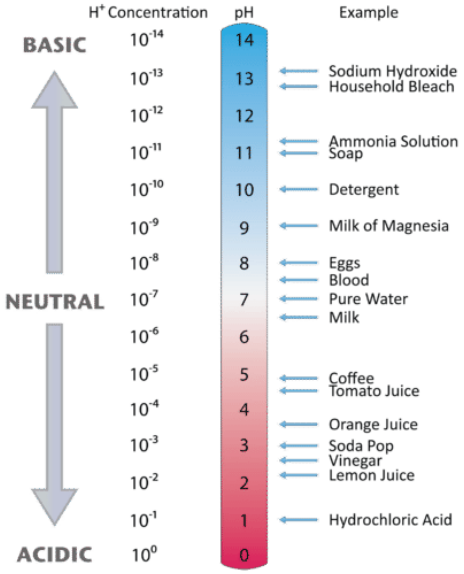


Figure $13.7.1$: The pH values for several common materials.

Example $13.7.1$

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7

c. wine, pH = 3.0

Answer

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

Exercise 13.7.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

a. human blood with pH = 7.4
b. household ammonia with pH = 11.0
c. cherries with pH = 3.6

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Calculating pH from Hydronium Concentration

The pH of solutions can be determined by using logarithms as illustrated in the next example for stomach acid. Stomach acid is a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}\ M$, what is the pH of the solution?

$$\begin{aligned} \text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) = 2.92 \end{aligned}$$

Logarithms

To get the **log value** on your calculator, enter the number (in this case, the hydronium ion concentration) first, then press the LOG key.

If the number is 1.0 x 10^{-5} (for $[H_3O^+]$ = 1.0 x 10^{-5} M) you should get an answer of "-5".

If you get a different answer, or an error, try pressing the LOG key before you enter the number.

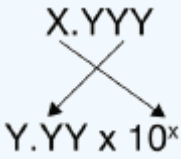
Example 13.7.2: Converting Ph to Hydronium Concentration

Find the pH, given the $[H_3O^+]$ of the following:

a. 1 ×10^{-3} M
b. 2.5 ×10^{-11} M
c. 4.7 ×10^{-9} M

Solution

Solutions to Example 14.9.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: a. $[H_3O^+]$ =1 × 10^{-3} M b. $[H_3O^+]$ =2.5 ×10^{-11} M c. $[H_3O^+]$ = 4.7 ×10^{-9} M Find: ? pH
Plan the problem.	Need to use the expression for pH (Equation 13.7.1). pH = - log $[H_3O^+]$
Calculate.	Now substitute the known quantity into the equation and solve. a. pH = - log [1 × 10^{-3}] = 3.0 (1 decimal places since 1 has 1 significant figure) b. pH = - log [2.5 ×10^{-11}] = 10.60 (2 decimal places since 2.5 has 2 significant figures) c. pH = - log [4.7 ×10^{-9}] = 8.30 (2 decimal places since 4.7 has 2 significant figures) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 13.7.2

Find the pH, given $[H_3O^+]$ of the following:

a. 5.8 ×10^{-4} M
b. 1.0×10^{-7}

Answer a available at the end of this section.

Answer b available at the end of this section.

Calculating Hydronium Concentration from pH

Sometimes you need to work "backwards"—you know the pH of a solution and need to find $[H_3O^+]$, or even the concentration of the acid solution. How do you do that? To convert pH into $[H_3O^+]$ we solve Equation 13.7.1 for $[H_3O^+]$. This involves taking the antilog (or inverse log) of the negative value of pH .

$$[H_3O^+] = \text{antilog}(-pH)$$

or

$$[H_3O^+] = 10^{-pH} \tag{13.7.2}$$

As mentioned above, different calculators work slightly differently—make sure you can do the following calculations using ***your*** calculator.

Calculator Skills

We have a solution with a pH = 8.3. What is $[H_3O^+]$?

With some calculators you will do things in the following order:

1. Enter 8.3 as a negative number (use the key with both the +/- signs, not the subtraction key).
2. Use your calculator's 2nd or Shift or INV (inverse) key to type in the symbol found above the LOG key. The shifted function should be 10^x.
3. You should get the answer **5.0×10^{-9}**.

Other calculators require you to enter keys in the order they appear in the equation.

1. Use the Shift or second function to key in the 10^x function.
2. Use the +/- key to type in a negative number, then type in 8.3.
3. You should get the answer **5.0×10^{-9}**.

If neither of these methods work, try rearranging the order in which you type in the keys. Don't give up—you must master your calculator!

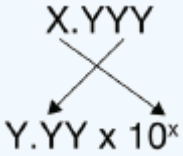
Example 13.7.3: Calculating Hydronium Concentration from pH

Find the hydronium ion concentration in a solution with a pH of 12.6. Is this solution an acid or a base? How do you know?

Solution

Solutions to Example 14.9.3

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: pH = 12.6 Find: $[H_3O^+]$ = ? M
Plan the problem.	Need to use the expression for $[H_3O^+]$ (Equation 13.7.2). $[H_3O^+]$ = antilog (-pH) or $[H_3O^+] = 10^{-pH}$

Steps for Problem Solving	
Calculate.	Now substitute the known quantity into the equation and solve. $[H_3O^+]$ = antilog (12.60) = 2.5 x 10^{-13} M (2 significant figures since 4.7 has 12.60 2 decimal places) or $[H_3O^+] = 10^{-12.60} = 2.5 \times 10^{-13}$ M (2 significant figures since 4.7 has 12.60 2 decimal places) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 13.7.3

If moist soil has a pH of 7.84, what is $[H_3O^+]$ of the soil solution?

Answer available at the end of this section.

The pOH scale

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration.

$$\mathrm{pOH} = -\log\left[\mathrm{OH}^-\right]$$

The relation between the hydronium and hydroxide ion concentrations expressed as p-functions is easily derived from the K_w expression:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH}^-] \tag{13.7.3}$$

$$-\log K_\mathrm{w} = -\log([\mathrm{H_3O^+}][\mathrm{OH}^-]) = -\log[\mathrm{H_3O^+}] + -\log[\mathrm{OH}^-] \tag{13.7.4}$$

$$\mathrm{p}K_\mathrm{w} = \mathrm{pH} + \mathrm{pOH} \tag{13.7.5}$$

At 25 °C, the value of K_w is 1.0×10^{-14}, and so:

$$14.00 = \mathrm{pH} + \mathrm{pOH} \tag{13.7.6}$$

The hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7}\ M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{13.7.7}$$

$$\mathrm{pOH} = -\log[\mathrm{OH}^-] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{13.7.8}$$

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And so, at this temperature, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities less than $1.0 \times 10^{-7}\ M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7}\ M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

Example 13.7.4:

Find the pOH of a solution with a pH of 4.42.

Solution

Solutions to Example 14.9.4

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: pH =4.42 Find: ? pOH
Plan the problem.	Need to use the expression pOH = 14 - pH
Calculate.	Now substitute the known quantity into the equation and solve. pOH=14−4.42=9.58
Think about your result.	The pH is that of an acidic solution, and the resulting pOH is the difference after subtracting from 14. The answer has two significant figures because the given pH has two decimal places.

Exercise 13.7.4

The pH of a solution is 8.22. What is the pOH?

Answer available at the end of this section.

The diagram below shows all of the interrelationships between [H3O+][H3O+], [OH−][OH−], pH, and pOH.

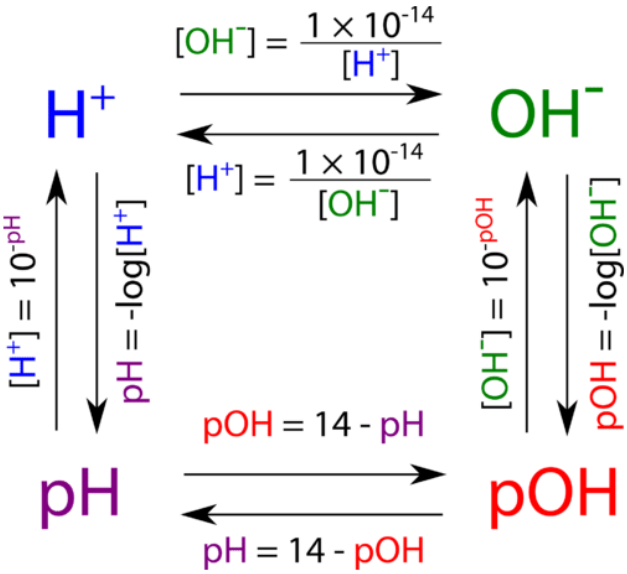


Figure 13.7.1: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH.

Contributions & Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
-
- Template:OpenStax

Answers

Exercise 13.7.1
basic

Exercise 13.7.1
basic

Exercise 13.7.1
acidic

Exercise 13.7.2
3.22

Exercise 13.7.2
7.00

Exercise 13.7.3
1.5×10^{-8} M

Exercise 13.7.4
5.78

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13.8: Acid-Base Reactions

Learning Objectives

- Write acid-base neutralization reactions.
- Write reactions of acids with metals.
- Write reactions of bases with metals.

Neutralization Reactions

The reaction that happens when an acid, such as HCl, is mixed with a base, such as NaOH:

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)}$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than $\mathrm{H^+}$ and a negative ion other than the hydroxide ion, $\mathrm{OH^-}$. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$\mathrm{H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow Na^+\,(aq) + Cl^-\,(aq) + H_2O\,(l)}$$

After removing the spectator ions, we get the net ionic equation:

$$\mathrm{H^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O\,(l)}$$

When a strong acid and a strong base are combined in the proper amounts—when $[\mathrm{H^+}]$ equals $[\mathrm{OH^-}]$\)—a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video: Equimolar (~0.01 M) and equivolume solutions of HCl and NaOH are combined to make salt water. https://youtu.be/TS-I9KrUjB0

Example 13.8.1: Propionic Acid + Calcium Hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$].

Solution

Solutions to Example 14.5.1

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$
Balance the equation. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$, to provide H^+ ions.	$\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$

Exercise 13.8.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer available at the end of this section.

Acids and Bases React with Metals

Acids react with most metals to form a salt and hydrogen gas. As discussed previously, metals that are more active than acids can undergo a **single displacement reaction**. For example, zinc metal reacts with hydrochloric acid, producing zinc chloride and hydrogen gas.

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Bases also react with certain metals, like zinc or aluminum, to produce hydrogen gas. For example, sodium hydroxide reacts with zinc and water to form sodium zincate and hydrogen gas.

$$\mathrm{Zn\,(s) + 2\,NaOH\,(aq) + 2\,H_2O\,(l) \longrightarrow Na_2Zn(OH)_4\,(aq) + H_2\,(g).}$$

Answers

Exercise 13.8.1

$$\mathrm{Ba(OH)_2\,(s) + 2\,CH_3CO_2H\ (aq) \longrightarrow Ba(CH_3CO_2)_2\,(aq) + 2\,H_2O\,(l)}$$

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13.9: Acid–Base Titrations

Learning Objectives

- Understand the basics of acid-base titrations.
- Understand the use of indicators.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base. The indicator will signal, by color change, when the base has been neutralized (when $[H^+] = [OH^-]$). At that point—called the **equivalence point,** or **end point**—the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

It is important to accurately measure volumes when doing titrations. The instrument you would use is called a burette (or buret).



Figure 13.9.1: Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\text{\# mol HCl} = (0.02566\ \text{L})(0.1078\ \text{M}) = 0.002766\ \text{mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:

$$\mathrm{HCl} + \mathrm{NaOH} \longrightarrow \mathrm{NaCl} + \mathrm{H_2O}$$

So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766\ \cancel{mol\ HCl} \times \frac{1\ mol\ NaOH}{1\ \cancel{mol\ HCl}} = 0.002766\ mol\ NaOH$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766\ \cancel{mol\ HCl} \times \frac{40.00\ g\ NaOH}{1\ \cancel{mol\ HCl}} = 0.1106\ g\ NaOH$$

This type of calculation is performed as part of a titration.

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Example 13.9.1: Equivalence Point

What mass of Ca(OH)2 is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO3? The balanced chemical equation is as follows:

$$2\,HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2\,H_2O$$

Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

moles HNO3 = (0.04402 L)(0.0885 M) = 0.00390 mol HNO3

Using the balanced chemical equation, we can determine the number of moles of Ca(OH)2 present in the analyte:

$$0.00390\ \cancel{mol\ HNO_3} \times \frac{1\,mol\,Ca(OH)_2}{2\ \cancel{mol\ HNO_3}} = 0.00195\,mol\,Ca(OH)_2$$

Then we convert this to a mass using the molar mass of Ca(OH)2:

$$0.00195\ \cancel{mol\,Ca(OH)_2} \times \frac{74.1\,g\,Ca(OH)_2}{\cancel{mol\,Ca(OH)_2}} = 0.144\,g\,Ca(OH)_2$$

Exercise 13.9.1

What mass of H2C2O4 is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:

$$H_2C_2O_4 + 2\,NaOH \longrightarrow Na_2C_2O_4 + 2\,H_2O$$

Answer available at the end of this section.

Exercise 13.9.2

If 25.00 mL of HCl solution with a concentration of 0.1234 M is neutralized by 23.45 mL of NaOH, what is the concentration of the base?

Answer available at the end of this section.

Exercise 13.9.3

A 20.0 mL solution of strontium hydroxide, $Sr(OH)_2$, is placed in a flask and a drop of indicator is added. The solution turns color after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the $Sr(OH)_2$ solution?

Answer available at the end of this section.

Indicator Selection for Titrations

The indicator used depends on the type of titration performed. The indicator of choice should change color when enough of one substance (acid or base) has been added to exactly use up the other substance. Only when a strong acid and a strong base are produced will the resulting solution be neutral. The three main types of acid-base titrations, and suggested indicators, are:

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The three main types of acid-base titrations, suggested indicators, and explanations

Titration between . . .	Indicator	Explanation
strong acid and strong base	any	
strong acid and weak base	methyl orange	changes color in the acidic range (3.2 - 4.4)
weak acid and strong base	phenolphthalein	changes color in the basic range (8.2 - 10.6)

Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all the analyte has reacted with the titrant.

Contributions & Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

Answers

Exercise 13.9.1

0.182 g

Exercise 13.9.2

0.1316 M NaOH

Exercise 13.9.3

3.12×10^{-2} M $Sr(OH)_2$

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13.10: Buffers- Solutions that Resist pH Change

Learning Objective

- Define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to life. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

This mechanism involves a buffer, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia (NH_3, a weak base) and ammonium chloride (NH_4Cl, a salt derived from that base).

Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base—a source of $\mathrm{OH^-\,(aq)}$ ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$\mathrm{HC_2H_3O_2\,(aq) + OH^-\,(aq) \rightarrow H_2O(\ell) + C_2H_3O_2^-\,(aq)} \tag{13.10.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$\mathrm{H^+\,(aq) + C_2H_3O_2^-\,(aq) \rightarrow HC_2H_3O_2\,(aq)} \tag{13.10.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure $13.10.1$ illustrates both actions of a buffer.

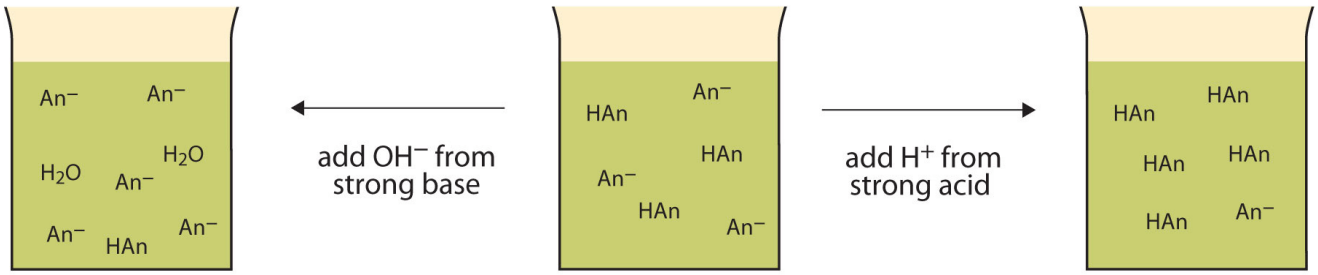


Figure $13.10.1$: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$\mathrm{NH_3\,(aq) + H^+\,(aq) \ \rightarrow NH_4^+\,(aq)} \tag{13.10.3}$$

while the **ammonium ion** ($\mathrm{NH_4{}^+\,(aq)}$) can react with any hydroxide ions introduced by strong bases:

$$\mathrm{NH_4^+\,(aq) + OH^-\,(aq) \ \rightarrow NH_3\,(aq) + H_2O(\ell)} \tag{13.10.4}$$

Example 13.10.1: Making Buffer Solutions

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $HCHO_2$ and $NaCHO_2$
b. HCl and NaCl
c. CH_3NH_2 and CH_3NH_3Cl
d. NH_3 and NaOH

Solution

a. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid—the formate ion (CHO_2^-). The combination of these two solutes would make a buffer solution.
b. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
c. Methylamine (CH_3NH_2) is like ammonia with one of its hydrogen atoms substituted with a CH_3 (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound CH_3NH_3Cl is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
d. Ammonia (NH_3) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

Exercise 13.10.1

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. H_3PO_4 and NaH_2PO_4
c. NH_3 and $(NH_4)_3PO_4$
d. NaOH and NaCl

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write CO_2(aq)]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering

system based on phosphate ions.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/– type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

Contributions & Attributions

Answers

Exercise 13.10.1

Yes.

Exercise 13.10.1

No. Need a weak acid or base and a salt of its conjugate base or acid.

Exercise 13.10.1

Yes.

Exercise 13.10.1

No. Need a weak base or acid.

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INDEX

A

absolute zero
10.4: Charles's Law- Volume and Temperature

acceleration
2.12: Derived Units

accuracy
2.2: Significant Figures - Writing Numbers to Reflect Precision

acetic acid
13.2: Acids- Properties and Examples

acid
13.8: Acid-Base Reactions

acid nomenclature
5: Chemical Bonding and Nomenclature
5.12: Naming Acids

acidity
13.7: The pH and pOH Scales

acids
13: Electrolytes, Acids and Bases
13.2: Acids- Properties and Examples
5.13: Nomenclature Summary
8.8: Acid–Base Reactions

actual yield
9.10: Yields

addition
2.3: Scientific Notation - Writing Large and Small Numbers
2.4: Significant Figures in Calculations

algebra
10.7: The Combined Gas Law

alkali metals
3.11: Looking for Patterns - The Periodic Table

alkaline earth metals
3.11: Looking for Patterns - The Periodic Table

anion
4.13: Ionic Bonding
5.2: Representing Valence Electrons with Dots
5.9: Writing Formulas for Ionic Compounds

anions
3.12: Ions - Losing and Gaining Electrons
3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character
4.11: Electron Configurations of Anions
5.10: Naming Ionic Compounds

aqueous solution
8.5: Aqueous Solutions and Solubility

aqueous solutions
11.6: Solution Stoichiometry

area
2.11: Units Raised to a Power
2.12: Derived Units

arithmetic
2.4: Significant Figures in Calculations

arithmetic operations
2.3: Scientific Notation - Writing Large and Small Numbers

Arrhenius acid
13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory

Arrhenius base
13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory

atmosphere
10.2: What is Gas Pressure?

Atom
3.2: Indivisible - The Atomic Theory

atomic elements
5.1: A Molecular View of Elements and Compounds

atomic mass
3.10: Atomic Mass - The Average Mass of an Element's Atoms
3.11: Looking for Patterns - The Periodic Table
3.9: Isotopes - When the Number of Neutrons Varies

atomic mass unit
3.7: The Properties of Protons, Neutrons, and Electrons
3.8: Elements- Defined by Their Number of Protons

atomic model
3.6: The Nuclear Atom
4: The Modern Model of the Atom

atomic number
3: The Evolution of Atomic Theory
3.10: Atomic Mass - The Average Mass of an Element's Atoms
3.8: Elements- Defined by Their Number of Protons

atomic size
3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character

atomic structure
3.6: The Nuclear Atom
3.7: The Properties of Protons, Neutrons, and Electrons
3.8: Elements- Defined by Their Number of Protons

atomic theory
3: The Evolution of Atomic Theory
3.2: Indivisible - The Atomic Theory
3.6: The Nuclear Atom

atoms
1.2: What is Matter?
3.1: Cutting Aluminum until you get Atoms
9.2: Conversions Between Moles and Particles

attractive
7: Intermolecular Forces

autoionization
13.6: Autodissociation of Water and Kw

average mass
3.10: Atomic Mass - The Average Mass of an Element's Atoms

Avogadro
9.1: Avogadro's Number
9.5: Conversions Between Mass and Number of Particles

avogadro's law
10.6: Avogadro's Law- Volume and Moles

Avogadro's number
9: Stoichiometry and the Mole

B

balanced chemical equation
8.3: How to Write Balanced Chemical Equations
9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!

balanced chemical equations
9.9: Mole-Mass and Mass-Mass Calculations

balanced equations
8.8: Acid–Base Reactions
9.7: Stoichiometry

balancing equations
8: Chemical Reactions

base
13.8: Acid-Base Reactions

base units
2.5: The Basic Units of Measurement

bases
13: Electrolytes, Acids and Bases
8.8: Acid–Base Reactions

basic units
2: The Numerical Side of Chemistry

basicity
13.7: The pH and pOH Scales

binary ionic compounds
5.9: Writing Formulas for Ionic Compounds

Bohr model
4.4: The Bohr Model - Atoms with Orbits

boiling point
1.5: Differences in Matter- Physical and Chemical Properties
2.6: Temperature - Random Motion of Molecules and Atoms

boiling points
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"

bond
4.13: Ionic Bonding

bond lengths
5.6: Resonance - Equivalent Lewis Structures for the Same Molecule

bond polarity
5.7: Bond Polarity
6.2: Molecule Polarity

bonding
5.4: Lewis Structures of Covalent Compounds- Electrons Shared
7: Intermolecular Forces

Boyle's law
10.3: Boyle's Law - Pressure and Volume

buffer
13.10: Buffers- Solutions that Resist pH Change

buffers
13: Electrolytes, Acids and Bases

C

calculation
2.10: Solving Multi-step Conversion Problems

calculations
2: The Numerical Side of Chemistry
2.4: Significant Figures in Calculations

cation
- 4.13: Ionic Bonding
- 5.2: Representing Valence Electrons with Dots
- 5.9: Writing Formulas for Ionic Compounds

cations
- 3.12: Ions - Losing and Gaining Electrons
- 3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character
- 4.10: Electron Configurations of Transition Metal Cations
- 4.9: Electron Configurations of Cations
- 5.10: Naming Ionic Compounds

Celsius
- 2.6: Temperature - Random Motion of Molecules and Atoms

changes
- 1.1: Scope of Chemistry

charge balance
- 5.9: Writing Formulas for Ionic Compounds

charles law
- 10.4: Charles's Law- Volume and Temperature

Chemical
- 13.8: Acid-Base Reactions

chemical bonding
- 5.3: Lewis Structures of Ionic Compounds- Electrons Transferred

chemical bonds
- 8.1: Evidence of a Chemical Reaction
- 8.2: The Chemical Equation

chemical change
- 1.6: Changes in Matter - Physical and Chemical Changes

chemical changes
- 1: What is Chemistry?

chemical compounds
- 3.2: Indivisible - The Atomic Theory
- 3.5: Law of Definite Proportions - Calculating Mass Percent
- 5.8: Chemical Nomenclature

chemical equations
- 8: Chemical Reactions
- 8.2: The Chemical Equation
- 8.3: How to Write Balanced Chemical Equations

chemical equilibrium
- 12.3: Heterogeneous Equilibria

chemical formula
- 5.8: Chemical Nomenclature
- 9.12: Mass Percent Composition from a Chemical Formula
- 9.14: Calculating Molecular Formulas

chemical identification
- 5.8: Chemical Nomenclature

chemical name
- 5.8: Chemical Nomenclature

chemical nomenclature
- 5: Chemical Bonding and Nomenclature

chemical properties
- 1: What is Chemistry?
- 1.5: Differences in Matter- Physical and Chemical Properties
- 3.11: Looking for Patterns - The Periodic Table

chemical reaction
- 1.6: Changes in Matter - Physical and Chemical Changes
- 3.3: Law of Conservation of Mass

chemical reactions
- 1.1: Scope of Chemistry
- 3: The Evolution of Atomic Theory
- 3.4: Law of Definite Proportions
- 8.1: Evidence of a Chemical Reaction
- 8.2: The Chemical Equation
- 9.7: Stoichiometry
- 9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!

chemical symbols
- 3.8: Elements- Defined by Their Number of Protons

chemistry
- 1: What is Chemistry?
- 1.1: Scope of Chemistry
- 1.8: A Beginning Chemist - How to Succeed
- 2.1: Taking Measurements
- 3.1: Cutting Aluminum until you get Atoms

classification of matter
- 1.3: Classifying Matter According to Its Composition

classification of solids
- 7.4: Types of Crystalline Solids

combined gas law
- 10: Gases
- 10.7: The Combined Gas Law

combustion
- 8.1: Evidence of a Chemical Reaction

combustion reactions
- 8.4: Classifying Chemical Reactions

complete ionic equation
- 8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

composition
- 1.1: Scope of Chemistry

compound
- 1.3: Classifying Matter According to Its Composition

concentrated solution
- 11.4: Solution Dilution (M1V1 = M2V2)

concentration
- 11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
- 11.5: Calculating Solution Concentration- Mass Percent
- 12.2: The Equilibrium Constant, Keq
- 13.6: Autodissociation of Water and Kw
- 2.12: Derived Units

conductivity
- 11.2: Solutions of Solids Dissolved in Water
- 8.5: Aqueous Solutions and Solubility

conservation of mass
- 3: The Evolution of Atomic Theory

constant
- 12.2: The Equilibrium Constant, Keq

conversion
- 10.2: What is Gas Pressure?

conversion factor
- 10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!
- 2.7: Density

conversion factors
- 2.10: Solving Multi-step Conversion Problems
- 2.11: Units Raised to a Power
- 2.8: Problem Solving and Unit Conversions
- 9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"
- 9.7: Stoichiometry
- 9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!

conversion methods
- 9: Stoichiometry and the Mole

conversions
- 2.3: Scientific Notation - Writing Large and Small Numbers
- 2.5: The Basic Units of Measurement
- 9.2: Conversions Between Moles and Particles
- 9.4: Conversions Between Moles and Mass
- 9.5: Conversions Between Mass and Number of Particles

counting
- 9.1: Avogadro's Number

covalent
- 5.13: Nomenclature Summary

covalent bond
- 5.4: Lewis Structures of Covalent Compounds- Electrons Shared

covalent bonding
- 5.1: A Molecular View of Elements and Compounds
- 5.6: Resonance - Equivalent Lewis Structures for the Same Molecule

covalent bonds
- 5.11: Naming Covalent Compounds

covalent compounds
- 5: Chemical Bonding and Nomenclature
- 5.11: Naming Covalent Compounds
- 5.5: Rules for Drawing Lewis Structures

covalent network crystals
- 7.4: Types of Crystalline Solids

cubic centimeters
- 2.11: Units Raised to a Power

D

Dalton's law
- 10.12: Dalton's Law of Partial Pressures

decomposition reactions
- 8.4: Classifying Chemical Reactions

definite proportions
- 3.4: Law of Definite Proportions

democritus
- 3.1: Cutting Aluminum until you get Atoms

density
- 2: The Numerical Side of Chemistry
- 2.12: Derived Units
- 2.7: Density

diatomic molecules
- 3.1: Cutting Aluminum until you get Atoms
- 5.1: A Molecular View of Elements and Compounds

dilute solution
- 11.4: Solution Dilution (M1V1 = M2V2)

dilution
- 11: Solutions

dimensional analysis
- 2: The Numerical Side of Chemistry
- 2.10: Solving Multi-step Conversion Problems
- 2.11: Units Raised to a Power
- 2.12: Derived Units
- 2.8: Problem Solving and Unit Conversions
- 2.9: Metric Unit Conversions

dissociation
- 13.1: Electrolytes
- 13.4: Strong and Weak Acids and Bases

distance
- 2.1: Taking Measurements
- 2.9: Metric Unit Conversions

division
2.3: Scientific Notation - Writing Large and Small Numbers
2.4: Significant Figures in Calculations
double replacement reactions
11.6: Solution Stoichiometry
8.4: Classifying Chemical Reactions

E

education
2.1: Taking Measurements
electrolyte
13.1: Electrolytes
electrolytes
11.2: Solutions of Solids Dissolved in Water
13: Electrolytes, Acids and Bases
13.3: Bases- Properties and Examples
8.5: Aqueous Solutions and Solubility
electromagnetic
4.1: Electromagnetic Spectrum
electromagnetic spectrum
4: The Modern Model of the Atom
electron behavior
4: The Modern Model of the Atom
electron configuration
4.10: Electron Configurations of Transition Metal Cations
4.7: Electron Configuration Shortcut - Noble Gas Configurations
4.9: Electron Configurations of Cations
5.2: Representing Valence Electrons with Dots
5.3: Lewis Structures of Ionic Compounds- Electrons Transferred
5.5: Rules for Drawing Lewis Structures
electron configurations
4.6: Electron Configurations
electron density
5.7: Bond Polarity
electron group geometry
6.1: VSEPR Theory
electron repulsion
6: The Shape of Molecules
electron_configuration
4.12: Octet Rule
electronegativity
5.7: Bond Polarity
6: The Shape of Molecules
6.2: Molecule Polarity
Electrons
3.7: The Properties of Protons, Neutrons, and Electrons
3.9: Isotopes - When the Number of Neutrons Varies
4.1: Electromagnetic Spectrum
4.12: Octet Rule
element
1.3: Classifying Matter According to Its Composition
elements
1.2: What is Matter?
3.1: Cutting Aluminum until you get Atoms
emission spectrum
4.4: The Bohr Model - Atoms with Orbits
empirical formula
9.13: Calculating Empirical Formulas
empirical formulas
9.14: Calculating Molecular Formulas
energy
2.12: Derived Units
4: The Modern Model of the Atom
4.3: Energy and Frequency Calculations
4.5: The Quantum-Mechanical Model- Atoms with Orbitals
energy levels
4.4: The Bohr Model - Atoms with Orbits
English units
2.8: Problem Solving and Unit Conversions
equation balancing
8.2: The Chemical Equation
equilibrium
12.1: The Idea of Dynamic Chemical Equilibrium
12.2: The Equilibrium Constant, Keq
13.6: Autodissociation of Water and Kw
equilibrium constant
12: Chemical Equilibrium
12.2: The Equilibrium Constant, Keq
12.4: Calculations with Equilibrium Constants
equivalence point
13.9: Acid–Base Titrations
evidence of change
1.6: Changes in Matter - Physical and Chemical Changes
exact numbers
2.2: Significant Figures - Writing Numbers to Reflect Precision
excess reactant
9.11: Limiting Reagents
excited state
4.4: The Bohr Model - Atoms with Orbits
exponents
2.3: Scientific Notation - Writing Large and Small Numbers

F

Fahrenheit
2.6: Temperature - Random Motion of Molecules and Atoms
families
3.11: Looking for Patterns - The Periodic Table
farming
2.12: Derived Units
flammability
1.5: Differences in Matter- Physical and Chemical Properties
flowchart
5.13: Nomenclature Summary
forces
7: Intermolecular Forces
freezing point
2.6: Temperature - Random Motion of Molecules and Atoms
7.1: Interactions between Molecules
frequency
4: The Modern Model of the Atom
4.2: Wavelength and Frequency Calculations
4.3: Energy and Frequency Calculations

G

g/cm3
2.7: Density
gas
1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas
7.2: Properties of Liquids and Solids
gas behavior
10.4: Charles's Law- Volume and Temperature
10.9: Molar Volume of a Gas at STP
gas collisions
10.1: Kinetic Molecular Theory- A Model for Gases
gas constant
10.8: The Ideal Gas Law
gas law
10.5: Gay-Lussac's Law- Temperature and Pressure
gas laws
10: Gases
10.3: Boyle's Law - Pressure and Volume
10.4: Charles's Law- Volume and Temperature
10.6: Avogadro's Law- Volume and Moles
gas mixture
10.12: Dalton's Law of Partial Pressures
gas pressure
10: Gases
10.12: Dalton's Law of Partial Pressures
gas properties
10.3: Boyle's Law - Pressure and Volume
10.7: The Combined Gas Law
gas stoichiometry
10.11: Gas Stoichiometry
gases
1.2: What is Matter?
get to moles
10.11: Gas Stoichiometry
grams
2.7: Density
Greek prefixes
5.11: Naming Covalent Compounds
ground state
4.4: The Bohr Model - Atoms with Orbits
groups
3.11: Looking for Patterns - The Periodic Table

H

halogens
3.11: Looking for Patterns - The Periodic Table
heat
2.6: Temperature - Random Motion of Molecules and Atoms
heterogeneous
12.3: Heterogeneous Equilibria
heterogeneous equilibria
12: Chemical Equilibrium
heterogeneous mixture
1.3: Classifying Matter According to Its Composition
homogeneous equilibria
12: Chemical Equilibrium
homogeneous mixture
1.3: Classifying Matter According to Its Composition
hydrochloric acid
13.2: Acids- Properties and Examples
hydrogen bonds
7.1: Interactions between Molecules
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"
hydronium ion
13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory

hydronium ion concentration
13.7: The pH and pOH Scales
hydroxide ion
13.5: Arrhenius vs. Bronsted-Lowry Acid-Base Theory
hypothesis
1.7: The Scientific Method - How Chemists Think

I

ideal gas
10.1: Kinetic Molecular Theory- A Model for Gases
ideal gas law
10: Gases
10.11: Gas Stoichiometry
10.8: The Ideal Gas Law
indicator
13.9: Acid–Base Titrations
indicators
13: Electrolytes, Acids and Bases
insoluble products
8.6: Precipitation Reactions
insoluble salts
8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations
interactions
7.2: Properties of Liquids and Solids
intermolecular
7: Intermolecular Forces
7.2: Properties of Liquids and Solids
intermolecular forces
7.1: Interactions between Molecules
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"
intermolecular interactions
11.1: Solutions - Homogeneous Mixtures
ion formation
4.11: Electron Configurations of Anions
ionic
4.13: Ionic Bonding
5.13: Nomenclature Summary
ionic bonds
7.1: Interactions between Molecules
ionic compounds
3.12: Ions - Losing and Gaining Electrons
5: Chemical Bonding and Nomenclature
5.1: A Molecular View of Elements and Compounds
8.5: Aqueous Solutions and Solubility
ionic nomenclature
5.10: Naming Ionic Compounds
ions
3.12: Ions - Losing and Gaining Electrons
isoelectronic
4.11: Electron Configurations of Anions
isotopes
3: The Evolution of Atomic Theory
3.10: Atomic Mass - The Average Mass of an Element's Atoms
3.8: Elements- Defined by Their Number of Protons
3.9: Isotopes - When the Number of Neutrons Varies

K

Kelvin
10.7: The Combined Gas Law
10.8: The Ideal Gas Law
2.6: Temperature - Random Motion of Molecules and Atoms
Kelvin scale
10.4: Charles's Law- Volume and Temperature
Keq
12: Chemical Equilibrium
12.1: The Idea of Dynamic Chemical Equilibrium
12.3: Heterogeneous Equilibria
kilograms
2.7: Density
kinetic energy
2.6: Temperature - Random Motion of Molecules and Atoms
kinetic molecular theory
10: Gases
kinetic theory of gases
10.1: Kinetic Molecular Theory- A Model for Gases
KMT
10.1: Kinetic Molecular Theory- A Model for Gases

L

law of conservation of mass
3.3: Law of Conservation of Mass
8.3: How to Write Balanced Chemical Equations
Le Chatelier's principle
12: Chemical Equilibrium
12.5: Le Chatelier's Principle
learning
2.1: Taking Measurements
learning objectives
11.6: Solution Stoichiometry
2.5: The Basic Units of Measurement
length
2.5: The Basic Units of Measurement
2.8: Problem Solving and Unit Conversions
length units
2.9: Metric Unit Conversions
Lewis diagrams
5: Chemical Bonding and Nomenclature
Lewis electron dot diagram
5.2: Representing Valence Electrons with Dots
5.4: Lewis Structures of Covalent Compounds- Electrons Shared
Lewis structures
5.3: Lewis Structures of Ionic Compounds- Electrons Transferred
5.5: Rules for Drawing Lewis Structures
5.6: Resonance - Equivalent Lewis Structures for the Same Molecule
light
4: The Modern Model of the Atom
4.1: Electromagnetic Spectrum
limiting reactant
9: Stoichiometry and the Mole
9.11: Limiting Reagents
limiting reagent
9.11: Limiting Reagents
linear
6.1: VSEPR Theory
liquid
1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas
7.2: Properties of Liquids and Solids
liquids
7.1: Interactions between Molecules
liters
2.9: Metric Unit Conversions

M

mass
1.2: What is Matter?
11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
2.5: The Basic Units of Measurement
2.7: Density
2.8: Problem Solving and Unit Conversions
9.4: Conversions Between Moles and Mass
9.5: Conversions Between Mass and Number of Particles
mass calculations
9.9: Mole-Mass and Mass-Mass Calculations
mass conservation
3.4: Law of Definite Proportions
mass percent
11.5: Calculating Solution Concentration- Mass Percent
9.12: Mass Percent Composition from a Chemical Formula
matter
1: What is Chemistry?
1.1: Scope of Chemistry
1.2: What is Matter?
3.3: Law of Conservation of Mass
matter classification
1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas
measurement
2.1: Taking Measurements
measurement systems
2.5: The Basic Units of Measurement
measurement uncertainty
2.2: Significant Figures - Writing Numbers to Reflect Precision
measurements
2.1: Taking Measurements
melting
7: Intermolecular Forces
melting point
1.5: Differences in Matter- Physical and Chemical Properties
7.1: Interactions between Molecules
melting points
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"
7.4: Types of Crystalline Solids
metallic character
3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character
metallic crystals
7.4: Types of Crystalline Solids
meters
2.9: Metric Unit Conversions
metric conversions
2.9: Metric Unit Conversions
metric prefixes
2.5: The Basic Units of Measurement
2.9: Metric Unit Conversions

metric system
- 2.10: Solving Multi-step Conversion Problems
- 2.8: Problem Solving and Unit Conversions

metric units
- 2.11: Units Raised to a Power

milliliters
- 2.9: Metric Unit Conversions

mixture
- 1.3: Classifying Matter According to Its Composition

mixtures
- 1: What is Chemistry?

mL
- 2.7: Density

mmHg
- 10.2: What is Gas Pressure?

molar mass
- 11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
- 9: Stoichiometry and the Mole
- 9.12: Mass Percent Composition from a Chemical Formula
- 9.4: Conversions Between Moles and Mass

molar volume
- 10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!

molar volume of gas
- 10.9: Molar Volume of a Gas at STP

molar_mass
- 9.3: Molar Mass

molarity
- 11: Solutions
- 11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
- 11.4: Solution Dilution (M1V1 = M2V2)

mole
- 9: Stoichiometry and the Mole
- 9.1: Avogadro's Number

mole calculations
- 9.9: Mole-Mass and Mass-Mass Calculations

molecular crystals
- 7.4: Types of Crystalline Solids

molecular elements
- 5.1: A Molecular View of Elements and Compounds

molecular formulas
- 9.14: Calculating Molecular Formulas

molecular geometry
- 6: The Shape of Molecules
- 6.1: VSEPR Theory

molecular polarity
- 6.2: Molecule Polarity

molecular relationships
- 9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"

Molecular Shapes
- 6.1: VSEPR Theory

molecule polarity
- 6: The Shape of Molecules

molecules
- 1.2: What is Matter?
- 3.1: Cutting Aluminum until you get Atoms
- 7: Intermolecular Forces
- 9.2: Conversions Between Moles and Particles

moles
- 9.2: Conversions Between Moles and Particles
- 9.3: Molar Mass
- 9.4: Conversions Between Moles and Mass
- 9.5: Conversions Between Mass and Number of Particles
- 9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"

moles of gas
- 10.6: Avogadro's Law- Volume and Moles

monatomic ions
- 5.10: Naming Ionic Compounds

multiplication
- 2.3: Scientific Notation - Writing Large and Small Numbers
- 2.4: Significant Figures in Calculations

N

naming acids
- 5.12: Naming Acids

net ionic equation
- 8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

neutralization
- 13: Electrolytes, Acids and Bases
- 13.3: Bases- Properties and Examples
- 13.8: Acid-Base Reactions
- 13.9: Acid–Base Titrations

neutralization reactions
- 8.8: Acid–Base Reactions

neutrons
- 3.7: The Properties of Protons, Neutrons, and Electrons
- 3.8: Elements- Defined by Their Number of Protons
- 3.9: Isotopes - When the Number of Neutrons Varies

nitric acid
- 13.2: Acids- Properties and Examples

noble gas configuration
- 4.10: Electron Configurations of Transition Metal Cations

noble gases
- 3.11: Looking for Patterns - The Periodic Table
- 4.7: Electron Configuration Shortcut - Noble Gas Configurations

noble_gases
- 4.12: Octet Rule

nomenclature
- 5.13: Nomenclature Summary
- 5.8: Chemical Nomenclature

nonelectrolyte
- 13.1: Electrolytes

nonpolar compounds
- 11.1: Solutions - Homogeneous Mixtures
- 6: The Shape of Molecules
- 6.2: Molecule Polarity

nonpolar covalent bond
- 5.7: Bond Polarity

nonpolar molecules
- 7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"

O

objectives
- 2.1: Taking Measurements

observations
- 1.7: The Scientific Method - How Chemists Think

octet
- 4.11: Electron Configurations of Anions

octet rule
- 5.3: Lewis Structures of Ionic Compounds- Electrons Transferred
- 5.5: Rules for Drawing Lewis Structures

octet_rule
- 4.12: Octet Rule

operations
- 2.4: Significant Figures in Calculations

orbital
- 4.5: The Quantum-Mechanical Model- Atoms with Orbitals

oxidation states
- 4.10: Electron Configurations of Transition Metal Cations

oxyacids
- 5.12: Naming Acids

P

partial pressure
- 10.12: Dalton's Law of Partial Pressures

Pascal
- 10.2: What is Gas Pressure?

percent by mass
- 3.5: Law of Definite Proportions - Calculating Mass Percent

percent composition
- 3.5: Law of Definite Proportions - Calculating Mass Percent
- 9.12: Mass Percent Composition from a Chemical Formula
- 9.13: Calculating Empirical Formulas
- 9.14: Calculating Molecular Formulas

percent yield
- 9.10: Yields

periodic table
- 3: The Evolution of Atomic Theory
- 3.11: Looking for Patterns - The Periodic Table
- 3.12: Ions - Losing and Gaining Electrons

periodic trends
- 3.13: Periodic Trends - Atomic Size, Ionization Energy, and Metallic Character

periodic_table
- 9.3: Molar Mass

pH
- 13.6: Autodissociation of Water and Kw
- 13.7: The pH and pOH Scales
- 13.8: Acid-Base Reactions

pH control
- 13.10: Buffers- Solutions that Resist pH Change

pH scale
- 13: Electrolytes, Acids and Bases

phase transition
- 7.1: Interactions between Molecules

photon
- 4.3: Energy and Frequency Calculations
- 4.4: The Bohr Model - Atoms with Orbits

physical change
- 1.6: Changes in Matter - Physical and Chemical Changes

physical changes
- 1: What is Chemistry?

physical properties
1: What is Chemistry?
1.5: Differences in Matter- Physical and Chemical Properties
physical property
2.7: Density
Planck's constant
4.3: Energy and Frequency Calculations
pOH
13.7: The pH and pOH Scales
polar compounds
11.1: Solutions - Homogeneous Mixtures
6: The Shape of Molecules
6.2: Molecule Polarity
polar covalent bond
5.7: Bond Polarity
Polar Molecules
7.3: Intermolecular Forces - London Dispersion Forces, Dipole–Dipole Forces, Hydrogen "Bonding"
polarity
11: Solutions
polyatomic ions
5.10: Naming Ionic Compounds
potassium
4.7: Electron Configuration Shortcut - Noble Gas Configurations
precipitation reaction
8.6: Precipitation Reactions
8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations
precipitation reactions
8: Chemical Reactions
precision
2.2: Significant Figures - Writing Numbers to Reflect Precision
pressure
10.2: What is Gas Pressure?
10.5: Gay-Lussac's Law- Temperature and Pressure
10.7: The Combined Gas Law
problem solving
2: The Numerical Side of Chemistry
2.11: Units Raised to a Power
product formation
9.11: Limiting Reagents
products
3.3: Law of Conservation of Mass
8.3: How to Write Balanced Chemical Equations
properties of atoms
3.2: Indivisible - The Atomic Theory
properties of matter
2.1: Taking Measurements
protons
3.6: The Nuclear Atom
3.7: The Properties of Protons, Neutrons, and Electrons
3.8: Elements- Defined by Their Number of Protons
3.9: Isotopes - When the Number of Neutrons Varies
pure substance
1.3: Classifying Matter According to Its Composition
pure substances
1: What is Chemistry?
PV=nRT
10.8: The Ideal Gas Law

Q

qualitative
1.7: The Scientific Method - How Chemists Think
quantitative
1.7: The Scientific Method - How Chemists Think
quantum mechanics
4.5: The Quantum-Mechanical Model- Atoms with Orbitals

R

rate
12.1: The Idea of Dynamic Chemical Equilibrium
ratio relationships
9.7: Stoichiometry
ratios
9.6: Chemical Formulas as Conversion Factors - "Big Particle/Little Particle"
reactants
3.3: Law of Conservation of Mass
8.3: How to Write Balanced Chemical Equations
reactivity
4.8: Valence Electrons
relative abundance
3.10: Atomic Mass - The Average Mass of an Element's Atoms
resonance
5.6: Resonance - Equivalent Lewis Structures for the Same Molecule
resonance structures
5.6: Resonance - Equivalent Lewis Structures for the Same Molecule
reversible
12.1: The Idea of Dynamic Chemical Equilibrium
reversible reactions
12: Chemical Equilibrium
rigid container
10.5: Gay-Lussac's Law- Temperature and Pressure
rounding
2.4: Significant Figures in Calculations

S

salts
8.8: Acid–Base Reactions
saturated solution
11.2: Solutions of Solids Dissolved in Water
scientific method
1: What is Chemistry?
1.7: The Scientific Method - How Chemists Think
3.2: Indivisible - The Atomic Theory
scientific notation
2: The Numerical Side of Chemistry
2.2: Significant Figures - Writing Numbers to Reflect Precision
2.3: Scientific Notation - Writing Large and Small Numbers
2.4: Significant Figures in Calculations
SI units
2.12: Derived Units
2.5: The Basic Units of Measurement
2.8: Problem Solving and Unit Conversions
significant figure rules
2.2: Significant Figures - Writing Numbers to Reflect Precision
significant figures
2: The Numerical Side of Chemistry
2.10: Solving Multi-step Conversion Problems
2.2: Significant Figures - Writing Numbers to Reflect Precision
2.4: Significant Figures in Calculations
2.8: Problem Solving and Unit Conversions
single replacement reactions
8.4: Classifying Chemical Reactions
solid
1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas
7.2: Properties of Liquids and Solids
solids
7.1: Interactions between Molecules
solubility
11: Solutions
11.1: Solutions - Homogeneous Mixtures
11.2: Solutions of Solids Dissolved in Water
8: Chemical Reactions
8.6: Precipitation Reactions
solubility rules
8.5: Aqueous Solutions and Solubility
8.6: Precipitation Reactions
solute
11.5: Calculating Solution Concentration- Mass Percent
solutes
11: Solutions
solutions
11: Solutions
11.1: Solutions - Homogeneous Mixtures
11.5: Calculating Solution Concentration- Mass Percent
solvent
11.5: Calculating Solution Concentration- Mass Percent
solvents
11: Solutions
spectator ions
8.7: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations
spectrum
4.1: Electromagnetic Spectrum
speed
2.12: Derived Units
speed of light
4.2: Wavelength and Frequency Calculations
square centimeters
2.11: Units Raised to a Power
square meters
2.11: Units Raised to a Power
states of matter
1: What is Chemistry?
1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas
stock solutions
11.4: Solution Dilution (M1V1 = M2V2)
stoichiometry
11.6: Solution Stoichiometry
12.4: Calculations with Equilibrium Constants
9: Stoichiometry and the Mole
9.11: Limiting Reagents
9.7: Stoichiometry
9.8: Mole-to-Mole Conversion - The Heart of Stoichiometry!
9.9: Mole-Mass and Mass-Mass Calculations

STP
10: Gases
10.10: Converting Between Moles and Gas Volume - STP is a Conversion Factor!
10.11: Gas Stoichiometry
10.9: Molar Volume of a Gas at STP
strong acid
13.10: Buffers- Solutions that Resist pH Change
strong acids
13.4: Strong and Weak Acids and Bases
strong base
13.10: Buffers- Solutions that Resist pH Change
strong bases
13.3: Bases- Properties and Examples
13.4: Strong and Weak Acids and Bases
strong electrolyte
13.1: Electrolytes
subshells
4.6: Electron Configurations
substances
2.7: Density
subtraction
2.3: Scientific Notation - Writing Large and Small Numbers
2.4: Significant Figures in Calculations
sulfuric acid
13.2: Acids- Properties and Examples
symmetry
6.2: Molecule Polarity
synthesis reactions
8.4: Classifying Chemical Reactions

T

taking measurements
2: The Numerical Side of Chemistry
temperature
2: The Numerical Side of Chemistry
2.1: Taking Measurements
2.5: The Basic Units of Measurement
2.6: Temperature - Random Motion of Molecules and Atoms
temperature change
8.1: Evidence of a Chemical Reaction
tetrahedral
6.1: VSEPR Theory
theoretical yield
9.10: Yields
thermal energy
2.6: Temperature - Random Motion of Molecules and Atoms
thermometers
2.6: Temperature - Random Motion of Molecules and Atoms
time
2.5: The Basic Units of Measurement
titration
13.9: Acid–Base Titrations
titrations
13: Electrolytes, Acids and Bases
track laps
2.9: Metric Unit Conversions
transition metals
4.10: Electron Configurations of Transition Metal Cations
trigonal planar
6.1: VSEPR Theory

U

unit conversion
11.3: Calculating Solution Concentration- Molarity (is a Conversion Factor!)
2.10: Solving Multi-step Conversion Problems
2.11: Units Raised to a Power
unit conversions
2: The Numerical Side of Chemistry
Units
2.1: Taking Measurements
units of measurement
2.8: Problem Solving and Unit Conversions

V

valence electrons
3.12: Ions - Losing and Gaining Electrons
4.6: Electron Configurations
4.7: Electron Configuration Shortcut - Noble Gas Configurations
4.8: Valence Electrons
4.9: Electron Configurations of Cations
5: Chemical Bonding and Nomenclature
5.2: Representing Valence Electrons with Dots
5.4: Lewis Structures of Covalent Compounds- Electrons Shared
5.5: Rules for Drawing Lewis Structures
valence_electrons
4.12: Octet Rule
volume
1.2: What is Matter?
2.11: Units Raised to a Power
2.12: Derived Units
2.7: Density
2.8: Problem Solving and Unit Conversions
volumetric flask
11.4: Solution Dilution (M1V1 = M2V2)
VSEPR
6.1: VSEPR Theory
VSEPR Theory
6: The Shape of Molecules

W

water
13.6: Autodissociation of Water and Kw
wavelength
4: The Modern Model of the Atom
4.2: Wavelength and Frequency Calculations
waves
4: The Modern Model of the Atom
4.1: Electromagnetic Spectrum
weak acid
13.10: Buffers- Solutions that Resist pH Change
weak acids
13.4: Strong and Weak Acids and Bases
weak bases
13.3: Bases- Properties and Examples
13.4: Strong and Weak Acids and Bases
weak electrolyte
13.1: Electrolytes
weighted average
3.10: Atomic Mass - The Average Mass of an Element's Atoms
whole number ratio
9.13: Calculating Empirical Formulas

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S

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    - 2.3: Scientific Notation - Writing Large and Small Numbers - *CK-12 License*
    - 2.4: Significant Figures in Calculations - *CK-12 License*
    - 2.5: The Basic Units of Measurement - *CK-12 License*
    - 2.6: Temperature - Random Motion of Molecules and Atoms - *CK-12 License*
    - 2.7: Density - *CK-12 License*
    - 2.8: Problem Solving and Unit Conversions - *CC BY-NC-SA 3.0*
    - 2.9: Metric Unit Conversions - *CK-12 License*
    - 2.10: Solving Multi-step Conversion Problems - *CC BY-NC-SA 3.0*
    - 2.11: Units Raised to a Power - *CK-12 License*
    - 2.12: Derived Units - *CK-12 License*
  - 3: The Evolution of Atomic Theory - *Public Domain*
    - 3.1: Cutting Aluminum until you get Atoms - *CK-12 License*
    - 3.2: Indivisible - The Atomic Theory - *CK-12 License*
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    - 3.4: Law of Definite Proportions - *CK-12 License*
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- 9.4: Conversions Between Moles and Mass - *CK-12 License*
- 9.5: Conversions Between Mass and Number of Particles - *CK-12 License*
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    - 1.2: What is Matter? - *CK-12 License*
    - 1.3: Classifying Matter According to Its Composition - *CK-12 License*
    - 1.4: Classifying Matter According to Its State—Solid, Liquid, and Gas - *CK-12 License*
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    - 1.6: Changes in Matter - Physical and Chemical Changes - *CK-12 License*
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    - 2.3: Scientific Notation - Writing Large and Small Numbers - *CK-12 License*
    - 2.4: Significant Figures in Calculations - *CK-12 License*
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    - 3.1: Cutting Aluminum until you get Atoms - *CK-12 License*
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