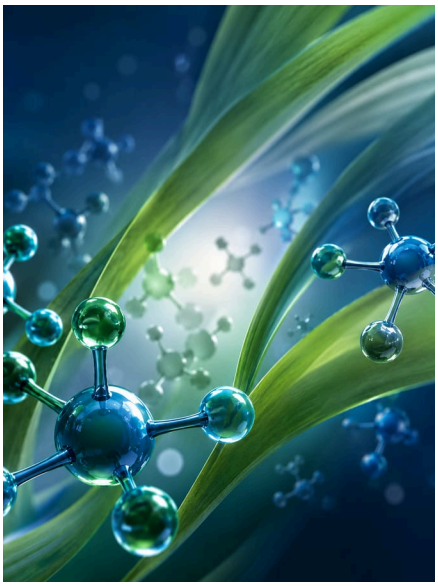
Foundations of General, Organic, and Biological Chemistry: An Integrated Approach

BurÁak Artun, Gorkem Ermut, and Semere Bairu

De Anza College

Chemistry





Foundations of General, Organic, and Biological Chemistry: An Integrated Approach

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Detailed Licensing

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Program Page

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1. **Template check.** The project coordinator and instructional designers checked the first chapter for accessibility as a model to ensure later chapters meet accessibility standards.
2. **Progress checks.** The manuscript was checked at approximately 40%, 60%, and 80% completion to catch and correct accessibility issues in production.
3. **Specialist review.** Once the manuscript was complete, an accessibility expert conducted an accessibility check.
4. **Remediation.** The author and the project manager who conducted the earlier progress checks worked together to remediate the issues identified by the accessibility expert.
5. **Final check.** The resource underwent another full check by the accessibility expert before the book was deemed ready for publication.

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This statement was last updated August 2026

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About the Book

Philosophy and Approach

Foundations of General, Organic, and Biological Chemistry: An Integrated Approach, was developed through a collaborative faculty effort to provide a modern, student-centered alternative to costly commercial textbooks. The text emphasizes:

- Clear explanations of foundational chemical principles
- Connections between chemistry and biological systems
- Real-world applications and everyday examples
- Visual learning and molecular representations
- Scientific literacy and critical thinking
- Affordable and accessible education

This textbook was created to make chemistry more accessible, engaging, and affordable for all learners. Chemistry is often called the *central science* because it connects physics, biology, medicine, environmental science, and countless aspects of everyday life. Our goal is to help students see chemistry not as a collection of isolated facts, but as a powerful way of understanding the world.

Designed for introductory chemistry courses, this text integrates concepts from general, organic, and biological chemistry with real-world applications relevant to health sciences, life sciences, and modern society. Emphasis is placed on conceptual understanding, problem-solving, scientific inquiry, and the molecular foundations of life.

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Online Homework

A ready-to-use set of **ADAPT** homework assignments aligned with this textbook is available. After logging into your instructor account, select **Import Course** and choose **Foundations of General, Organic, and Biological Chemistry: An Integrated Approach**.

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About the Authors

Information About the Authors

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	**Burcak Artun, Ph.D.,** is a continuing lecturer in the Department of Microbiology and Environmental Toxicology at the University of California, Santa Cruz. She also teaches chemistry at Cabrillo College and De Anza College. She earned her Ph.D. in Chemistry from Brown University and completed postdoctoral research at the Massachusetts Institute of Technology. Dr. Artun has experience teaching general chemistry, organic chemistry, and biochemistry for allied health and science majors in both in-person and online formats. In this book, she authored Chapters 1, 2, 8, and 11. She is committed to supporting students from a wide range of backgrounds and helping them build confidence in science. She also advocates for Open Educational Resources (OER) and contributes to the development of affordable, equitable learning materials. Before teaching, Dr. Artun conducted research in genetic medicine and environmental health. Outside of work, she enjoys traveling and spending time with her family.
	**Gorkem Ermut** is a chemistry educator who is passionate about making chemistry logical, approachable, and relevant. She teaches college-level general chemistry and develops learning materials that simplify complex topics through clear explanations and visual supports. She earned a B.S. in Chemistry from Yildiz Technical University and an M.S. in Chemistry from Istanbul University in Türkiye. She also completed doctoral coursework in medicinal chemistry and conducted advanced study in that field. Her research on the synthesis and antibacterial activity of novel heterocyclic compounds has been published in international journals. In this book, she authored Chapters 5, 6, 7, and 10. These chapters focus on writing and balancing equations, reaction types, stoichiometry, and solutions. Her goal is to help students build curiosity, confidence, and a deeper understanding of chemistry.

Author Images	Author Bios
	**Dr. Semere Bairu** is a part-time instructor at De Anza College with 10 years of teaching experience in higher education. He holds degrees from the University of Asmara, the University of Pretoria, and Western Michigan University. He teaches General Chemistry, Organic Chemistry, GOB, and Biochemistry, including courses for pre-med and post-baccalaureate students. His academic interests include metabolism and organ biochemistry. As a LibreTexts contributor, he develops and adapts content for Chapters 3, 4, 9, and 12. Dr. Bairu cares deeply about the power of education to change lives and hopes to share that legacy with future generations.

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1: CHEMISTRY, MATTER, MEASUREMENTS

Introduction

The study of chemistry will open your eyes to a fascinating world. Chemical processes are continuously at work all around us. They happen as you cook and eat food, strike a match, shampoo your hair, and even read this page. Chemistry is called the central science because a knowledge of chemical principles is essential for other sciences. You might be surprised at the extent to which chemistry pervades your life.

Learning Objectives

- Define Chemistry in relation to other sciences
- Identify steps in the scientific method
- Use physical and chemical properties to describe matter, and identify as element, compound or mixture
- Express a large or a small number in scientific notation
- Identify the number of significant figures in a reported value and use significant figures correctly in arithmetical operations
- Express quantities properly as number and unit, recognizing the SI base units with the system of prefixes used with them
- Convert a value reported in one unit to a corresponding value in a different unit

1.1: Prelude to Chemistry, Matter, and Measurement

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1.1: Prelude to Chemistry, Matter, and Measurement

Introduction

In April 2003, the US Pharmacopeia, a national organization that establishes quality standards for medications, reported a case in which a physician ordered “morphine [a powerful painkiller] 2–3 mg IV [intravenously] every 2–3 hours for pain.” A nurse misread the dose as “23 mg” and thus administered approximately 10 times the proper amount to an 8-year-old boy with a broken leg. The boy stopped breathing but was successfully resuscitated and left the hospital three days later.

Quantities and measurements are as important in our everyday lives as they are in medicine. The posted speed limits on roads and highways, such as 55 miles per hour (mph), are quantities we might encounter all the time. Both parts of a quantity, the amount (55) and the unit (mph), must be properly communicated to prevent potential problems. In chemistry, as in any technical endeavor, the proper expression of quantities is a necessary fundamental skill. As we begin our journey into chemistry, we will learn this skill so that errors—from homework mistakes to traffic tickets to more serious consequences—can be avoided.

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1.2: What is Chemistry

What is Chemistry?

The study of chemistry will open your eyes to a fascinating world. Chemical processes are continuously at work all around us. They happen as you cook and eat food, strike a match, shampoo your hair, and even read this page. Chemistry is called the central science because a knowledge of chemical principles is essential for other sciences. You might be surprised at the extent to which chemistry pervades your life.

Chemistry is the study of matter and the changes that material substances undergo, what it consists of, what its properties are. Being able to describe the ingredients in a cake and how they change when the cake is baked is called chemistry.

Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter.

Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations.

Because the physical universe is so vast, there are many different branches of science. Chemistry is one branch of science. For example, chemistry is the study of matter, whereas biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry. Figure 1.2.1 illustrates the relation of chemistry with different branches of science.

> Of all the scientific disciplines, Chemistry is perhaps the most extensively connected to other fields of study.

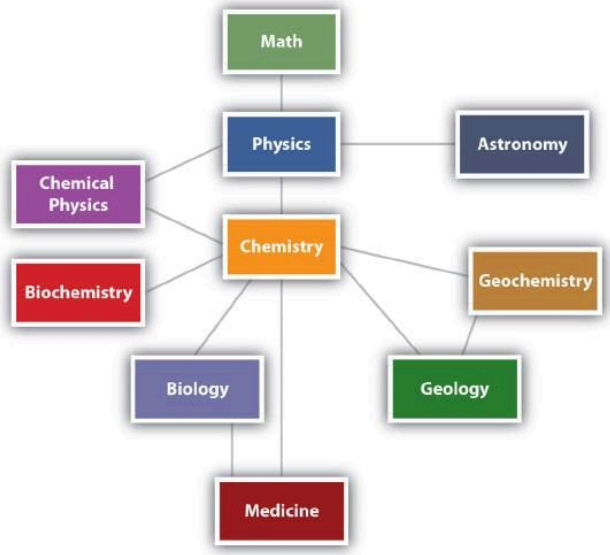


Figure 1.2.1: Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 1.2.1.

Geologists who want to locate new mineral or oil deposits use chemical techniques to analyze and identify rock samples. Oceanographers use chemistry to track ocean currents, determine the flux of nutrients into the sea, and measure the rate of exchange of nutrients between ocean layers. Engineers need to understand the chemical properties of the substances when designing biologically compatible implants for joint replacements or designing roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Physicists take advantage of the properties

of substances to detect new subatomic particles. Astronomers use chemical signatures to determine the age and distance of stars and thus answer questions about how stars form and how old the universe is. The entire subject of environmental science depends on chemistry to explain the origin and impacts of phenomena such as air pollution, ozone layer depletion, and global warming.

The disciplines that focus on living organisms and their interactions with the physical world rely heavily on biochemistry, the application of chemistry to the study of biological processes. A living cell contains a large collection of complex molecules that carry out thousands of chemical reactions, including those that are necessary for the cell to reproduce. Biological phenomena such as vision, taste, smell, and movement result from numerous chemical reactions. Fields such as medicine, pharmacology, nutrition, and toxicology focus specifically on how the chemical substances that enter our bodies interact with the chemical components of the body to maintain our health and well-being. For example, in the specialized area of sports medicine, a knowledge of chemistry is needed to understand why muscles get sore after exercise as well as how prolonged exercise produces the euphoric feeling known as “runner’s high.”

Examples of the practical applications of chemistry are everywhere (Figure 1.2.2). Archaeology 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.

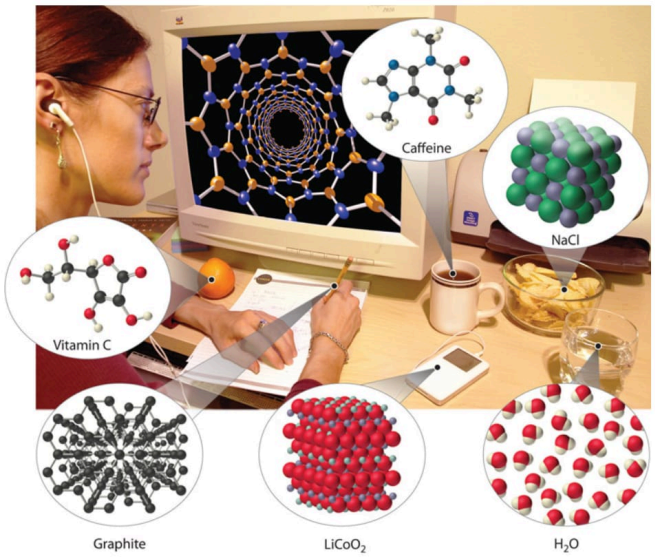


Figure 1.2.2: Although most people do not recognize it, chemistry and chemical compounds are crucial ingredients in almost everything we eat, wear, and use. (CC BY-NC-SA 3.0; Kristin Piljay via General Chemistry: Principles, Patterns, and Applications). Detailed Description of Figure 1.2.2.

Link to Learning: Alchemy - Understanding History of Chemistry as the Central Science

As our understanding of the universe has changed over time, so has the practice of science. Chemistry in its modern form, based on principles that we consider valid today, was developed in the 1600s and 1700s. Before that, the study of matter was known as alchemy and was practiced mainly in China, Arabia, Egypt, and Europe.

Alchemy was a somewhat mystical and secretive approach to learning how to manipulate matter. Practitioners, called alchemists, thought that all matter was composed of different proportions of the four basic elements—fire, water, earth, and air—and believed that if you changed the relative proportions of these elements in a substance, you could change the substance. The long-standing attempts to “transmute” common metals into gold represented one goal of alchemy. Alchemy’s other major goal was to synthesize the philosopher’s stone, a material that could impart long life—even immortality.

Alchemists used symbols to represent substances, some of which are shown in the accompanying figure below, Figure 1.2.3. This was not done to better communicate ideas, as chemists do today, but to maintain the secrecy of alchemical knowledge, keeping others from sharing in it.

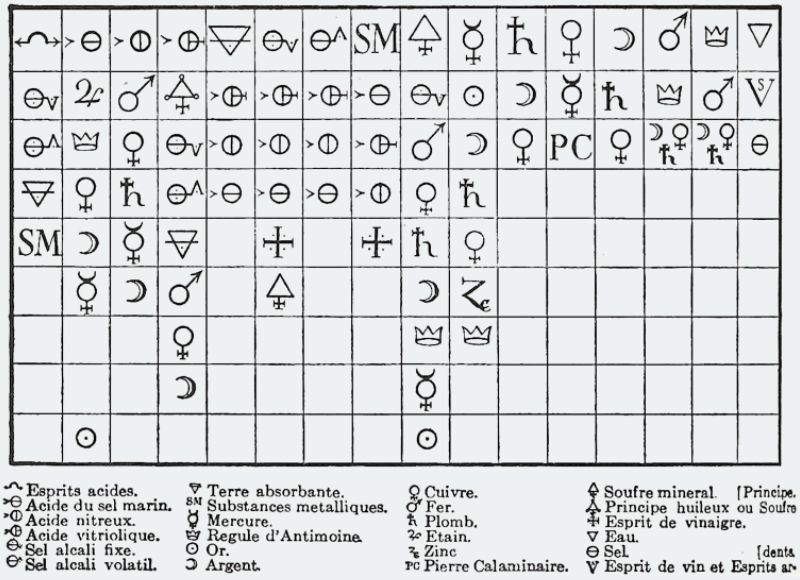


Figure 1.2.3: Table of different relations observed in chemistry between different substances. Alchemists used symbols like these to represent substances. (Public domain, Étienne-François Geoffroy via Wikimedia). Detailed Description of Figure 1.2.3.

In spite of this secrecy, in its time alchemy was respected as a serious, scholarly endeavor. Isaac Newton, the great mathematician and physicist, was also an alchemist.

We will also consider ways in which chemistry affects our daily lives: the addition of iodine to table salt; the development of more effective drugs to treat diseases such as cancer, AIDS (acquired immunodeficiency syndrome), and arthritis; the retooling of industry to use nonchlorine-containing refrigerants, propellants, and other chemicals to preserve Earth's ozone layer; the use of modern materials in engineering; current efforts to control the problems of acid rain and global warming; and the awareness that our bodies require small amounts of some chemical substances that are toxic when ingested in larger doses. By the time you finish this text, you will be able to discuss these kinds of topics knowledgeably, either as a beginning scientist who intends to spend your career studying such problems or as an informed observer who is able to participate in public debates that will certainly arise as society grapples with scientific issues.

The Scientific Method

How do scientists work? Generally, they follow a process called the **scientific method.** The scientific method is an organized procedure for learning answers to questions and making explanations for observations. Figure 1.2.4 illustrates the steps a scientist must take in the process for the scientific method.

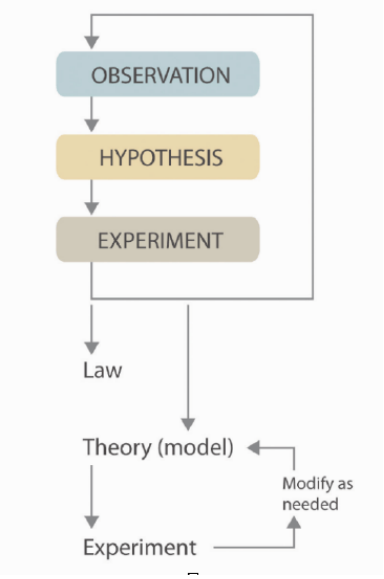


Figure 1.2.4: The Scientific Method Flowchart. As depicted in this flowchart, the scientific method consists of making observations, formulating hypotheses, and designing experiments. A scientist may enter the cycle at any point. (CC BY-NC-SA 3.0; anonymous via General Chemistry: Principals, Patterns, Applications). Detailed Description of Figure 1.2.4.

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 degrees Celsius, and 35.9 grams of table salt—whose chemical name is sodium chloride—dissolve in 100 grams of water at 20 degrees Celsius.
2. Propose 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 hypothesis** is 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 of two hypotheses: (1) Earth rotates on its axis every 24 hours, alternately exposing one side to the sun, or (2) the sun revolves around Earth every 24 hours.
3. Test the hypothesis through experimentation. 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. A properly designed and executed experiment enables a scientist to determine whether the original hypothesis is valid. Experiments often demonstrate that the hypothesis is incorrect or that it must be modified. Suitable experiments can be designed to choose between these two alternatives. If the hypothesis passes the test, it may be a proper answer to the question. If the hypothesis does not pass the test, it may not be a good answer.
4. Refine the hypothesis if necessary and test again. Depending on the results of experiments, a scientist may want to modify the hypothesis and then test it again. Sometimes the results show the original hypothesis to be completely wrong, in which case a scientist will have to devise a new hypothesis. Experiments often demonstrate that the hypothesis is incorrect or that it must be modified. 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).
5. Come up with a theory. Once sufficiently reproducible data is collected, one or more well established hypothesis may form the basis for a scientific theory.

Link to Learning: Law vs Theory

Some people use the word theory in place of hypothesis, but the word hypothesis is the proper word in science. For scientific applications, the word theory is a general statement that describes a large set of observations and data. A theory represents the highest level of scientific understanding.

A **law** is a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply says what happens; it does not address the question of why.

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.

Not all scientific investigations are simple enough to be separated into these discrete steps. But these steps represent the general method by which scientists learn about our natural universe.

Example 1.2.1: Use of Scientific Method

To find the answer to a question (for example, "Why do birds fly toward Earth's equator during the cold months?"), a scientist goes through the following steps, which are also illustrated in Figure 1.2.5.

Solution

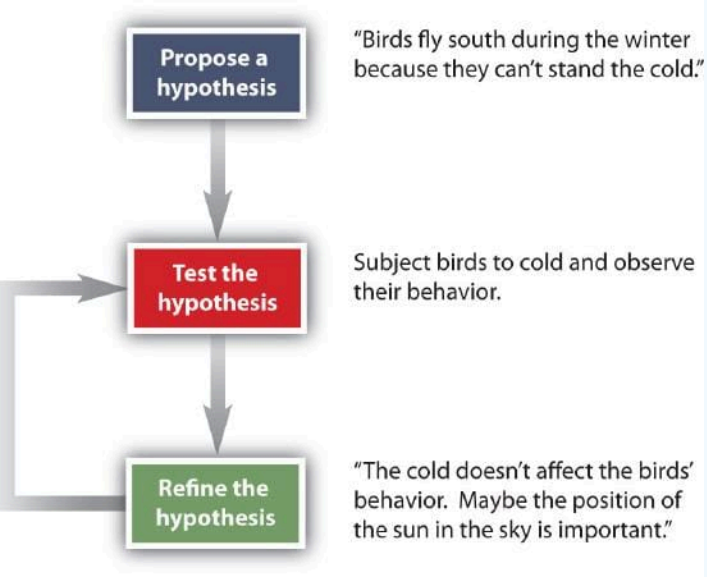


Figure 1.2.5: After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.(CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 1.2.5.

Example 1.2.2: Use of Scientific Method

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

1. Ice always floats on liquid water.
2. Birds evolved from dinosaurs.
3. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
4. 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.
5. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Analysis

Given: components of the scientific method

Asked for: statement classification

Strategy:

Refer to the definitions in this section to determine which category best describes each statement.

Solution:

1. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
2. This is a possible explanation for the origin of birds, so it is a hypothesis.
3. 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.
4. The temperature is measured before and after a change is made in a system, so these are quantitative observations.
5. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

It is important to remember that scientists have a tendency to formulate hypotheses in familiar terms simply because it is difficult to propose something that has never been encountered or imagined before. As a result, scientists sometimes discount or overlook unexpected findings that disagree with the basic assumptions behind the hypothesis or theory being tested.

Fortunately, truly important findings are immediately subject to independent verification by scientists in other laboratories, so science is a self-correcting discipline. This is called **peer review.** Peer review helps make sure that academic and professional work is trustworthy, correct, and meets high standards, building confidence in the field.

Key Takeaways

- Chemistry is the study of matter and how it behaves.
- The scientific method is the general process by which we learn about the natural universe.

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1.3: Classification of Matter

Properties of Matter

Matter is anything that occupies space and has mass. Chemists study the structures, physical properties, and chemical properties of material substances. Part of understanding matter is being able to describe it. One way chemists describe matter is to assign different kinds of properties to different categories. For example, Gold and Iridium are matter, as are peanuts, people, and postage stamps. Smoke, smog, and laughing gas are matter. Energy, light, and sound, however, are not matter; ideas and emotions are also not matter. Here we will learn about how we classify matter and study and differentiate its physical and chemical properties. The properties that chemists use to describe matter fall into two general categories.

Physical Properties

Physical properties are characteristics that describe matter. They include characteristics such as size, shape, color, and mass and density. These characteristics can be observed or measured without changing the *identity* of the matter in question.

Physical properties can be extensive or intensive. **Extensive properties** vary with the amount of the substance and include mass, weight, and volume. **Intensive properties**, in contrast, do not depend on the amount of the substance; they include color, melting point, boiling point, electrical conductivity, and physical state at a given temperature. For example, elemental sulfur is a yellow crystalline solid that does not conduct electricity and has a melting point of 115.2°C, no matter what amount is examined Figure 1.3.1 is an illustration showing the difference between extensive and intensive properties of Sulfur. Scientists commonly measure intensive properties to determine a substance's identity, whereas extensive properties convey information about the amount of the substance in a sample.

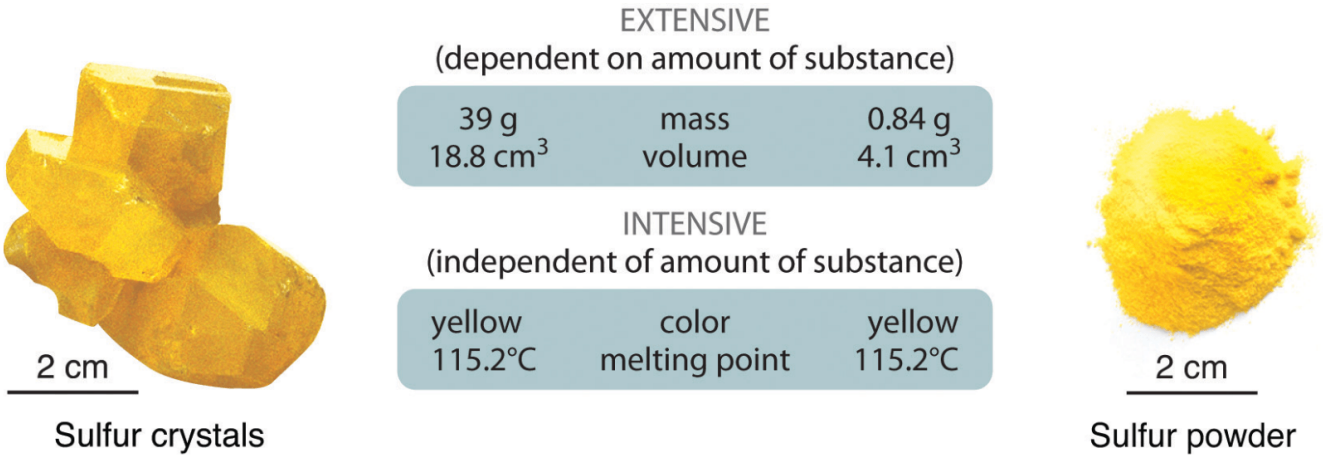


Figure 1.3.1: Because they differ in size, the two samples of sulfur have different extensive properties, such as mass and volume. In contrast, their intensive properties, including color, melting point, and electrical conductivity, are identical. (CC BY-NC-SA 3.0; Dorling Kindersley via General Chemistry: Principles, Patterns, and Applications). Detailed Description of Figure 1.3.1.

The **mass** of an object is the quantity of matter it contains. Do not confuse an object's mass with its **weight**, which is a force caused by the gravitational attraction that operates on the object. Mass is a fundamental property of an object that does not depend on its location.In physical terms, the mass of an object is directly proportional to the force required to change its speed or direction. Weight, on the other hand, depends on the location of an object. An astronaut whose mass is 95 kg weighs about 210 lb on Earth but only about 35 lb on the moon because the gravitational force he or she experiences on the moon is approximately one-sixth the force experienced on Earth. For practical purposes, weight and mass are often used interchangeably in laboratories. Because the force of gravity is considered to be the same everywhere on Earth's surface, 2.2 lb (a weight) equals 1.0 kg (a mass), regardless of the location of the laboratory on Earth.

Although mass and volume are both extensive properties, their ratio is an important intensive property called density (d).

Density (d) is defined as mass per unit volume and is usually expressed in grams per cubic centimeter, g/cm^3 (or grams per milliliter, g/mL).

As mass increases in a given volume, density also increases. For example, lead, with its greater mass, has a far greater density than the same volume of air, just as a brick has a greater density than the same volume of Styrofoam. At a given temperature and pressure, the density of a pure substance is a constant (Eqn 1.3.1):

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{v} \tag{1.3.1}$$

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 1.3.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float." Example 1.3.1 shows how density measurements can be used to identify pure substances.

Table 1.3.1 Densities of Common Substances

Substance	Density at 25°C (g/cm^3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.42

Example 1.3.1: Calculation of Density of Liquids through Experimentation

Imagine you have five bottles containing colorless liquids (labeled A–E).

Masses of samples: A: 17.72 g; B: 19.75 g; C: 24.91 g; D: 19.65 g; E: 27.80 g

You must identify them by measuring the density of each.

1. Using a pipette, a laboratory instrument for accurately measuring and transferring liquids, you carefully measure 25.00 mL of each liquid into five beakers of known mass (1 mL = 1 cm^3). You then weigh each sample on a laboratory balance.
2. Calculate the density of each sample. Based solely on your results, can you unambiguously identify all five liquids? If necessary, review the use of significant figures in calculations (see section 1.4) prior to working this example.
3. Use the tabulated data to identify the liquids.

Table 1.3.2 Densities of Liquids in Example 3

Substance	Density at 25°C (g/cm^3)
water	0.998
methanol (wood alcohol)	0.792
isopropanol (rubbing alcohol)	0.785
ethylene glycol (used in antifreeze)	1.113

Substance	Density at 25°C (g/cm^3)
ethanol (the alcohol in beverages)	0.789
diethyl ether (“ether”; once widely used as an anesthetic)	0.708

Analysis

Given: volume and mass.

Asked for: density.

Strategy:

1. Calculate the density of each liquid from the volumes and masses given.
2. Check to make sure that your answer makes sense.
3. Compare each calculated density with those given in Table 1.3.2 "Densities of Liquids in Example 3". If the calculated density of a liquid is not significantly different from that of one of the liquids given in the table, then the unknown liquid is most likely the corresponding liquid.
4. If none of the reported densities corresponds to the calculated density, then the liquid cannot be unambiguously identified.

Solution

1. Density is mass per unit volume and is usually reported in grams per cubic centimeter (or grams per milliliter because 1 mL = 1 cm^3). The masses of the samples are given in grams, and the volume of all the samples is 25.00 mL (= 25.00 cm^3). The density of each sample is calculated by dividing the mass by its volume See Equation 1.3.1 .

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{v}$$

17.72 g / 25.00 cm^3 = 0.7088 g/cm^3

Both the volume and the mass are given to four significant figures, so four significant figures are permitted in the result.

When calculated the same way, the densities of the other samples (in grams per cubic centimeter) are as follows:

B: 0.7900; C: 0.9964; D: 0.7860; and E: 1.112

2. Except for sample *E*, the calculated densities are slightly less than 1 g/cm^3. This makes sense because the masses (in grams) of samples *A*–*D* are all slightly less than the volume of the samples, 25.00 mL. In contrast, the mass of sample *E* is slightly greater than 25 g, so its density must be somewhat greater than 1 g/cm^3.

3. Comparing these results with the data given in Table 1.3.2 "Densities of Liquids in Example 3" shows that sample A is probably diethyl ether (0.708 g/cm^3 and 0.7088 g/cm^3 are not substantially different), sample C is probably water (0.998 g/cm^3 in the table versus 0.9964 g/cm^3 measured), and sample E is probably ethylene glycol (1.113 g/cm^3 in the table versus 1.112 g/cm^3 measured).

4. Samples B and D are more difficult to identify for two reasons: (1) Both have similar densities (0.7900 and 0.7860 g/cm^3), so they may or may not be chemically identical. (2) Within experimental error, the measured densities of B and D are indistinguishable from the densities of ethanol (0.789 g/cm^3), methanol (0.792 g/cm^3), and isopropanol (0.785 g/cm^3). Thus some property other than density must be used to identify each sample.

Chemical Properties

Chemical properties are characteristics that describe how matter changes its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. The observation of chemical properties involves a *chemical change* of the matter in question, resulting in matter with a different *identity* and different physical and chemical properties. *The conversion of two or more elements, such as sodium and chlorine, to a chemical compound, sodium chloride, is an example of a chemical change, often called a* ***chemical reaction.***

Figure 1.3.2 illustrates the difference between physical and chemical changes beautifully. When ice melts into liquid water (H_2O), it appears changed; however, this change is only physical as the the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (right) Burning of wax to generate water and carbon dioxide is a chemical reaction.(left) Together, the two images contrast physical and chemical changes, showcasing the differences between reversible changes, like melting, and irreversible transformations, like combustion.



Figure 1.3.2: (left) Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Anonymous, Introduction to Chemistry: General, Organic, and Biological). Detailed Description of Figure 1.3.2.

Substances and Mixtures

Substances

A **substance** is any matter that has the *same physical and chemical properties throughout* the sample. A *pure chemical substance* will have a fixed chemical composition and characteristic properties.

> 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). Oxygen, for example, is a **pure substance** that is a colorless, odorless gas at 25°C.

Elements and Compounds

There are two types of substances, elements and compounds.

An element is a substance that cannot be broken down into chemically simpler components. Aluminum, which is used in soda cans and is represented by the symbol Al, is an element.

A compound is a substance that can be broken down into chemically simpler components (because it consists of more than one element). Water is a compound composed of the elements hydrogen and oxygen and is described by the chemical formula, H_2O.

Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date. Sometimes the word *pure* is used to describe a substance, but this is not absolutely necessary. By definition, any single substance, element or compound is *pure*.

The smallest part of an element that maintains the identity of that element is called an **atom**. Atoms are extremely tiny; to make a line of iron atoms that is 1 inch long, you would need approximately 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a **molecule**. Molecules are composed of two or more atoms that are attached together and behave as a unit. Figure 1.3.3 illustrates the difference between elements and compounds as illustrated by the process of electrolysis, in which water, a compound, is broken down to its elements that make it up: hydrogen and oxygen.

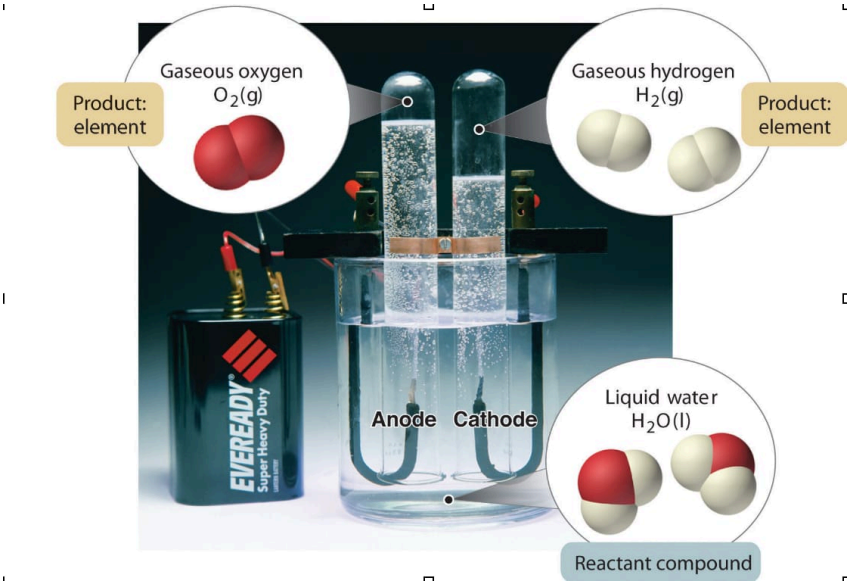


Figure 1.3.3: Water is decomposed to the elements hydrogen and oxygen by *Electrolysis*. Water is a chemical compound; hydrogen and oxygen are elements. (CC BY-NC-SA 3.0; Charles D. Winters/Photo Researchers via General Chemistry: Principles, Patterns, and Applications). Detailed Description of Figure 1.3.3.

Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the **macroscopic viewpoint** of the universe. However, scientists can also describe chemical events on the level of individual atoms or molecules, which is referred to as the **microscopic viewpoint,** Figure 1.3.4. We will see examples of both macroscopic and microscopic viewpoints throughout this book.

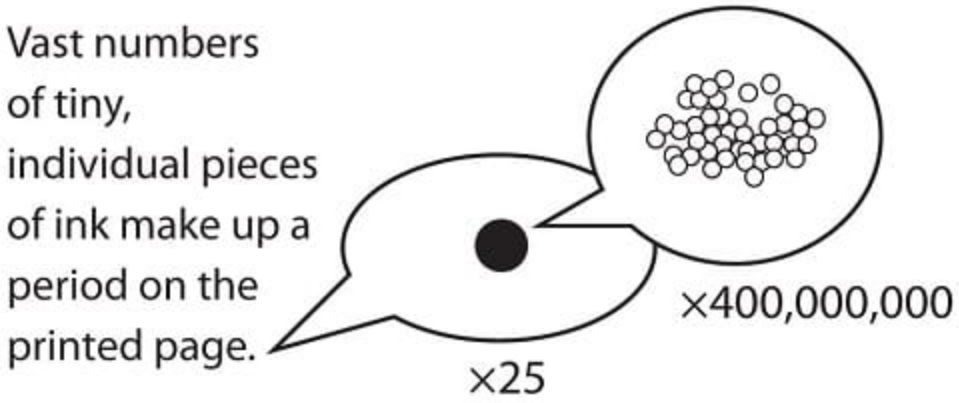


Figure 1.3.4: How Many Particles Are Needed for a Period in a Sentence? Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The times 25 and times 400,000,000 indicate the number of times the image is magnified. (CC-SA-BY-3.0; Andrikkos, Introduction to Chemistry: General, Organic, and Biological). Detailed Description of Figure 1.3.4.

Mixtures

A **mixture** is a material which is composed of the combination of two or more substances. Very few samples of matter consist of pure substances.

> In a mixture, the individual substances maintain their chemical identities.

For example, air, tap water, milk, blue cheese, bread, and dirt are all mixtures. The type of mixture depends on whether they have visible boundaries and/or if they have a uniform composition.

Homogeneous Mixtures

Homogeneous mixtures are mixtures with a consistent or uniform composition throughout.

In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Examples of homogeneous mixtures are the air we breathe and the tap water we drink. Homogeneous mixtures are also called solutions. Thus air is a solution of nitrogen, oxygen, water vapor, carbon dioxide, and several other gases; tap water is a solution of small amounts of several substances in water.

The specific compositions of both of these solutions are not fixed, however, but depend on both source and location; for example, the composition of tap water in Boise, Idaho, is *not* the same as the composition of tap water in Buffalo, New York. Although most solutions we encounter are liquid, solutions can also be solid. The gray substance still used by some dentists to fill tooth cavities is a complex solid solution that contains 50% mercury and 50% of a powder that contains mostly silver, tin, and copper, with small amounts of zinc and mercury. Solid solutions of two or more metals are commonly called **alloys.**

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.

Heterogenous Mixtures

Heterogeneous mixtures are obvious combinations of two or more substances, such as a mixture of sand and water.

Mixtures that appear to be homogeneous are often found to be heterogeneous after microscopic examination. Milk, for example, appears to be homogeneous, but when examined under a microscope, it clearly consists of tiny globules of fat and protein

dispersed in water. The components of heterogeneous mixtures can usually be separated by simple means. Solid-liquid mixtures such as sand in water or tea leaves in tea are readily separated by *filtration*, which consists of passing the mixture through a barrier, such as a strainer, with holes or pores that are smaller than the solid particles. More complex methods may be necessary when separating intricate mixtures.

The overall organization of matter and the methods used to separate mixtures are summarized in Figure 1.3.5.

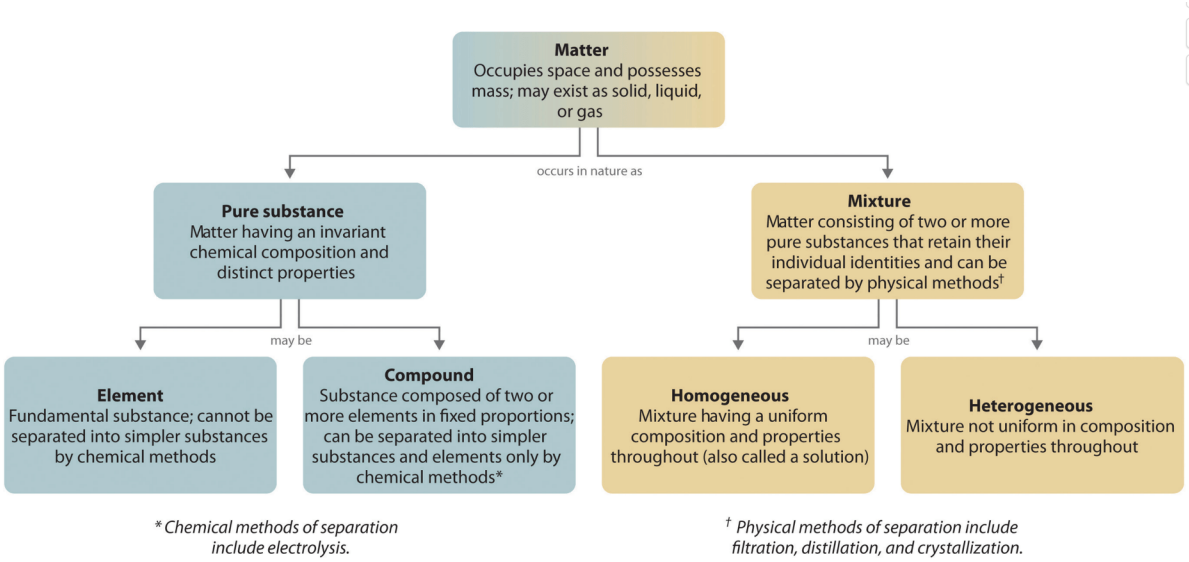


Figure 1.3.5: Relationships between the types of matter and the methods used to separate mixtures is presented in the form of a flowchart. (CC BY-NC-SA 3.0; anonymous via General Chemistry: Principles, Patterns, and Applications). Detailed Description of Figure 1.3.5.

Example 1.3.2: Classification of Matter

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

1. filtered tea
2. freshly squeezed orange juice
3. a compact disc
4. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
5. selenium

Analysis

Given: a chemical substance

Asked for: its classification

Strategy:

- 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.

- 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.

Filtered Tea

- Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
- Because the composition of the solution is uniform throughout, it is a homogeneous mixture.

Orange Juice

- Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
- Because its composition is not uniform throughout, orange juice is a heterogeneous mixture.

Compact Disc

- 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.
- The regions of different composition indicate that a compact disc is a heterogeneous mixture.

Aluminium oxide

- Aluminum oxide is a single, chemically pure compound.

Selenium

- Selenium is one of the known elements.

Example 1.3.3: Classification of Matter

How would a chemist categorize each example of matter?

1. saltwater
2. soil
3. water
4. oxygen

Solution

1. Saltwater is a *homogeneous mixture*, or a solution. Saltwater acts as if it were a single substance even though it contains two substances—salt and water.
2. Soil is a *heterogeneous mixture*. It is composed of small pieces of a variety of materials.
3. Water is a *substance;* more specifically, because water is composed of hydrogen and oxygen, it is a *compound*.
4. Oxygen, a *substance*, is an *element*.

Answer available at the end of this section.

Phases or Physical States of Matter

Under normal conditions all matter can be further classified by one of three **physical states** or **phases**: solid, liquid or gas. These three descriptions each imply that the matter has certain physical properties when in these states.

- **Solids** are relatively rigid and have fixed shapes and volumes. A rock, for example, is a solid.
- **Liquids,** in contrast, have fixed volumes but flow to assume the shape of their containers, such as a beverage in a can.
- **Gases**, such as air in an automobile tire, have neither fixed shapes nor fixed volumes and expand to completely fill their containers. Whereas the volume of gases strongly depends on their temperature and pressure (the amount of force exerted on a given area), the volumes of liquids and solids are virtually independent of temperature and pressure.

Figure 1.3.6 is an illustration showing these states.

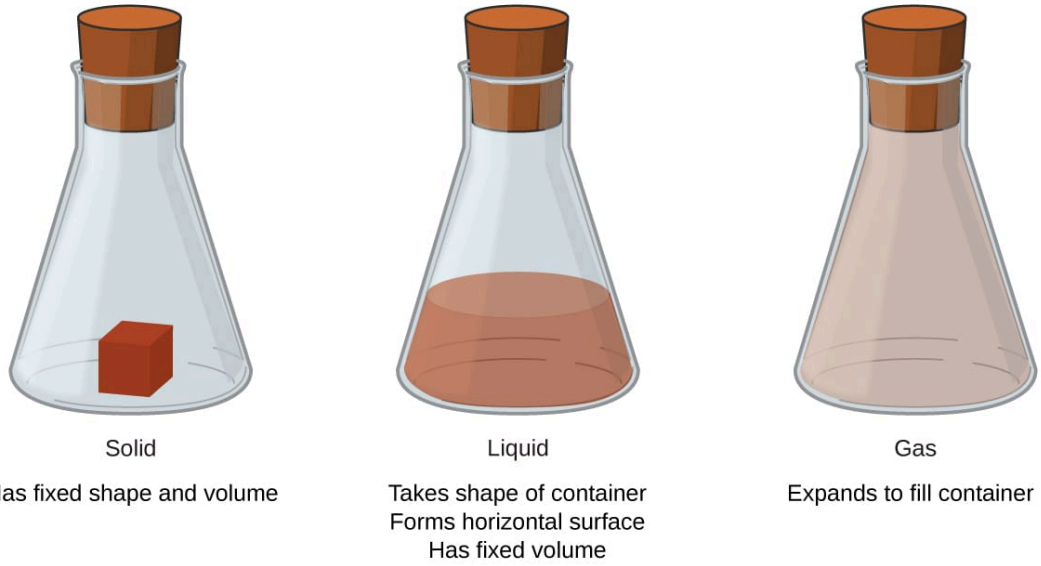


Figure 1.3.6: The three most common states or phases of matter are solid, liquid, and gas. (CC BY 4.0; anonymous, Openstax). Detailed Description of Figure 1.3.6.

We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas).



Figure 1.3.7: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. (CC BY-SA 3.0 ; Markus Schweiss via Wikimedia Commons).

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a **phase change**, a physical process in which a substance changes from one phase to another. Figure 1.3.7 show one such change: water going from liquid to gas phase. Phase changes are identified by particular names depending on what phases are involved, as summarized in Table 1.3.2.

Table 1.3.2: Phase Changes

Change	Name
solid to liquid	**melting, fusion**

Change	Name
solid to gas	**sublimation**
liquid to gas	**boiling, evaporation**
liquid to solid	**solidification, freezing**
gas to liquid	**condensation**
gas to solid	**deposition**

Figure 1.3.8 illustrates the relationships between the different ways matter can be classified.

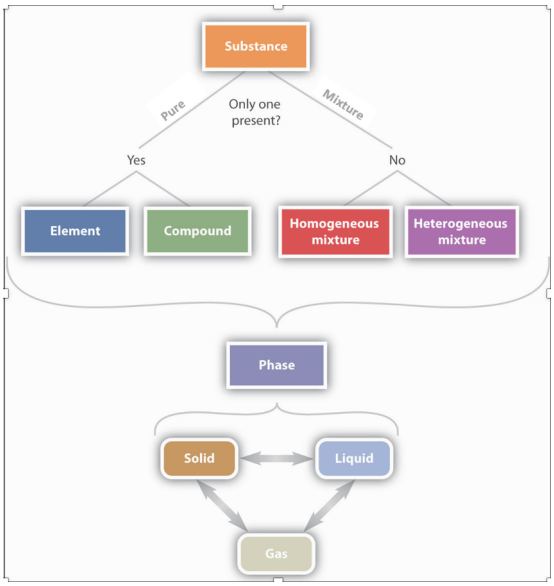


Figure 1.3.8: Matter can be classified in a variety of ways, depending on its properties. This table starts with a substance. If there is only one present, it can either be an element or a compound. If there is more than one present, it can either be a homogeneous mixture or a heterogeneous mixture. From these choices, the substance can be in a certain phase. It can either be solid, liquid, or gas, and each phase can go from one to another. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 1.3.8.

Key Takeaways

- Matter can be described with both physical properties and chemical properties.
- Matter can be identified as an element, a compound, or a mixture.

Answers

Example 1.3.3: Classification of Matter

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1.4: Measurements

Introduction

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.

Measurements

A coffee maker’s instructions tell you to fill the coffeepot with 4 cups of water and 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 1.4.1); the nurse is also measuring.



Figure 1.4.1: Blood pressure is measured using an inflatable cuff, pressure gauge, and tubing to assess arterial pressure. (GNU Free Documentation License; Pia von Lützau via Wikimedia).

Chemists measure the properties of matter using a variety of devices or measuring tools, many of which are similar to those used in everyday life. Rulers are used to measure length, balances (scales) are used to measure mass (weight), and graduated cylinders or pipettes are used to measure volume. Measurements made using these devices are expressed 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 “5.2 kilometers,” we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5.2.

$$\underbrace{5.2}_{\text{number}}\ \underbrace{\text{kilometers}}_{\text{unit}}$$

If you ask a friend how far he or she walks from home to school, and the friend answers “12” without specifying a unit, you do not know whether your friend walks—for example, 12 miles, 12 kilometers, 12 furlongs, or 12 yards.

Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient’s seizures and states a dosage of “100” without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

> 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.

Key Takeaway

Identify a quantity properly with a number and a unit.

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1.5: Expressing Numbers-Scientific Notation

Introduction

The instructions for making a pot of coffee specifies 3 scoops (rather than 12,000 grounds) because any measurement is expressed more efficiently with units that are appropriate in size. In science, however, we often must deal with quantities that are extremely small or incredibly large. For example, you may have 5,000,000,000,000 red blood cells in a liter of blood, and the diameter of an iron atom is 0.000000014 inches. Numbers with many zeros can be cumbersome to work with, so scientists use scientific notation.

Scientific Notation

Scientific notation is a system for expressing very large or very small numbers in a compact manner. It uses the idea that such numbers can be rewritten as a simple number multiplied by 10 raised to a certain exponent, or power.

Let us look first at very large numbers. Suppose a spacecraft is 1,500,000 miles from Mars. The number 1,500,000 can be thought of as follows:

$$1.5 \times \underbrace{1,000,000}_{10^6} = 1.5 \times 10^6 \qquad (1.5.1)$$

$$\underbrace{10 \times 10 \times 10 \times 10 \times 10 \times 10}_{10^6}$$

That is, 1,500,000 is the same as 1.5 times 1 million, and 1 million is 10 × 10 × 10 × 10 × 10 × 10, or 10^6 (which is read as "ten to the sixth power"). Therefore, 1,500,000 can be rewritten as 1.5 times 10^6, or 1.5 × 10^6. See Equation 1.5.1 The distance of the spacecraft from Mars can therefore be expressed as 1.5 × 10^6 miles.

Expressing Numbers Greater than 1 in Scientific Notation

In writing scientific notations, the convention is to have only one digit before the decimal point.

Recall that:

$10^0 = 1$

$10^1 = 10$

$10^2 = 100$

$10^3 = 1,000$

$10^4 = 10,000$ and so forth

As illustrated in Figure 1.5.1, the standard convention for expressing numbers in scientific notation is to

1. Write a single *nonzero* first digit, a decimal point, and the rest of the digits, excluding any trailing zeros (see rules for significant figures in the next section for more details on what to exclude).
2. This number is followed by a multiplication sign and then by 10 raised to the power necessary to reproduce the original number.

$$1{,}500{,}000 = 1.5 \times 10^6$$

Figure 1.5.1: Scientific Notation for a Number Greater Than Zero. For example, although 1,500,000 can also be written as 15. × 10^6 (which would be 15. × 100,000), the convention is to have only one digit before the decimal point. How do we know to what power 10 is raised? The power is the number of places you have to move the decimal point to the *left* to place it after the *first digit*, so that the number being multiplied is between 1 and 10. (CC BY-NC-SA 3.0; Anonymous via Introduction to Chemistry: General, Organic and Biological).

Expressing Numbers Less Than 1 in Scientific Notation

Step 1: We also use scientific notation to express numbers whose magnitudes are less than 1.

For example, the quantity 0.006 centimeters can be expressed as follows:

$$6 \times \underbrace{\frac{1}{1,000}}_{} = 6 \times 10^{-}3$$

$$\underbrace{\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}}_{10^{-3}} \tag{1.5.2}$$

That is, 0.006 centimeters is the same as 6 *divided by* one thousand, which is the same as 6 *divided* by 10 x 10 x 10 or 6 *times* 10^{-3} (which is read as "ten to the negative third power"). See Equation 1.5.2. Therefore, 0.006 centimeters can be rewritten as 6 times 10^{-3}, or 6 × 10^{-3} centimeters.

Recall that:

10^{-1} = 1/10

10^{-2} = 1/100

10^{-3} = 1/1,000

10^{-4} = 1/10,000

10^{-5} = 1/100,000 and so forth

Step 2: We use a negative number as the power to indicate the number of places we have to move the decimal point to the right to make it follow the first nonzero digit so that the number is between 1 and 10.

This is illustrated as follows as demonstrated in (Figure 1.5.2):

$$0.006 = 6 \times 10^{-3}$$

Figure 1.5.2: Scientific Notation for a Number Less Than Zero. We use a negative number as the power to indicate the number of places we have to move the decimal point to the right to make it follow the first nonzero digit so that the number is between 1 and 10. (CC BY-NC-SA 3.0; Anonymous via Introduction to Chemistry: General Organic Biological).

Step 3: As with numbers with positive powers of 10, when changing from scientific notation to standard or conventional format, we reverse the process.

The process is reversed as follows:

$$6 \times 10^{-3} = .006 = 0.006$$

Figure 1.5.3: Changing to Standard Format. (CC BY-NC-SA 3.0; Anonymous via Introduction to Chemistry: General Organic Biological).

Example 1.5.1: Expressing Numbers in Scientific Notation

Express each number in scientific notation.

1. 0.000006567
2. −0.0004004
3. 0.000000000000123

Solution

1. Move the decimal point six places to the right to get 6.567×10^{-6}.
2. Move the decimal point four places to the right to get -4.004×10^{-4}.
3. The negative sign on the number itself does not affect how we apply the rules of scientific notation. Move the decimal point 13 places to the right to get 1.23×10^{-13}.

Converting Numbers to Standard Form

To change a number in scientific notation to standard form, we reverse the process, moving the decimal point to the right.

1. Add zeros to the end of the number being converted, if necessary, to produce a number of the proper magnitude.
2. Drop the number 10 and its power.
3. To convert 1.5×10^6 back to standard form, we move the decimal place same number of places as the power. We add zeros to the end of the number. Keep adding appropriate number of zeros. Then remove the number 10 and the power from the number.

$$1.5 \times 10^6 = 1.500000. = 1{,}500{,}000$$

Figure 1.5.4: Conversion from Scientific Notation to Standard Form. (CC BY-NC-SA 3.0; Anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)).

- If the scientific notation has a positive power, the standard number is greater than one. Example: $8 \times 10^4 = 80{,}000$
- If the scientific notation has a negative power, then the standard number is less than one. Example: $8 \times 10^{-2} = 0.08$

Example 1.5.2: Converting Numbers from Scientific to in Standard Notation-1

Express each number in standard, or conventional notation.

1. 5.27×10^4
2. 1.0008×10^6

Solution

1. Moving the decimal four places to the right and adding zeros give 52,700.
2. Moving the decimal six places to the right and adding zeros give 1,000,800.

Answer available at the end of this section.

Example 1.5.3: Converting Numbers from Scientific to in Standard Notation-2

Change the number in scientific notation to standard form.

1. 6.22×10^{-2}
2. 9.9×10^{-9}

Solution

1. 0.0622
2. 0.0000000099

Link to Learning: Calculators and Scientific Notation

Although calculators can show 8 to 10 digits in their display windows, that is not always enough when working with very large or very small numbers. For this reason, many calculators are designed to handle scientific notation. The method for entering scientific notation differs for each calculator model, so take the time to learn how to do it properly on your calculator, *asking your instructor for assistance if necessary*. Figure 1.5.5 shows one way to enter very large numbers, such as Avogadro's Number, using "E" symbol to express the scientific notation. If you do not learn to enter scientific notation into your calculator properly, you will not get the correct final answer when performing a calculation.

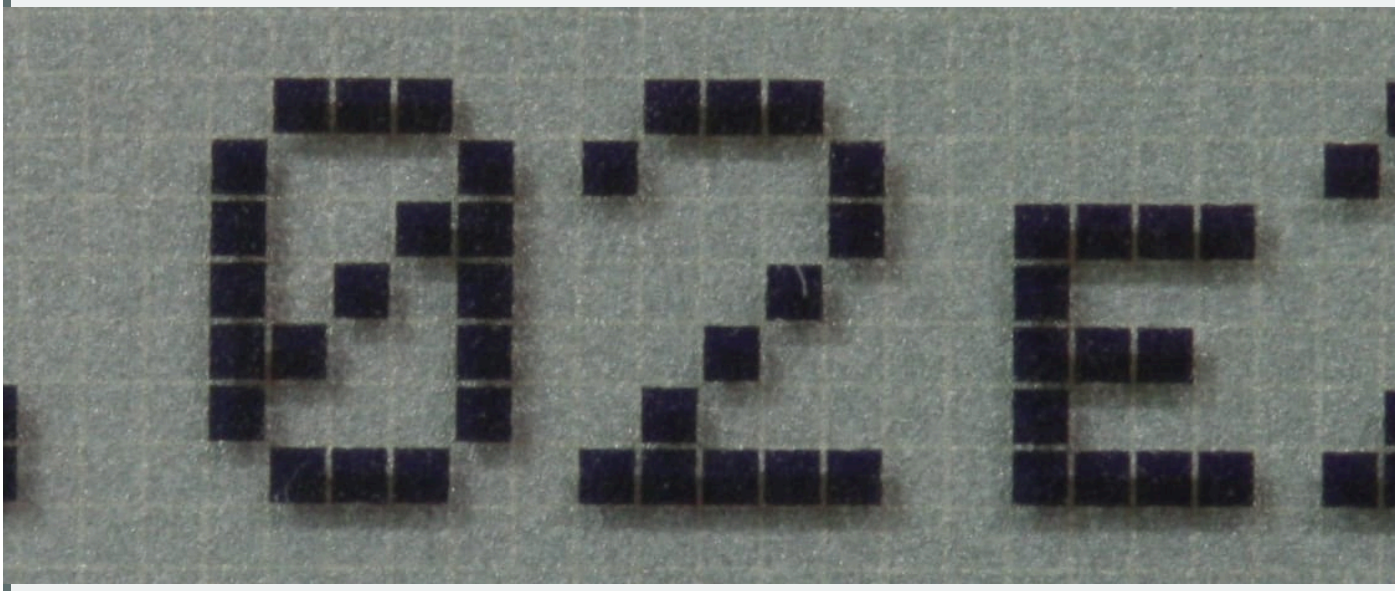


Figure 1.5.5: Figure 1.5.5: **Avogadro's number** in scientific notation on a calculator. In scientific notation, "E" represents "times ten to the power of" so when you see 6.02**E**23 on a calculator, it means 6.02 x 10^{23} . (CC BY-SA 3.0; PRHaney via Wikimedia).

Key Takeaway

Large or small numbers are expressed in scientific notation, which use powers of 10.

Answers

Example 1.5.2: Converting Numbers from Scientific to in Standard Notation-1

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1.6: Expressing Numbers - Significant Figures

Significant Figures in Measurements

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

Quantities derived from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used. **Significant figures** properly report the number of measured and estimated digits in a measurement. There are rules for applying significant figures in calculations. The numbers of measured quantities, unlike defined or directly counted quantities, are not exact.

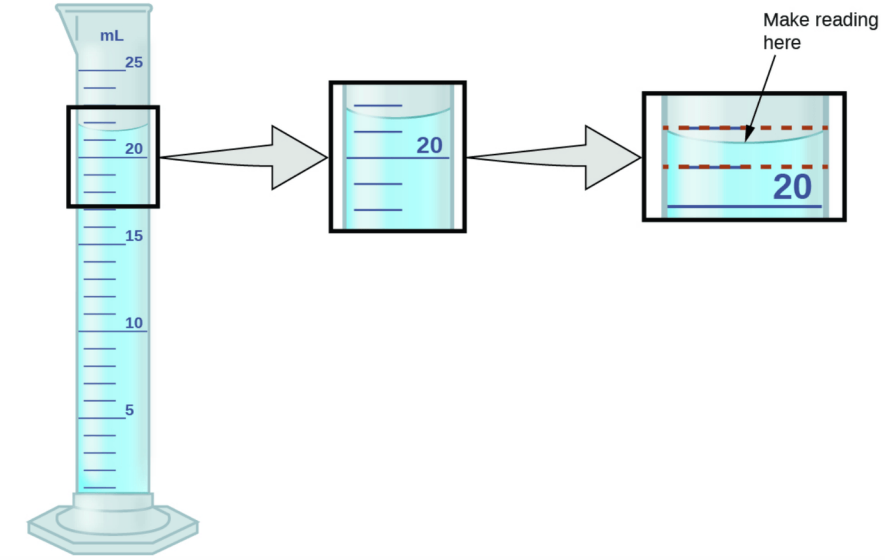


Figure 1.6.1: To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax). Detailed description of Figure 1.6.1.

To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid. In Figure 1.6.1, showing a liquid in a graduated cylinder, the bottom of the meniscus clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL as shown. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain.

In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs *about* 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If the coin is weighed on a more sensitive balance, the mass might be 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram.

> Every measurement has some **uncertainty**, which depends on the device used and the person using it. All of the digits in a measurement, including the uncertain last digit are significant figures.

Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

Reporting Measured Values

A measurement result is properly reported when its significant digits accurately represent the certainty of the measurement process. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Below are the rules to follow when determining if a number is significant or not:

Rule 1: All nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them. This is illustrated in Figure 1.6.2.

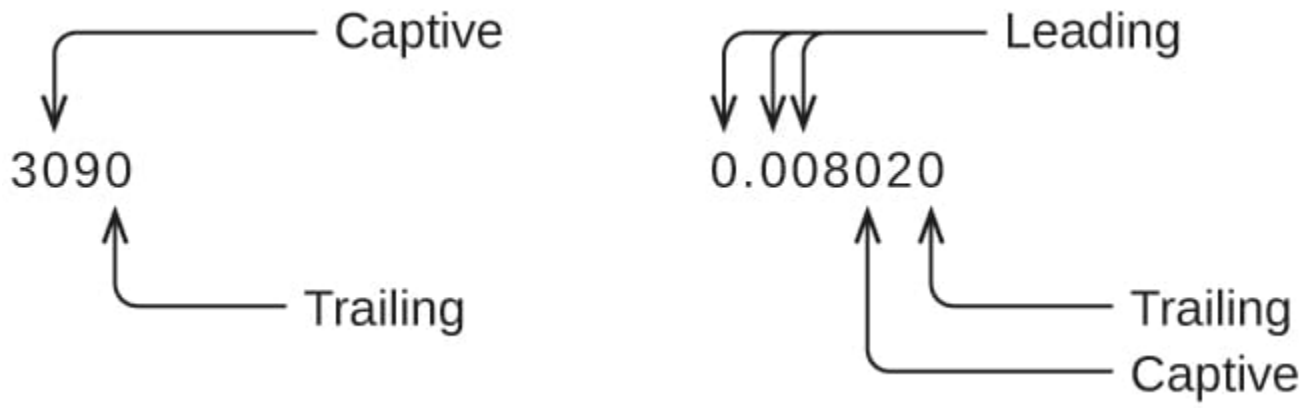


Figure 1.6.2: The left diagram uses the example of 3090. The zero in the hundreds place is labeled "captive" and the zero in the ones place is labeled trailing. The right diagram uses the example 0.008020. The three zeros in the ones, tenths, and hundredths places are labeled "leading." The zero in the ten-thousandths place is labeled "captive" and the zero in the millionths place is labeled "trailing." (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax)

Rule 2: Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point. This is illustrated in Figure 1.6.3.



Figure 1.6.3: The left diagram uses the example of 1267 meters. The number 1 is the first nonzero figure on the left. 1267 has 4 significant figures in total. The right diagram uses the example of 55.0 grams. The number 5 in the tens place is the first nonzero figure on the left. 55.0 has 3 significant figures. Note that the 0 is to the right of the decimal point and therefore is a significant figure. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax).

Rule 3: Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located. The leading zeros in this example are not significant. We could use scientific notation and express the number as 8.32407×10^{-3}; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point. This is illustrated in Figure 1.6.4.

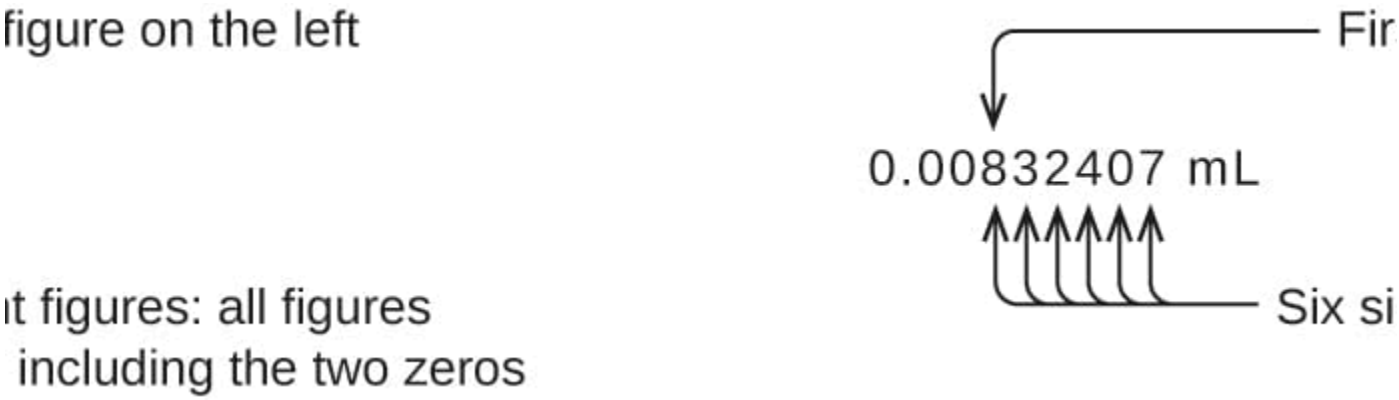


Figure 1.6.4: The left diagram uses the example of 70.607 milliliters. The number 7 is the first nonzero figure on the left. 70.607 has 5 significant figures in total, as all figures are measured including the 2 zeros. The right diagram uses the example of 0.00832407 M L. The number 8 is the first nonzero figure on the left. 0.00832407 has 6 significant figures. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax).

Rule 5: The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant.

The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant. This is illustrated in Figure 1.6.5.

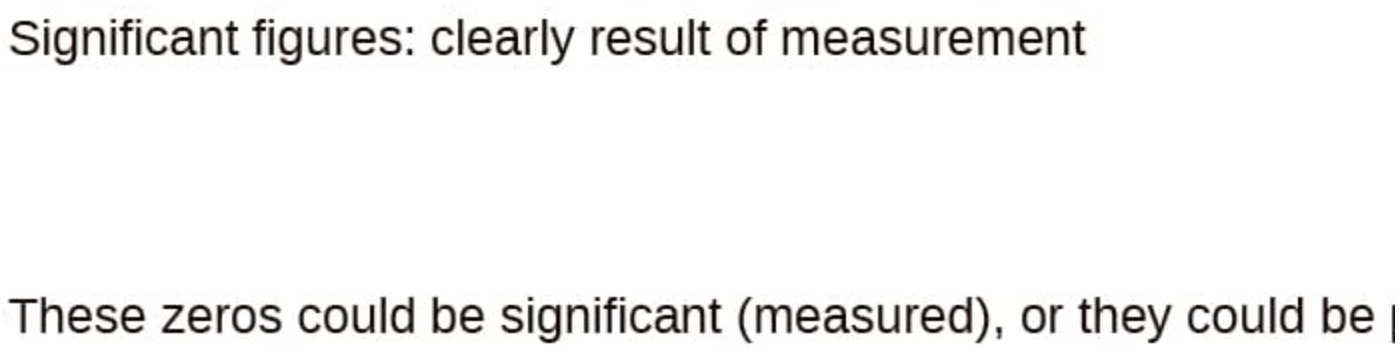


Figure 1.6.5: This figure uses the example of 1300 grams. The one and the 3 are significant figures as they are clearly the result of measurement. The 2 zeros could be significant if they were measured or they could be placeholders. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax).

Significant Figures in Calculations

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

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

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

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

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

Example 1.6.1: Rounding Numbers

Round the following to the indicated number of significant figures:

1. 1.57 (to two significant figures)
2. 8.1649 (to three significant figures)
3. 0.051065 (to four significant figures)
4. 0.90275 (to four significant figures)

Solution

1. 31.57 rounds “up” to 32 (the dropped digit is 5, and the retained digit is even)
2. 8.1649 rounds “down” to 8.16 (the dropped digit, 4, is less than 5)
3. 0.051065 rounds “down” to 0.05106 (the dropped digit is 5, and the retained digit is even)

4. 0.90275 rounds “up” to 0.9028 (the dropped digit is 5, and the retained digit is even)

Addition and Subtraction with Significant Figures

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

Example 1.6.2: Addition and Subtraction with Significant Figures

1. Add 1.0023 g and 4.383 g.
2. Subtract 421.23 g from 486 g.

Solution

1. 1.0023 g + 4.383 g = 5.3853 g

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

2. 486 g – 421.23 g = 64.77 g

Answer is 65 g (round to the ones place; no decimal places) as explained in Figure 1.6.6.

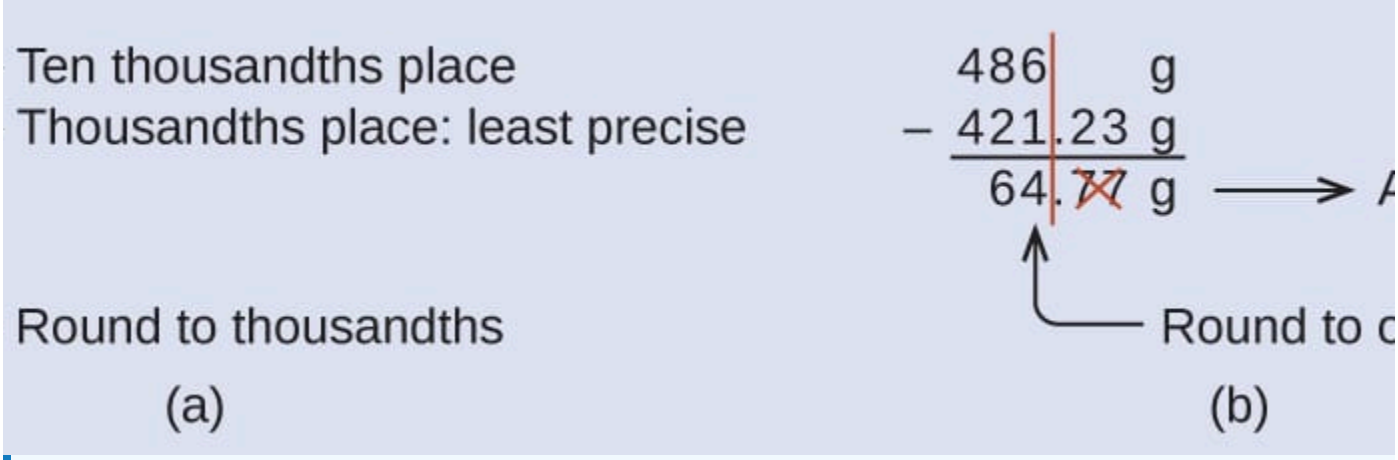


Figure 1.6.6: Figure (a) shows 1.0023 being added to 4.383 to yield the answer 5.385. 1.0023 goes to the ten thousandths place, but 4.383 goes to the thousandths place, making it the less precise of the two numbers. Therefore the answer, 5.3853, should be rounded to the thousandths, to yield 5.385. Figure (b) shows 486 grams minus 421.23 grams, which yields the answer 64.77 grams. This answer should be round to the ones place, making the answer 65 grams. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax).

Multiplication and Division with Significant Figures

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

Example 1.6.3: Multiplication and Division with Significant Figures

1. Multiply 0.6238 cm by 6.6 cm.

2. Divide 421.23 g by 486 mL.

Solution

1. 0.6238 cm × 6.6cm = 4.11708 cm^2 result is 4.1cm^2 (round to two significant figures).

four significant figures × two significant figures results is two significant figures

2. 421.23 g / 486 mL = 0.866728... g/mL result is 0.867 g/mL (round to three significant figures)

five significant figures / three significant figures result is three significant figures

Calculation with Significant Figures

Example 1.6.3: Calculations with Significant Figures

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

Solution

 = × ×

= 13.44 dm × 5.920 dm × 2.54 dm

= 202.09459...dm^3 (value from calculator)

= 202 dm^3, or 202 L (answer rounded to three significant figures)

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to evaluate both the **precision** and the **accuracy** of their results.

Measurements are said to be precise if they yield very similar results when repeated in the same manner.

A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value.

These characterizations can be extended to other contexts, such as the results of an archery competition Figure 1.6.7, which is used to illustrate the difference between precision and accuracy.

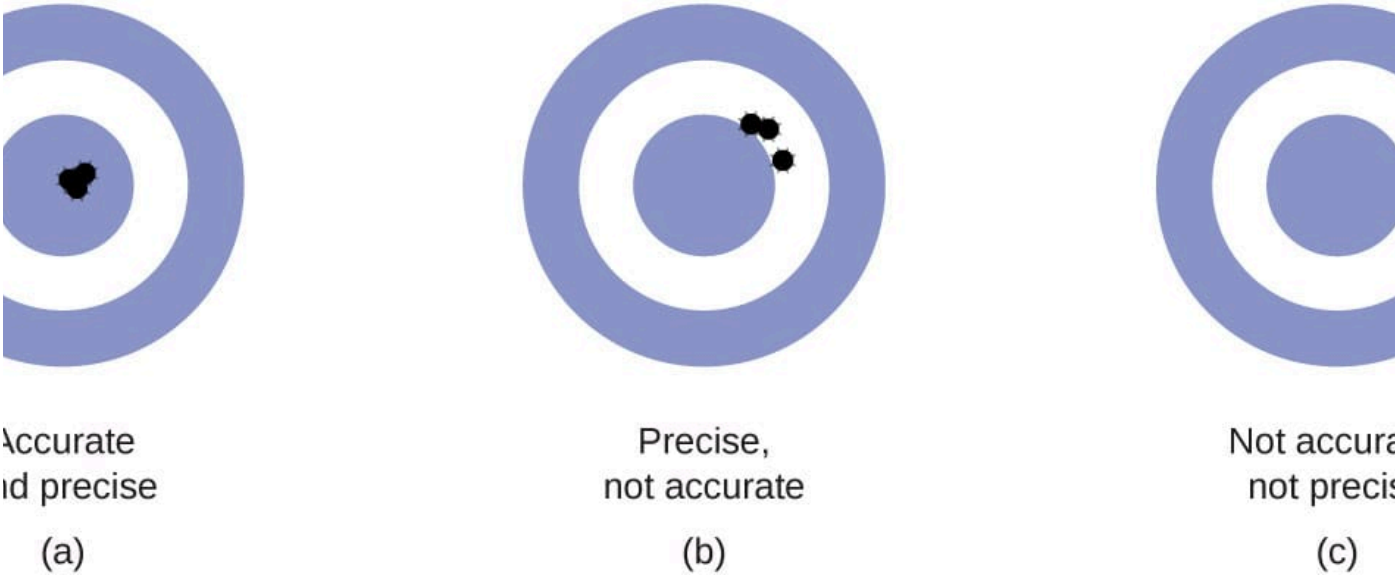


Figure 1.6.7: (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson, PhD via Openstax). Detailed description of Figure 1.6.7.

Chemistry in Everyday Life: Job of a Quality Control Chemist

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

Table 1.6.1 Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers

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

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

Key Takeaways

- Significant figures properly report the number of measured and estimated digits in a measurement.
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the value with the fewest significant figures.

- The rule in addition and subtraction is that the final answer should have the same number of decimal places as the term with the fewest decimal places.

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1.7: The International System of Units (SI)

Base SI Units

People who live in the United States measure weight in pounds, height in feet and inches, and a car's speed in miles per hour. In contrast, chemistry and other branches of science use the International System of Units (also known as **SI Units**, after Système Internationale d'Unités), which was established so that scientists around the world could communicate efficiently with each other. Many countries have also adopted SI units for everyday use as well. The United States is one of the few countries that has not.

Base (or basic) units, are the fundamental units of SI. There are seven base units, which are listed in Table 1.7.1.

Note: Base Units

Chemistry uses five of the base units:

- the mole for amount,
- the kilogram for mass,
- the meter for length,
- the second for time and
- the kelvin for temperature. The degree Celsius (°C) is also commonly used for temperature. The numerical relationship between kelvins and degrees Celsius is as follows: K = °C + 273.

Table 1.7.1: The Seven Base SI Units.

Property	Unit	Abbreviation
length	meter	m
mass	kilogram	kg
time	second	s
amount	mole	mol
temperature	kelvin	K
electrical current	ampere	amp
luminous intensity	candela	cd

The United States uses the English (sometimes called Imperial) system of units for many quantities. Inches, feet, miles, gallons, pounds, and so forth, are all units connected with the English system of units. There have been many mistakes due to the improper conversion of units between the SI and English systems. The size of each base unit is defined by international convention.

For example, the SI unit for **mass,** the kilogram (kg)**,** is defined as the quantity of mass of a special metal cylinder kept in a vault in France (Figure 1.7.1). The other base units have similar definitions and standards.

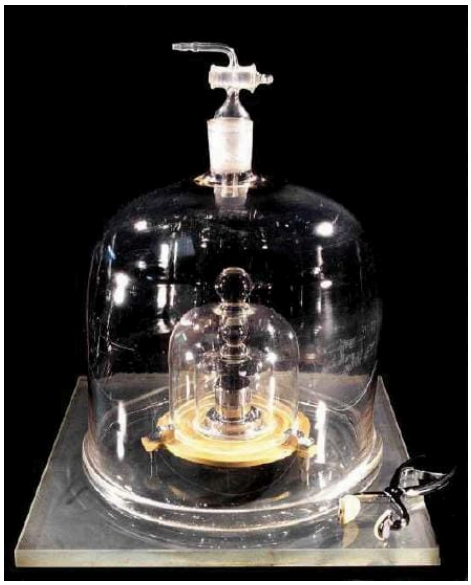


Figure 1.7.1: The replica prototype kilogram as previously defined is housed at the National Institute of Standards and Technology (NIST) in Maryland.The standard for the kilogram is a platinum-iridium cylinder is kept in a special vault in France. (CC BY 4.0, anonymous, Openstax). Detailed description of Figure 1.7.1.

The standard unit of **length** in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd in Figure 1.7.2. A meter is about 3 inches longer than a yard; one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

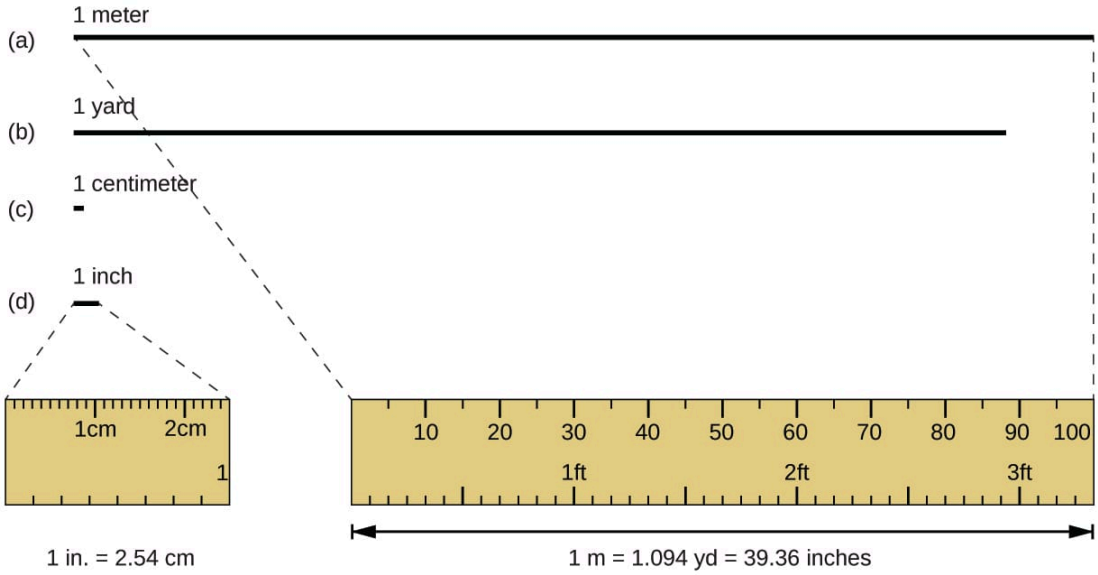


Figure 1.7.2: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd. (CC BY 4.0, anonymous, Openstax). Detailed description of Figure 1.7.2.

The SI unit of temperature is the **Kelvin (K)**. The IUPAC convention is to use kelvin (all lowercase) for the word, K (uppercase) for the unit symbol, and neither the word “degree” nor the degree symbol (°). The degree **Celsius (°C)** is also allowed in the SI system, with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C), and normal human body temperature is approximately 310 K (37 °C).

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

Prefixes for Units

The sizes of the base units are not always convenient for all measurements. For example, a meter is a rather large unit for describing the width of something as narrow as human hair. Instead of reporting the diameter of hair as 0.00012 m or as 1.2 × 10^{-4} m using scientific notation as discussed in section 1.4, SI also provides a series of prefixes that can be attached to the units, creating units that are larger or smaller by powers of 10. Common prefixes and their multiplicative factors are listed in Table 1.7.2. (Perhaps you have already noticed that the base unit kilogram is a combination of a prefix, kilo- meaning 1,000 ×, and a unit of mass, the gram.) Some prefixes create a multiple of the original unit: 1 kilogram equals 1,000 grams, and 1 mega meter equals 1,000,000 meters. Other prefixes create a fraction of the original unit. Thus, 1 centimeter equals 1/100 of a meter, 1 millimeter equals 1/1,000 of a meter, 1 microgram equals 1/1,000,000 of a gram, and so forth.

Table 1.7.2: Prefixes Used with SI Units.

Prefix	Abbreviation	Multiplicative Factor	Multiplicative Factor in Scientific Notation
giga-	G	1,000,000,000 ×	10^{9} ×
mega-	M	1,000,000 ×	10^{6} ×
kilo-	k	1,000 ×	10^{3} ×
deca-	D	10 ×	10^{1} ×
deci-	d	1/10 ×	10^{-1} ×
centi-	c	1/100 ×	10^{-2} ×
milli-	m	1/1,000 ×	10^{-3} ×
micro-	μ*	1/1,000,000 ×	10^{-6} ×
nano-	n	1/1,000,000,000 ×	10^{-9} ×
***The letter μ is the Greek lowercase letter for m and is called "mu," which is pronounced "myoo."**			

Both SI units and prefixes have abbreviations, and the combination of a prefix abbreviation with a base unit abbreviation gives the abbreviation for the modified unit. For example, kg is the abbreviation for kilogram. We will be using these abbreviations throughout this book.

Link to Learning: The Difference Between Mass and Weight

The mass of a body is a measure of its inertial property or how much matter it contains.

The weight of a body is a measure of the force exerted on it by gravity or the force needed to support it.

Gravity on earth gives a body a downward acceleration of about 9.8 m/s^2. In common parlance, weight is often used as a synonym for mass in weights and measures. For instance, the verb "to weigh" means "to determine the mass of" or "to have a mass of." The incorrect use of weight in place of mass should be phased out, and the term mass used when mass is meant.

In science and technology, the weight of a body in a particular reference frame is defined as the force that gives the body an acceleration equal to the local acceleration of free fall in that reference frame.

Thus, the SI unit of the quantity weight defined in this way (force) is the newton (N).

Derived SI Units

Derived units are combinations of SI base units. Units can be multiplied and divided, just as numbers can be multiplied and divided. For example, the area of a square having a side of 2 cm is 2 cm × 2 cm, or 4 cm^2 (read as “four centimeters squared” or “four square centimeters”). Notice that we have squared a length unit, the centimeter, to get a derived unit for area, the square centimeter.

Volume is an important quantity that uses a derived unit. It is the amount of space that a given substance occupies and is defined geometrically as length × width × height. Each distance can be expressed using the meter unit, so volume has the derived unit m × m × m, or m^3 (read as “meters cubed” or “cubic meters”). A cubic meter is a rather large volume, so scientists typically express volumes in terms of 1/1,000 of a cubic meter. This unit has its own name—the liter (L). A liter is a little larger than 1 US quart in volume.

As shown in Figure Figure 1.7.3, a liter is also 1,000 cm^3. By definition, there are 1,000 mL in 1 L, so 1 milliliter and 1 cubic centimeter represent the same volume.

$$1\ mL = 1\ cm^3 \tag{1.7.1}$$

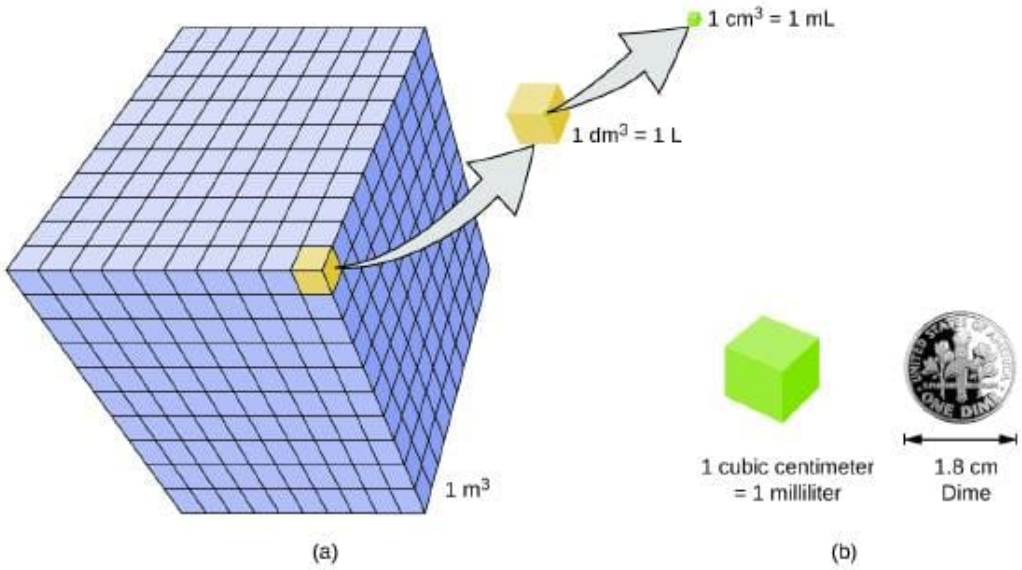


Figure 1.7.3: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). A liter (L) is defined as a cube 1 dm (1/10th of a meter) on a side. A milliliter (mL), 1/1,000th of a liter, is equal to 1 cubic centimeter. (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube. (CC BY 4.0, anonymous, Openstax). Detailed description of Figure 1.7.3.

Below are approximate equivalents for some of the units used in chemistry. We will learn about the converting units in the next section.

Approximate Equivalents to Some SI Units

1 m ≈ 39.36 in. ≈ 3.28 ft ≈ 1.09 yd

1 in. ≈ 2.54 cm

1 km ≈ 0.62 mi

1 kg ≈ 2.20 lb

1 lb ≈ 454 g

1 L ≈ 1.06 qt

1 qt ≈ 0.946 L

Example 1.7.1: Abbreviations of Units

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

1. kiloliter
2. microsecond
3. decimeter
4. nanogram

Solution

1. The abbreviation for a kiloliter is kL. Because kilo means "1,000 ×," 1 kL equals 1,000 L.
2. The abbreviation for microsecond is µs. Micro implies 1/1,000,000th of a unit, so 1 µs equals 0.000001 s.
3. The abbreviation for decimeter is dm. Deci means 1/10th, so 1 dm equals 0.1 m.
4. The abbreviation for nanogram is ng and equals 0.000000001 g.

Energy, another important quantity in chemistry, is the ability to perform work, such as moving a box of books from one side of a room to the other side. It has a derived unit of kg•m^2/s^2. (The dot between the kg and m units implies the units are multiplied together.) Because this combination is cumbersome, this collection of units is redefined as a joule (J). An older unit of energy, but likely more familiar to you, the calorie (cal), is also widely used. There are 4.184 J in 1 cal. Energy changes occur during all chemical processes and will be discussed in a later chapter.

Chemistry in Everyday Life: Energy and Food

The food in our diet provides the energy our bodies need to function properly. The energy contained in food could be expressed in joules or calories, which are the conventional units for energy, but the food industry prefers to use the kilocalorie and refers to it as the Calorie (with a capital C). The average daily energy requirement of an adult is about 2,000–2,500 Calories, which is 2,000,000–2,500,000 calories (with a lowercase c).

If we expend the same amount of energy that our food provides, our body weight remains stable. If we ingest more Calories from food than we expend, however, our bodies store the extra energy in high-energy-density compounds, such as fat, and we gain weight. On the other hand, if we expend more energy than we ingest, we lose weight. Other factors affect our weight as well—genetic, metabolic, behavioral, environmental, cultural factors—but dietary habits are among the most important.

In 2008 the US Centers for Disease Control and Prevention issued a report stating that 73% of Americans were either overweight or obese. More alarmingly, the report also noted that 19% of children aged 6–11 and 18% of adolescents aged 12–19 were overweight—numbers that had tripled over the preceding two decades. Two major reasons for this increase are excessive calorie consumption (especially in the form of high-fat foods) and reduced physical activity. Partly because of that report, many restaurants and food companies are working to reduce the amounts of fat in foods and provide consumers with more healthy food options.

Density is defined as the mass of an object divided by its volume; it describes the amount of matter contained in a given amount of space.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.7.2}$$

Thus, the units of density are the units of mass divided by the units of volume: g/cm^3 or g/mL (for solids and liquids), g/L (for gases), kg/m^3, and so forth.

For example, the density of water is about 1.00 g/cm^3, while the density of mercury is 13.6 g/mL. (Remember that 1 mL equals 1 cm^3.) Mercury is over 13 times as dense as water, meaning that it contains over 13 times the amount of matter in the same amount of space. The density of air at room temperature is about 1.3 g/L. Table 1.7.3 shows the densities of some common substances.

Table 1.7.3: Densities of Common Substances.

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

Example 1.7.2: Density of Bone

What is the density of a section of bone if a 25.3 cm^3 sample has a mass of 27.8 g?

Solution

Because density is defined as the mass of an object divided by its volume, we can set up the following relationship:

$$\begin{aligned} \text{density} &= \frac{mass}{volume} \\ &= \frac{27.8\ g}{25.3\ cm^3} \\ &= 1.10\ g/cm^3 \end{aligned}$$

Note that we have limited our final answer to three significant figures.

Key Takeaways

- Recognize the SI base units and derived units.
- Combining prefixes with base units creates new units of larger or smaller sizes.

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1.8: Converting Units

Conversion Factors

The ability to convert from one unit to another is an important skill. For example, a nurse with 50 mg aspirin tablets who must administer 0.2 g of aspirin to a patient needs to know that 0.2 g equals 200 mg, so 4 tablets are needed. Fortunately, there is a simple way to convert from one unit to another.

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

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

or

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

Suppose we divide both sides of the equation by 1 m (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 (Figure 1.8.1.

A **conversion factor** is a fraction that has equivalent quantities in the numerator and the denominator but expressed in different units.

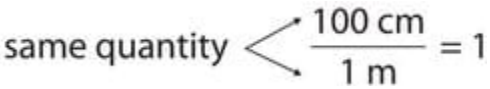


Figure 1.8.1: A fraction showing the conversion factor between cm and m. We know that 100 cm is 1 m, so we have the same quantity on the top and the bottom of our fraction. (CC BY-NC-SA 3.0, anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0))

Example 1.8.1: Conversion of Meters to Centimeters

How many centimeters are there in 3.55 m?

Solution

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:

$$\frac{3.55}{1} \times \frac{100\ \text{cm}}{1} = 355\ \text{cm} \tag{1.8.1}$$

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:

$$\text{quantity (in old units)} \times \text{conversion factor} = \text{quantity (in new units)}$$

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

In the previous example (Equation 1.8.1), 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 1.8.2 shows a **concept map** for constructing a proper conversion.

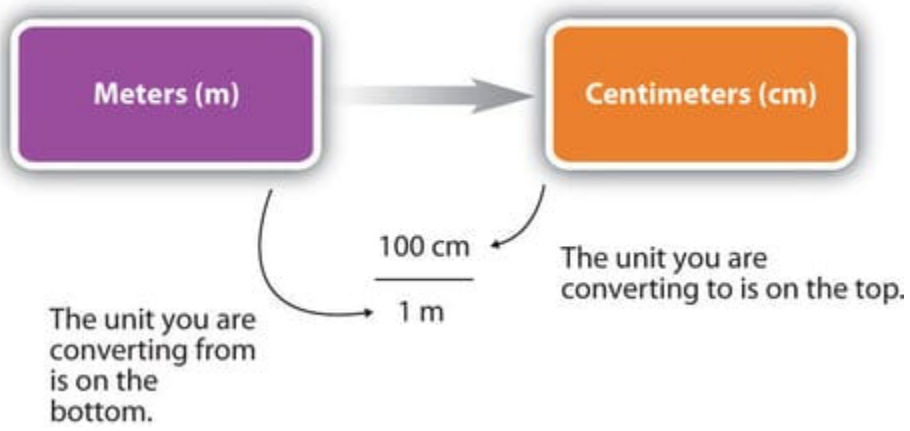


Figure 1.8.2: This is how you construct a conversion factor to convert from one unit to another. Meters converted to centimeters. A fraction showing 100 cm over 1 m. The unit you convert to is on top, and the unit you convert from is on bottom. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed description of Figure 1.8.2.

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 are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

Example 1.8.2: More Conversions

a. The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?
b. A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?

Solution

a. We start with what we are given, 4.7 L. We want to change the unit from liters to milliliters. There are 1,000 mL in 1 L. From this relationship, we can construct two conversion factors:

$$\frac{1\text{ L}}{1,000\text{ mL}} \text{ or } \frac{1,000\text{ mL}}{1\text{ L}}$$

We use the conversion factor that will cancel out the original unit, liters, and introduce the unit we are converting to, which is milliliters. The conversion factor that does this is the one on the right, Figure 1.8.3.



$$4.7\cancel{\text{L}} \times \frac{1,000\text{ mL}}{1\cancel{\text{L}}} = 4,700\text{ mL}$$

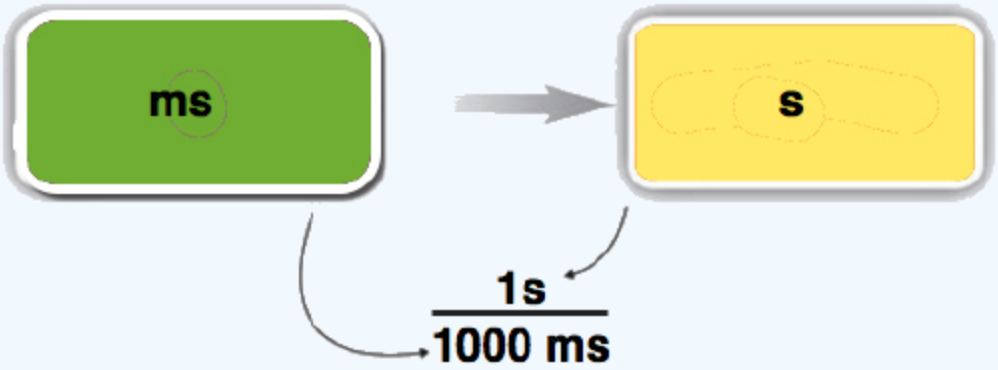
Figure 1.8.3: Liters, converted to milliliters. A fraction showing 1000 mL over 1 L. The unit you convert to is on top, and the unit you convert from is on bottom. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)).

Because the numbers in the conversion factor are exact, we do not consider them when determining the number of significant figures in the final answer. Thus, we report two significant figures in the final answer.

b. We can construct two conversion factors from the relationships between milliseconds and seconds:

$$\frac{1,000\text{ ms}}{1\text{ s}} \text{ or } \frac{1\text{ s}}{1,000\text{ ms}}$$

To convert 18 ms to seconds, we choose the conversion factor that will cancel out milliseconds and introduce seconds. The conversion factor on the right is the appropriate one. We set up the conversion as follows, Figure 1.8.4:



$$18 \cancel{\text{ms}} \times \frac{1 \text{ s}}{1,000 \cancel{\text{ms}}} = 0.018 \text{ s}$$

Figure 1.8.4: A fraction showing 1 s over 1000 ms. The unit you convert to is on top, and the unit you convert from is on bottom. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)).

Conversion Factors From Different Units

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:

$$\frac{13.6 \text{ g}}{1 \text{ mL}} \text{ and } \frac{1 \text{ mL}}{13.6 \text{ g}}$$

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 16 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:

$$16 \cancel{\text{mL}} \times \frac{13.6 \text{ g}}{1 \cancel{\text{mL}}} = 217.6 \text{ g}$$

$$\approx 220 \text{ g}$$

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 1.8.3: Mercury Thermometer

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

Because we are starting with grams, we want to use the conversion factor that has grams in the denominator. The gram unit will cancel algebraically, and milliliters will be introduced in the numerator.

$$0.750 \cancel{\text{ g}} \times \frac{1 \text{ mL}}{13.6 \cancel{\text{ g}}} = 0.055147\ldots \text{ mL}$$

$$\approx 0.0551 \text{ mL}$$

We have limited the final answer to three significant figures.

Chemistry in Everyday Life

The densities of many components and products of the body have a bearing on our health.

Bones

Bone density is important because bone tissue of lower-than-normal density is mechanically weaker and susceptible to breaking. The density of bone is, in part, related to the amount of calcium in one's diet; people who have a diet deficient in calcium, which is an important component of bones, tend to have weaker bones. Dietary supplements or adding dairy products to the diet seems to help strengthen bones. As a group, women experience a decrease in bone density as they age. It has been estimated that fully half of women over age 50 suffer from excessive bone loss, a condition known as osteoporosis. Exact bone densities vary within the body, but for a healthy 30-year-old female, it is about 0.95–1.05 g/cm^3. Osteoporosis is diagnosed if the bone density is below 0.6–0.7 g/cm^3.

Urine

The density of urine can be affected by a variety of medical conditions. Sufferers of diabetes produce an abnormally large volume of urine with a relatively low density. In another form of diabetes, called diabetes mellitus, there is excess glucose dissolved in the urine, so that the density of urine is abnormally high. The density of urine may also be abnormally high because of excess protein in the urine, which can be caused by congestive heart failure or certain renal (kidney) problems. Thus, a urine density test can provide clues to various kinds of health problems. The density of urine is commonly expressed as a specific gravity, which is a unitless quantity defined as

$$\frac{\text{density of some material}}{\text{density of water}}$$

Normal values for the specific gravity of urine range from 1.002 to 1.028.

Body Fat

The overall density of the body is one indicator of a person's total body fat. Fat is less dense than muscle and other tissues, so as it accumulates, the overall density of the body decreases. Measurements of a person's weight and volume provide the overall body density, which can then be correlated to the percentage of body fat. (The body's volume can be measured by immersion in a large tank of water. The amount of water displaced is equal to the volume of the body.)

Problem Solving With 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. You can either memorize the relationship between kilometers and millimeters, or you can do the conversion in two steps. Most people prefer to convert in steps.

To do a stepwise conversion, we first convert the given amount to the base unit. In this example, the base unit is meters. We know that there are 1,000 m in 1 km:

$$54.7 \cancel{\text{ km}} \times \frac{1,000 \text{ m}}{1 \cancel{\text{ km}}} = 54,700 \text{ m}$$

Then we take the result (54,700 m) and convert it to millimeters, remembering that there are $1,000\ \text{mm}$ for every $1\ \text{m}$:

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

We have expressed the final answer in scientific notation.

As a shortcut, both steps in the conversion can be combined into a single, multistep expression, Figure 1.8.5.

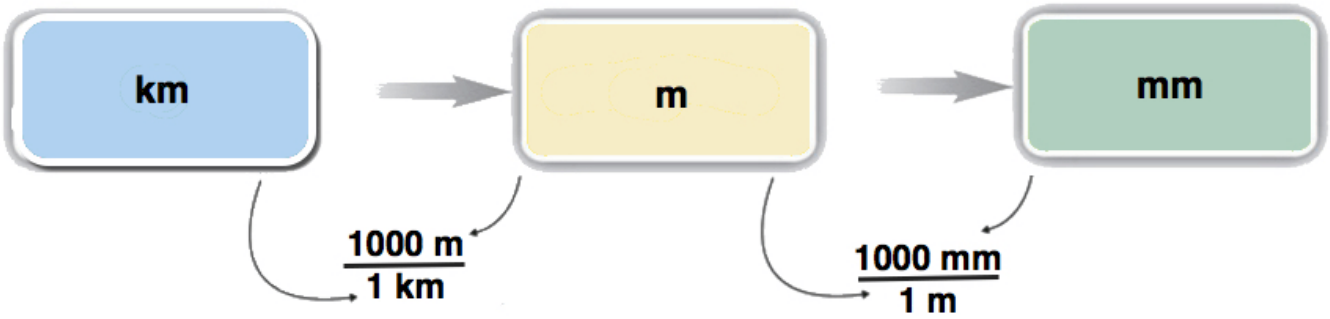


Figure 1.8.5: Concept Map for Conversions. A fraction showing 1000 m over 1km, then 1000 mm over 1 m. The unit you convert to is on top, and the unit you convert from is on bottom. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)).

Calculation

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

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.

Either method—one step at a time or all the steps together—is acceptable. If you do all the steps together, the restriction for the proper number of significant figures should be done after the last step. As long as the math is performed correctly, you should get the same answer no matter which method you use.

Example 1.8.4: Multiple Conversions

Convert 58.2 ms to mega seconds in one multistep calculation.

Solution

First, convert the given unit (ms) to the base unit—in this case, seconds—and then convert seconds to the final unit, mega seconds, Figure 1.8.6.

Concept Map

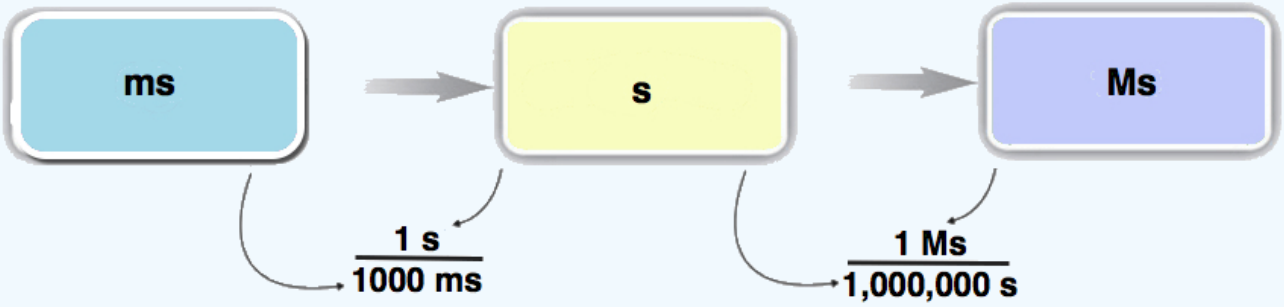


Figure 1.8.6: Concept Map for Conversions. A fraction showing 1 s over 1000 ms, then 1 Ms over 1,000,000 s. The unit you convert to is on top, and the unit you convert from is on bottom. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)).

Calculation

$$58.2 \cancel{\text{ ms}} \times \frac{1 \cancel{\text{s}}}{1,000 \cancel{\text{ ms}}} \times \frac{1 \text{ Ms}}{1,000,000 \cancel{\text{ s}}} = 0.0000000582 \text{ Ms}$$

$$= 5.82 \times 10^{-8} \text{ Ms}$$

Neither conversion factor affects the number of significant figures in the final answer.

Link to Learning: 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.

Key Takeaway

A unit can be converted to another unit of the same type with a conversion factor.

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1.9: Dosage Calculations

Introduction

A medicine can be more harmful than helpful if it is not taken in the proper dosage. **A dosage (or dose)** is the specific amount of a medicine that is known to be therapeutic for an ailment in a patient of a certain size.

Dosages of the active ingredient in medications are usually described by units of mass, typically grams or milligrams, and generally are equated with a number of capsules or teaspoonfuls to be swallowed or injected. The amount of the active ingredient in a medicine is carefully controlled so that the proper number of pills or spoonfuls contains the proper dose.

Most drugs must be taken in just the right amount. If too little is taken, the desired effects will not occur (or will not occur fast enough for comfort); if too much is taken, there may be potential side effects that are worse than the original ailment. Some drugs are available in multiple dosages. For example, tablets of the medication levothyroxine sodium, a synthetic thyroid hormone for those suffering from decreased thyroid gland function, are available in 11 different doses, ranging from 25 micrograms (µg) to 300 µg. It is a doctor's responsibility to prescribe the correct dosage for a patient, and it is a pharmacist's responsibility to provide the patient with the correct medicine at the dosage prescribed. Thus, proper quantities—which are expressed using numbers and their associated units—are crucial for keeping us healthy.

Effects are dose-dependency

Chemicals are the most common things for which doses are measured, but there are others, such as radiation exposure. For humans, most doses of micronutrients and medications are measured in milligrams (mg), but some are measured in micrograms because of their potency. Nonmedicinal poisons span the measurement scale; some poisons are so dangerous that a single microgram of it could be deadly, whereas other substances take much more. For example, even water is toxic when consumed in large enough quantities.

Dosage (the size of each dose) determines the strength and duration of the health benefits of nutrients, and also of the therapeutic effects of medical treatments. Dosage also determines the severity of adverse effects of treatments and toxins.

Duration of exposure, that is, the period of time over which the dose was received (all at once or gradually) also determines its effects (the body may build tolerance to gradual exposure to a drug, while a large immediate dose could be deadly).

The route by which a dose is exposed to, may affect the outcome, because some medications have different effects depending on whether they are inhaled, ingested, taken transdermally, injected, or inserted.

The dosage, route, concentration, and division over time may all be critical considerations in the administering of drugs, or in responding to exposure to a toxin. In nutrition, the route is usually a given, as nutrients are generally eaten; while dosage and the frequency of ingestion of nutrients are very important variables in preventing disease and promoting overall health.

Calculation of dose

Calculating drug dosages for humans based on the doses used in animal studies can be based on weight (e.g., mg/kg) or surface area (e.g., mg/m^2). [1]

Link to Learning: Dosages

Drug dosage calculations

Drug dosage calculation is required if the physician's order is different from what is available, illustrated in Example 1.9.1).

Drug Dosage Calculation based on Body Weight

Many drugs (especially in children) are dosed according to body weight (mg/kg). These calculations are carried in 3-step conversions. The first step is to convert the body weight from pounds (lbs) to kg. The second step is to convert kg to mg (the total mg dose calculated based on body weight). And, finally, the mg dose is converted to the number of tablets, illustrated in Example 1.9.2).

Example 1.9.1: Drug Dosage Calculations

a. A physician ordered 100 mg of Demerol. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?
b. The doctor's order is 1.2 g of Folic Acid. Folic Acid is available as 800 mg per tablet. How many tablets should be taken?

Solution

a. We start with the given, 100 mg. We want to change the unit from mg to tablets. There are 50 mg in 1 tablet (Remember that per tablet means one tablet. From this relationship, we can construct the conversion factor. We use the conversion factor that will cancel out the original unit, mg, and introduce the unit we are converting to, which is tablet.

$$100 \cancel{\text{mg}} \times \frac{1 \text{ tablet}}{50 \cancel{\text{ mg}}} = 2 \text{ tablets}$$

Hence, the nurse should administer 2 tablets.

b. We start with the given, 1.2 g and we want to change grams to number of tablets. First, we convert 1.2 g to mg and then convert mg to tablets. We need a conversion factor for each step.

$$1.2 \cancel{g} \times \frac{1{,}000 \cancel{\text{mg}}}{1 \cancel{g}} \times \frac{1 \text{ tablet}}{800 \cancel{\text{ mg}}} = 1.5 \text{ tablets}$$

Hence, 1.5 tablets should be taken.

Example 1.9.2: Drug Dosage Calculations Based on Body weight

a. Demerol is ordered 1.5mg/kg for a patient that is 220 lbs. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?
b. A doctor prescribes amoxicillin 30mg/kg to a child weighing 73.5 lbs. Amoxicillin is available as 500 mg tablets. How many tablets should the nurse administer?

Solution

a. We start with the given, 220 lbs. We want to change the unit from lbs to kg, and then, from kg to total dose (mg) and then the mg dose to tablets. The first conversion factor will cancel out the original unit, lbs, and introduce the unit we are converting to, which is kg. The second conversion factor will cancel out kg, and introduce the unit of the dose (usually mg) and then mg to tablet.

$$220 \cancel{\text{lbs}} \times \frac{1 \cancel{\text{kg}}}{2.2 \cancel{\text{ lbs}}} \times \frac{1.5 \cancel{\text{mg}}}{1 \cancel{\text{ kg}}} \times \frac{1 \text{ tablet}}{50 \cancel{\text{ mg}}} = 3 \text{ tablets}$$

Hence, the nurse administers 3 tablets.

b. Start with the given, 73.5 lbs. We want to change the unit from lbs to kg, and then, from kg to total dose (mg) and then mg to tablets. The first conversion factor will cancel out the original unit, lbs, and convert to kg. The second conversion factor will cancel out kg, and convert to the total mg dose and the final conversion will cancel mg to introduce the final unit, tablet.

$$73.5 \cancel{\text{ lbs}} \times \frac{1 \cancel{\text{ kg}}}{2.2 \cancel{\text{ lbs}}} \times \frac{30 \cancel{\text{ mg}}}{1 \cancel{\text{ kg}}} \times \frac{1 \text{ tablet}}{500 \cancel{\text{ mg}}} = 2 \text{ tablets}$$

Hence, the nurse administers 2 tablets.

Key Takeaways

Drug dosages can be calculated using conversion factors.

References

1. Pinkel, D. (1958). The use of body surface area as a criterion of drug dosage in cancer chemotherapy. Cancer Research, 18 (7), 853–856. (opens in new window)

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1.10: Summary

Chemistry is the study of matter, which is anything that has mass and takes up space.

Chemistry is one branch of science, which is the study of the natural universe. Like all branches of science, chemistry relies on the scientific method, which is a process of learning about the world around us. In the scientific method, a guess or hypothesis is tested through experiment and measurement.

Matter can be described in a number of ways. Physical properties describe characteristics of a sample that do not change the chemical identity of the material (size, shape, color, and so on), while chemical properties describe how a sample of matter changes its chemical composition. A substance is any material that has the same physical and chemical properties throughout. An element is a substance that cannot be broken down into chemically simpler components. The smallest chemically identifiable piece of an element is an atom. A substance that can be broken down into simpler chemical components is a compound. The smallest chemically identifiable piece of a compound is a molecule. Two or more substances combine physically to make a mixture. If the mixture is composed of discrete regions that maintain their own identity, the mixture is a heterogeneous mixture. If the mixture is so thoroughly mixed that the different components are evenly distributed throughout, it is a homogeneous mixture. Another name for a homogeneous mixture is a solution.

Substances can also be described by their phase: solid, liquid, or gas.

Scientists learn about the universe by making measurements of quantities, which consist of numbers (how many) and units (of what). The numerical portion of a quantity can be expressed using scientific notation, which is based on powers, or exponents, of 10. Large numbers have positive powers of 10, while numbers less than 1 have negative powers of 10. The proper reporting of a measurement requires proper use of significant figures, which are all the known digits of a measurement plus the first estimated digit. The number of figures to report in the result of a calculation based on measured quantities depends on the numbers of significant figures in those quantities. For addition and subtraction, the number of significant figures is determined by position; for multiplication and division, it is decided by the number of significant figures in the original measured values. Nonsignificant digits are dropped from a final answer in accordance with the rules of rounding.

Chemistry uses SI, a system of units based on seven basic units. The most important ones for chemistry are the units for length, mass, amount, time, and temperature. Basic units can be combined with numerical prefixes to change the size of the units. They can also be combined with other units to make derived units, which are used to express other quantities such as volume, density, or energy. A formal conversion from one unit to another uses a conversion factor, which is constructed from the relationship between the two units. Numbers in conversion factors may affect the number of significant figures in a calculated quantity, depending on whether the conversion factor is exact. Conversion factors can be applied in separate computations, or several can be used at once in a single, longer computation. Conversion factors are very useful in calculating dosages.

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1.11: Key Terms

- **Accuracy**: The degree to which a measured value matches the true or accepted value.
- **Alchemy**: An ancient practice that combined elements of chemistry, physics, medicine, astrology, mysticism, and art.
- **Alloy**: A homogeneous mixture of two or more metals, or a metal and a nonmetal, that has metallic properties.
- **Atom**: The smallest unit of an element that retains the chemical properties of that element.
- **Avogadro's Number**: The number of particles in a mole, approximately 6.022×10^{23} particles per mole.
- **Basic Units**: The fundamental units of measurement in the SI system, including meter, kilogram, second, mole, ampere, kelvin, and candela.
- **Boiling**: The process by which a liquid changes to a gas when it reaches its boiling point.
- **Celcius**: A temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water.
- **Chemical Properties**: Characteristics of a substance that can only be observed during a chemical change, such as reactivity and flammability.
- **Chemical Reaction:**A chemical reaction is a process in which one or more substances (often called reactants) are transformed into new substances (known as products) with different compositions and properties.
- **Chemistry**: The study of matter, its properties, how it interacts with other matter, and the changes it undergoes.
- **Compound:** A substance formed when two or more elements chemically combine in a fixed ratio.
- **Concept Map**: A visual representation of the relationships between different concepts, often used as a learning tool.
- **Condensation**: The process by which a gas turns into a liquid when cooled.
- **Conversion Factor**: A ratio used to convert one unit of measurement to another.
- **Density**: The mass of a substance per unit volume, often expressed in g/cm^3 or kg/m^3.
- **Deposition**: The process by which a gas changes directly into a solid without passing through the liquid state.
- **Derived Units**: Units of measurement derived from the seven base SI units, such as meters per second for speed.
- **Dosage**: The amount of a substance to be administered, often used in medical contexts.
- **Element**: A pure substance made of only one kind of atom, represented by a chemical symbol.
- **Evaporation**: The process by which molecules at the surface of a liquid absorb enough energy to change into a gaseous state.
- **Experiment**: A procedure carried out to test a hypothesis and gather data.
- **Extensive Properties**: Properties that depend on the amount of matter present, such as mass and volume.
- **Fusion**: The process of changing from a solid to a liquid, also known as melting.
- **Gas**: A state of matter with no fixed shape or volume, particles move freely.
- **Heterogenous Mixture**: A mixture in which the composition is not uniform throughout, such as a salad.
- **Homogenous Mixture**: A mixture that has a uniform composition throughout, such as salt water.
- **Hypothesis**: A testable prediction or explanation for an observation or phenomenon.
- **Intensive Properties**: Properties that do not depend on the amount of matter present, such as density and boiling point.
- **Kelvin**: The SI base unit for temperature, where 0 K is absolute zero.
- **Law**: A statement that describes a consistent natural phenomenon, often in mathematical form.
- **Length**: The measurement of distance between two points, measured in meters in the SI system.
- **Liquid**: A state of matter with a definite volume but no fixed shape.
- **Macroscopic Viewpoint**: Observing matter and its changes on a scale visible to the naked eye.
- **Mass**: A measure of the amount of matter in an object, typically measured in grams or kilograms.
- **Matter**: Matter can be characterized as any substance that has mass and occupies space, encompassing solids, liquids, and gases. In contrast, forms of energy such as heat and light do not possess mass or volume, so they are not considered matter.
- **Melting**: The change of state from a solid to a liquid by the addition of heat.
- **Microscopic Viewpoint**: Observing matter at a molecular or atomic level, often requiring a microscope.
- **Mixture**: A physical combination of two or more substances where each retains its own properties.
- **Molecule**: A group of atoms bonded together, representing the smallest fundamental unit of a chemical compound.
- **Peer Review**: The process of evaluating scientific work by others in the same field to ensure accuracy and validity.
- **Phase**: A distinct state of matter with uniform properties, such as solid, liquid, or gas.
- **Phase Change**: The transition of matter from one state to another, such as from solid to liquid.
- **Physical Properties**: Characteristics of a substance that can be observed without changing its chemical composition, such as color, odor, and melting point.
- **Physical State**: The form in which matter exists, including solid, liquid, and gas.
- **Precision**: The degree to which repeated measurements under unchanged conditions show the same results.
- **Pure Substance**: A material with a uniform and definite composition, either an element or a compound.

- **Science:** Science is the systematic process of investigating the universe and all its phenomena through observation, experimentation, and analysis. It aims to build a reliable foundation of knowledge, explain how the natural world operates, and enable us to make predictions or develop new applications based on that understanding.
- **Scientific Method**: A systematic approach to research and experimentation used to answer scientific questions.
- **Scientific Notation**: A method of expressing very large or very small numbers in the form of a decimal multiplied by a power of ten.
- **Second**: The SI base unit for time.
- **SI Units**: The International System of Units, the standard set of units used in science, including meter, kilogram, second, ampere, kelvin, mole, and candela.
- **Significant Figures**: Digits in a measurement that are known with certainty plus one final digit that is estimated.
- **Solid**: A state of matter with a definite shape and volume, with particles tightly packed in a fixed position.
- **Solution**: A homogeneous mixture where one substance is completely dissolved in another.
- **Sublimation**: The process by which a solid changes directly into a gas without passing through the liquid state.
- **Substance**: A form of matter with a consistent composition and distinct properties.
- **Theory**: A well-substantiated explanation of an aspect of the natural world based on a body of evidence.
- **Uncertainty**: The range within which the true value of a measurement lies, often due to limitations in measurement precision.
- **Unit**: A standard quantity used to specify measurements.
- **Volume**: The amount of space occupied by a substance, typically measured in liters or cubic meters.
- **Weight**: The force exerted on an object due to gravity, often measured in newtons.

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1.12: Alternative text for Chapter 1

Figure 1.2.1: Relationship Between Branches of Science

The image shows a concept map illustrating the relationships between chemistry and various scientific disciplines. At the center of the diagram is a rectangular box labeled "Chemistry," which serves as the central science connecting multiple fields.

Above the chemistry box, a box labeled "Math" is positioned at the top of the hierarchy, connected directly to a box labeled "Physics," which in turn is connected to the central chemistry box. To the right of physics, a box labeled "Astronomy" is also linked to chemistry.

To the left of the central chemistry box, two rectangular boxes are connected, representing specialized fields that integrate chemistry and physics. These boxes include a box labeled "Chemical Physics" and a red box labeled "Biochemistry," highlighting areas where chemistry overlaps with other disciplines.

Below the central chemistry box, a box labeled "Biology" is connected, representing the intersection of chemical and biological sciences. A box labeled "Medicine" is positioned beneath biology, showing how chemistry underpins the medical field.

To the right of the central chemistry box, two additional fields are displayed. A box labeled "Geochemistry" is connected, representing the application of chemistry to geological processes. A box labeled "Geology" is connected beneath geochemistry, illustrating the integration of chemistry in understanding Earth's physical structure and processes.

The boxes and connecting lines emphasize the interdisciplinary nature of chemistry and its foundational role in a variety of scientific fields, including physics, biology, geology, and their specialized subfields.

Figure 1.2.2: Chemistry is Everywhere

A woman sits at a desk in a lab or study setting and looks at a computer screen showing a 3D model of a carbon nanotube. She is wearing glasses and earphones. The image connects the molecular structure on the screen to common materials and substances found around her.

Several call out circles show molecular models tied to everyday items. These include:

- **Graphite**, shown as layered hexagonal carbon structures
- **Vitamin C**, shown as a molecule made of carbon, hydrogen, and oxygen atoms
- **Caffeine**, shown as a molecule containing carbon, hydrogen, nitrogen, and oxygen
- **Sodium chloride (table salt)**, shown as a crystal lattice of sodium and chloride ions
- **Water**, shown as a bent molecule with one oxygen atom and two hydrogen atoms
- **Lithium cobalt oxide**, shown as a material used in lithium-ion batteries

Objects on the desk visually match some of these substances, including a salt shaker, a glass of water, snacks, and an electronic device. The image illustrates how chemistry connects atomic and molecular structures to familiar everyday materials.

Figure 1.2.3: The First Affinity Table

The image is a historical chart displaying alchemical symbols used to represent various substances and compounds. The chart is organized as a grid with rows and columns of symbols, each corresponding to different alchemical materials. The symbols are a mixture of geometric shapes, planetary symbols, and stylized markings, each designed to convey a specific substance.

The chart includes a legend at the bottom that provides descriptions of the symbols. For example, circular symbols with additional markings represent acids, salts, and alkalis, such as "Esprits acides" (acid spirits), "Acide du sel marin" (sea salt acid), and "Sel alcalin volatil" (volatile alkali salt). Other symbols correspond to metals and elements, including "Mercure" (mercury), "Or" (gold), and "Argent" (silver). Additional symbols denote materials like "Terre absorbante" (absorbent earth) and "Soufre mineral" (mineral sulfur).

The grid itself is a structured matrix that shows relationships or combinations between these substances, though the exact interpretation is unclear without additional context. Some of the symbols are more complex, such as "PC" for "Pierre Calaminaire" (calamine stone), while others are simple geometric figures like triangles or circles.

The chart is a historical artifact that reflects the symbolic language of alchemy, used prior to the development of modern chemistry. The intricate symbols and their associated materials offer a glimpse into how early scientists and alchemists conceptualized and recorded their understanding of substances and their interactions.

Figure 1.2.4: Scientific Method Flowchart

The image depicts a flowchart representing the scientific method. At the top of the chart, the process begins with a rectangle labeled "Observation," which represents the initial step of identifying phenomena or events to study. An arrow points downward to a rectangle labeled "Hypothesis," indicating that observations lead to the formulation of a testable explanation or prediction.

From the hypothesis, another arrow points downward to a rectangle labeled "Experiment," signifying that hypotheses are tested through experiments designed to gather data and validate or refute the initial hypothesis. The experiment box has an arrow looping back to the hypothesis box, indicating that if experimental results do not support the hypothesis, it may need to be revised or modified.

Below the experiment box, the flowchart splits into two branches. One branch points to "Law," representing a concise statement or equation summarizing observed phenomena if the results are consistent and widely accepted. The second branch leads to "Theory (model)," representing a comprehensive explanation of phenomena based on extensive evidence. From the theory box, an arrow points back to "Experiment," indicating that theories can generate predictions that are tested through further experiments.

Additionally, a secondary arrow loops back from the "Theory (model)" box to itself, with a label stating "Modify as needed." This highlights that theories are continually refined or adjusted based on new evidence or experimental results.

The flowchart is enclosed in a rectangular frame, emphasizing the iterative and self-correcting nature of the scientific method. It visually represents the dynamic and interconnected process of observation, hypothesis, experimentation, theory, and law in scientific inquiry.

Figure 1.2.5: General Steps of the Scientific Method

The image is a flowchart demonstrating the iterative process of forming, testing, and refining a hypothesis, with an example about bird migration. At the top, the first step is represented by a box labeled "Propose a hypothesis." An arrow points downward to the next step, labeled "Test the hypothesis," which is displayed in another box. From this box, another arrow points downward to the step labeled "Refine the hypothesis," which is shown in another box. A horizontal arrow loops back from the "Refine the hypothesis" box to the "Test the hypothesis" box, indicating the iterative nature of this process.

Each step includes a specific example in text on the right side of the flowchart. In the first step, "Propose a hypothesis," the hypothesis is stated as, "Birds fly south during the winter because they can't stand the cold." In the second step, "Test the hypothesis," the proposed method is described as, "Subject birds to cold and observe their behavior." Finally, in the third step, "Refine the hypothesis," the example response to experimental findings is provided: "The cold doesn't affect the birds' behavior. Maybe the position of the sun in the sky is important."

The flowchart is designed to visually convey the scientific process of inquiry and experimentation, emphasizing that hypotheses are tested and adjusted based on experimental evidence. The boxes and the arrows clearly guide the viewer through the cyclical nature of refining scientific understanding.

Figure 1.3.1: Extensive and Intensive Properties of Sulfur

The image illustrates the differences between extensive and intensive properties of a substance using sulfur as an example. The visual is divided into two sections, each highlighting sulfur in different physical forms: sulfur crystals on the left and sulfur powder on the right. Both forms are accompanied by measurements and properties categorized as either extensive or intensive.

On the left, sulfur crystals are depicted as irregularly shaped, and faceted. A scale below the crystals indicates their approximate size of 2 centimeters. The label beneath the image reads "Sulfur crystals." On the right, sulfur powder is shown as a fine, loose

substance, with a similar scale below it indicating a size of 2 centimeters. The label beneath this image reads "Sulfur powder."

Between the two images, the properties of sulfur are grouped into two categories: "Extensive" and "Intensive." Extensive properties are described as "dependent on the amount of substance" and include measurements for both sulfur forms. For sulfur crystals, the mass is 39 grams, and the volume is 18.8 cubic centimeters. For sulfur powder, the mass is 0.84 grams, and the volume is 4.1 cubic centimeters. These values emphasize that extensive properties vary with the quantity of the material.

Intensive properties, described as "independent of the amount of substance," are also listed for both forms. The color is noted as yellow for both sulfur crystals and sulfur powder, and the melting point is 115.2°C in both cases. This highlights that intensive properties remain constant regardless of the sample size or form of the material.

The diagram effectively compares the extensive and intensive properties of sulfur, demonstrating that some properties depend on the sample's quantity, while others remain consistent across different forms and sizes.

Figure 1.3.2: Physical vs Chemical Change

The image is divided into two sections, each depicting a different example of matter in transformation or interaction with energy.

On the left side, there is a glass filled with water and large ice cubes floating on the surface. The glass is clear, revealing the contrast between the solid ice and liquid water. The water is colorless and reflects light, while the ice cubes are transparent with sharp edges, showing intricate internal patterns created by the freezing process. The scene is illuminated with a warm light, and the glass is placed on a wooden surface, creating a cozy and natural setting. This side represents a physical change, as the ice melts into water, illustrating the transition between solid and liquid states.

On the right side, a lit candle is shown against a dark background. The candle has cylindrical wax body that is partially melted, with a flame burning steadily at the top. The flame emits a bright light with a faint blue base near the wick, demonstrating the combustion process. The wick is visibly darkened where it burns, and the melted wax has pooled at the candle's top, emphasizing the dynamic interaction between heat and material. This side represents a chemical change, as the wax combusts and transforms into gases, heat, and light energy.

Together, the two images contrast physical and chemical changes, showcasing the differences between reversible changes, like melting, and irreversible transformations, like combustion.

Figure 1.3.3: Electrolysis of Water

The image depicts the electrolysis of water, showing the setup and the products of the reaction. At the center is a clear, transparent container filled with liquid water (H_2O). Two vertical test tubes, partially submerged in the water, are positioned over electrodes labeled "Anode" on the left and "Cathode" on the right. Both electrodes are connected to a electrical circuit powered by an "Everready Super Heavy Duty" battery placed to the left of the container. The battery has wires connected to the anode and cathode, supplying electricity for the electrolysis process.

Bubbles of gas are visibly forming at both electrodes, collecting inside the test tubes. The gas at the anode is oxygen, while the gas at the cathode is hydrogen. Two circular insets at the top of the image provide detailed molecular representations of the products. On the left, a red sphere representing gaseous oxygen (O_2) is labeled "Product: element" and "Gaseous oxygen $O_2(g)$." On the right, two white spheres representing gaseous hydrogen (H_2) are labeled "Product: element" and "Gaseous hydrogen $H_2(g)$."

At the bottom right, another circular inset shows the molecular structure of liquid water, consisting of one red oxygen atom and two smaller white hydrogen atoms arranged in a bent geometry. This inset is labeled "Reactant compound" and "Liquid water $H_2O(l)$."

The image visually demonstrates the decomposition of water into its elemental components, hydrogen and oxygen, using electricity. The labeled components emphasize the role of water as the reactant and the formation of hydrogen and oxygen gases as the products, showcasing the principle of electrolysis and the conservation of matter.

Figure 1.3.4: Macroscopic vs Microscopic View

The image is a conceptual illustration explaining the microscopic nature of printed ink on a page. On the left, there is a block of

text that reads: "Vast numbers of tiny, individual pieces of ink make up a period on the printed page." To the right of this text is a visual representation of the concept.

At the center of the image is a large circular dot, representing a period from printed text magnified 25 times. A speech bubble extends from this dot to the right, pointing to another larger circular bubble that shows an even more magnified view, labeled "×400,000,000." Inside this highly magnified bubble, individual dots are displayed as small circles grouped together, illustrating the individual microscopic particles of ink that collectively form the larger printed period.

The image emphasizes the immense number of tiny ink particles required to create what appears to the human eye as a single dot, highlighting the complexity and precision involved in printing at the microscopic level. The labels "×25" and "×400,000,000" underscore the progressively increasing magnification levels used to visualize the ink particles.

Figure 1.3.5: Classification of Matter Flowchart

The image is a hierarchical flowchart that categorizes and defines different types of matter. At the top of the chart, a box labeled "Matter" provides a broad definition: "Occupies space and possesses mass; may exist as solid, liquid, or gas." Two arrows extend downward from this box, leading to two primary categories: "Pure substance" on the left and "Mixture" on the right.

The "Pure substance" category, displayed in a box, is defined as "Matter having an invariant chemical composition and distinct properties." Two arrows extend downward from this box, dividing pure substances into "Element" and "Compound." The "Element" category is described as "Fundamental substance; cannot be separated into simpler substances by chemical methods." The "Compound" category is described as "Substance composed of two or more elements in fixed proportions; can be separated into simpler substances and elements only by chemical methods." A footnote adds that chemical methods of separation include electrolysis.

On the right side of the flowchart, the "Mixture" category, shown in a box, is defined as "Matter consisting of two or more pure substances that retain their individual identities and can be separated by physical methods." Two arrows extend downward, dividing mixtures into "Homogeneous" and "Heterogeneous." The "Homogeneous" category is defined as "Mixture having a uniform composition and properties throughout (also called a solution)." The "Heterogeneous" category is defined as "Mixture not uniform in composition and properties throughout." A footnote specifies that physical methods of separation include filtration, distillation, and crystallization.

The chart visually organizes the classification of matter, highlighting distinctions between pure substances and mixtures, as well as the methods used to separate or identify these categories. The use of color-coded boxes and hierarchical structure makes it easy to follow the relationships and characteristics of each type of matter.

Figure 1.3.6: States of Matter

The three most common states or phases of matter are solid, liquid, and gas. A flask labeled solid contains a cube of matter and says has fixed shape and volume. A flask labeled liquid contains a colored liquid. This flask says takes shape of container, forms horizontal surfaces, has fixed volume. The flask labeled gas is filled with a colored gas. This flask says expands to fill container.

Figure 1.3.8: Classification of Matter

The image is a flowchart that categorizes substances based on their composition and physical state. At the top of the chart, the term "Substance" is displayed in a box, representing the starting point. Two branches extend downward from this box, labeled with the question, "Only one present?" This question divides substances into two pathways: "Yes" and "No."

The "Yes" pathway leads to two boxes, one labeled "Element" and the other labeled "Compound". These represent pure substances, where an element contains only one type of atom, and a compound consists of two or more types of atoms chemically combined.

The "No" pathway leads to two additional boxes, representing mixtures. A box labeled "Homogeneous mixture" indicates a mixture with uniform composition throughout, while another box labeled "Heterogeneous mixture" represents a mixture where the composition is not uniform.

All four categories—elements, compounds, homogeneous mixtures, and heterogeneous mixtures—are connected by arrows that converge into a single pathway leading to a box labeled "Phase". This indicates the physical state of matter.

From the "Phase" box, arrows extend downward to three boxes labeled "Solid", "Liquid", and "Gas". Arrows connect these three states in a triangular pattern, representing the transitions between them. The arrows illustrate how matter can change phases, such as from solid to liquid, liquid to gas, or gas to solid.

The flowchart visually organizes the classification of substances based on their composition and physical state, emphasizing the relationships between pure substances, mixtures, and the phases of matter. The color-coded boxes and directional arrows provide clarity to the hierarchical and interconnected concepts.

Figure 1.6.1: Measuring Volume of a Liquid in a Graduated Cylinder

The image illustrates the proper technique for reading the liquid level in a graduated cylinder. On the left side of the image, there is a full view of a transparent graduated cylinder standing upright. The cylinder is filled with a liquid up to the 20-milliliter (mL) mark, which is clearly labeled on the scale running vertically along the side of the cylinder. The scale also includes additional markings at intervals of 5 mL, starting from 0 mL at the bottom and increasing up to 25 mL at the top. A rectangular box is drawn around the area near the liquid's meniscus at the 20 mL level to focus attention on this part of the cylinder.

An arrow points from the boxed area to the center of the image, where a zoomed-in view of the liquid level is displayed. In this enlarged section, the meniscus is visible as a curved surface at the top of the liquid column. The curved surface dips slightly downward in the middle due to the cohesive and adhesive properties of the liquid. The scale marking "20" is labeled in blue next to the meniscus for clarity.

Another arrow points from this zoomed-in view to the far right of the image, where the meniscus is shown in an even closer magnification. This detailed view highlights the curve of the meniscus with dashed red lines indicating the upper and lower edges of the curve. A black arrow points to the bottom of the meniscus, where a note reads, "Make reading here." This emphasizes that the correct way to measure liquid volume is by aligning the measurement with the bottom of the meniscus at eye level.

The image provides a step-by-step visual guide for accurately measuring liquid volume in a graduated cylinder, emphasizing the importance of focusing on the meniscus and taking the reading at its lowest point.

Figure 1.6.7: Accuracy and Precision

The image illustrates the concepts of accuracy and precision using three target diagrams labeled (a), (b), and (c). Each diagram depicts a circular target with a concentric ring pattern, where the innermost circle represents the center of the target, the "bullseye." Dots indicate the positions of hits on the target, representing measurements or outcomes.

In diagram (a), labeled "Accurate and precise," all the dots are clustered tightly together at the center of the target, within the innermost circle. This demonstrates that the measurements are both accurate (close to the true value, represented by the bullseye) and precise (closely grouped together with minimal variation).

In diagram (b), labeled "Precise, not accurate," the dots are tightly clustered but located in the upper-right region of the target, outside the bullseye. This shows that the measurements are precise, as they are closely grouped together, but not accurate, as they are far from the true value.

In diagram (c), labeled "Not accurate, not precise," the dots are scattered across the upper-right quadrant of the target, far from the bullseye and not closely grouped together. This represents measurements that are neither accurate (close to the true value) nor precise (closely grouped).

The image visually conveys the difference between accuracy and precision, highlighting the importance of both qualities in reliable measurements or outcomes. The concentric rings of the target serve as a visual metaphor for closeness to the true value, while the grouping of dots represents the consistency of measurements.

Figure 1.7.1: The Kilogram

The image shows a historical scientific apparatus known as the International Prototype of the Kilogram (IPK), which was the standard for defining the kilogram. The centerpiece is a cylindrical platinum-iridium artifact housed inside a series of protective enclosures to prevent contamination or environmental interference.

The IPK is situated at the center of a glass bell jar, resting on a circular platform. The bell jar has a rounded top with a ground-glass seal and is equipped with a valve at the top, used to create a vacuum or ensure the enclosure remains sealed. Surrounding the inner jar is a larger outer glass bell jar, which further shields the artifact from dust, air, or other contaminants. The outer bell jar is larger in size, with a similar valve at its top and a wide base sealed to the platform.

The platform is made of a transparent material, likely glass or acrylic, and is positioned on a metallic base. Clips and clasps are visible, securing the glass enclosures in place. The entire assembly is illuminated under focused light, reflecting off the glossy surfaces of the glass jars, creating a visually striking contrast against the dark background.

This apparatus represents the meticulous care taken to preserve the kilogram standard and highlights the importance of maintaining exacting conditions for scientific measurement standards. It serves as a symbol of precision and the legacy of the metric system.

Figure 1.7.2: Measurement of Length

The image is a diagram illustrating the relative lengths of various units of measurement, including meters, yards, centimeters, and inches. At the top of the diagram, a horizontal line labeled "(a) 1 meter" represents a one-meter length. Below it, another horizontal line labeled "(b) 1 yard" represents a yard, which is slightly shorter than the meter. These two lines are aligned horizontally for comparison, with the meter line extending further to the right than the yard line.

Below these two lines, there is a much shorter line labeled "(c) 1 centimeter," demonstrating the relative smallness of a centimeter compared to a meter and a yard. The centimeter line is visually isolated and magnified for clarity in subsequent portions of the diagram.

On the bottom left, a zoomed-in section labeled "(d)" shows a ruler divided into smaller intervals. This ruler highlights the relationship between centimeters and inches. Each centimeter is marked and labeled, with "1 cm" and "2 cm" indicated. Above the ruler, a horizontal line labeled "1 inch" is shown as slightly longer than 2.54 centimeters. A dashed line connects this zoomed-in section to the centimeter measurement in the top portion of the diagram.

On the bottom right, another ruler represents larger measurements. This ruler is marked in inches, with labels at intervals such as "10," "20," "30," and so on, up to "100." These markings also include "1 ft," "2 ft," and "3 ft," indicating one, two, and three feet, respectively. The ruler spans 36 inches, equivalent to three feet, visually connecting the inch and foot measurements to the earlier yard and meter comparisons.

A note at the bottom of the diagram provides the conversion factors: "1 in. = 2.54 cm" and "1 m = 1.094 yd = 39.36 inches." These annotations offer numerical context for the visual comparisons, emphasizing the relationships between metric and imperial measurement systems. The diagram uses clean lines and clear labels to facilitate understanding of how these units relate in size and scale.

Figure 1.7.3: Units of volume

The image visually demonstrates the relationship between different units of volume: cubic meters, cubic decimeters, and cubic centimeters. It also provides a real-world comparison using a U.S. dime.

On the left, a large cube is shown, subdivided into smaller, equal-sized cubic units. The entire cube represents a volume of 1 cubic meter (1 m^3) and is labeled "(a)" at the bottom. A single smaller cube, representing 1 cubic decimeter (1 dm^3 or 1 liter), is highlighted at the corner of the large cube. An arrow extends from the smaller cube, pointing to an enlarged view of the cube outside the large cube, emphasizing its size relative to the entire structure.

From the second cube, another arrow points to a much smaller cube, representing 1 cubic centimeter (1 cm^3 or 1 milliliter). This illustrates the relationship that 1 dm^3 contains 1,000 cm^3.

On the right side of the image, labeled "(b)," the smallest cube is depicted again, accompanied by the text "1 cubic centimeter = 1 milliliter." Below it, a U.S. dime is shown for scale, with a label indicating that its diameter is 1.8 centimeters. This comparison provides a tangible reference for the size of a cubic centimeter.

The arrows and labeled volumes emphasize the hierarchical relationship between these units of volume, illustrating how larger cubic measurements are composed of smaller ones. The inclusion of the dime as a reference highlights the practical scale of a

cubic centimeter in everyday terms. This visualization effectively bridges abstract measurements with real-world understanding.

Figure 1.8.2: Concept Map for Conversions

The image demonstrates the process of converting units from meters (m) to centimeters (cm) using a conceptual flow diagram. On the left, there is a rectangular box labeled "Meters (m)," which represents the starting unit. An arrow points to the right, leading to another rectangular box labeled "Centimeters (cm)," which represents the target unit of conversion.

Below the arrow, a mathematical representation of the conversion is shown. It reads "100 cm / 1 m," indicating that one meter is equivalent to 100 centimeters. An explanatory note on the left side of the conversion equation states, "The unit you are converting from is on the bottom," referring to the placement of the original unit (meters) in the denominator of the fraction. On the right side of the equation, another explanatory note states, "The unit you are converting to is on the top," highlighting that the desired unit (centimeters) is placed in the numerator.

The diagram uses clear labels, directional arrows, and text annotations to visually and textually explain the process of unit conversion. It emphasizes the relationship between meters and centimeters and the proper placement of units in the conversion equation to achieve accurate results. The use of dfferent sizes for the boxes and the inclusion of guiding text enhances the clarity of the explanation.

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2: ELEMENTS, ATOMS, AND THE PERIODIC TABLE

Introduction

Chemistry is the study of matter, and matter is made up of compounds. Compounds can further be broken down to chemical elements that make them up. Elements can be thought to resemble an alphabet that make up a language. Just as any language is created from words that consist of the combination of letters, elements create the language of chemistry. They combine to make compounds. The chemical alphabet from which matter is formed is much larger than the one we use for spelling. Hence the role of elements is central to chemistry. They combine to form the millions and millions of compounds known to us. In this Chapter we will dive into the world of the elements, the atoms that make them up including the particles that make up the atom, the periodic table and its properties.

Learning Objectives

- Define what an element is and give examples of the abundance of different elements and their chemical symbols
- Explain how all matter is composed of atoms and which elements exist as diatomic molecules
- Describe the modern atomic theory and learn about the three main subatomic particles
- Describe how electrons are grouped within atoms
- Define and differentiate between the atomic number and the mass number of an element
- Explain how isotopes differ from one another
- Define atomic mass unit, atomic mass and calculate atomic mass
- Explain how elements are organized into the periodic table how some properties of elements relate to their positions on the periodic table
- Describe how electrons are grouped within atoms into shells, subshells, and orbitals using quantum numbers

2.1: Prelude to Elements, Atoms, and the Periodic Table

2.2: The Elements

2.3: The Atomic Theory

2.4: The Structure of Atoms

2.5: The Nuclei of Atoms

2.6: Atomic Masses

2.7: The Periodic Table

2.8: Arrangement of Electrons

2.9: Periodic Table Trends

2.10: Summary

2.11: Key Terms

2.12: Alternative Text for Chapter 2

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2.1: Prelude to Elements, Atoms, and the Periodic Table

Introduction

The hardest material in the human body is tooth enamel. It has to be hard so that our teeth can serve us for a lifetime of biting and chewing; however, tough as it is, tooth enamel is susceptible to chemical attack. Acids found in some foods or made by bacteria that feed on food residues on our teeth are capable of dissolving enamel. Unprotected by enamel, a tooth will start to decay, thus developing cavities and other dental problems.

In the early 1900s, a dentist in Colorado Springs, Colorado, noted that many people who lived in the area had brown-stained teeth that, while unsightly, were surprisingly resistant to decay. After years of study, excess fluorine compounds in the drinking water were discovered to be the cause of both these effects. Research continued, and in the 1930s, the US Public Health Service found that low levels of fluorine in water would provide the benefit of resisting decay without discoloring teeth.

The protective effects of fluorine have a simple chemical explanation. Tooth enamel consists mostly of a mineral called hydroxyapatite, which is composed of calcium, phosphorus, oxygen, and hydrogen. We know now that fluorine combines with hydroxyapatite to make fluorapatite, which is more resistant to acidic decay than hydroxyapatite is. Currently about 50% of the US population drinks water that has some fluorine added (in the form of sodium fluoride, NaF) to reduce tooth decay. This intentional fluoridation, coupled with the use of fluoride-containing toothpastes and improved oral hygiene, has reduced tooth decay by as much as 60% in children. The nationwide reduction of tooth decay has been cited as an important public health advance in history. (Another important advance was the eradication of polio.)

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2.2: The Elements

Introduction

An element is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. Figure 2.2.1 shows some of the chemical elements.



Figure 2.2.1: a) Scandium is the first transition metal and the first rare earth element (CC BY 3.0; Hi-Res Images of Chemical Elements, via Wikimedia Commons). b) Close-up photograph of sulfur showing its bright color and brittle, broken texture. (Public Domain; Benjah-bmm27, via Wikimedia Commons). c) Calcium is a very abundant element, which in nature above all occurs as calcium carbonate (CaCO3, lime) and calcium sulfate (CaSO4, gypsum). Elemental calcium is a grey metal, that slowly reacts with air and fiercely reacts with water. (CC BY 3.0; Hi-Res Images of Chemical Elements, via Wikimedia Commons). Detailed Description of Figure 2.2.1.

Abundance

The elements vary widely in abundance. In the universe as a whole, the most common element is hydrogen (about 90% of atoms), followed by helium (most of the remaining 10%). All other elements are present in relatively minuscule amounts, as far as we can detect.

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.2%. Hydrogen, the most abundant element in the universe, makes up only 0.14% of Earth's crust. Table 2.2.1 lists the relative abundances of elements on Earth as a whole and in Earth's crust.

Table 2.2.1: Elemental Composition of Earth

Earth's Crust Element	Earth's Crust Percentage	Earth (overall) Element	Earth (overall) Percentage
oxygen	46.1	iron	34.6
silicon	28.2	oxygen	29.5
aluminum	8.23	silicon	15.2
iron	5.53	magnesium	12.7

Earth's Crust Element	Earth's Crust Percentage	Earth (overall) Element	Earth (overall) Percentage
calcium	4.15	nickel	2.4
sodium	2.36	sulfur	1.9
magnesium	2.33	all others	3.7
potassium	2.09		
titanium	0.565		
hydrogen	0.14		
phosphorus	0.105		
all others	0.174		
Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, Internet Version 2005, CRC Press, Boca Raton, FL, 2005			

Table 2.2.2 lists the relative abundances of elements in the human body. If you compare Table 2.2.1 and Table 2.2.2, you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in Table 2.2.1; carbon is part of the 0.174% representing "other" elements.

Table 2.2.2: Elemental Composition of a Human Body

Element	Percentage by Mass
oxygen	61
carbon	23
hydrogen	10
nitrogen	2.6
calcium	1.4
phosphorus	1.1
sulfur	0.20
potassium	0.20
sodium	0.14
chlorine	0.12
magnesium	0.027
silicon	0.026
iron	0.006
fluorine	0.0037
zinc	0.0033
all others	0.174
Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, Internet Version 2005, CRC Press, Boca Raton, FL, 2005	

How does the human body concentrate so many apparently rare elements? The relative amounts of elements in the body have less to do with their abundances on Earth than with their availability in a form we can assimilate. We obtain oxygen from the air we breathe and the water we drink. We also obtain hydrogen from water. On the other hand, although carbon is present in the atmosphere as carbon dioxide, and about 80% of the atmosphere is nitrogen, we obtain those two elements from the food we eat, not the air we breathe.

Link to Learning: Phosphorus, The Chemical Bottleneck

There is an element that we need more of in our bodies than is proportionately present in Earth's crust, and *this* element is not easily accessible. Phosphorus makes up 1.1% of the human body but only 0.105% of Earth's crust. We need phosphorus for our bones and teeth, and it is a crucial component of all living cells. Unlike carbon, which can be obtained from carbon dioxide through photosynthesis, there is no phosphorus compound present in our surroundings that can serve as a convenient source. Phosphorus, then, is nature's bottleneck. Its availability limits the amount of life our planet can sustain.

Higher forms of life, such as humans, can obtain phosphorus by selecting a proper diet (plenty of protein); but lower forms of life, such as algae, must absorb it from the environment. When phosphorus-containing detergents were introduced in the 1950s, wastewater from normal household activities greatly increased the amount of phosphorus available to algae and other plant life. Lakes receiving this wastewater experienced sudden increases in growth of algae. When the algae died, concentrations of bacteria that ate the dead algae increased. Because of the large bacterial concentrations, the oxygen content of the water dropped, causing fish to die in large numbers. This process, called *eutrophication*, is considered a negative environmental impact.

Today, many detergents are made without phosphorus so the detrimental effects of eutrophication are minimized. You may even see statements to that effect on detergent boxes. It can be sobering to realize how much impact a single element can have on life—or the ease with which human activity can affect the environment.

Names and Symbols

Each element has a name. Some of these names date from antiquity, while others are quite new. Today, the names for new elements are proposed by their discoverers but must be approved by the International Union of Pure and Applied Chemistry (IUPAC), an international organization that makes recommendations concerning all kinds of chemical terminology.

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter chemical symbol. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. Table 2.2.3 lists the names and symbols of some of the most familiar elements. In the table *the symbol comes from the Latin name of element.

Table 2.2.3: Element Names and Symbols

Elements	Symbols	Elements	Symbols
aluminum	Al	magnesium	Mg
argon	Ar	manganese	Mn
arsenic	As	mercury	Hg*
barium	Ba	neon	Ne
bismuth	Bi	nickel	Ni
boron	B	nitrogen	N
bromine	Br	oxygen	O
calcium	Ca	phosphorus	P
carbon	C	platinum	Pt
chlorine	Cl	potassium	K*
chromium	Cr	silicon	Si
copper	Cu*	silver	Ag*
fluorine	F	sodium	Na*

Elements	Symbols	Elements	Symbols
gold	Au*	strontium	Sr
helium	He	sulfur	S
hydrogen	H	tin	Sn*
iron	Fe	tungsten	W†
iodine	I	uranium	U
lead	Pb*	zinc	Zn
lithium	Li	zirconium	Zr

Element names in languages other than English are often close to their Latin names. For example, gold is *oro* in Spanish and *or* in French (close to the Latin *aurum*), tin is *estaño* in Spanish (compare to *stannum*), lead is *plomo* in Spanish and *plomb* in French (compare to *plumbum*), silver is *argent* in French (compare to *argentum*), and iron is *fer* in French and *hierro* in Spanish (compare to *ferrum*). The closeness is even more apparent in pronunciation than in spelling.

Example 2.2.1: Element Names and Symbols

Write the chemical symbol for each element without consulting Table 2.2.3 "Element Names and Symbols".

a. bromine
b. boron
c. carbon
d. calcium
e. gold

Answer a available at the end of this section.

Example 2.2.2: Element Names and Symbols

Write the chemical symbol for each element without consulting Table 2.2.3.

a. manganese
b. magnesium
c. neon
d. nitrogen
e. silver

Solutions

Answer a available at the end of this section.

Example 2.2.3-Element Names and Symbols

What element is represented by each chemical symbol?

a. Na
b. Hg
c. P
d. K
e. I

Solutions

Answer a available at the end of this section.

Example 2.2.4-Element Names and Symbols

What element is represented by each chemical symbol?

a. Pb
b. Sn
c. U
d. O
e. F

Solutions

Answer a available at the end of this section.

Example 2.2.5-Concept Review

1. What is an element?
2. Give some examples of how the abundance of elements varies.
3. Why are chemical symbols so useful? What is the source of the letter(s) for a chemical symbol?

Solutions

1. An element is the basic chemical building block of matter; it is the simplest chemical substance.
2. Elements vary from being a small percentage to more than 30% of the atoms around us.
3. Chemical symbols are useful to concisely represent the elements present in a substance. The letters usually come from the name of the element.

Answer available at the end of this section.

Key Takeaways

- All matter is composed of elements.
- Chemical elements are represented by a one- or two-letter symbol.

Answers

Example 2.2.1: Element Names and Symbols

Br
B
C
Ca
Au

Example 2.2.2: Element Names and Symbols

Mn
Mg
Ne
N
Ag

Example 2.2.3-Element Names and Symbols

sodium
mercury
phosphorus
potassium

iodine

Example 2.2.4**-Element Names and Symbols**

lead
tin
uranium
oxygen
fluorine

Example 2.2.5**-Concept Review**

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2.3: The Atomic Theory

Introduction

Although no one has actually seen the inside of an atom, experiments have demonstrated much about atomic structure. Take some aluminum foil. Cut it in half. Now you have 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 you take this exercise, at least in theory? Can you 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.)

The Modern Atomic Theory

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 2.3.1), is a fundamental concept that states that all elements are composed of atoms. 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.*

These concepts form the basis of chemistry. 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).

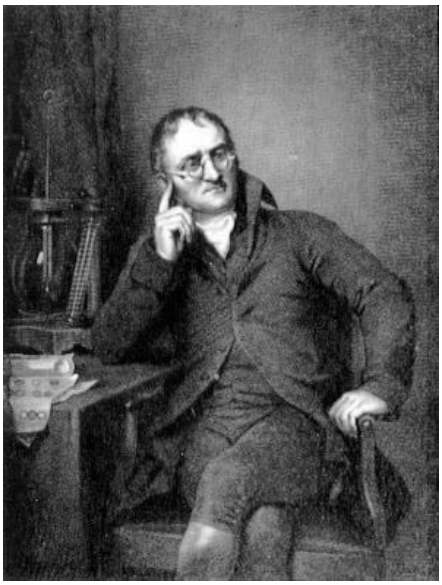


Figure 2.3.1: John Dalton was an English scientist who enunciated the modern atomic theory. (CC BY 4.0; Wellcome Collection via Wikimedia Commons). Detailed Description of Figure 2.3.1.

Looking Closer: Atomic Theory

Dalton's ideas are called the *modern* atomic theory because the concept of atoms is very old. The Greek philosophers Leucippus and Democritus originally introduced atomic concepts in the fifth century BC. (The word *atom* comes from the Greek word *atomos*, which means "indivisible" or "uncuttable.") Dalton had something that the ancient Greek philosophers didn't have, however; he had experimental evidence, such as the formulas of simple chemicals and the behavior of gases. In the 150 years or so before Dalton, natural philosophy had been maturing into modern science, and the scientific method was being used to study nature. So when Dalton announced a modern atomic theory, he was proposing a fundamental theory to describe many previous observations of the natural world; he was not just participating in a philosophical discussion.

Representation of Atoms

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron (Fe) metal is composed, microscopically, of individual atoms. 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 and their representations are:

- hydrogen (H_2)
- nitrogen (N_2)
- fluorine (F_2)
- oxygen (O_2)
- iodine (I_2)
- chlorine (Cl_2) and
- bromine (Br_2)

> Memorizing the diatomic molecules is worthwhile for our future endeavors. A mnemonic device to help in the memorization of the diatomic elements is as follows: **H**ave **N**o **F**ear **O**f **I**ce **C**old **B**eer. (hydrogen, nitrogen, fluorine, oxygen, iodine, chlorine and bromine).

A few other elements can exist as 3 and 4 atom molecules, and the most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur:

- ozone (O_3)
- phosphorus (P_4)
- sulfur (S_8)

There are different ways to represent atoms and molecules. We can use various models to showcase how atoms combine to make molecules. Figure 2.3.2 shows different model representations of S_8. Figure 2.3.2 (a) shows eight sulfur atoms, symbolized with the letter S, that are bonded to each other to form an octagon. Figure 2.3.2 (b) shows a 3-D, ball-and-stick model of the arrangement of the sulfur atoms. The shape is clearly not octagonal as it is represented in the structural formula. Figure 2.3.2 (c) is a space-filling model that shows each sulfur atom is partially embedded into the sulfur atom it bonds with.

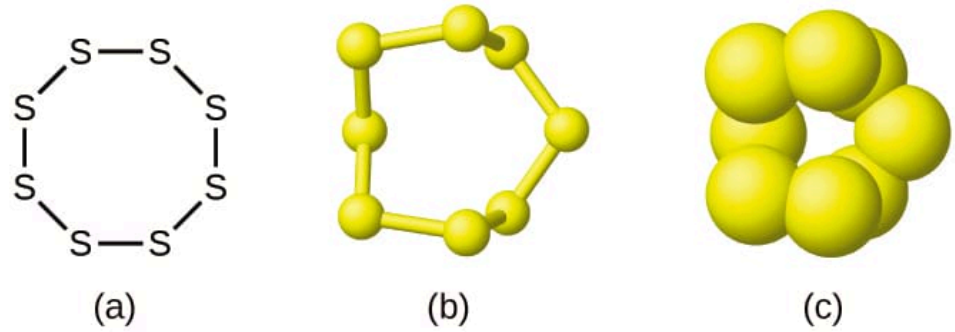


Figure 2.3.2: A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres. (CC BY 4.0; Marion Martin and Ana Ison via Introductory Chemistry-CH111). Detailed Description of Figure 2.3.2.

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and 2H represent distinctly different species.

- H_2 is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together.
- The expression 2H, on the other hand, indicates two separate hydrogen atoms that are not combined as a unit.
- The expression $2H_2$ represents two molecules of diatomic hydrogen.

Figure 2.3.3 shows four diagrams. The diagram for H shows a single, white sphere and is labeled one H atom. The diagram for 2H shows two white spheres that are not bonded together. It is labeled 2H atoms. The diagram for H subscript 2 shows two white spheres bonded together. It is labeled one H subscript 2 molecule. The diagram for 2H subscript 2 shows two sets of bonded, white spheres. It is labeled 2H subscript 2 molecules.

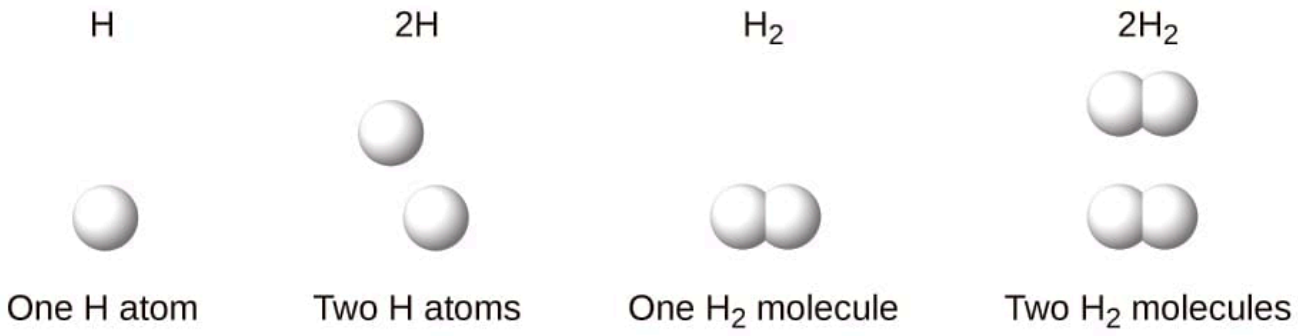


Figure 2.3.3: The symbols H, 2H, H_2, and $2H_2$ represent very different entities. (CC BY 4.0; Marion Martin and Ana Ison via Introductory Chemistry-CH111). Detailed Description of Figure 2.3.3.

Example 2.3.1

Write the chemical formula of the following elements:

1. oxygen

2. carbon
3. potassium

Solution

1. Oxygen is one of the seven diatomic molecular elements. Its chemical formula is O_2.
2. Carbon is monatomic, hence its formula is C.
3. Potassium is monatomic hence its formula is K.

Key Takeaways

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- The diatomic elements are hydrogen, nitrogen, fluorine, oxygen, iodine, chlorine and bromine.

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2.4: The Structure of Atoms

Subatomic Particles

If matter is composed of atoms, what are atoms composed of? Are they the smallest particles, or is there something smaller? There have been several minor but important modifications to Dalton's atomic theory that shed light to the inner structure of atoms answering just these questions. For one thing, Dalton considered atoms to be indivisible. We know now that atoms not only can be divided, but also are composed of three different kinds of particles, called **subatomic particles**, with their own properties that are different from the chemical properties of atoms.

It was learned that an atom contains a very small **nucleus** composed of positively charged **protons** and uncharged **neutrons,** surrounded by a much larger volume of space containing negatively charged **electrons.**

The Electron

The first subatomic particle was identified in in the late 1800s. A number of scientists who were interested in the structure of the atom investigated the electrical discharges that could be produced in low-pressure gases. The most significant discovery was made by an English physicist, J. J. Thomson, using a cathode ray tube in a magnetic field. This apparatus consisted of a sealed glass tube from which almost all the air had been removed; the tube contained two metal electrodes. When high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them. This beam was deflected toward the positive charge and away from the negative charge, and the result did not change when different metals were used for the electrodes. His experiments with magnetic fields further revealed the existence of an extremely tiny subatomic particle which was negatively charged. As a result of these experiments, Thomson was able to calculate the charge-to-mass ratio of the cathode ray particles. This subatomic particle was later called the electron, by Irish physicist George Stoney, derived from "*electr*ic i*on*.".

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

Since the charge of an electron was now known due to Millikan's research, and the charge-to-mass ratio was already known due to Thomson's research it was very easy to calculate the mass of an electron, and it was determined that it had a mass of about 9.109×10^{-31} kg.

The Nucleus

Scientists had now established that the atom was not indivisible as Dalton had believed, and due to the work of Thomson, Millikan, and others, the charge and mass of the negative subatomic particles—the electrons—were known.

Between 1909 and 1911, Ernest Rutherford, a Cambridge physicist, and his associates Hans Geiger and Ernest Marsden performed experiments that provided strong evidence concerning the internal structure of an atom. They took a very thin metal foil, such as gold or platinum, and aimed a beam of positively charged particles (called alpha particles, which are combinations of two protons and two neutrons) from a radioactive source toward the foil. Surrounding the foil was a detector—either a scintillator (a material that glows when hit by such particles) or some unexposed film (which is exposed where the particles hit it). The detector allowed the scientists to determine the distribution of the alpha particles after they interacted with the foil. Figure $2.4.1$ shows a diagram of the experimental setup. The experiment showed that most particles passed right through the foil without being deflected at all. However, some were diverted slightly, and a very small number were deflected almost straight back toward the source.

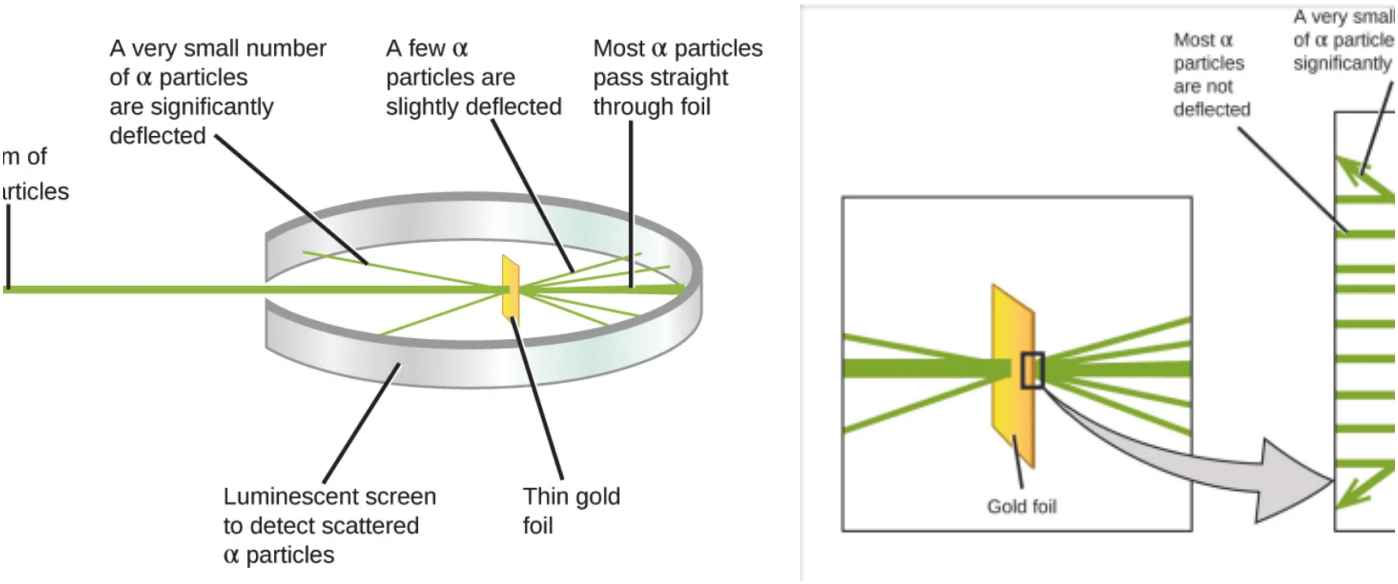


Figure 2.4.1: a) Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. b) Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected.(CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson via OpenStax). Detailed Description of Figure 2.4.1.

Analyzing a series of such experiments in detail, Rutherford drew two conclusions:

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

This analysis led Rutherford to propose a model in which an atom consists of a very small, positively charged nucleus, in which most of the mass of the atom is concentrated, surrounded by the negatively charged electrons, so that the atom is electrically neutral. Protons and neutrons are concentrated in a central region he called the nucleus (plural, *nuclei*) of the atom. Electrons are outside the nucleus and orbit about it because they are attracted to the positive charge in the nucleus. Most of the mass of an atom is in the nucleus, while the orbiting electrons account for an atom's size. As a result, an atom consists largely of empty space. Rutherford called his description the "planetary model" of the atom, Figure 2.4.2.

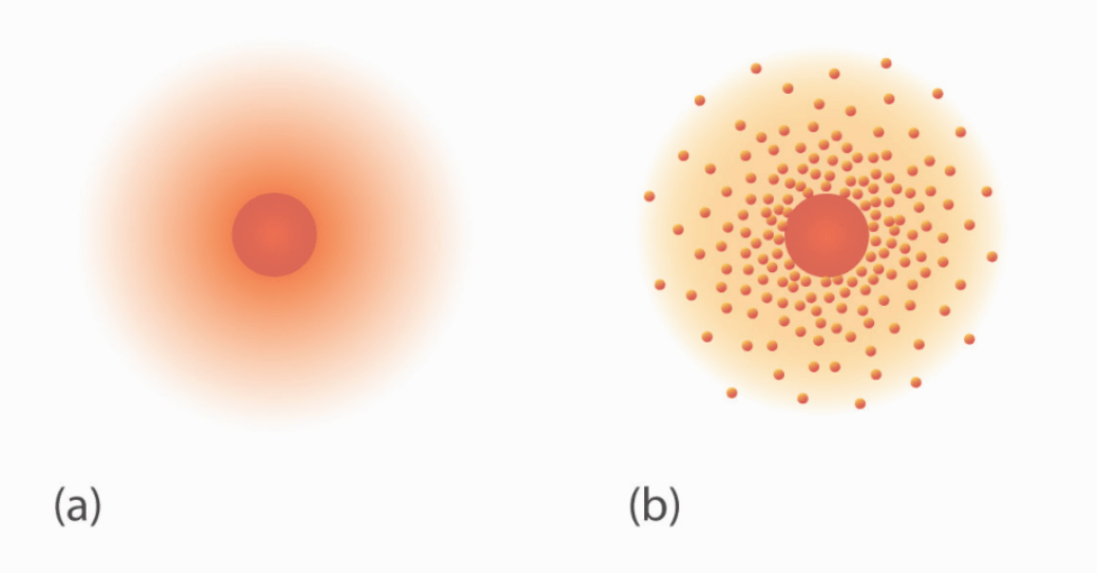


Figure 2.4.2: A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. In both (a) and (b), the nucleus is in the center of the diagram. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 2.4.2

Isotopes

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

The Proton and the Neutron

By 1920, experimental evidence indicated the existence of a second particle. Rutherford also discovered that the nuclei of other elements contain the hydrogen nucleus as a "building block," and he named this more fundamental particle the proton, the positively charged, subatomic particle found in the nucleus. A proton has the same amount of charge as an electron, but its charge is positive, not negative. Another major difference between a proton and an electron is mass. Although still incredibly small, the mass of a proton is 1.673×10^{-27} kg, which is almost 2,000 times greater than the mass of an electron. Because opposite charges attract each other (while like charges repel each other), protons attract electrons (and vice versa).

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

Key Takeaways

- Atoms are composed of three main subatomic particles: protons, neutrons, and electrons.
- Protons and neutrons are grouped together in the nucleus of an atom, while electrons orbit about the nucleus.
- Isotopes differ in mass because they have different numbers of neutrons, but they are chemically identical because they have the same number of protons.

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2.5: The Nuclei of Atoms

Introduction

We understand now that all atoms can be broken down into subatomic particles: protons, neutrons, and electrons. Table 2.5.1 lists some of their important characteristics and the symbols used to represent each particle.

Table 2.5.1: Properties of the Subatomic Particles

Particle	Symbol	Mass (kg)	Relative Mass (proton = 1)	Relative Charge
proton	p^+	1.673×10^{-27}	1	+1
neutron	n^0	1.675×10^{-27}	1	0
electron	e^-	9.109×10^{-31}	0.00055	–1

Now that we know how atoms are generally constructed, what do atoms of any particular element look like? How many protons, neutrons, and electrons are in a specific kind of atom? First, if an atom is electrically neutral overall, then the number of protons equals the number of electrons. Because these particles have the same but opposite charges, equal numbers cancel out, producing a neutral atom.

Atomic Number

The number of protons in the nucleus of an atom is its **atomic number (Z)**.

> Atomic Number (Z) is the defining trait of an element: Its value determines the identity of the atom.

For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons (p^+) equals the number of electrons (e^-). Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number (A)**. The number of neutrons is therefore the difference between the mass number and the atomic number.

- number of protons (p^+) = number of electrons (e^-)
- atomic number (Z) = number of protons = p^+
- mass number (A) = number of protons (p^+) + number of neutrons (n^0)
- A–Z = number of neutrons = (n^0)

For example, hydrogen has the atomic number 1; all hydrogen atoms have 1 proton in their nuclei. Helium has the atomic number 2; all helium atoms have 2 protons in their nuclei. There is no such thing as a hydrogen atom with 2 protons in its nucleus; a nucleus with 2 protons would be a helium atom. The atomic number *defines* an element. The largest atoms have over 100 protons in their nuclei. A **chemical symbol** is an abbreviation that we use to indicate an element or an atom of an element. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112. Table 2.5.1 lists some common elements and their atomic numbers as well as their chemical symbols.

Table 2.5.1: Some Common Elements and Their Atomic Numbers

Element	Atomic Number	Element	Atomic Number	Element	Atom
bromine (Br)	35	nitrogen (N)	7	lithium (Li)	
calcium (Ca)	20	oxygen (O)	8	fluorine (F)	
barium (Ba)	56	neon (Ne)	10	iron (Fe)	
gallium (Ga)	31	sodium (Na)	11	chlorine (Cl)	
aluminum (Al)	13	magnesium (Mg)	12	argon (Ar)	
carbon (C)	6	phosphorus (P)	15	cesium (Cs)	
helium (He)	2	sulfur (S)	16	iodine (I)	
chromium (Cr)	24	potassium (K)	19	mercury (Hg)	
hydrogen (H)	1	titanium (Ti)	22	radon (Rn)	
zirconium (Zr)	40	manganese (Mn)	25	uranium (U)	

Ions

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

Atomic charge = number of protons (p^+) - number of electrons (e^{-1})

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

Isotopes

Atoms of the same element always have the same number of protons, same *Z*, but often have different numbers of neutrons, therefore, different mass numbers. These atoms are called isotopes. For example, 99% of the carbon atoms on Earth have 6 neutrons and 6 protons in their nuclei; about 1% of the carbon atoms have 7 neutrons in their nuclei. Naturally occurring carbon on Earth, therefore, is actually a mixture of isotopes, albeit a mixture that is 99% carbon with 6 neutrons in each nucleus, and only 1% carbon with 7 neutrons in each nucleus.

Another An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus.

Most elements exist as mixtures of isotopes. In fact, there are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. As we discussed, the mass number of an atom is the sum of the numbers of protons and neutrons in the nucleus. Given the mass number for a nucleus (and knowing the atomic number of that particular atom), you can determine the number of neutrons by subtracting the atomic number from the mass number.

A simple way of indicating the mass number of a particular isotope is to list it as a superscript on the left side of an element's symbol. Atomic numbers are often listed as a subscript on the left side of an element's symbol (Figure 2.5.2).

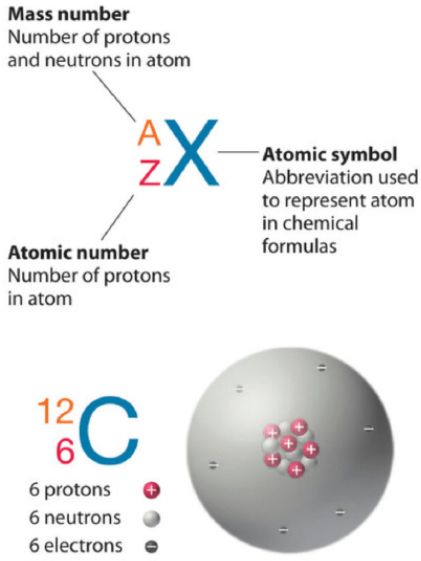


Figure 2.5.1: The atomic symbol, X, is the abbreviation used to represent an atom in chemical formulas. The mass number, A, is the number of protons and neutrons in the atom, which is to the left of the X. The atomic number, Z, is the number of protons in the atom, which is to the left of the X and below the A. Below is the example for the Carbon atom. (CC BY-NC-SA; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 2.5.1.

Example 2.5.1

What is the number of protons in the nucleus of each element?

a. aluminum
b. iron

Solution

Answer a available at the end of this section.

Example 2.5.2

How many electrons are present in the atoms of each element?

a. sulfur
b. tungsten
c. argon

Solution

How many electrons are in an atom? Previously we said that for an electrically neutral atom, *the number of electrons equals the number of protons*, so the total opposite charges cancel. Thus, the atomic number of an element also gives the number of electrons in an atom of that element. According to Table 2.5.1:

Answer a available at the end of this section.

Key Takeaways

- Each element is identified by its atomic number. The atomic number provides the element's location on the periodic table
- The isotopes of an element have different masses and are identified by their mass numbers.

Answers

Example 2.5.1

According to Table 2.5.1, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.

Iron has an atomic number of 26. Therefore, every iron atom has 26 protons in its nucleus.

Example 2.5.2

The atomic number of sulfur is 16. Therefore, in a neutral atom of sulfur, there are 16 electrons.

The atomic number of tungsten is 74. Therefore, in a neutral atom of tungsten, there are 74 electrons.

The atomic number of argon is 18. Therefore, in a neutral atom of argon, there are 18 electrons.

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2.6: Atomic Masses

Atomic Mass Unit

Even though atoms are very tiny pieces of matter, they have mass. Their masses are so small, however, that chemists often use a unit other than grams to express them—the atomic mass unit.

The atomic mass unit (abbreviated u, although amu is also used) is defined as 1/12 of the mass of a ^{12}C atom:

$$1\ \text{u} = \frac{1}{12}\ \text{the mass of}\ ^{12}\text{C atom} \tag{2.6.1}$$

It is equal to 1.661×10^{-24} g.

Masses of other atoms are expressed with respect to the atomic mass unit. For example, the mass of an atom of ^{1}H is 1.008 u, the mass of an atom of ^{16}O is 15.995 u, and the mass of an atom of ^{32}S is 31.97 u. Note, however, that these masses are for particular isotopes of each element. Because most elements exist in nature as a mixture of isotopes, any sample of an element will actually be a mixture of atoms having slightly different masses (because neutrons have a significant effect on an atom's mass). How, then, do we describe the mass of a given element? By calculating an average of an element's atomic masses, weighted by the natural abundance of each isotope, we obtain a weighted average mass called the atomic mass (also commonly referred to as the *atomic weight*) of an element.

Atomic Mass is the Weighted Average Mass of Isotopes

As stated above, most elements occur naturally as a mixture of two or more isotopes. Listed below (Table 2.6.1) are the naturally occurring isotopes of selected elements along with the percent natural abundance of each.

Table 2.6.1: Atomic Masses and Percent Abundances of Some Natural Isotopes

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Hydrogen	$^{1}_{1}H$	99.985	1.0078	1.0079
	$^{2}_{1}H$	0.015	2.0141	
	$^{3}_{1}H$	negligible	3.0160	
Carbon	$^{12}_{6}C$	98.89	12.000	12.011
	$^{13}_{6}C$	1.11	13.003	
	$^{14}_{6}C$	trace	14.003	
Oxygen	$^{16}_{8}O$	99.759	15.995	15.999
	$^{17}_{8}O$	0.037	16.995	
	$^{18}_{8}O$	0.204	17.999	
Chlorine	$^{35}_{17}Cl$	75.77	34.969	35.453
	$^{38}_{17}Cl$	24.23	36.966	
Copper	$^{63}_{29}Cu$	69.17	62.930	63.546
	$^{65}_{29}Cu$	30.83	64.928	

For some elements, one particular isotope is much more abundant than any other isotopes. For example, naturally occurring hydrogen is nearly all hydrogen-1, and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in substantial quantities. Chlorine (atomic number 17) is yellowish-green toxic gas. About three

quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise atomic masses, you would get approximately 36.

$$\frac{34.969\,u + 36.966\,u}{2} = 35.968\,u$$

As you can see, the average atomic mass given in the last column of the table above (35.453) is significantly lower. Why? The reason is that we need to take into account the natural abundance percentages of each isotope in order to calculate what is called the **weighted average.**

> The **atomic mass** *of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element.*
>
> $$0.7577\,(34.969\,u) + 0.2423\,(36.966\,u) = 35.453\,u$$

The weighted average is determined by multiplying the percent of natural abundance by the actual mass of the isotope. This is repeated until there is a term for each isotope. For chlorine, there are only two naturally occurring isotopes so there are only two terms.

Atomic mass = (%1)(mass 1) + (%2)(mass 2) + ⋯

Another example: oxygen exists as a mixture that is 99.759% ^{16}O, 0.037% ^{17}O and 0.204% ^{18}O. The atomic mass of oxygen (use percent natural abundance data from Table 2.5.1) would be calculated as follows:

Atomic mass = (%1)(mass 1) + (%2)(mass 2) + (%3)(mass 3)

$$0.99759\,(15.995u) + 0.00037\,(16.995u) + 0.00204\,(17.999u) = 15.999u$$

To confirm your answer, compare the calculated value to the weighted mass displayed on the periodic table.

Example 2.6.1

Calculate the atomic mass of oxygen. Oxygen exists as a mixture of 3 isotopes. Their respective masses and natural abundance are shown below.

- ^{16}O: 15.995 u (99.759%)
- ^{17}O: 16.995 u (0.037%)
- ^{18}O: 17.999 u (0.204%)

Solution

Multiply the isotope abundance by the actual mass of the isotope, and then sum up the products.

$$0.99759\,(15.995\,u) + 0.00037\,(16.995\,u) + 0.00204\,(17.999\,u) = 15.999\,u$$

The atomic mass of each element is found under the element symbol in the **Periodic Table.** The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. Examples are shown below. The atomic mass of tin (Sn) is 118.71 u while the atomic mass of carbon (C) is 12.011 u. On the other hand, the atomic number (Z) of each element is found above the atomic symbol (Figure 2.6.1). We will learn about the periodic table in detail in the next section.

Link to Learning

Use the Periodic Table of Elements with Atomic Mass from PubChem to look up each element's symbol, atomic number, and atomic mass while completing this activity.

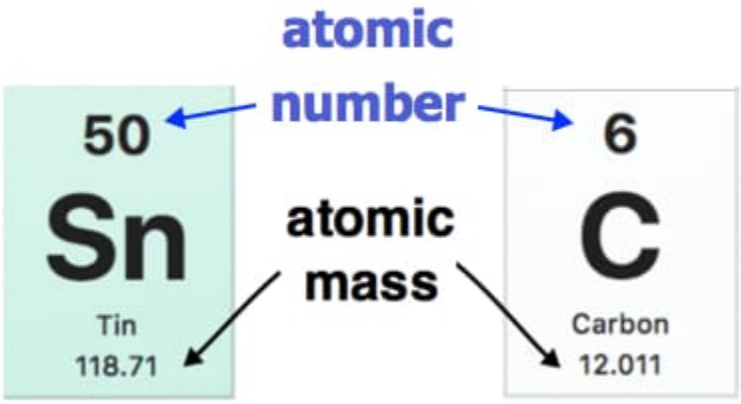


Figure 2.6.1: Atomic mass indicated on entries of the Periodic Table. (Public Domain; anonymous via Pubchem). Detailed Description of Figure 2.6.1.

Example 2.6.2: Mass of Carbon

What is the average mass of a carbon atom in grams? The atomic mass is found in the Periodic Table. Please use two decimal places.

Solution

This is a simple one-step conversion, similar to conversions we did in **Chapter 1.** We use the fact that 1 u = 1.661×10^{-24} g:

$$12.01 \cancel{\text{u}} \times \frac{1.661 \times 10^{-24}\ \text{g}}{1\ \cancel{\text{u}}} = 1.995 \times 10^{-23}\ \text{g}$$

This is an extremely small mass, which illustrates just how small individual atoms are.

Key Takeaways

- Atoms have a mass that is based largely on the number of protons and neutrons in their nucleus.
- The atomic mass of each element in the Periodic Table is the weighted average of the mass of all its isotopes.

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2.7: The Periodic Table

Introduction

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. 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.

The Periodic Table

The **periodic table** is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is illustrated in Figure 2.7.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's chemical symbol. The elements are listed in order of atomic number.

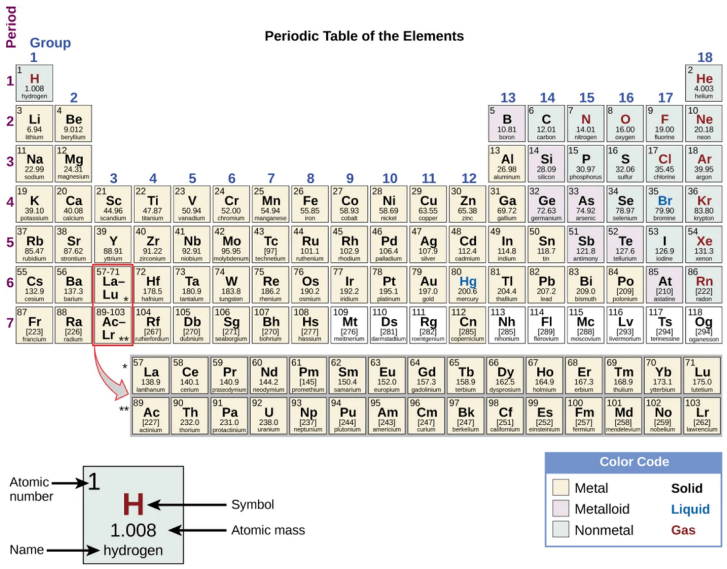


Figure 2.7.1: Elements in the periodic table are organized by atomic number, with groups, periods, and separate lanthanide and actinide rows. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson via OpenStax). Detailed Description of Figure 2.7.1.

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).

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.

Metals, Nonmetals, and Metalloids

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 (semimetal) , as shown in Figure 2.7.2.

- 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.
- A **metalloid** is an element with properties intermediate between those of metals and nonmetals. They are sometimes called are called **semimetals**. Elements adjacent to the bold zigzag line in the right-hand portion of the periodic table have semimetal properties.

As illustrated in Figure 2.7.2, 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. Elements adjacent to the bold zigzag line in the right-hand portion of the periodic table have semimetal properties.

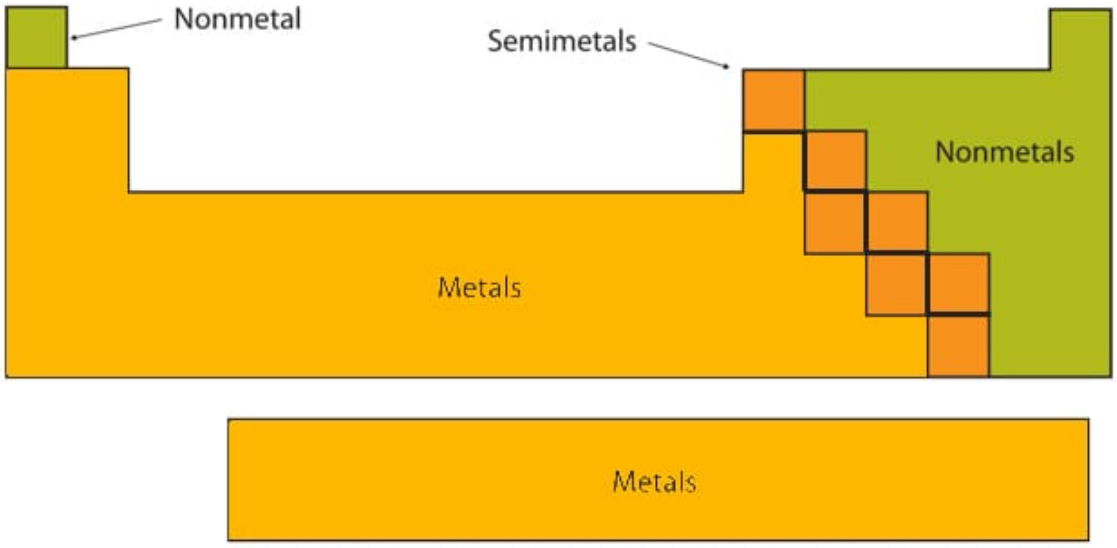


Figure 2.7.2: Elements are either metals, nonmetals, or semimetals. Each group is located in a different part of the periodic table. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 2.7.2.

Example 2.7.1

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. The atomic number of selenium is 34, which places it in period 4 and group 16. In Figure 2.7.2, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal. Note, however, that because selenium is close to the metal-nonmetal dividing line, it would not be surprising if selenium were similar to a semimetal in some of its properties.
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.

Representative, Transition, and Inner-transition

Another way to categorize the elements of the periodic table is shown in Figure 2.7.3. The first two columns on the left (groups 1 and 2) and the last six columns on the right (groups 13-19) are called the **main group elements**. The ten-column block between these columns (groups 3-12) 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 **lanthanides** and the **actinides**.

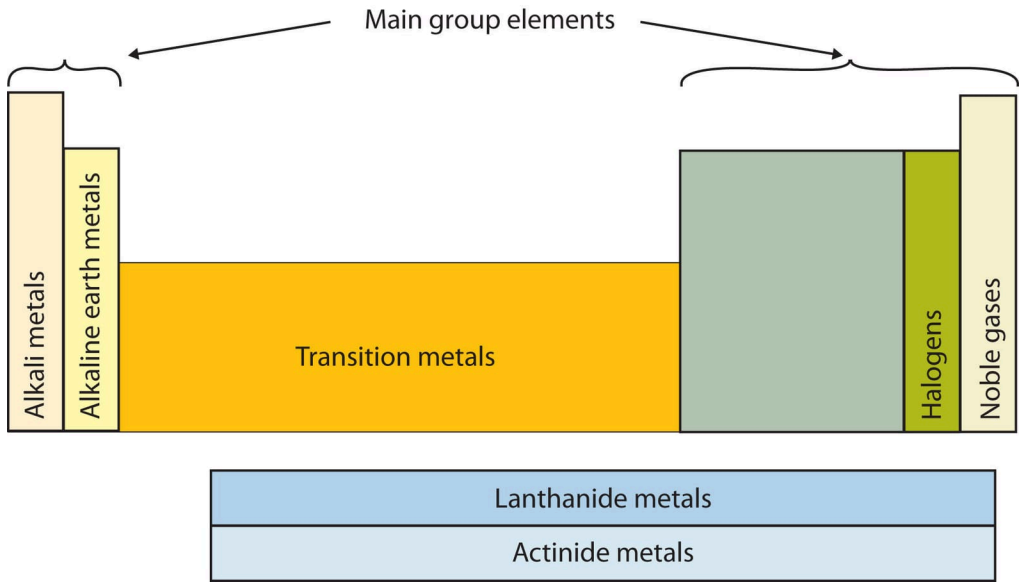


Figure 2.7.3: Some sections of the periodic table have special names. The elements lithium, sodium, potassium, rubidium, cesium, and francium are collectively known as alkali metals. (CC BY-NC-SA 3.0; anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 2.7.3.

Link to Learning: Transition Metals in the Body

Most of the elemental composition of the human body consists of main group elements. The most abundant non-main group element is iron, at 0.006 percentage by mass. Because iron has relatively massive atoms, it would appear even lower on a list organized in terms of percent by *atoms* rather than percent by mass.

Iron is a transition metal and the chemistry of iron makes it a key component in the proper functioning of red blood cells.

Red blood cells are cells that transport oxygen from the lungs to cells of the body and then transport carbon dioxide from the cells to the lungs. Without red blood cells, animal respiration as we know it would not exist. The critical part of the red blood cell is a protein called *hemoglobin*. Hemoglobin combines with oxygen and carbon dioxide, transporting these gases from one location to another in the body. Hemoglobin is a relatively large molecule, with a mass of about 65,000 u.

The crucial atom in the hemoglobin protein is iron. Each hemoglobin molecule has four iron atoms, which act as binding sites for oxygen. It is the presence of this particular transition metal in your red blood cells that allows you to use the oxygen you inhale.

Other transition metals have important functions in the body, despite being present in low amounts. Zinc is needed for the body's immune system to function properly, as well as for protein synthesis and tissue and cell growth. Copper is also needed for several proteins to function properly in the body. Manganese is needed for the body to metabolize oxygen properly. Cobalt is a necessary component of vitamin B-12, a vital nutrient. These last three metals are present in the body in very small quantities. However, even these small quantities are required for the body to function properly.

Key Takeaways

- All known elements are organized into the periodic table on the basis of their chemical properties.

- Elements that have similar chemical properties are grouped in columns called groups, and the rows of elements are called periods.
- Every element can be classified as either a metal, a nonmetal, or a metalloid.

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2.8: Arrangement of Electrons

Quantum-Mechanical Model of the Atom

You now know that the periodic table is arranged in groups and periods (columns and rows) based on chemical and physical properties of the different elements. The first element, hydrogen (Z=1) has one proton and one electron and as you move right across the rows, each subsequent element has one additional proton and electron. You may have asked yourself, why are periodic trends observed across the rows and down the groups? Or, why do the rows have different numbers of elements, giving the table a unique shape?

These questions can be answered by learning more about the electrons in atoms. Although we have discussed the general arrangement of subatomic particles in atoms, we have said little about how electrons occupy the space around the nucleus. Do they move around the nucleus at random, or do they exist in some ordered arrangement?

In 1913, the Danish scientist Niels Bohr suggested that the electron in a hydrogen atom could not have any random energy, having *only* certain fixed values of energy that were indexed by the number n (now called a **quantum number**). Bohr suggested that the energy of the electron in hydrogen was **quantized** because it was in a specific orbit; much like the steps on a staircase does not have half or quarter stairs or the keys on a piano don't have notes in between, there are no energy levels in between each orbit. Bohr's ideas were useful, but were applicable only to the hydrogen atom. Figure 2.8.1 shows a model of the atom which is based on Bohr's ideas.

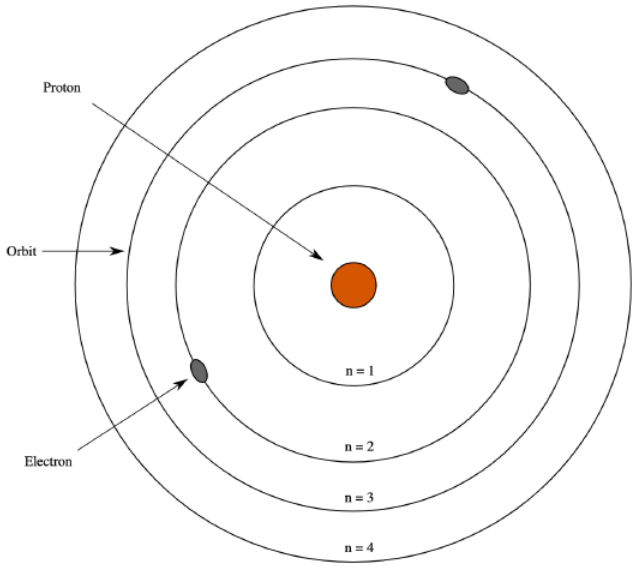


Figure 2.8.1: Bohr's description had specific orbits for the electron, which had quantized energies (CC BY-SA 3.0; Sharon Bewick via Wikimedia Commons). Detailed Description of Figure 2.8.1.

However, later researchers generalized Bohr's ideas into a new theory called **quantum mechanics**, which explains the behavior of electrons as if they were acting as a wave, not as particles. Quantum mechanics predicts two major things: quantized energies for electrons of all atoms (not just hydrogen) and an organization of electrons within atoms. Electrons are no longer thought of as being randomly distributed around a nucleus or restricted to certain orbits (in that regard, Bohr was wrong). Instead, electrons are collected into groups (*shells*) and subgroups (*subshells*) that explain much about the chemical behavior of the atom.Bohr's ideas were useful, but were applicable only to the hydrogen atom.

In the quantum-mechanical model of an atom, which is the modern and currently accepted model, the location of electrons in the atom are described by *four quantum numbers*, not just the one predicted by Bohr. Much like your home address can be used to locate you in a specific state, city, street, and house number, the first three quantum numbers identify approximately where electrons are in an atom. The fourth quantum number describes the electron and whether it is spin up or down (clockwise or counterclockwise).

The theory and mathematics behind these four quantum numbers are well beyond the scope of this textbook, however, it is useful to learn some of the basics in order to understand how atoms behave and interact with (react) with other atoms.

Electron Arrangements: Shells, Subshells, and Orbitals

Electrons are organized according to their energies into sets called **shells** (labeled by the principle quantum number, ***n***).

- Generally the higher the energy of a shell, the farther it is (on average) from the nucleus.
- Shells do not have specific, fixed distances from the nucleus, but an electron in a higher-energy shell will spend more time farther from the nucleus than does an electron in a lower-energy shell.

Shells are further divided into subsets of electrons called **subshells**, labeled by type as *s*, *p*, *d*, or *f*.

- The first shell has only one subshell, s.
- The second shell has two subshells, *s* and *p;*
- the third shell has three subshells, s, *p*, and *d,* and
- the fourth shell has four subshells, s, *p*, *d*, and *f*.

Within each subshell, electrons are arranged into different numbers of **orbitals**,

- an *s* subshell is made up of one *s* orbital,
- a *p* subshell has two *p* orbitals,
- a *d* subshell, five *d* orbitals, and
- an *f* subshell, seven *f* orbitals.
- Each orbital has a different shape and orientation around the nucleus (Figure $2.8.2$, however, rather than representing an orbit, as the name suggests, orbitals define a boundary for the region of space where a given electron is most likely to be found.

Lastly, a single orbital can hold up to *two electrons* each with a different **spin**.

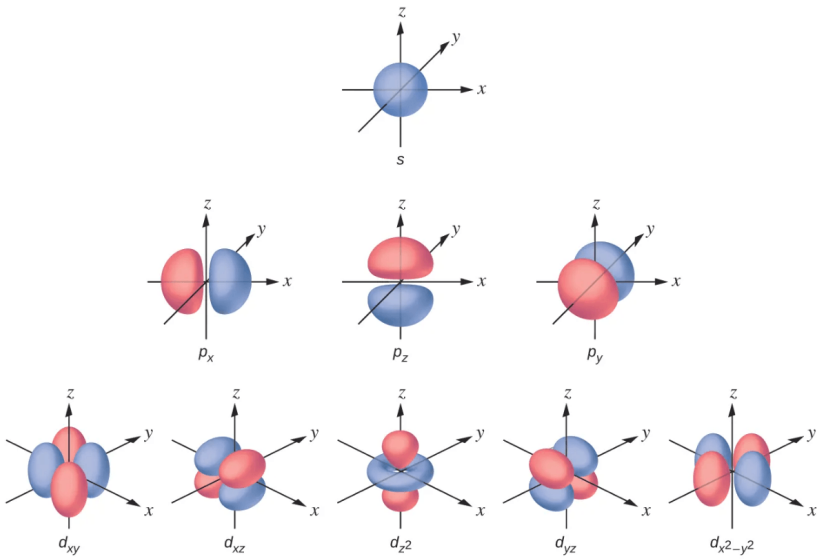


Figure $2.8.2$: (a) The lone s orbital in an s subshell is spherical in distribution. (b) The three *p* orbitals have two lobes, shaped kind of like dumbbells, each is oriented around the nucleus along a different axis. (c) The five *d* orbitals have four lobes, except for the d_{z^2} orbital, which is a "dumbbell + torus" combination. They are all oriented in different directions. (CC BY 4.0; Paul Flowers, William R. Robinson, PhD, Richard Langley, Klaus Theopold via OpenStax). Detailed Description of Figure $2.8.2$.

It is important to note that according to quantum theory, there are specific *allowed* combinations of quantum numbers and others that are not allowed. For example, shell two can only have two subshells, *s* with one orbital and *p* with 3 orbitals, therefore, this shell can hold a maximum of eight electrons (four orbitals times two electrons each). It takes practice to learn the allowed combinations as shown in Table $2.8.1$ but it is helpful to visualize the atom as a sphere with the nucleus in the center. Close to the nucleus, there is a smaller amount of space for electrons – a smaller shell. As the number of electrons increases, the shells

that hold the electrons get larger and thus further away from the nucleus.

Table 2.8.1*: Shells and Subshells*

Shell (*n*)	Number of Subshells	Names of Subshells	Number of Orbitals (*per Subshell*)	Number of Electrons (*per Subshell*)	Total Electrons (*per Shell*)
1	1	*1s*	1	2	2
2	2	*2s* and *2p*	1, 3	2, 6	8
3	3	*3s*, *3p,* and *3d*	1, 3, 5	2, 6, 10	18
4	4	*4s*, *4p*, *4d,* and *4f*	1, 3, 5, 7	2, 6, 10, 14	32

All of this information about the shell, subshell, and orbital is put together to make up the "address" for an electron and all of the addresses for all the electrons in an atom make up the **electron configuration.**

Electron Configurations

As mentioned above, the arrangement of electrons in the orbitals of an atom is called the electron configuration of the atom. We describe an electron configuration with a symbol that contains three pieces of information

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

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

To start, we use numbers to indicate which shell an electron is in. The first shell, closest to the nucleus and with the lowest-energy electrons, is shell 1. This first shell has only one subshell, which is labeled 1s and can hold a maximum of 2 electrons. We combine the shell and subshell labels when referring to the organization of electrons about a nucleus and use a superscript to indicate how many electrons are in a subshell.

A shell diagram shows the arrangement of electrons in an atom as dots placed on concentric circles around a central nucleus, with each circle representing an energy level or shell. The number of electrons on each shell reflects how electrons are distributed, helping to visualize valence electrons and atomic structure as seen in Figure 2.8.3.

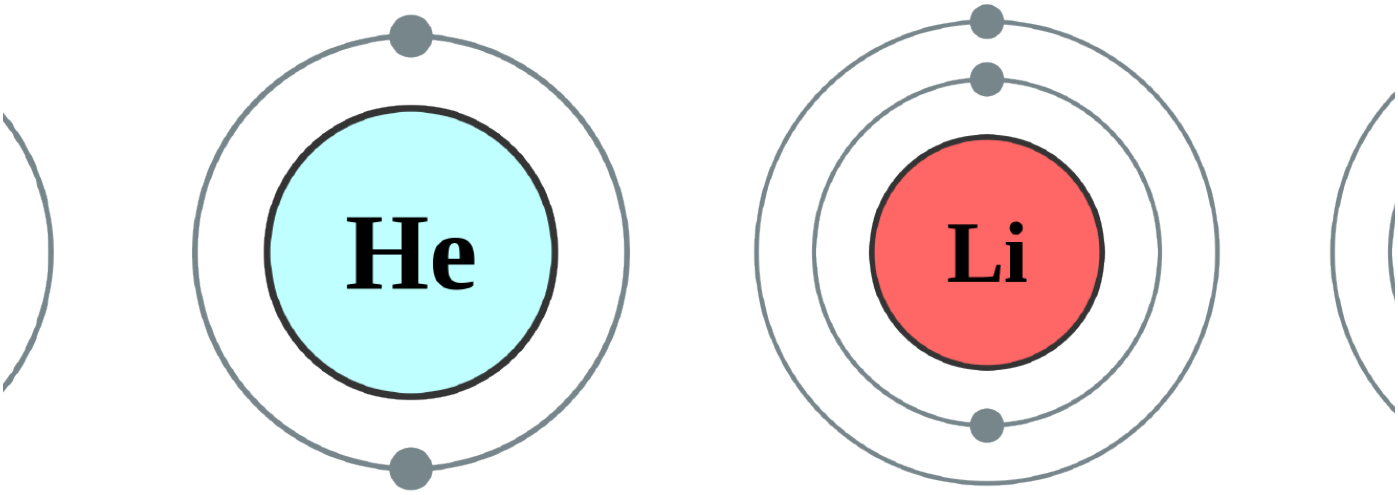


Figure 2.8.3: Shell Diagrams of Various Atoms. (CC BY-SA 2.0 UK; Greg Robson modified by Pumbaa via Wikipedia). Detailed Description of Figure 2.8.3.

Thus, because a hydrogen atom has its single electron in the *s* subshell of the first shell, we use $1s^1$ (spoken as “one-ess-one”) to describe the electron arrangement or distribution of electrons in hydrogen. This electron configuration is unique to hydrogen (Figure 2.8.4).

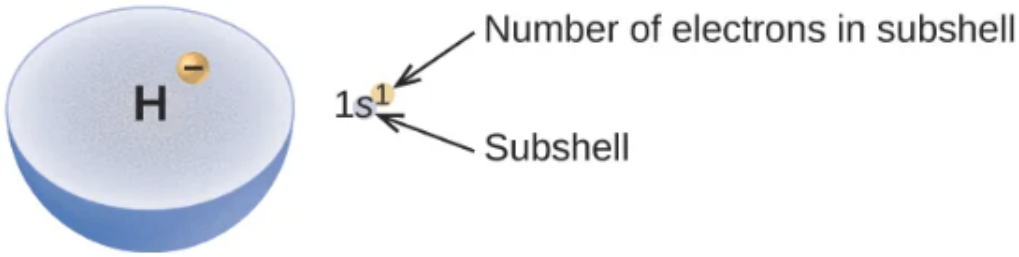


Figure 2.8.4: The Diagram specifies the subshell (*n* and *l* value, with letter symbol) and superscript number of electrons.. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson via OpenStax). Detailed Description of Figure 2.8.4.

Helium atoms have 2 electrons. Both electrons fit into the 1s subshell because *s* subshells contain one *s* orbital which can hold up to 2 electrons; therefore, the electron configuration for helium atoms is $1s^2$ (spoken as “one-ess-two”).

The 1s subshell can hold a maximum of 2 electrons, so the electron configuration for a lithium atom, which has three electrons, cannot be $1s^3$. Two of the lithium electrons can fit into the 1s subshell, but the third electron must go into the second shell and the lower energy orbital, which is the 2*s* orbital. Therefore, we write the electron configuration of a lithium atom as $1s^22s^1$ (spoken as “one-ess-two two-ess-one”).

Rules for Electron Configurations

There are a set of general rules that are used to figure out the electron configuration of an atomic species: **Aufbau Principle**, **Hund's Rule** and the **Pauli-Exclusion Principle**.

- *Rule 1 (Aufbau Principle)*: Electrons occupy the lowest-energy orbitals (closest to the nucleus) possible, starting with 1s, then 2*s*, 2*p*, and continuing on to higher energy (further away from the nucleus). Shells increase in energy in order from 1 to 2 to

3, and so on. Within these shells, an *s* subshell is the lowest energy followed by *p*, then *d*, then *f*.

- *Rule 2 (Hund's Rule):* When electrons occupy degenerate orbitals (i.e. same shell and subshell), they must first singly occupy (half-fill) each empty orbital in a subshell before double occupying (completely filling) them. Furthermore, the most stable configuration results when the spins are parallel (i.e. all spin up or all spin down). For example, all three *p* orbitals in a *p* subshell will have one electron before a single *p* orbital contains two electrons.
- *Rule 3 (Pauli-Exclusion Principle):* Each electron is described with a unique set of four quantum numbers (a unique address). Therefore, if two electrons occupy the same orbital, they must have different spins. This is the reason all orbitals can hold a maximum of two electrons.

Continuing on the periodic table to the next largest atom, beryllium, with 4 electrons, the electron configuration is $1s^2 2s^2$. Now that the 2*s* subshell is filled, electrons in larger atoms, starting with boron, begin filling the 2*p* subshell, which can hold a maximum of six electrons. The next six elements progressively fill up the 2p subshell:

- B: $1s^2 2s^2 2p^1$
- C: $1s^2 2s^2 2p^2$
- N: $1s^2 2s^2 2p^3$
- O: $1s^2 2s^2 2p^4$
- F: $1s^2 2s^2 2p^5$
- Ne: $1s^2 2s^2 2p^6$

At the end of the period the element neon, has filled the 2s, and 2*p* subshells, which completes the second shell. Now atoms with more electrons now must begin the third shell starting with the 3*s* subshell. The first two subshells of the third shell are filled in order—for example, the electron configuration of aluminum, with 13 electrons, is $1s^2 2s^2 2p^6 3s^2 3p^1$. However, a curious thing happens after the 3*p* subshell is filled: the 4*s* subshell begins to fill before the 3*d* subshell does. In fact, the exact ordering of subshells becomes more complicated at this point (after argon, with its 18 electrons), so we will not consider the electron configurations of larger atoms.

Another consideration when figuring out electron configurations is that the energy increases as we move up to the 2*s* and then 2*p*, 3*s*, and 3*p* orbitals, showing that the increasing *n* value has more influence on energy than the increasing *l* value for small atoms. However, this pattern does not hold for larger atoms. The 3*d* orbital is higher in energy than the 4*s* orbital. Such overlaps continue to occur frequently as we move up the chart. The electron filling diagram shown below in Figure 2.8.5 is useful in remembering the order for electrons to occupy shells and subshells.

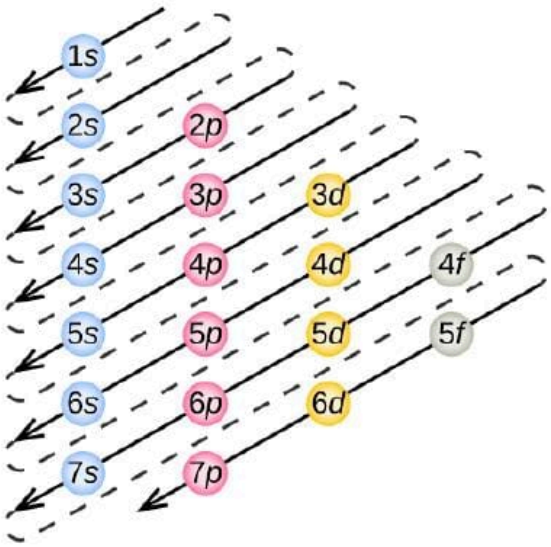


Figure 2.8.5: Follow each arrow in order from top to bottom. The subshells you reach along each arrow give the ordering of filling of subshells in larger atoms. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson via OpenStax). Detailed Description of Figure 2.8.5.

Valence Electrons and the Noble Gas Configuration

In the study of chemical reactivity, we will find that the electrons in the outermost principal energy level are very important and so they are given a special name. **Valence electrons** are the electrons in the *highest occupied principal energy level of an atom*.

For Example, the alkali metal sodium (atomic number 11) has one more electron than the neon atom. This electron must go into the lowest-energy subshell available, the 3s orbital, giving a $1s^2 2s^2 2p^6 3s^1$ configuration. The electrons occupying the outermost shell orbital(s) (highest value of *n*) are called valence electrons, and those occupying the inner shell orbitals are called **core electrons** (Figure 2.8.6.

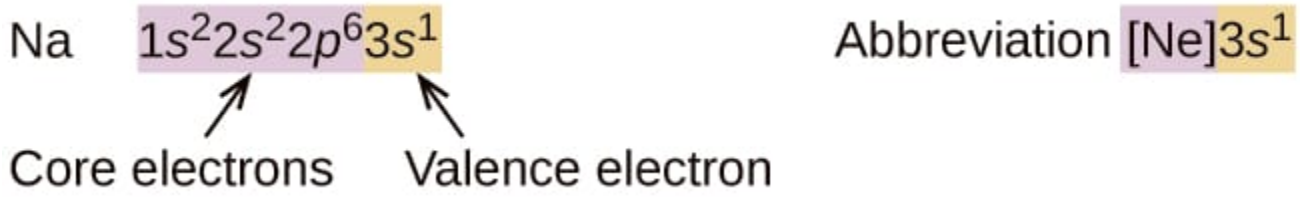


Figure 2.8.6: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.(CC BY 4.0; Gregory Anderson; Caryn Fahey; Jackie MacDonald; Adrienne Richards; Samantha Sullivan Sauer; J.R. van Haarlem; and David Wegman via Enhanced Introductory College Chemistry). Detailed Description of Figure 2.8.6.

A shorthand notation for electron configurations is called the **noble gas configuration**, which atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons.

For Example, as seen above in Figure 2.8.6, the electron configuration of sodium 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 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 $[\mathrm{Ne}]\, 3s^1$.

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

$$\mathrm{Li} : [\mathrm{He}]\, 2s^1$$
$$\mathrm{Na} : [\mathrm{Ne}]\, 3s^1$$

Chemical Reactivity

A chemical reaction results from electron removal, electron addition, or electron sharing of the valence electrons of the different atoms. The path a specific element will take depends on where the electrons are in the atom and how many there are. Thus, it is convenient to separate electrons into two groups.

Valence shell electrons (or, more simply, the *valence electrons*) are the electrons in the highest-numbered shell, or valence shell, while core electrons are the electrons in lower-numbered shells. We can see from the electron configuration of a carbon atom—$1s^2 2s^2 2p^2$—that it has 4 valence electrons ($2s^2 2p^2$) and 2 core electrons ($1s^2$). You will see in the next chapters that the chemical properties of elements are determined by the number of valence electrons.

Table 2.8.2 lists the electron configurations and the noble gas configurations of the first 20 elements of the periodic table.

Table 2.8.2: Electron Configurations of the First 20 Elements

Atomic Number	Element Symbol	Outermost Shell	Electron Configuration	Noble Gas Configuration
1	H	1	$1s^1$	$1s^1$

Atomic Number	Element Symbol	Outermost Shell	Electron Configuration	Noble Gas Configuration
2	He	1	$1s^2$	$1s^2$
3	Li	2	$1s^2 2s^1$	[He] $2s^1$
4	Be	2	$1s^2 2s^2$	[He] $2s^2$
5	B	2	$1s^2 2s^2 2p^1$	[He] $2s^2 2p^1$
6	C	2	$1s^2 2s^2 2p^2$	[He] $2s^2 2p^2$
7	N	2	$1s^2 2s^2 2p^3$	[He] $2s^2 2p^3$
8	O	2	$1s^2 2s^2 2p^4$	[He] $2s^2 2p^4$
9	F	2	$1s^2 2s^2 2p^5$	[He] $2s^2 2p^5$
10	Ne	2	$1s^2 2s^2 2p^6$	[He] $2s^2 2p^6$
11	Na	3	$1s^2 2s^2 2p^6 3s^1$	[Ne] $3s^1$
12	Mg	3	$1s^2 2s^2 2p^6 3s^2$	[Ne] $3s^2$
13	Al	3	$1s^2 2s^2 2p^6 3s^2 3p^1$	[Ne] $3s^2 3p^1$
14	Si	3	$1s^2 2s^2 2p^6 3s^2 3p^2$	[Ne]$3s^2 3p^2$
15	P	3	$1s^2 2s^2 2p^6 3s^2 3p^3$	[Ne] $3s^2 3p^3$
16	S	3	$1s^2 2s^2 2p^6 3s^2 3p^4$	[Ne] $3s^2 3p^4$
17	Cl	3	$1s^2 2s^2 2p^6 3s^2 3p^5$	[Ne] $3s^2 3p^5$
18	Ar	3	$1s^2 2s^2 2p^6 3s^2 3p^6$	[Ne] $3s^2 3p^6$
19	K	4	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	[Ar] $4s^1$
20	Ca	4	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	[Ar] $4s^2$

Example 2.8.1: Electronic Configuration of Phosphorus Atoms

Using Figure 2.8.5 as your guide, write the electron configuration of a neutral phosphorus atom. The atomic number of P is 15.

Solution

A neutral phosphorus atom has 15 electrons. Two electrons can go into the 1s subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 5 electrons. Of those 5 electrons, 2 can go into the 3s subshell, and the remaining 3 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral phosphorus atoms is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Example 2.8.2

Examine the electron configuration of neutral phosphorus atoms in Example 2.8.1, $1s^2 2s^2 2p^6 3s^2 3p^3$ and write the abbreviated notation.

Solution

Phosphorus has electron configuration, $1s^2 2s^2 2p^6 3s^2 3p^3$.

The highest-numbered shell is the third shell ($3s^2 3p^3$): 2 electrons in the 3s subshell and 3 electrons in the 3*p* subshell. That gives a total of 5 valence electrons.

The 10 inner shell (core) electrons, $1s^2 2s^2 2p^6$ can be replaced by [Ne] (see Figure 2.8.3). Abbreviated notation is : $[Ne]3s^2 3p^3$

Orbital Diagrams

An **orbital diagram** is the more visual way to represent the arrangement of all the electrons in a particular atom. In an orbital diagram, the individual orbitals are shown as squares and orbitals within a sublevel are drawn next to each other horizontally. Each sublevel is labeled by its shell and sublevel. Electrons are indicated by arrows inside of the squares. An arrow pointing upwards indicates one spin direction, while a downward pointing arrow indicates the other direction. The orbital filling diagrams for hydrogen, helium, and lithium are shown in the figure below in Figure 2.8.7.

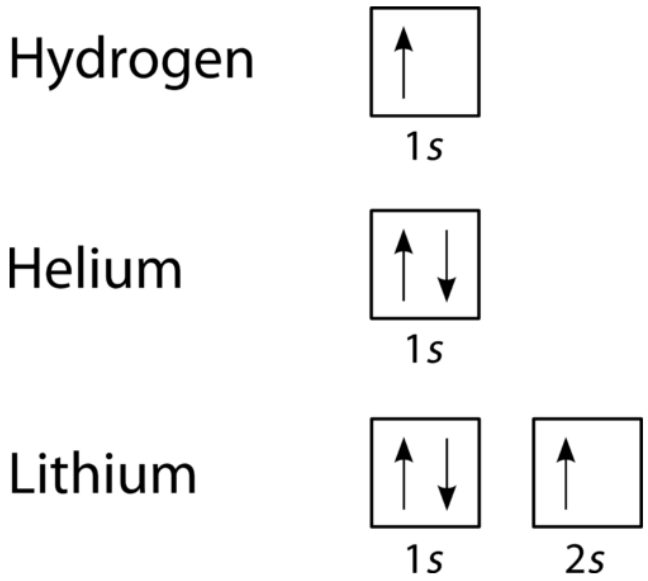


Figure 2.8.7: Electrons are indicated by arrows inside of the squares. An arrow pointing upwards indicates one spin direction, while a downward pointing arrow indicates the other direction (CC BY-NC-SA 3.0; anonymous via Open Oregon / Chemistry LibreTexts). Detailed Description of Figure 2.8.7.

According to the Aufbau Principle, sublevels and orbitals are filled with electrons in order of increasing energy. Since the s sublevel consists of just one orbital, the second electron simply pairs up with the first electron as in helium. The next element is lithium and necessitates the use of the next available sublevel, the $2s$.

The orbital diagram for carbon is shown in Figure 2.8.8. There are two $2p$ electrons for carbon and each occupies its own $2p$ orbital (Hund's Rule).

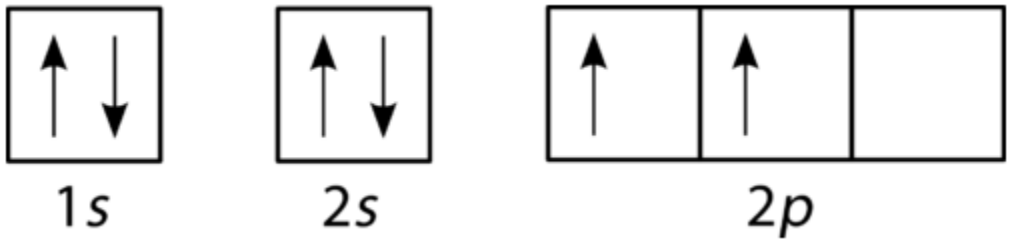


Figure 2.8.8: Orbital box diagram showing paired electrons in 1s and 2s and two unpaired electrons in separate 2p orbitals. (CC BY-NC-SA 3.0; anonymous via Open Oregon / Chemistry LibreTexts). Detailed Description of Figure 2.8.8.

Oxygen has four $2p$ electrons. After each $2p$ orbital has one electron in it, the fourth electron can be placed in the first $2p$ orbital with a spin opposite that of the other electron in that orbital. Figure 2.8.9.

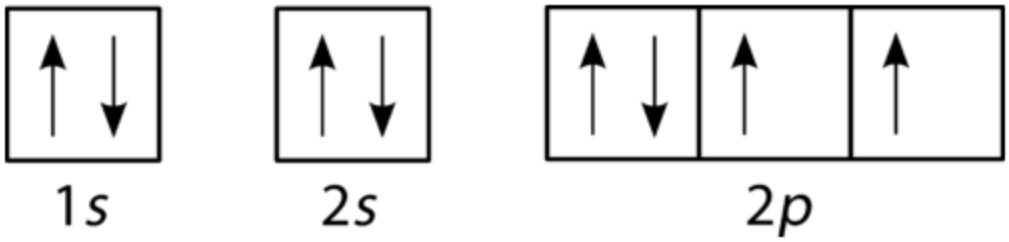


Figure 2.8.9: Box diagram showing the electron arrangement in 1s, 2s, and 2p orbitals for a 2p4 configuration. (CC BY-NC-SA 3.0; anonymous via Open Oregon / Chemistry LibreTexts). Detailed Description of Figure 2.8.9.

Electron Configurations and the Periodic Table

Since the arrangement of the periodic table is based on the electron configurations, Figure 2.8.10 provides an alternative method for determining the electron configuration.

- The filling order simply *begins at the top left,* with hydrogen (Z=1) and includes each subshell as you proceed in increasing atomic number (*Z)* order.
- For example, the first row (Period 1) contains H and He only, because only two electrons are required to fill the 1s subshell.
- The second row s-block, contains only two elements, Li and Be, to fill the 2s subshell. This is followed by the second row p-block, containing 6 elements (B through Ne) since six electrons are required to fill the 2p subshell.
- The third row is similar to the second row elements. Two electrons are needed (Na and Mg) to fill the 3s subshell and six electrons are required (Al through Ar) to complete the 3p subshell.
- After filling the 3*p* block up to Ar, we see the next subshell will be 4s (K, Ca), followed by the 3*d* subshell, which are filled by ten electrons (Sc through Zn).
- The 4p subshell is filled next by six electrons (Ga through Kr).

As you can see, the periodic table shown in Figure 2.8.10 provides a simple way to remember the order of filling the subshells in determining the electron configuration. The order of filling subshells is the same: 1s, 2s, 2*p*, 3s, 3*p*, 4s, 3*d*, 4*p*, 5s, 4d, 5p, 6s**,** etc.

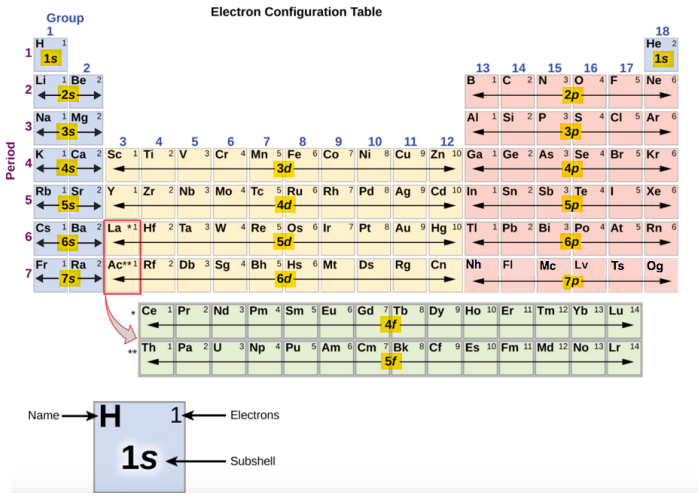


Figure 2.8.10: Electron Configuration Table. This periodic table shows the electron configuration for each subshell. By “building up” from hydrogen, this table can be used to determine the electron configuration for any atom on the periodic table. (CC BY 4.0; Paul Flowers, Klaus Theopold, Richard Langley, William R. Robinson via OpenStax). Detailed Description of Figure 2.8.10.

Example 2.8.3: Aluminum

Using Figure 2.8.10 as your guide, write the electron configuration of neutral aluminum atom. The atomic number of Al is 13.

Solution

Aluminum has 13 electrons.

Start at Period 1 of the periodic table, Figure 2.8.10. Place two electrons in the 1s subshell ($1s^2$).

Proceed to Period 2 (left to right direction). Place the next two electrons in the 2s subshell ($2s^2$) and the next six electrons in the 2p subshell ($2p^6$).

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell ($3s^2$) and the last one electron in the 3p subshell ($3p^1$).

The electron configuration of Aluminum is $1s^2 2s^2 2p^6 3s^2 3p^1$

Example 2.8.4

Based on their respective locations in the periodic table (use Figure 2.8.10), determine the number of valence electrons and the valence shell configuration of elements A, B and C as shown in Figure 2.8.11.

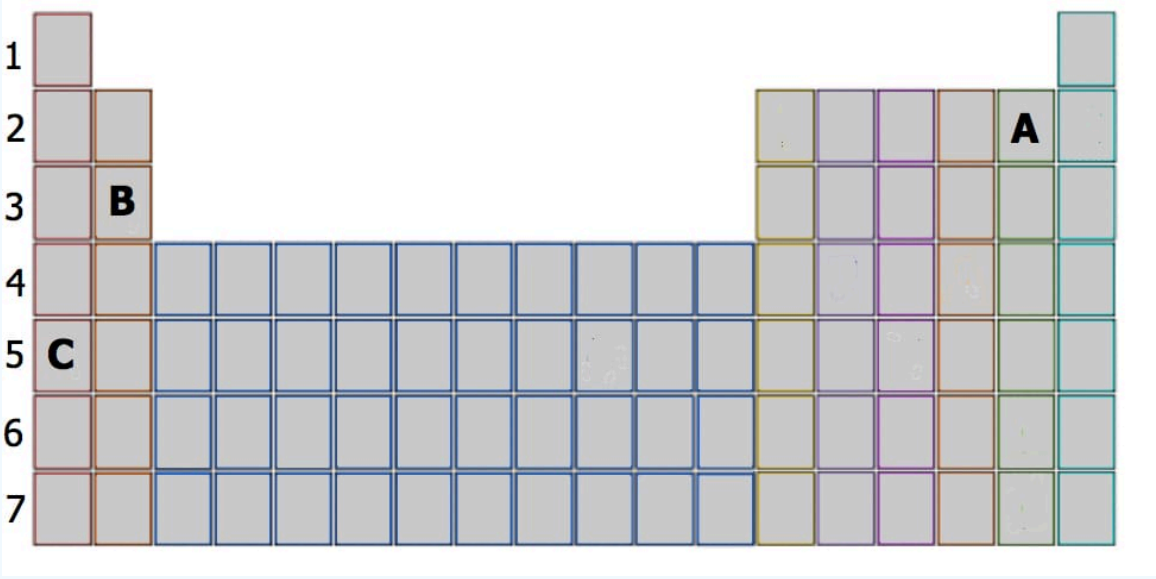


Figure 2.8.11: Respective locations of elements A, B, and C in the periodic table. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 2.8.11.

Element A is located in Period 2, the *5th position* in *2p*-block. Before the electrons are placed in *2p* subshell, the *2s* subshell must be filled first. This means that A has *two valence electrons* in 2s ($2s^2$) and *five valence electrons* in 2p ($2p^5$). Answer: $2s^2 2p^5$. It has *2 + 5 = 7* valence electrons.

Element B is located in Period 3, the *2nd position* in *3s*-block. This means that B has *two valence electrons* in 3s ($3s^2$). Answer: $3s^2$.

Element C is located in Period 5, the *1st position* in *5s*-block). This means that there is only one valence electron in 5s ($5s^1$). Answer: $5s^1$.

Key Takeaway

- Electrons are organized into shells and subshells about the nucleus of an atom.
- The valence electrons determine the reactivity of an atom.

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2.9: Periodic Table Trends

Introduction

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. For example, the elements of Group 1 are known as the alkali metals, Group 2 are the alkaline earth metals, Group 17 are the halogens, and Group 18 are the noble gases.

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.

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 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. 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.

Link to Learning: 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.

Why do elements in a given group have similar properties?

The periodic table is organized on the basis of similarities in elemental properties, but what explains these similarities? It turns out that the shape of the periodic table reflects the filling of subshells with electrons, as shown in Figure 2.9.1. Starting with the first period and going from left to right, the table reproduces the order of filling of the electron subshells in atoms. Furthermore, elements in the same group share the same valence shell electron configuration. For example, all elements in the first column have a single *s* electron in their valence shells, so their electron configurations can be described as ns^1 (where *n* represents the shell number). This last observation is crucial. Chemistry is largely the result of interactions between the valence electrons of different atoms.

Thus, atoms that have the same valence shell electron configuration will have similar chemistry.

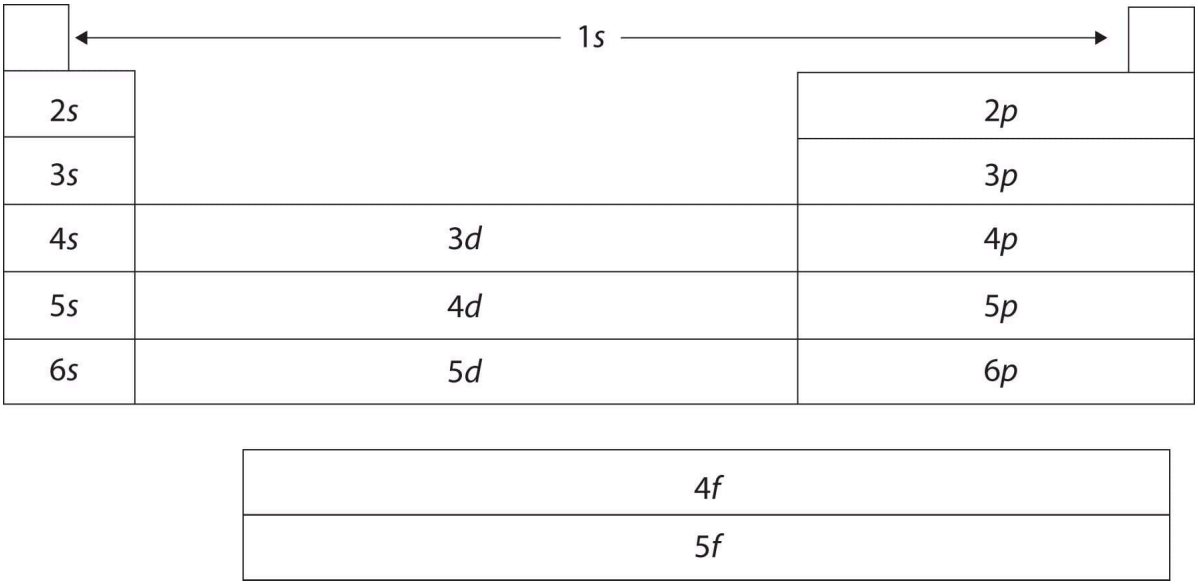


Figure 2.9.1: The shape of the periodic table reflects the order in which electron shells and subshells fill with electrons. (CC BY-NC-SA 3.0; Andy Schmitz via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 2.9.1.

Example 2.9.1

Using the variable *n* to represent the number of the valence electron shell, write the valence shell electron configuration for each group.

a. the alkaline earth metals
b. the column of elements headed by carbon

Answer a available at the end of this section.

Valence Electrons and Group Number

The number of valence electrons of an element can be determined by the periodic table group (vertical column) in which the element is categorized. With the exception of groups 3–12 (the transition metals), the units digit of the group number identifies how many valence electrons are associated with a neutral atom of an element listed under that particular column.

Table 2.9.1. The Group number and the number of valence electrons.

Periodic table group	Valence electrons
Group 1 (I) (alkali metals)	1
Group 2 (II) (alkaline earth metals)	2
Groups 3-12 (transition metals)	2*
Group 13 (III) (boron group)	3
Group 14 (IV) (carbon group)	4
Group 15 (V) (pnictogens)	5
Group 16 (VI) (chalcogens)	6
Group 17 (VII) (halogens)	7
Group 18 (VIII or 0) (noble gases)	8**

* The general method for counting valence electrons is generally not useful for transition metals.

** Except for helium, which has only two valence electrons.

Atomic Radius

The periodic table is useful for understanding atomic properties that show periodic trends. One such property is the **atomic radius** (Figure 2.9.2). The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be $74\ \text{pm}$. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = 37\ \text{pm}$.

As mentioned earlier, the higher the shell number, the farther from the nucleus the electrons in that shell are likely to be. In other words, the size of an atom is generally determined by the number of the valence electron shell. Therefore, as we go down a column on the periodic table, the atomic radius increases. As we go *across* a period on the periodic table, however, electrons are being added to the *same* valence shell; meanwhile, more protons are being added to the nucleus, so the positive charge of the nucleus is increasing. The increasing positive charge attracts the electrons more strongly, pulling them closer to the nucleus. Consequently, as we go across a period, from left to right, the atomic radius decreases. These trends are seen clearly in Figure 2.9.2.

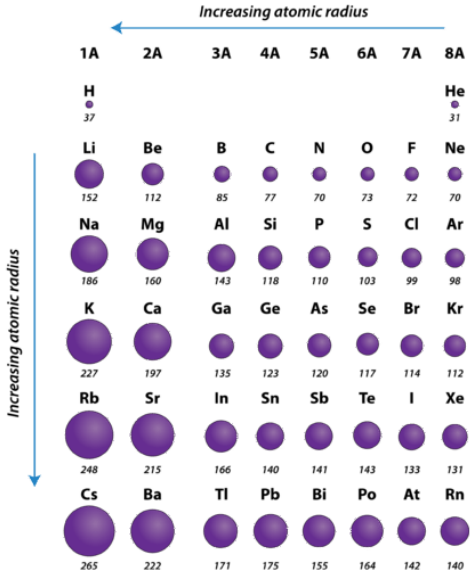


Figure 2.9.2: Atomic radii of the representative elements measured in picometers. The relative sizes of the atoms show several trends with regard to the structure of the periodic table. Atoms become larger going down a group and going from right to left across a period. (CK-12 ; Christopher Auyeung via CK-12 Foundation). Detailed Description of Figure 2.9.2.

Example 2.9.2

Using the periodic table (rather than Figure 2.9.2), which atom is larger?

a. N or Bi
b. Mg or Cl

Answer a available at the end of this section.

Link to Learning: Career Focus-Clinical Chemist

Clinical chemistry is the area of chemistry concerned with the analysis of body fluids to determine the health status of the human body. Clinical chemists measure a variety of substances, ranging from simple elements such as sodium and potassium to complex molecules such as proteins and enzymes, in blood, urine, and other body fluids. The absence or presence, or abnormally low or high amounts, of a substance can be a sign of some disease or an indication of health. Many clinical chemists use sophisticated equipment and complex chemical reactions in their work, so they not only need to understand basic chemistry, but also be familiar with special instrumentation and how to interpret test results.

Key Takeaways

- The chemical elements are arranged in a chart called the periodic table.
- Some characteristics of the elements are related to their position on the periodic table.
- The number of valence electrons of an element can be determined by the group (vertical column) number in the Periodic Table. Elements with the same number of valence electrons have similar chemical properties.

Answers

Example 2.9.1

The alkaline earth metals are in the second column of the periodic table. This column corresponds to the s subshell being filled with 2 electrons. Therefore, the valence shell electron configuration is ns^2.

The electron configuration of carbon is $1s^2 2s^2 2p^2$. Its valence shell electron configuration is $2s^2 2p^2$. Every element in the same column should have a similar valence shell electron configuration, which we can represent as $ns^2 np^2$.

Example 2.9.2

Bi is below N in Group 5A in the periodic table and has electrons in higher-numbered shells, hence we expect that Bi atoms are larger than N atoms.

Both Mg and Cl are in period 3 of the periodic table, but Cl lies farther to the right. Therefore we expect Mg atoms to be larger than Cl atoms.

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2.10: Summary

An element is a substance that cannot be broken down into simpler chemical substances. Only about 90 naturally occurring elements are known. They have varying abundances on Earth and in the body. Each element has a one- or two-letter chemical symbol.

The modern atomic theory states that the smallest piece of an element is an atom. Individual atoms are extremely small, on the order of 10–10 m across. Most elements exist in pure form as individual atoms, but some exist as diatomic molecules. Atoms themselves are composed of subatomic particles. The electron is a tiny subatomic particle with a negative charge. The proton has a positive charge and, while small, is much larger than the electron. The neutron is also much larger than an electron but has no electrical charge.

Protons, neutrons, and electrons have a specific arrangement in an atom. The protons and neutrons are found in the center of the atom, grouped together into a nucleus. The electrons are found in fuzzy clouds around the nucleus.

Each element has a characteristic number of protons in its nucleus. This number of protons is the atomic number of the element. An element may have different numbers of neutrons in the nuclei of its atoms; such atoms are referred to as isotopes. Two isotopes of hydrogen are deuterium, with a proton and a neutron in its nucleus, and tritium, with a proton and two neutrons in its nucleus. The sum of the numbers of protons and neutrons in a nucleus is called the mass number and is used to distinguish isotopes from each other.

Masses of individual atoms are measured in atomic mass units. An atomic mass unit is equal to 1/12th of the mass of a single carbon-12 atom. Because different isotopes of an element have different masses, the atomic mass of an element is a weighted average of the masses of all the element's naturally occurring isotopes.

The modern theory of electron behavior is called quantum mechanics. According to this theory, electrons in atoms can only have specific, or quantized, energies. Electrons are grouped into general regions called shells, and within these into more specific regions called subshells. There are four types of subshells, and each type can hold up to a maximum number of electrons. The distribution of electrons into shells and subshells is the electron configuration of an atom. Chemistry typically occurs because of interactions between the electrons of the outermost shell of different atoms, called the valence shell electrons. Electrons in inner shells are called core electrons.

Elements are grouped together by similar chemical properties into a chart called the periodic table. Vertical columns of elements are called groups or families. Some of the groups of elements have names, like the alkali metals, the alkaline earth metals, the halogens, and the noble gases. A horizontal row of elements is called a period. Periods and groups have differing numbers of elements in them. The periodic table separates elements into metals, nonmetals, and semimetals. The periodic table is also separated into main group elements, transition metals, lanthanide elements, and actinide elements. The lanthanide and actinide elements are also referred to as inner transition metal elements. The shape of the periodic table reflects the sequential filling of shells and subshells in atoms.

The periodic table helps us understand trends in some of the properties of atoms. One such property is the atomic radius of atoms. From top to bottom of the periodic table, atoms get bigger because electrons are occupying larger and bigger shells. From left to right across the periodic table, electrons are filling the same shell but are being attracted by an increasing positive charge from the nucleus, and thus the atoms get smaller.

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2.11: Key Terms

- **Atom:** The smallest unit of an element that retains its chemical properties.
- **Diatomic molecule:** A molecule consisting of two atoms, often of the same element (e.g., O_2, H_2).
- **Subatomic particles:** Particles that make up atoms: protons, neutrons, and electrons.
- **Nucleus:** The dense central part of an atom containing protons and neutrons.
- **Protons:** Positively charged subatomic particles found in the nucleus.
- **Neutrons:** Neutral subatomic particles found in the nucleus.
- **Electrons:** Negatively charged subatomic particles found outside the nucleus.
- **Isotopes:** Atoms of the same element with different numbers of neutrons.
- **Atomic number (Z):** The number of protons in the nucleus of an atom.
- **Mass number (A):** The total number of protons and neutrons in an atom.
- **Chemical symbol:** A one- or two-letter abbreviation for an element (e.g., H for hydrogen).
- **Ion:** An atom or molecule with a net electric charge due to the loss or gain of electrons.
- **Anion:** A negatively charged ion.
- **Cation:** A positively charged ion.
- **Atomic mass unit (amu):** A unit of mass equal to one-twelfth the mass of a carbon-12 atom.
- **Weighted average:** An average that takes into account the relative abundances of isotopes.
- **Atomic mass:** The weighted average mass of an element's atoms based on isotopic abundance.
- **Periodic table:** A chart organizing elements by increasing atomic number and similar properties.
- **Groups:** Vertical columns in the periodic table; elements in a group have similar properties.
- **Alkali metals:** Group 1 elements, highly reactive, especially with water.
- **Alkaline earth metals:** Group 2 elements, reactive metals but less so than alkali metals.
- **Halogens:** Group 17 elements, very reactive nonmetals.
- **Noble gases:** Group 18 elements, inert gases with full outer electron shells.
- **Period:** A horizontal row in the periodic table.
- **Metal:** An element that is typically shiny, conductive, malleable, and ductile.
- **Non-metal:** An element that lacks metallic properties; usually poor conductors.
- **Metalloid:** An element with properties intermediate between metals and non-metals.
- **Semimetals:** Another term for metalloids.
- **Main group elements:** Elements in groups 1, 2, and 13–18 of the periodic table.
- **Transition metals:** Elements in groups 3–12, known for forming colorful compounds and variable oxidation states.
- **Inner transition metals:** Elements in the lanthanide and actinide series.
- **Lanthanides:** The 14 elements following lanthanum (atomic numbers 58–71).
- **Actinides:** The 14 elements following actinium (atomic numbers 90–103).
- **Quantum number:** A number that describes the energy level and properties of an electron.
- **Quantized:** Existing only in specific, discrete values.
- **Quantum mechanics:** The branch of physics that describes the behavior of very small particles.
- **Shells:** Energy levels where electrons reside around the nucleus.
- **Subshells:** Divisions within electron shells (s, p, d, f).
- **Orbitals:** Regions in an atom where there is a high probability of finding electrons.
- **Spin:** A quantum property of electrons; can be +½ or −½.
- **Electron configuration:** The arrangement of electrons in an atom's orbitals.
- **Aufbau Principle:** Electrons fill orbitals starting with the lowest energy levels first.
- **Hund's Rule:** Electrons occupy separate orbitals in the same subshell before pairing up.
- **Pauli Exclusion Principle:** No two electrons in the same atom can have the same set of quantum numbers.
- **Valence electrons:** Electrons in the outermost shell that determine chemical reactivity.
- **Core electrons:** Electrons in inner shells that are not involved in bonding.
- **Noble gas configuration:** A shorthand for electron configuration using the nearest noble gas.
- **Orbital diagram:** A visual representation of electron arrangement using arrows in boxes.
- **Atomic radius:** The distance from the nucleus to the outermost electron of an atom.

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2.12: Alternative Text for Chapter 2

Figure 2.2.1: Images of Various Elements

A composite image showing three pure elemental samples, each in a separate container, representing different physical appearances and properties of elements. The left panel displays a silver metallic solid with a shiny, crystalline structure, likely a transition metal such as titanium or zinc. The middle panel shows a bright, non-metallic solid in the form of jagged, irregularly shaped flakes, characteristic of elemental sulfur. The right panel features a sample of a solid composed of small, granular chunks, resembling powdered carbon or iodine. These images illustrate the wide variety of physical forms and appearances that elements can take in their pure states.

Figure 2.3.1: Picture of John Dalton

An engraving of a seated man in early 19th-century attire, resting his head thoughtfully on his right hand with his elbow on a table. He wears round spectacles and a high-collared shirt with a waistcoat and long coat. On the table to his left are papers and scientific instruments, including what appears to be a barometer or early laboratory apparatus, suggesting his involvement in scientific work. His expression is contemplative, and the setting is a dimly lit study or laboratory. This is a historical portrait, often used to depict John Dalton, the English chemist known for developing the atomic theory.

Figure 2.3.2: Model Representations of S_8

A three-part image showing different molecular representations of a cyclooctasulfur (S_8) molecule, which consists of eight sulfur atoms arranged in a ring.

a. The first panel shows a line-angle structural formula, where each vertex and endpoint is labeled with an "S" representing a sulfur atom, connected in a closed eight-membered ring with single bonds.
b. The second panel displays a ball-and-stick model of the same S_8 molecule, where each yellow sphere represents a sulfur atom and each stick represents a covalent bond, forming a puckered ring shape to reflect the actual 3D geometry.
c. The third panel presents a space-filling model, emphasizing the relative size and proximity of the atoms. The sulfur atoms are shown as large yellow spheres tightly packed into a circular arrangement, illustrating the actual occupied space of the molecule.

This figure demonstrates how chemical structures can be visualized in different ways to highlight bonding, shape, and volume.

Figure 2.3.3: Representation of Chemical Symbols and Formulas

A four-part diagram showing the difference between individual hydrogen atoms and hydrogen molecules, using both symbolic notation and visual representations.

From left to right:

1. The first section is labeled “H” and shows a single sphere representing one hydrogen atom. Below it, the caption reads “One H atom.”
2. The second section is labeled “2H” and shows two separate spheres, each representing an individual hydrogen atom. The caption below reads “Two H atoms.”
3. The third section is labeled “H_2” and shows two spheres joined together, representing a single hydrogen molecule (H_2) composed of two bonded hydrogen atoms. The caption reads “One H_2 molecule.”
4. The fourth section is labeled “$2H_2$” and shows two sets of paired spheres, representing two hydrogen molecules. The caption reads “Two H_2 molecules.”

This figure illustrates the distinction between atomic and molecular hydrogen in both notation and particle models, helping clarify the meaning of coefficients and subscripts in chemical formulas.

Figure 2.4.1: Gold Foil Experiment

A two-part scientific diagram illustrating Rutherford's gold foil experiment, which led to the discovery of the atomic nucleus.

Left Diagram:

A schematic setup shows a beam of alpha (α) particles being emitted from a radium source on the left. The beam travels toward a thin sheet of gold foil at the center of a circular luminescent screen used to detect scattered particles. Most alpha particles pass straight through the foil, as indicated by arrows continuing in a straight path. A few particles are slightly deflected at small angles, and a very small number are significantly deflected backward, shown by arrows bouncing off at sharp angles. Labels indicate the three paths taken by the particles: (1) "Most α particles pass straight through foil," (2) "A few α particles are slightly deflected," and (3) "A very small number of α particles are significantly deflected."

Right Diagram:

A magnified cross-sectional view of the gold foil shows the arrangement of gold atoms as orange-yellow spheres packed in a grid. The nucleus of each gold atom is depicted as a small dark spot at the center of each sphere. Arrows representing α particles pass through the array: most go undisturbed between atoms, some are deflected slightly by passing near the nuclei, and a few bounce back sharply after direct collisions with nuclei. Labels indicate: "Most α particles are not deflected," "A few α particles are slightly deflected," and "A very small number of α particles are significantly deflected."

This image visually explains how the results of the gold foil experiment demonstrated that atoms are mostly empty space with a small, dense, positively charged nucleus.

Figure 2.4.2: A Modern Depiction of Atomic Structure

A side-by-side comparison of two visual representations of atomic structure based on modern quantum mechanical understanding.

a. The image on the left shows a large central sphere surrounded by a diffuse, hazy cloud that gradually fades outward. This represents the probabilistic electron cloud model. The central sphere is the nucleus of the atom, while the surrounding gradient indicates the regions of highest to lowest probability of finding an electron. The electron is not shown as a discrete particle but rather as a cloud-like distribution.
b. The image on the right presents a similar central nucleus but is surrounded by numerous tiny dots forming a spherical pattern around the nucleus. These dots represent possible electron positions at different times, consistent with the quantum mechanical model. This image conveys the idea that electrons do not orbit in fixed paths but exist in orbitals—regions where they are likely to be found.

Together, these images illustrate the shift from classical particle-based atomic models to the modern, probabilistic view of electron behavior in atoms.

Figure 2.5.1: Formalism Used for Identifying Specific Nuclide

An educational diagram explaining atomic notation and the structure of a specific nuclide. At the top, a generic format is presented for identifying an atom using symbolic notation:

- The atomic symbol X represents the element (e.g., "C" for carbon).
- The atomic number Z is written as a subscript on the lower left of the symbol, indicating the number of protons in the atom.
- The mass number A is shown as a superscript on the upper left, representing the total number of protons and neutrons.

Below the symbolic explanation is a specific example:

- The carbon atom is shown as ^{12}C with a 6 as the atomic number and 12 as the mass number.
- A 3D model of the carbon atom is displayed. The nucleus, located in the center, contains six spheres labeled with a "+" sign (protons) and six spheres (neutrons). Surrounding the nucleus are six smaller spheres labeled with a "–" sign, representing electrons orbiting in the electron cloud.

A key beside the image indicates:

- 6 protons (positive charge)
- 6 neutrons (neutral)
- 6 electrons (negative charge)

The image demonstrates how atomic number and mass number relate to the subatomic composition of an atom and how symbolic notation represents this information.

Figure 2.6.1: Representation of Atomic Mass and Atomic Number

An annotated image showing two element blocks from the periodic table to explain atomic number and atomic mass.

On the left is the block for tin (Sn):

- The atomic number 50 is displayed at the top, with an arrow and label pointing to it, indicating it represents the number of protons in the atom.
- The atomic symbol Sn is prominently shown in large, bold letters.
- The name Tin is written below the symbol.
- The atomic mass 118.71 is at the bottom, with an arrow and label pointing to it, representing the average mass of the atom in atomic mass units (amu), accounting for all naturally occurring isotopes.

On the right is the block for carbon (C):

- The atomic number 6 is at the top, also highlighted by an arrow and labeled “atomic number.”
- The atomic symbol C is in large font.
- The name Carbon is shown underneath.
- The atomic mass 12.011 is at the bottom, indicated by an arrow and labeled “atomic mass.”

This image visually distinguishes between atomic number (top of the box) and atomic mass (bottom of the box) in periodic table entries for chemical elements.

Figure 2.7.1: The Modern Periodic Table

This image shows a complete periodic table of the elements arranged by increasing atomic number in 18 vertical columns (called groups) and 7 horizontal rows (called periods). Each element is displayed in a rectangular box that contains the following:

- Atomic number in the upper left corner,
- Chemical Symbol in large bold letters in the center,
- Element Name underneath the symbol,
- Atomic mass at the bottom of the box.

At the bottom left corner, an example of how to read an element square is provided using hydrogen:

- Atomic number: 1
- Symbol: H
- Name: hydrogen
- Atomic mass: 1.008

Elements Listed by Period and Group

Period 1

- Group 1: 1 – H – Hydrogen – 1.008 (gas)
- Group 18: 2 – He – Helium – 4.003 (gas)

Period 2

- Group 1: 3 – Li – Lithium – 6.94 (solid)
- Group 2: 4 – Be – Beryllium – 9.012 (solid)
- Group 13: 5 – B – Boron – 10.81 (solid) – metalloid
- Group 14: 6 – C – Carbon – 12.01 (solid) – nonmetal
- Group 15: 7 – N – Nitrogen – 14.01 (gas) – nonmetal
- Group 16: 8 – O – Oxygen – 16.00 (gas) – nonmetal
- Group 17: 9 – F – Fluorine – 19.00 (gas) – nonmetal
- Group 18: 10 – Ne – Neon – 20.18 (gas)

Period 3

- Group 1: 11 – Na – Sodium – 22.99 (solid)
- Group 2: 12 – Mg – Magnesium – 24.31 (solid)
- Group 13: 13 – Al – Aluminum – 26.98 (solid)
- Group 14: 14 – Si – Silicon – 28.09 (solid) – metalloid
- Group 15: 15 – P – Phosphorus – 30.97 (solid)
- Group 16: 16 – S – Sulfur – 32.06 (solid)
- Group 17: 17 – Cl – Chlorine – 35.45 (gas)
- Group 18: 18 – Ar – Argon – 39.95 (gas)

Period 4

- Group 1: 19 – K – Potassium – 39.10 (solid)
- Group 2: 20 – Ca – Calcium – 40.08 (solid)
- Group 3: 21 – Sc – Scandium – 44.96 (solid)
- Group 4: 22 – Ti – Titanium – 47.87 (solid)
- Group 5: 23 – V – Vanadium – 50.94 (solid)
- Group 6: 24 – Cr – Chromium – 52.00 (solid)
- Group 7: 25 – Mn – Manganese – 54.94 (solid)
- Group 8: 26 – Fe – Iron – 55.85 (solid)
- Group 9: 27 – Co – Cobalt – 58.93 (solid)
- Group 10: 28 – Ni – Nickel – 58.69 (solid)
- Group 11: 29 – Cu – Copper – 63.55 (solid)
- Group 12: 30 – Zn – Zinc – 65.38 (solid)
- Group 13: 31 – Ga – Gallium – 69.72 (solid)
- Group 14: 32 – Ge – Germanium – 72.63 (solid) – metalloid
- Group 15: 33 – As – Arsenic – 74.92 (solid) – metalloid
- Group 16: 34 – Se – Selenium – 78.97 (solid)
- Group 17: 35 – Br – Bromine – 79.90 (liquid)
- Group 18: 36 – Kr – Krypton – 83.80 (gas)

Period 5

- Group 1: 37 – Rb – Rubidium – 85.47 (solid)
- Group 2: 38 – Sr – Strontium – 87.62 (solid)
- Group 3: 39 – Y – Yttrium – 88.91 (solid)
- Group 4: 40 – Zr – Zirconium – 91.22 (solid)
- Group 5: 41 – Nb – Niobium – 92.91 (solid)
- Group 6: 42 – Mo – Molybdenum – 95.95 (solid)
- Group 7: 43 – Tc – Technetium – [98] (solid, radioactive)
- Group 8: 44 – Ru – Ruthenium – 101.1 (solid)
- Group 9: 45 – Rh – Rhodium – 102.9 (solid)
- Group 10: 46 – Pd – Palladium – 106.4 (solid)
- Group 11: 47 – Ag – Silver – 107.9 (solid)
- Group 12: 48 – Cd – Cadmium – 112.4 (solid)
- Group 13: 49 – In – Indium – 114.8 (solid)
- Group 14: 50 – Sn – Tin – 118.7 (solid)
- Group 15: 51 – Sb – Antimony – 121.8 (solid) – metalloid
- Group 16: 52 – Te – Tellurium – 127.6 (solid) – metalloid

- Group 17: 53 – I – Iodine – 126.9 (solid)
- Group 18: 54 – Xe – Xenon – 131.3 (gas)

Period 6

- Group 1: 55 – Cs – Cesium – 132.9 (solid)
- Group 2: 56 – Ba – Barium – 137.3 (solid)
- f-block (Lanthanides): 57–71 – La to Lu
- Group 3: 72 – Hf – Hafnium – 178.5
- Group 4: 73 – Ta – Tantalum – 180.9
- Group 5: 74 – W – Tungsten – 183.8
- Group 6: 75 – Re – Rhenium – 186.2
- Group 7: 76 – Os – Osmium – 190.2
- Group 8: 77 – Ir – Iridium – 192.2
- Group 9: 78 – Pt – Platinum – 195.1
- Group 10: 79 – Au – Gold – 197.0
- Group 11: 80 – Hg – Mercury – 200.6 (liquid)
- Group 12: 81 – Tl – Thallium – 204.4
- Group 13: 82 – Pb – Lead – 207.2
- Group 14: 83 – Bi – Bismuth – 208.9
- Group 15: 84 – Po – Polonium – [209]
- Group 16: 85 – At – Astatine – [210]
- Group 17: 86 – Rn – Radon – [222] (gas)

Lanthanide Series (Period 6, f-block)

- 57 – La – Lanthanum – 138.9
- 58 – Ce – Cerium – 140.1
- 59 – Pr – Praseodymium – 140.9
- 60 – Nd – Neodymium – 144.2
- 61 – Pm – Promethium – [145]
- 62 – Sm – Samarium – 150.4
- 63 – Eu – Europium – 152.0
- 64 – Gd – Gadolinium – 157.3
- 65 – Tb – Terbium – 158.9
- 66 – Dy – Dysprosium – 162.5
- 67 – Ho – Holmium – 164.9
- 68 – Er – Erbium – 167.3
- 69 – Tm – Thulium – 168.9
- 70 – Yb – Ytterbium – 173.1
- 71 – Lu – Lutetium – 175.0

Period 7

- Group 1: 87 – Fr – Francium – [223]
- Group 2: 88 – Ra – Radium – [226]
- f-block (Actinides): 89–103 – Ac to Lr – see below
- Group 3: 104 – Rf – Rutherfordium – [267]
- Group 4: 105 – Db – Dubnium – [270]
- Group 5: 106 – Sg – Seaborgium – [271]
- Group 6: 107 – Bh – Bohrium – [270]
- Group 7: 108 – Hs – Hassium – [277]
- Group 8: 109 – Mt – Meitnerium – [276]
- Group 9: 110 – Ds – Darmstadtium – [281]
- Group 10: 111 – Rg – Roentgenium – [282]
- Group 11: 112 – Cn – Copernicium – [285]
- Group 12: 113 – Nh – Nihonium – [286]
- Group 13: 114 – Fl – Flerovium – [289]
- Group 14: 115 – Mc – Moscovium – [288]

- Group 15: 116 – Lv – Livermorium – [293]
- Group 16: 117 – Ts – Tennessine – [294]
- Group 17: 118 – Og – Oganesson – [294] (gas)

Actinide Series (Period 7, f-block)

- 89 – Ac – Actinium – [227]
- 90 – Th – Thorium – 232.0
- 91 – Pa – Protactinium – 231.0
- 92 – U – Uranium – 238.0
- 93 – Np – Neptunium – [237]
- 94 – Pu – Plutonium – [244]
- 95 – Am – Americium – [243]
- 96 – Cm – Curium – [247]
- 97 – Bk – Berkelium – [247]
- 98 – Cf – Californium – [251]
- 99 – Es – Einsteinium – [252]
- 100 – Fm – Fermium – [257]
- 101 – Md – Mendelevium – [258]
- 102 – No – Nobelium – [259]
- 103 – Lr – Lawrencium – [262]

Figure 2.7.2: Types of Elements

This image shows a color-coded schematic of the periodic table that classifies elements into three categories based on their chemical and physical properties. The table is outlined without individual element names or numbers but uses solid color regions to indicate metals, nonmetals, and semimetals.

The majority of the table is filled with a golden-yellow color representing metals. This includes all elements on the left side of the table, most of the central block, and all of the separate block at the bottom, which represents the lanthanide and actinide series. These metals are known for being conductive, malleable, and lustrous.

A bright green color occupies the upper right portion of the table and marks the nonmetals. This region includes a block of elements that typically have poor conductivity and are found as gases or brittle solids at room temperature. In addition, the green square in the top-left corner represents hydrogen, which, although positioned in Group 1, is chemically classified as a nonmetal.

A staircase of small orange-red blocks runs diagonally between the metal and nonmetal regions, beginning under boron and stepping down to the right. These blocks represent the semimetals, also called metalloids, which have properties intermediate between metals and nonmetals. They act as a bridge between the two types, both in position and behavior.

The visual layout communicates how element properties vary across the periodic table, with metals occupying the broadest region, nonmetals clustered in the upper right, and semimetals forming a boundary between them. The use of clear color blocks allows easy identification of each category without the need for text or element symbols.

Figure 2.7.3: Special Names for Sections of the Periodic Table

This image shows a simplified outline of the periodic table divided into large color-coded regions based on the special names given to particular groups or blocks of elements.

On the far left side are two vertical columns: the first is labeled "alkali metals" and the second "alkaline earth metals." Both are part of the main group elements and are colored in pale shades of cream and yellow, respectively. These columns represent Group 1 and Group 2 elements, which are highly reactive metals typically found in nature as compounds rather than in pure form.

At the center of the table is a large horizontal block shaded in golden-yellow labeled "transition metals." This region includes the elements in Groups 3 through 12. These metals are known for forming colored compounds, having variable oxidation states, and being useful as catalysts and conductors.

To the right of the transition metals is a wide rectangular block shaded in muted green. This area includes multiple groups and contains many of the p-block elements. It is not labeled as a whole but includes two distinct vertical bands: one labeled “halogens” and the other “noble gases.” The halogens are colored green and represent Group 17; they are reactive nonmetals including fluorine and chlorine. The noble gases, shaded in pale cream, occupy Group 18 and include inert gases like helium, neon, and argon.

Above both ends of the table are curved arrows pointing from the left and right regions labeled together as "main group elements." This designation refers to the elements in Groups 1, 2, and 13–18, which typically display predictable periodic trends and valence behavior.

Below the main table are two horizontal blue bars. The top bar is labeled "lanthanide metals" and the bottom "actinide metals." These two rows represent the f-block elements and are shown separately to keep the main table compact. The lanthanides include elements such as cerium and europium, known for their magnetic and phosphorescent properties. The actinides include radioactive elements like uranium and plutonium.

Overall, this diagram presents the periodic table as a set of clearly defined regions to help learners recognize functional categories: alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanides, actinides, and the broader distinction of main group elements. Each region is color-coded and boxed for clarity, without including specific element names or symbols.

Figure 2.8.1: Bohr's Model of the Atom

This diagram represents Bohr’s model of the atom, showing electrons in fixed circular orbits around a central nucleus. At the center is a solid sphere labeled "proton," representing the nucleus of the atom. Surrounding the nucleus are four concentric circular paths labeled as energy levels: n = 1, n = 2, n = 3, and n = 4. These circles represent quantized electron orbits, or shells, where electrons are allowed to exist without radiating energy.

Along two of the orbits (n = 2 and n = 4), small spheres are shown and labeled "electron." These represent negatively charged electrons revolving around the nucleus in their respective energy levels. Each electron is located directly on its circular orbit, following Bohr's proposal that electrons move in fixed, defined paths rather than in a cloud or probabilistic region.

Arrows and labels identify the key components:

- The "proton" is the positively charged particle at the center.
- The "electron" is a small, negatively charged particle moving in an orbit.
- "Orbit" refers to the specific circular path the electron follows, with the principal quantum number “n” denoting the energy level of each orbit.

This classical atomic model emphasizes the quantized nature of electron energies and introduces the idea of electrons transitioning between energy levels by absorbing or emitting energy. The simplicity of this depiction makes it especially useful for early chemistry education, despite being superseded by more accurate quantum mechanical models.

Figure 2.8.2: Electron Orbitals

This image shows a complete set of visual representations for atomic orbitals in three dimensions, including s, p, and d orbitals. Each orbital is displayed with respect to labeled X, Y, and Z axes. The orbitals are color-coded with red and blue to indicate opposite phases of the electron wave function. The shapes reflect the regions of space where electrons are most likely to be found.

At the top center is the s orbital. It is a single sphere centered at the origin, uniformly distributed around the nucleus with no nodal planes. It appears as a solid blue sphere, showing spherical symmetry around all three axes.

In the middle row are three p orbitals. Each consists of two lobes with a nodal plane passing through the nucleus. The px orbital is oriented along the X-axis, with one red lobe to the left and one blue lobe to the right. The pz orbital is vertically oriented along the Z-axis, with the red lobe on top and blue on the bottom. The py orbital is aligned along the Y-axis, with the red lobe in front and the blue lobe behind. All three p orbitals are centered at the nucleus and extend symmetrically in opposite directions along their respective axes.

The bottom row displays the five d orbitals, which have more complex shapes. The dxy orbital has four lobes situated between the X and Y axes in the XY plane, arranged diagonally with alternating red and blue colors. The dxz orbital also has four lobes, arranged between the X and Z axes. The dz^2 orbital has two lobes along the Z-axis with a torus, or doughnut-shaped ring, around the center in the XY plane. The dyz orbital resembles dxz but lies between the Y and Z axes. The $dx^2–y^2$ orbital has four lobes oriented directly along the X and Y axes, unlike the dxy orbital where lobes are between the axes.

The orbitals are drawn to emphasize their directional nature and angular momentum characteristics, with the axes intersecting at the origin of each orbital. The coloring highlights the mathematical sign of the wavefunction in each region, and the spatial arrangements show how different orbitals occupy three-dimensional space in an atom.

Figure 2.8.3: Shell Diagrams of Various Atoms

This image presents shell (Bohr) diagrams of four different atoms—hydrogen (H), helium (He), lithium (Li), and beryllium (Be)—to show how electrons are arranged in energy levels or shells around each atomic nucleus. Each atom is depicted inside a colored circle at the center, labeled with its element symbol, and surrounded by one or more concentric circular lines representing electron shells.

Starting from the top left:

1. Hydrogen (H): A green circle at the center with the symbol "H" represents the nucleus of a hydrogen atom. It has one circular orbit (n = 1) surrounding it, with one small gray dot on the shell representing its single electron.
2. Helium (He): A light blue circle at the center labeled "He" shows the nucleus of a helium atom. It also has one electron shell (n = 1), but this time with two gray dots on the shell, evenly spaced, indicating two electrons filling the first energy level.
3. Lithium (Li): A red circle in the top right labeled "Li" represents lithium. This atom has two shells: the first (innermost) shell has two electrons, and the second shell (n = 2) has one electron, for a total of three electrons. This illustrates lithium's electron configuration of 2-1.
4. Beryllium (Be): At the bottom is a peach-colored circle labeled "Be," representing a beryllium atom. It has two electron shells: the first shell contains two electrons, and the second shell contains two more, totaling four electrons. This shows a 2-2 electron configuration.

Each gray dot represents an individual electron, and the circular lines represent fixed energy levels or electron shells. The number and distribution of electrons in these diagrams demonstrate the structure of neutral atoms from atomic numbers 1 through 4 and highlight periodic trends in electron configuration.

Figure 2.8.4: The Diagram of an Electron Configuration of Hydrogen Atom

This image illustrates the electron configuration of a hydrogen atom using both a visual and symbolic representation.

On the left side of the image is a stylized atom, labeled "H" to represent hydrogen. The atom is shown as a shaded sphere, with a single small black dot carrying a negative sign (–) orbiting near the surface. This dot represents hydrogen’s one electron.

To the right of the atom is the electron configuration notation written as $1s^1$. Annotations clarify this notation:

- The "1" refers to the principal energy level or shell (n = 1), the closest shell to the nucleus.
- The "s" refers to the subshell, which defines the shape of the orbital; s orbitals are spherical in shape.
- The superscript "1" indicates the number of electrons in the 1s subshell—in this case, just one electron.

This depiction combines the visual model of an atom with the symbolic shorthand used in electron configuration, emphasizing that hydrogen’s single electron occupies the lowest energy orbital, the 1s orbital.

Figure 2.8.5: The Order of Electron Filling in an Atom

This image is a diagram of the Aufbau Principle, which shows the order in which electrons fill atomic orbitals from lower energy to higher energy.

The diagram contains labeled orbitals arranged in rows by principal energy level and subshell type. The orbital labels include 1s,

2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, and 7p. Different colors are used to distinguish the orbital families s, p, d, and f, but the labels on each orbital identify the subshells directly.

Diagonal arrows indicate the sequence in which electrons occupy the orbitals. Reading the filling order from lowest energy to higher energy, the sequence is:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, and 7p.

The diagram emphasizes that electrons do not fill orbitals simply by increasing principal energy number. Instead, they occupy the available orbitals in order of increasing energy. For example, 4s fills before 3d, and 6s fills before 4f and 5d.

The purpose of the image is to help students visualize the electron filling sequence used when writing electron configurations for atoms.

Figure 2.8.6: Electron Configuration of Sodium

This image displays the electron configuration of a sodium (Na) atom and highlights the distinction between core electrons and the valence electron.

On the left, the full electron configuration of sodium is written as:

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

- The first part, $1s^2\ 2s^2\ 2p^6$ is labeled as core electrons. These represent the electrons in the inner, fully filled shells that are not involved in bonding.
- The last part, $3s^1$, is labeled as valence electron, representing the single outermost electron in the third energy level, which is chemically reactive and available for bonding.

To the right, the configuration is expressed in its abbreviated (noble gas) form:

[Ne] $3s^1$

- The abbreviation [Ne] represents the electron configuration of neon ($1s^2\ 2s^2\ 2p^6$), a noble gas with a closed shell of 10 electrons.
- The notation emphasizes that sodium has the same core configuration as neon, with one additional electron in the 3s orbital.

This image helps clarify how electron configurations are structured and shows the importance of valence electrons in determining an atom's chemical properties.

Figure 2.8.7: Orbital diagrams for Hydrogen, Helium, and Lithium

This image shows the orbital diagrams for the elements hydrogen, helium, and lithium using box and arrow notation to represent the occupancy and spin of electrons in atomic orbitals.

Each box represents a single orbital, and arrows inside the boxes represent electrons. The direction of the arrow indicates the electron's spin (up or down).

- Hydrogen: One box is labeled 1s It contains a single upward-pointing arrow, representing one electron with spin-up occupying the 1s orbital. This reflects hydrogen's electron configuration of $1s^1$.
- Helium: A single box labeled 1s contains two arrows: one pointing up and one pointing down, indicating two electrons with opposite spins paired in the same orbital. This reflects helium's configuration of $1s^2$, with a completely filled 1s orbital.
- Lithium: Two boxes are shown: one labeled 1s, the other 2s.
  - The 1s box has two arrows (one up, one down), indicating two paired electrons in the 1s orbital.
  - The 2s box contains one upward-pointing arrow, showing a single unpaired electron in the second energy level's s orbital. This represents lithium's electron configuration of $1s^2\ 2s^1$.

These diagrams visually illustrate the Pauli Exclusion Principle (no two electrons in the same orbital can have the same spin) and the Aufbau Principle (electrons fill the lowest energy orbitals first).

Figure 2.8.8: Orbital Diagram for Carbon

This image shows the orbital diagram for a carbon atom using box-and-arrow notation to represent its electron configuration and spin states.

There are three distinct groups of boxes, each labeled with the corresponding atomic orbital:

- 1s orbital: One box is labeled “1s” and contains two arrows, one pointing up and one down. These represent two electrons of opposite spin occupying the 1s orbital, forming a complete pair.
- 2s orbital: he next box is labeled “2s” and also contains two oppositely directed arrows, indicating a second pair of electrons occupying the 2s orbital.
- 2p orbital: Three adjacent boxes represent the three degenerate 2p orbitals ($2p_x$, $2p_y$, and $2p$◈). The first two boxes each contain a single upward arrow, representing two unpaired electrons with parallel spins occupying separate 2p orbitals. The third box is empty, reflecting that carbon only has two electrons in the 2p subshell.

This configuration demonstrates both the Aufbau Principle (electrons fill lower energy orbitals first) and Hund's Rule which states that electrons will occupy separate orbitals with parallel spins before pairing up. Carbon has six electrons in total, with an electron configuration of $1s^2\ 2s^2\ 2p^2$.

Figure 2.8.9: Orbital Diagram for Oxygen

This image shows the orbital diagram for an oxygen atom, using box-and-arrow notation to represent the distribution and spin of electrons in atomic orbitals.

There are three groups of boxes, each labeled to indicate the orbital:

- 1s orbital: One box labeled “1s” contains two arrows—one pointing up and one down—representing a pair of electrons with opposite spins fully occupying the 1s orbital.
- 2s orbital: The second box labeled “2s” also contains a pair of oppositely spinning electrons, filling the 2s orbital.
- 2p orbitals: Three adjacent boxes labeled collectively as “2p” represent the three degenerate p orbitals ($2p_x$, $2p_y$, $2p$◈).
  - The first box contains two arrows (one up, one down), showing two paired electrons.
  - The second box has a single upward arrow, representing one unpaired electron.
  - The third box also has one upward arrow, another unpaired electron.

This distribution reflects Hund’s rule, which states that electrons occupy orbitals singly with parallel spins before pairing. Oxygen has a total of eight electrons, with an electron configuration of $1s^2\ 2s^2\ 2p^4$, and this orbital diagram accurately shows two unpaired electrons in the 2p subshell.

Figure 2.8.10: Electron Configuration Table

This image is a modified periodic table titled Electron Configuration Table organized by periods (horizontal rows) and groups (vertical columns), and color-coded by orbital block:

- Blue: s block
- Pink: p block
- Yellow: d block
- Gray: f block

Each element’s square displays:

- The chemical symbol
- A superscripted number showing the number of electrons in the final subshell
- The subshell label (e.g., 2p, 4d, 6s)

Arrows across the table indicate the direction of electron filling order according to the Aufbau Principle. A legend at the bottom left illustrates the format:

- Hydrogen (H), atomic number 1, has one electron in the 1s orbital → shown as H 1s

s-block Elements (blue section):

Period 1:

- H: $1s^1$
- He: $1s^2$

Period 2:

- Li: $2s^1$
- Be: $2s^2$

Period 3:

- Na: $3s^1$
- Mg: $3s^2$

Period 4:

- K: $4s^1$
- Ca: $4s^2$

Period 5:

- Rb: $5s^1$
- Sr: $5s^2$

Period 6:

- Cs: $6s^1$
- Ba: $6s^2$

Period 7:

- Fr: $7s^1$
- Ra: $7s^2$

p-block Elements (red section):

Period 2:

- B: $2p^1$
- C: $2p^2$
- N: $2p^3$
- O: $2p^4$
- F: $2p^5$
- Ne: $2p^6$

Period 3:

- Al: $3p^1$
- Si: $3p^2$
- P: $3p^3$
- S: $3p^4$
- Cl: $3p^5$

- Ar: $3p^6$

Period 4:

- Ga: $4p^1$
- Ge: $4p^2$
- As: $4p^3$
- Se: $4p^4$
- Br: $4p^5$
- Kr: $4p^6$

Period 5:

- In: $5p^1$
- Sn: $5p^2$
- Sb: $5p^3$
- Te: $5p^4$
- I: $5p^5$
- Xe: $5p^6$

Period 6:

- Tl: $6p^1$
- Pb: $6p^2$
- Bi: $6p^3$
- Po: $6p^4$
- At: $6p^5$
- Rn: $6p^6$

Period 7:

- Nh: $7p^1$
- Fl: $7p^2$
- Mc: $7p^3$
- Lv: $7p^4$
- Ts: $7p^5$
- Og: $7p^6$

d-block Elements (yellow section):

Period 4 (3d):

- Sc: $3d^1$
- Ti: $3d^2$
- V: $3d^3$
- Cr: $3d^5$
- Mn: $3d^5$
- Fe: $3d^6$
- Co: $3d^7$
- Ni: $3d^8$
- Cu: $3d^{10}$
- Zn: $3d^{10}$

Period 5 (4d):

- Y: $4d^1$
- Zr: $4d^2$
- Nb: $4d^4$
- Mo: $4d^5$

- Tc: $4d^5$
- Ru: $4d^7$
- Rh: $4d^8$
- Pd: $4d^{10}$
- Ag: $4d^{10}$
- Cd: $4d^{10}$

Period 6 (5d):

- Hf: $5d^2$
- Ta: $5d^3$
- W: $5d^4$
- Re: $5d^5$
- Os: $5d^6$
- Ir: $5d^7$
- Pt: $5d^9$
- Au: $5d^{10}$
- Hg: $5d^{10}$

Period 7 (6d):

- Rf: $6d^2$
- Db: $6d^3$
- Sg: $6d^4$
- Bh: $6d^5$
- Hs: $6d^6$
- Mt: $6d^7$
- Ds: $6d^8$
- Rg: $6d^9$
- Cn: $6d^{10}$

f-block Elements (green section):

4f orbitals (lanthanides, Period 6):

- Ce: $4f^1$
- Pr: $4f^2$
- Nd: $4f^3$
- Pm: $4f^4$
- Sm: $4f^5$
- Eu: $4f^6$
- Gd: $4f^7$
- Tb: $4f^8$
- Dy: $4f^9$
- Ho: $4f^{10}$
- Er: $4f^{11}$
- Tm: $4f^{12}$
- Yb: $4f^{13}$
- Lu: $4f^{14}$

5f orbitals (actinides, Period 7):

- Th: $5f^0$
- Pa: $5f^2$
- U: $5f^3$
- Np: $5f^4$
- Pu: $5f^6$
- Am: $5f^7$
- Cm: $5f^7$

- Bk: $5f^9$
- Cf: $5f^{10}$
- Es: $5f^{11}$
- Fm: $5f^{12}$
- Md: $5f^{13}$
- No: $5f^{14}$
- Lr: $5f^{14}$

Arrows point from La and Ac in the d-block to the start of the f-block rows, indicating the continuation of electron filling into the 4f and 5f subshells.

This table enables the user to determine the ending subshell for each element's electron configuration based on its position in the periodic table. It visualizes the order in which orbitals are filled as atomic number increases.

Figure 2.8.11: Example Image

This image is a simplified blank periodic table of the elements, with labeled regions marked A, B, and C to highlight specific categories or element families. The table is arranged by period (horizontal rows, numbered 1 through 7 on the left) and group (vertical columns).

Each region is indicated as follows:

- A is located in the top right section of the table, in Group 17, Period 3.
- B is in the upper left section, in Group 2, Period 3.
- C is below B in Group 1, Period 5.

All squares in the table are unmarked except for A, B, and C, and the boxes are outlined in soft, multicolored borders suggesting the separation of element groups by category. The image refers to the question in Example 2.8.4

Figure 2.9.1: The Shape of the Periodic Table

This image is a stylized outline of the periodic table, organized by orbital types and energy levels to illustrate how electron subshells correspond to regions of the table.

The layout is presented in a grid-like format, where horizontal rows represent periods (principal energy levels) and blocks correspond to electron subshells. Labels such as "1s," "2p," "3d," and "4f" appear inside specific regions, connecting the table's structure to the atomic orbital filling sequence.

- The leftmost columns in each period are labeled with s orbitals:
  - 1s, 2s, 3s, 4s, 5s, and 6s
- These columns represent Groups 1 and 2—the alkali and alkaline earth metals.
- The rightmost columns are labeled with p orbitals:
  - 2p, 3p, 4p, 5p, and 6p
- These correspond to Groups 13 through 18, which include the main group nonmetals, metalloids, halogens, and noble gases.
- The middle block is labeled with d orbitals:
  - 3d (Period 4), 4d (Period 5), 5d (Period 6)
- This central area represents the transition metals (Groups 3 to 12).
- A separate block at the bottom contains two horizontal rows labeled:
  - 4f (top row)
  - 5f (bottom row)

These rows represent the lanthanide and actinide series, corresponding to elements in Periods 6 and 7, which fill the f orbitals.

An arrow is drawn across the top row labeled "1s," pointing from left to right, indicating the direction in which elements are read or filled in the Aufbau process. This structure shows that the periodic table's shape is inherently connected to the quantum mechanical structure of atoms and how electrons occupy orbitals across different energy levels.

Figure 2.9.2: Trends on the Periodic Table for Atomic Radii

This image shows a visual representation of the trend in atomic radius across the periodic table. Elements are arranged according to their group (vertical columns, labeled 1A to 8A) and period (horizontal rows), and each element is shown as a circle whose size is proportional to atomic radius in picometers (pm). A numeric value representing the atomic radius in pm is shown below each element symbol.

Two arrows indicate the overall periodic trends:

- A horizontal arrow across the top points leftward, labeled "Increasing Atomic Radius", showing that atomic radius increases as you move from right to left across a period.
- A vertical arrow along the left margin points downward, also labeled "Increasing Atomic Radius", indicating that atomic radius increases from top to bottom within a group

For example:

- Hydrogen (H) in the top left has a small radius of 37 pm, while Cesium (Cs) at the bottom of Group 1 has the largest radius of 265 pm, shown by a much larger circle.
- In contrast, Helium (He) in the top right has the smallest radius on the table, 31 pm.
- As you move left to right across a period (e.g., from Na to Cl in Period 3), the circles decrease in size, and atomic radii decrease (Na = 186 pm, Cl = 99 pm).
- As you move down a group (e.g., Group 1A: Li → Na → K → Rb → Cs), the circles become progressively larger.

This image effectively illustrates the periodic trend that:

- Atomic radius decreases across a period (left to right),
- Atomic radius increases down a group (top to bottom).

These changes are due to increasing nuclear charge (which pulls electrons closer across a period) and the addition of electron shells (which expands the atomic size down a group).

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3: IONIC COMPOUNDS AND BONDING

Introduction

Ionic compounds form when atoms transfer electrons, resulting in the creation of positively charged cations and negatively charged anions that are held together by strong electrostatic forces. The formula of these compounds reflects the simplest whole-number ratio of the ions, and their names are derived by combining the name of the cation (often the metal) with the name of the anion (often the non-metal or polyatomic ion). For example, in sodium chloride (NaCl), sodium (Na) loses an electron to become Na^{+}, while chlorine (Cl) gains an electron to become Cl^{-}. These ions form a stable crystal lattice, giving ionic compounds distinct physical properties, such as high melting and boiling points, solubility in water, and the ability to conduct electricity when molten or dissolved. Additionally, ionic compounds are brittle and hard due to their rigid structure. Beyond their chemical significance, compounds like NaCl are essential in biological processes, such as nerve function and fluid balance, and have practical uses in cooking, food preservation, de-icing, and various industrial applications.

Learning Objectives

- Understand ionic bond formation through electron transfer between metals and nonmetals.
- Identify and differentiate between monatomic and polyatomic ions.
- Predict and write formulas for ionic compounds based on ion charges.
- Apply systematic rules to name ionic compounds.
- Calculate the formula mass of ionic compounds.
- Understand the characteristics of ionic compounds, such as high melting points, electrical conductivity, and solubility.

3.1: Types of Bonding

3.2: Ionic Bonding

3.3: Formulas for Ionic Compounds

3.4: Formula Mass

3.5: Ionic Nomenclature

3.6: Characteristics of Ionic Compounds

3.7: Summary

3.8: Key Terms

3.9: Alternative text for Chapter 3

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3.1: Types of Bonding

Ions, Ionic Bonds and Octet Rule

Atoms can join together by forming a **chemical bond**, which is a very strong attraction between two atoms. Chemical bonds are formed when electrons in different atoms interact with each other to make an arrangement that is more stable than when the atoms are apart.

What causes atoms to make a chemical bond with other atoms, rather than remaining as individual atoms? A clue comes by considering the noble gas elements, the rightmost column of the periodic table. These elements—helium, neon, argon, krypton, xenon, and radon—do not form compounds very easily, which suggests that they are especially stable as lone atoms. What else do the noble gas elements have in common? Except for helium, they all have eight valence electrons. Chemists have concluded that atoms are especially stable if they have eight electrons in their outermost shell. This useful rule of thumb is called the **octet rule**, and it is a key to understanding why compounds form.

> Some elements break the octet rule by having fewer (like H, Be, B) or more (like P, S, Cl) than 8 electrons. Molecules with an odd number of electrons (e.g., NO) also cannot satisfy the octet rule.

There are two ways for an atom that does not have an octet of valence electrons to obtain an octet in its outer shell. One way is the transfer of electrons between two atoms until all atoms have octets. Because some atoms will lose electrons and some atoms will gain electrons, there is no overall change in the number of electrons, but individual atoms acquire a nonzero electric charge. Those that lose electrons become positively charged, and those that gain electrons become negatively charged. Charged atoms are called **ions**. Because opposite charges attract (while like charges repel), these oppositely charged ions attract each other, forming **ionic bonds**. The resulting compounds are called ionic compounds and are the primary subject of this chapter.

Links to Health

Ions are charged particles formed when atoms gain or lose electrons and are vital for many physiological functions. For example, ions like sodium (Na^+), potassium (K^+), and calcium (Ca^{2+}) help regulate nerve signals, muscle contractions, and fluid balance. The balance of these ions is crucial for heart health, muscle function, and maintaining hydration.

The second way for an atom to obtain an octet of electrons is by sharing electrons with another atom. These shared electrons simultaneously occupy the outermost shell of more than one atom. The bond made by electron sharing is called a covalent bond (discussed in more detail in next chapter).

Example 3.1.1

A sodium atom has one valence electron. Do you think it is more likely for a sodium atom to lose one electron or gain seven electrons to obtain an octet?

Solution

Although either event is possible, a sodium atom is more likely to lose its single valence electron. When that happens, it becomes an ion with a net positive charge. This can be illustrated as follows:

Table 3.1.1: Table explaining the solution to the example.

Sodium atom (Na)		Sodium ion (Na^+)	
11 protons	11+	11 protons	11+

Sodium atom (Na)		Sodium ion (Na^+)	
11 electrons	11–	10 electrons	10–
	0 overall charge		+1 overall charge

Example 3.1.2

A sulfur atom has six valence electron. Do you think it is more likely for a sulfur atom to lose six electrons or gain two electrons to obtain an octet?

Solution

Although either event is possible, a sulfur atom is more likely to gain two electrons. When that happens, it becomes an ion with a 2- net negative charge. The formula of the resulting ion is S^{2-}.

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3.2: Ionic Bonding

Electron Transfer

Most atoms do not have eight electrons in their **valence electron** shell. Some atoms have only a few electrons in their outer shell, while some atoms 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 as a result 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.

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.

> Atoms with few valence electrons lose them to form cations (metals), while atoms with nearly full valence shells gain electrons to form anions (nonmetals).

We can use electron configurations to illustrate the **electron transfer** process between sodium atoms and chlorine atoms.

$$\text{Na: } 1s^2 2s^2 2p^6 3s^1$$

As demonstrated here, a sodium atom (Na) has one valence electron in the third principal energy level. It is likely to achieve an octet in its outermost shell by losing its one valence electron. The **cation** produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The sodium ion, Na^+, has the electron configuration with an octet of electrons from the second principal energy level ($2s^2 2p^6$). It 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, each with 10 electrons. The equation below illustrates this process.

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

Figure 3.2.1 is a graphical depiction of this process.

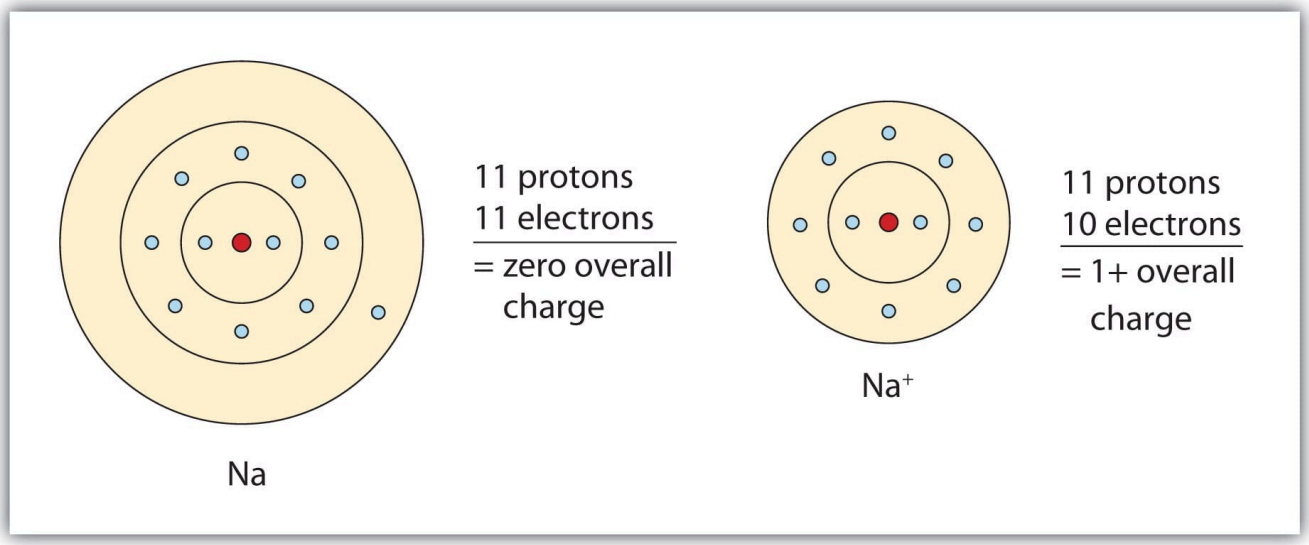


Figure 3.2.1: On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.1.

Now, let's consider chlorine atom, Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

Only one more electron is needed to achieve an octet in chlorine's valence shell. When a chlorine atom gains an electron, its outermost principal energy level achieves an octet. 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. This process is illustrated below. (In table salt, this electron comes from the sodium atom.)

$$\begin{array}{lcl} \mathrm{Cl} + \mathrm{e}^- & \rightarrow & \mathrm{Cl}^- \\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5 & & 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 (\text{octet}) \end{array}$$

Figure 3.2.2 is a graphical depiction of this process.

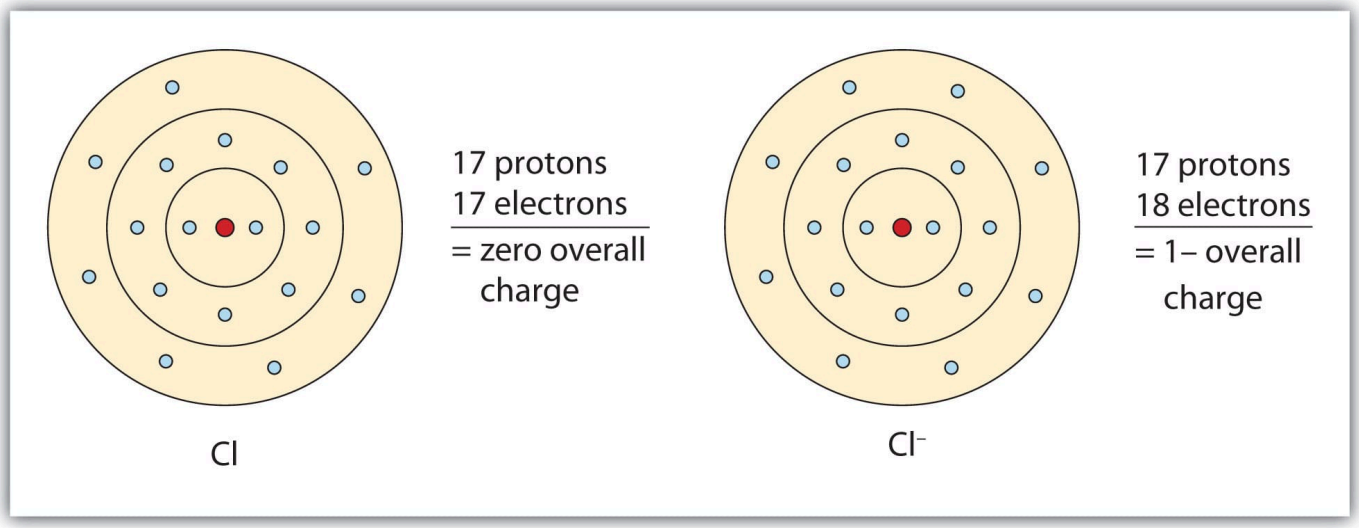


Figure 3.2.2: On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1– charge. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.2.

With two oppositely charged ions, there is an electrostatic attraction between them because opposite charges attract. The

resulting combination is the compound sodium chloride. Notice that there are no leftover electrons. The number of electrons lost by the sodium atom (one) equals the number of electrons gained by the chlorine atom (one), so the compound is electrically neutral. In macroscopic samples of sodium chloride, there are billions and billions of sodium and chloride ions, although there is always the same number of cations and anions.

Example 3.2.1

Write the electron configuration of aluminum atom (Z=13). How many electrons must Al lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Solution

The electron configuration of Al atom is $1s^2 2s^2 2p^6 3s^2 3p^1$. The second shell has octet ($2s^2 2p^6$) while the valence shell has 3 electrons ($3s^2 3p^1$). Al can achieve octet by losing the 3 valence electrons. The resulting cation is Al^{3+} with electron configuration, $1s^2 2s^2 2p^6$.

Example 3.2.2

Write the electron configuration of magnesium atom (Z=12). How many electrons must Mg lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Solution

The electron configuration of Mg atom is $1s^2 2s^2 2p^6 3s^2$. The second shell has octet ($2s^2 2p^6$) while the valence shell has 2 electrons ($3s^2$). Mg can achieve octet by losing the 2 valence electrons. The resulting cation is Mg^{2+} with electron configuration, $1s^2 2s^2 2p^6$.

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 3.2.3 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 multiply charged ion. The barium cation is written Ba^{2+}, not Ba^{+2}.

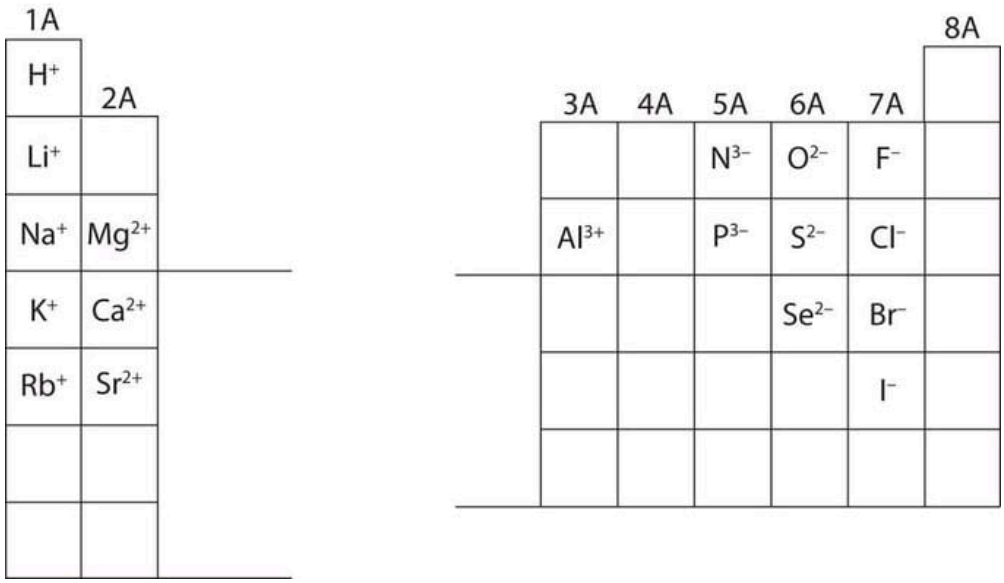


Figure 3.2.3: The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.3.

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Example 3.2.3

Which of these ions is *not likely* to form?

a. Mg^+
b. K^+

Solution

a. Mg is in Group 2A and has two valence electrons. It achieves octet by losing two electrons to form Mg^{2+} cation. Losing only one electron to form Mg^+ does not make an octet, hence, Mg^+ is not likely to form.
b. K is in Group 1A and has one valence electron. It achieves octet by losing one electron to form K^+ cation. Losing only one electron to form K^+ does make an octet, hence, K^+ is likely to form.

Chemists use simple diagrams to show an atom's valence electrons and how they transfer. These diagrams have two advantages over the electron shell diagrams. First, they show only valence electrons. Second, instead of having a circle around the chemical symbol to represent the electron shell, they have up to eight dots around the symbol; each dot represents a valence electron. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. For example, the representation for sodium is Na·

and the representation for chlorine is as follows:



For the above diagrams, it does not matter what sides the dots are placed on in Lewis diagrams as long as each side has a maximum of two dots.

These diagrams are called **Lewis electron dot diagrams**, or simply Lewis diagrams, after Gilbert N. Lewis, the American chemist

who introduced them. To write an element's Lewis dot symbol, place the dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. In other words, place the dots singly on each side before pairing them.

Figure 3.2.4 shows the electron configurations and Lewis diagrams of the elements lithium through neon, which is the entire second period of the periodic table. For the **main group elements**, the number of valence electrons is the same as the group number listed at the top of the periodic table.



Figure 3.2.4: Lewis Electron Dot Diagrams of the Elements Lithium through Neon. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.4.

In Figure 3.2.5 below, the sodium atom (Na) is shown with a single dot, indicating its one valence electron. The chlorine atom (Cl) is depicted with seven dots around it, representing its seven valence electrons. Sodium transfers its one valence electron to chlorine, resulting in the formation of two ions: a positively charged sodium ion (Na$^+$) with no valence electrons (now resembling neon), and a negatively charged chloride ion (Cl$^-$) with a complete octet of eight valence electrons (now resembling argon). The Lewis diagram clearly shows this electron transfer by using an arrow pointing from the single dot on sodium to the incomplete octet on chlorine. This transfer results in the formation of an **ionic bond** between Na$^+$ and Cl$^-$, producing the stable compound sodium chloride (NaCl).

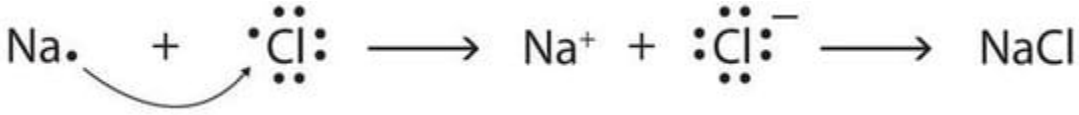


Figure 3.2.5: Electron Transfer and Ionic Bond Formation in NaCl. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.5.

In representing the final formula, the dots are omitted.

Links to Learning: The Role of Ions in Human Physiology

In the human body, ionic compounds such as sodium chloride (NaCl), common table salt, are essential for various physiological processes. Sodium (Na$^+$) and chloride (Cl$^-$) ions help maintain fluid balance regulate blood pressure, and support nerve function. The movement of these ions across cell membranes is vital for nerve signal transmission and muscle contraction, including the heart muscle.

Other ions, like calcium (Ca^{2+}), are crucial for bone health and muscle function. Calcium ions help in muscle contractions and the formation of bone structures. An imbalance of ionic compounds in the body can lead to health problems such as dehydration, high blood pressure, and muscle cramps.

Example 3.2.4

Starting with lithium and bromine atoms, use Lewis diagrams to show the formation of the ionic compound LiBr.

Solution

From the periodic table, we see that lithium is in the same column as sodium, so it will have the same valence shell electron configuration. That means that the neutral lithium atom will have the same Lewis diagram that the sodium atom has. Similarly, bromine is in the same column as chlorine, so it will have the same Lewis diagram that chlorine has. Therefore,

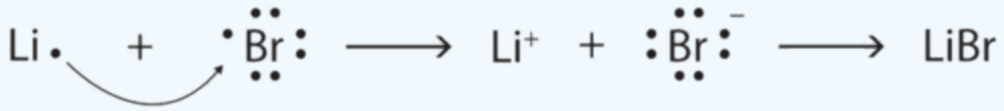


Figure 3.2.6: Electron Transfer and Ionic Bond Formation in LiBr. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.6.

Some ionic compounds contain different numbers of cations and anions, requiring electron transfer between multiple atoms. For example, Figure 3.2.7 illustrates the formation of magnesium bromide ($MgBr_2$). Magnesium (Mg), with two valence electrons represented by two dots, transfers one electron to each of the two bromine (Br) atoms, each shown with seven valence electrons. This electron transfer results in the formation of one Mg^{2+} ion, having lost both valence electrons, and two Br^- ions, each gaining one electron to complete its octet. The Lewis diagram uses arrows to show the movement of electrons from magnesium to each bromine atom, forming the stable ionic compound $MgBr_2$, consisting of one Mg^{2+} and two Br^- ions.

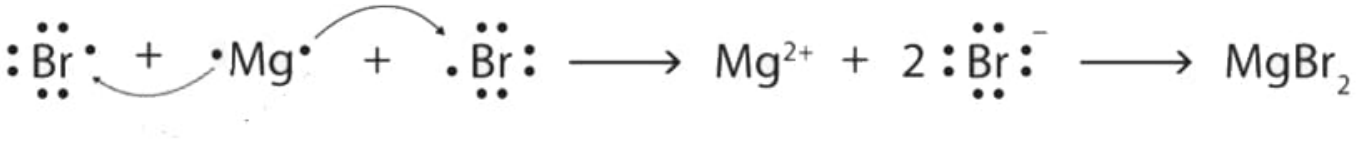


Figure 3.2.7: Electron Transfer and Ionic Bond Formation in MgBr2. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.7.

Notice that in this example there are two bromide ions (1– charge) needed for every one magnesium ion (2+ charge) in order for the overall charge of the compound to equal zero. This is called **charge balance**. The number of each type of ion is indicated in the formula by the subscript.

Common Ion Charges in Main Group and Transition Metals

Most of the elements that make ionic compounds form an ion that has a characteristic charge. For example, sodium makes ionic compounds in which the sodium ion always has a 1+ charge. Chlorine makes ionic compounds in which the chloride ion always has a 1– charge. Some elements, especially transition metals, can form ions of multiple charges. Figure 3.2.8 shows the characteristic charges for some of these ions. As illustrated in Figure 3.2.3, there is a pattern to the charges on many of the main group ions, but there is no simple pattern for transition metal ions (or for the larger main group elements).

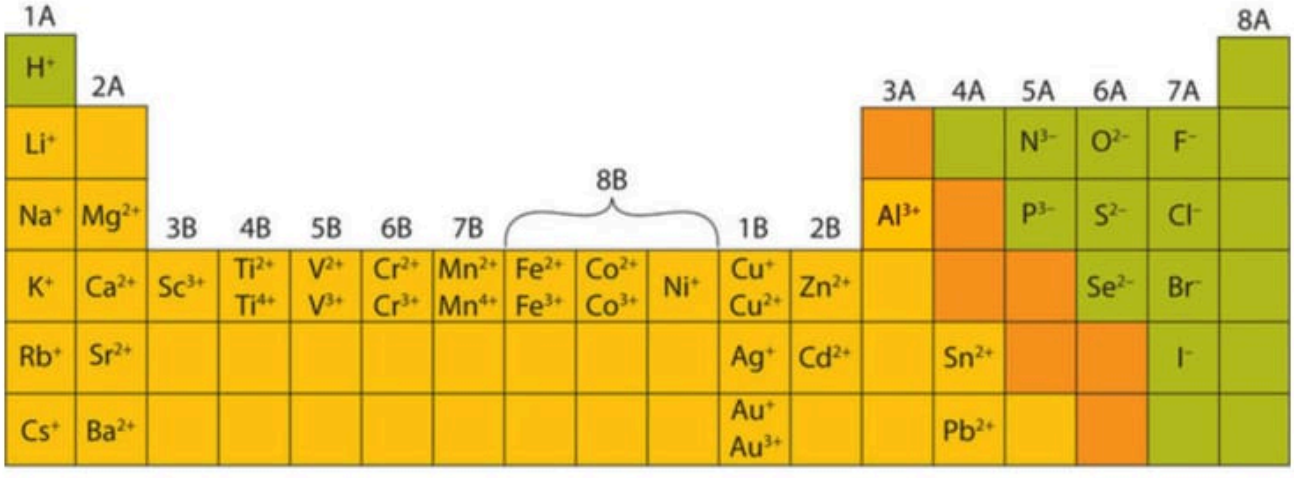


Figure 3.2.8: Note that some atoms commonly form ions of variable charges. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.2.8.

Key Takeaways

- Ions can be positively charged or negatively charged.
- A Lewis diagram is used to show how electrons are transferred to make ions and ionic compounds.

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3.3: Formulas for Ionic Compounds

Chemical Formula

We have already encountered some chemical formulas for simple ionic compounds. A **chemical formula** is a concise list of the elements in a compound and the ratios of these elements. To better understand what a chemical formula means, we must consider how an ionic compound is constructed from its ions.

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure $3.3.1$). As you can observe, there are no individual NaCl "particles" in the array; instead, there is a continuous lattice of alternating sodium and chloride ions. However, we can use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers, as a way of describing the compound. In the case of sodium chloride, the ratio of sodium ions to chloride ions, expressed in lowest whole numbers, is 1:1, so we use NaCl (one Na symbol and one Cl symbol) to represent the compound. Thus, NaCl is the chemical formula for sodium chloride, which is a concise way of describing the relative number of different ions in the compound. A macroscopic sample is composed of myriads of NaCl pairs; each individual pair called a **formula unit**. Although it is convenient to think that NaCl crystals are composed of individual NaCl units, Figure $3.3.1$ shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.

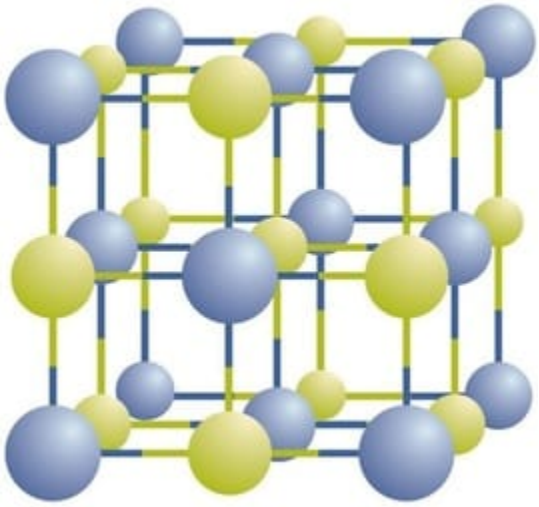


Figure $3.3.1$: A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure $3.3.1$.

The formula for an ionic compound follows several conventions. First, the cation is written before the anion. Because most metals form cations and most nonmetals form anions, formulas typically list the metal first and then the nonmetal. Second, charges are not written in a formula. Remember that in an ionic compound, the component species are ions, not neutral atoms, even though the formula does not contain charges. Finally, the proper formula for an ionic compound always has a **net zero charge**, meaning the total positive charge must equal the total negative charge. To determine the proper formula of any combination of ions, determine how many of each ion is needed to balance the total positive and negative charges in the compound.

Steps to Write an Ionic Compound Formula

1. Write the symbols of the cation and anion. Example: Na^{+} (sodium) and Cl^{-} (chloride)
2. Determine the charges of each ion. Na = +1, Cl = –1
3. Balance the total charges so the compound has a net zero charge. +1 (Na) + –1 (Cl) = 0 → already balanced
4. Write using the lowest whole number ratio. Result: NaCl

> By convention, assume that there is only one atom if a subscript is not present. We do not use 1 as a subscript.

If we look at the ionic compound consisting of lithium ions and bromide ions, we see that the lithium ion has a 1+ charge and the bromide ion has a 1– charge. Only one ion of each is needed to balance these charges. The formula for lithium bromide is LiBr.

When an ionic compound is formed from magnesium and oxygen, the magnesium ion has a 2+ charge, and the oxygen atom has a 2– charge. Although both of these ions have higher charges than the ions in lithium bromide, they still balance each other in a one-to-one ratio. Therefore, the proper formula for this ionic compound is MgO.

Now consider the ionic compound formed by magnesium and chlorine. A magnesium ion has a 2+ charge, while a chlorine ion has a 1– charge:

$$\mathrm{Mg^{2+}Cl^{-}}$$

Combining one ion of each does not completely balance the positive and negative charges. The easiest way to balance these charges is to assume the presence of *two* chloride ions for each magnesium ion:

$$\mathrm{Mg^{2+}\ Cl^{-}\ Cl^{-}}$$

Now the positive and negative charges are balanced. We could write the chemical formula for this ionic compound as MgClCl, but the convention is to use a numerical subscript when there is more than one ion of a given type—$\mathrm{MgCl_2}$. This chemical formula says that there are one magnesium ion and *two* chloride ions in this formula.

For compounds in which the ratio of ions is not as obvious, the subscripts in the formula can be obtained by crossing charges: use the absolute value of the charge on one ion as the subscript for the other ion. This method is shown schematically in Figure 3.3.2.

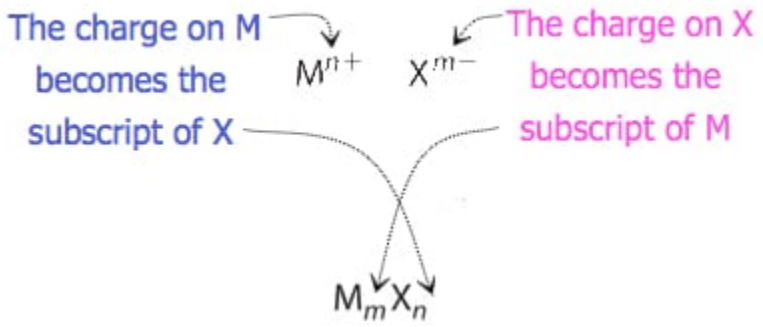


Figure 3.3.2: Subscripts come from crossing charges: M^{n} becomes X's subscript, X^{m} becomes M's subscript. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.3.2.

In Figure 3.3.3, the charges on aluminum (Al) and oxygen (O) are shown as +3 for aluminum (Al^{3}) and -2 for oxygen (O^{2}). To determine the ionic compound formula using the cross-charge method, we take the absolute value of the charge on each ion and use it as the subscript for the other ion. Therefore, the subscript for oxygen becomes 3 (from the +3 charge on aluminum), and the subscript for aluminum becomes 2 (from the -2 charge on oxygen). This gives the formula Al_2O_3, where two aluminum ions combine with three oxygen ions to balance the charges.

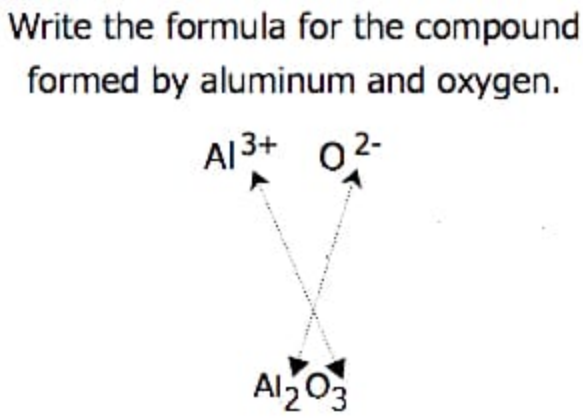


Figure 3.3.3: The image shows Al^{3+} and O^{2-} ions, and by crossing charges, the formula is Al_2O_3. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.3.3.

Simplifying Ionic Formulas Using the Cross-Charge Method

When crossing charges, it is sometimes necessary to reduce the subscripts to their simplest ratio to write the **empirical formula**. Consider, for example, the compound formed by Pb^{4+} and O^{2-}. Using the absolute values of the charges on the ions as subscripts gives the formula Pb_2O_4. This simplifies to its correct empirical formula PbO_2. The empirical formula has one Pb^{4+} ion and two O^{2-} ions. By convention, the empirical formula of a compound represents the **lowest whole number ratio** of the elements present. By convention, the lowest whole-number ratio of the ions is used in ionic formulas. There are exceptions for certain ions, such as Hg_2^{2+}.

By convention, the lowest whole-number ratio of the ions is used in ionic formulas. There are exceptions for certain ions, such as Hg_2^{2+}.

Example 3.3.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the sodium ion and the sulfur ion
b. the aluminum ion and the fluoride ion
c. the 3+ iron ion and the oxygen ion
d. the 2+ iron and the oxygen ion

Solution

a. To obtain a valence shell octet, sodium forms an ion with a 1+ charge, while the sulfur ion has a 2– charge. Two sodium 1+ ions are needed to balance the 2– charge on the sulfur ion. Rather than writing the formula as NaNaS, we shorten it by convention to Na_2S.
b. The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1– charge. Three fluorine 1– ions are needed to balance the 3+ charge on the aluminum ion. This combination is written as AlF_3.
c. Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxygen atom has a 2– charge as an ion. To balance the positive and negative charges, we look to the least common multiple—6: two iron 3+ ions will give 6+, while three 2– oxygen ions will give 6–, thereby balancing the overall positive and negative charges. Thus, the formula for this ionic compound is Fe_2O_3.
d. If the iron ion has a 2+ charge (Fe^{2+}) and the oxygen ion has a 2- charge (O^{2-}), the formula for the ionic compound can be determined by balancing the charges. Since the charges of iron and oxygen are both 2 (one being positive and the other negative), they naturally balance each other out. This means that only one iron ion is needed for every one oxygen ion to form a neutral compound. Therefore, the formula for this ionic compound is FeO. Alternatively, using the crossing

charges method, the absolute value of the charge on iron (2) becomes the subscript for oxygen, and the absolute value of the charge on oxygen (2) becomes the subscript for iron. This also results in the formula FeO, which is the correct empirical formula for iron(II) oxide.

Polyatomic Ions

Some ions consist of groups of atoms covalently bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called **polyatomic ions**. Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table 3.3.1 lists the ion names and ion formulas of the most common polyatomic ions. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1– charge.

Table 3.3.1: Ion Names and Ion Formulas of Common Polyatomic Ions

Ion Name	Ion Formula
ammonium ion	NH_4^+
hydroxide ion	OH^-
cyanide ion	CN^-
carbonate ion	CO_3^{2-}
bicarbonate or hydrogen carbonate	HCO_3^-
acetate ion	$C_2H_3O_2^{-1}$ **or** $CH_3CO_2^{-1}$
nitrate ion	NO_3^-
nitrite ion	NO_2^-
sulfate ion	SO_4^{2-}
sulfite ion	SO_3^{2-}
phosphate ion	PO_4^{3-}
phosphite ion	PO_3^{3-}

Note that only one polyatomic ion in this Table, the ammonium ion (NH_4^+), is a cation. This polyatomic ion contains one nitrogen and four hydrogens that collectively bear a +1 charge. The remaining polyatomic ions are all negatively-charged and, therefore, are classified as anions. However, only two of these, the hydroxide ion and the cyanide ion, are named using the "-ide" suffix that is typically indicative of negatively-charged particles. The remaining polyatomic anions, which all contain oxygen, in combination with another non-metal, exist as part of a series in which the number of oxygens within the polyatomic unit can vary. As has been repeatedly emphasized in several sections of this text, no two chemical formulas should share a common chemical name. A single suffix, "-ide," is insufficient for distinguishing the names of the anions in a related polyatomic series. Therefore, "-ate" and "-ite" suffixes are employed, in order to denote that the corresponding polyatomic ions are part of a series. Additionally, these suffixes also indicate the relative number of oxygens that are contained within the polyatomic ions. Note that all of the polyatomic ions whose names end in "-ate" contain one more oxygen than those polyatomic anions whose names end in "-ite." Unfortunately, much like the common system for naming transition metals, these suffixes only indicate the relative number of oxygens that are contained within the polyatomic ions. For example, the nitrate ion, which is symbolized as NO_3^-, has one more oxygen than the nitrite ion, which is symbolized as NO_2^-. However, the sulfate ion is symbolized as SO_4^{2-}. While both the nitrate ion and the sulfate ion share an "-ate" suffix, the former contains three oxygens, but the latter contains four. Additionally, both the nitrate ion and the sulfite ion contain three oxygens, but these polyatomic ions do not share a common suffix. Unfortunately, the relative nature of these suffixes mandates that the ion formula/ion name combinations of the polyatomic ions must simply be memorized.

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. Two examples are shown below:

Example 3.3.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the barium ion and the nitrate ion
b. the ammonium ion and the phosphate ion

Solution

a. the barium ion and the nitrate ion

- Barium ion = Ba^{2+}
- Nitrate ion = NO_3^-

Cross the charges

- Ba → 1 (from NO_3)
- NO_3 → 2 (from Ba^2)

b. the ammonium ion and the phosphate ion

- Ammonium ion = NH_4^+
- Phosphate ion = PO_4^{3-}

Cross charges

- NH_4 → 3 (from PO_4^3)
- PO_4 → 1 (from NH_4)

Figure 3.3.4 illustrates this concept more clearly

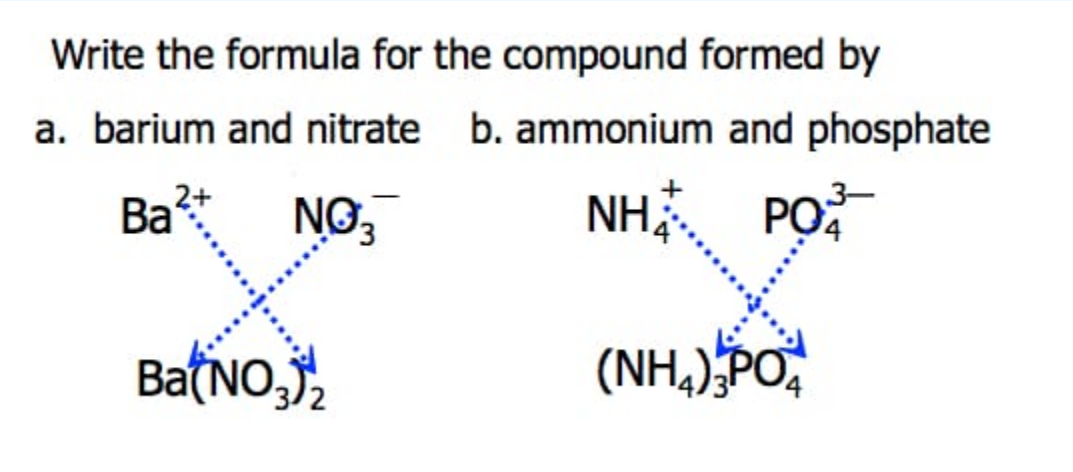


Figure 3.3.4: The image asks a two-part question: Write the formula for the compound formed by (a) barium and nitrate, and (b) ammonium and phosphate. For (a), Ba has a charge of 2+ and NO_3 has a charge of -1. By crossing charges, the formula is $Ba(NO_3)_2$. For (b), NH_4 has a charge of +1 and PO_4 has a charge of 3-, so by crossing charges the formula is $(NH_4)_3PO_4$. (CC BY-NC-SA 3.0; anonymous, via LibreTexts). Detailed Description of Figure 3.3.4.

When there is more than one polyatomic ion, parentheses and a subscript are used, as the ion acts as a unit (e.g., $Ca(NO_3)_2$).

Example 3.3.3

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the potassium ion and the sulfate ion
b. the calcium ion and the nitrate ion
c. the iron(II) and sulfate ion

Solution

a. 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.
b. Calcium ions have a charge of 2+, while nitrate ions have a charge of 1–. We will need two nitrate ions to balance the charge on each calcium ion. The formula for nitrate must be enclosed in parentheses. Thus, we write $Ca(NO_3)_2$ as the formula for this ionic compound.
c. Using the cross-charge method for Iron(II) (Fe^{2}◈) and sulfate (SO_4^{2}◈), the charges are both 2. The absolute values of the charges become the subscripts for the opposite ion. Since both charges are 2, the formula is $FeSO_4$ as one Fe^{2}◈ ion combines with one SO_4^{2}◈ ion to balance the charges.

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2A metal) and a nonmetallic element (bromine, a group 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 5A, and oxygen, from group 6A). It is not an ionic compound; it belongs to the category of covalent compounds discuss in the next chapter. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, 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 Ba^{2+} ion, with the 2+ charge balancing the overall 2– charge from the two nitrate ions. Thus, this compound is also ionic.

Example 3.3.4

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. NH_4Cl
d. OF_2

Solution

a. Sodium is a metal, and oxygen is a nonmetal; therefore, Na_2O is expected to be ionic.

b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic.
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.

Links to Learning: Blood vs. Seawater Composition

Science has long recognized that blood and seawater have similar compositions. After all, both liquids have ionic compounds dissolved in them. The similarity may be more than mere coincidence; many scientists think that the first forms of life on Earth arose in the oceans. A closer look, however, shows that blood and seawater are quite different. A 0.9% solution of sodium chloride approximates the salt concentration found in blood. In contrast, seawater is principally a 3% sodium chloride solution, over three times the concentration in blood. Table 3.3.2 compares the levels of various ions present in blood and seawater.

Table 3.3.2: A comparison of the amounts of ions in blood and seawater.

Ion	Percent in Seawater	Percent in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002
SO_4^{2-}	0.09	—
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
HPO_4^{2-}, $H_2PO_4^-$	—	0.01

Most ions are more abundant in seawater than they are in blood, with some important exceptions. There are far more hydrogen carbonate ions (HCO_3^-) in blood than in seawater. This difference is significant because the hydrogen carbonate ion and some related ions have a crucial role in controlling the acid-base properties of blood. The amount of hydrogen phosphate ions—HPO_4^{2-} and $H_2PO_4^-$—in seawater is very low, but they are present in higher amounts in blood, where they also affect acid-base properties. Another notable difference is that blood does not have significant amounts of the sulfate ion (SO_4^{2-}), but this ion is present in seawater.

Key Takeaways

- Ionic compounds form by balancing positive and negative charges to create neutral compounds.
- Polyatomic ions are charged groups of atoms acting as single ions.
- Formulas show the simplest ratio of ions needed to balance charges.
- Subscripts indicate how many ions are required.

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3.4: Formula Mass

Understanding Formula Mass

Many argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Dalton's atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances.

Formula Mass

An earlier chapter of this text described the development of the atomic mass unit (amu), the concept of average atomic masses, and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the **formula mass** of a substance by summing the average atomic masses of all the atoms represented in the substance's formula.

> The molar mass of an element, as found on the periodic table, is equivalent to its atomic mass (amu).

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is calculated to be 58.44 amu (as shown in Figure 3.4.1).

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

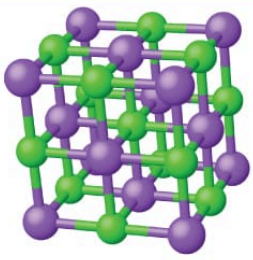


Figure 3.4.1: Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu. (CC BY 4.0; source via OpenStax). Detailed Description of Figure 3.4.1.

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded.

Example 3.4.1: Computing Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Using the method described above, the formula mass of this compound is determined as illustrated in Figure 3.4.2 below.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Molecular mass	342.14

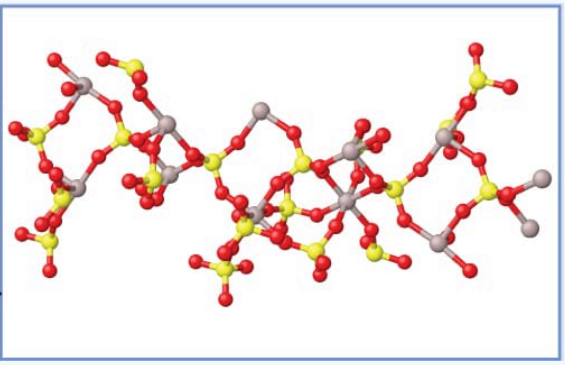


Figure 3.4.2: Aluminum sulfate, $Al_2(SO_4)_3$, contains an array of aluminium and sulfate ions combined in a 2:3 ratio respectively. Its formula mass is 342.14 amu. (CC BY 4.0; source via OpenStax). Detailed Description of Figure 3.4.2.

Example 3.4.2

Calculate the formula mass of calcium phosphate. Write the correct chemical formula for calcium phosphate and solve.

Solution

The chemical formula of calcium phosphate is $Ca_3(PO_4)_2$. This formula contains three calcium (Ca) atoms, two phosphorus (P) atoms, and eight oxygen (O) atoms. The subscript "3" indicates three calcium atoms, while the subscript "2" outside the phosphate group (PO_4) means there are two phosphate units. Each phosphate unit consists of one phosphorus atom and four oxygen atoms, contributing a total of two phosphorus atoms and eight oxygen atoms. The formula mass is calculated by summing the atomic masses of all atoms in the formula:

120.24 (3 Ca) + 61.94 (2 P) + 128.00 (8 O) = 310.18 amu

Key Takeaway

- Formula mass for ionic compounds is the sum of atomic masses in one formula unit.
- Formula mass represents the mass of the smallest repeating unit in the crystal lattice.
- Formula mass is expressed in atomic mass units (amu) or grams per mole (molar mass).

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3.5: Ionic Nomenclature

Naming Ions

In this section, we will begin the formal study of **nomenclature**, the systematic method used to name chemical compounds. As we dive into this important topic, you'll learn how to name and recognize various ions, which are key building blocks in the formation of ionic compounds. By understanding the names and formulas of individual ions, you will be just one step away from mastering the process of naming ionic compounds. This skill is essential in chemistry, as it allows you to communicate the identity and composition of substances clearly and consistently. Let's get started on this exciting journey into chemical naming.

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. 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.

The systematic name (Stock name) of cations is based on the charge of the metal ion. For metals that can form more than one cation, the charge is indicated in parentheses using Roman numerals. Here's the general rule:

- For monatomic cations (single atom ions), the name of the element is used followed by the word "ion." For example, Na⃟ is called sodium ion.
- For metals that form multiple cations, the charge is indicated using Roman numerals. For example, Fe²⃟ is iron(II) ion, and Fe³⃟ is iron(III) ion.

Table 3.5.1 lists cations with more than one charge, it directly refers to elements that can form ions with different charges, and these are typically represented with their systematic names using Roman numerals. These cations often come from transition metals.

Table 3.5.1: The systematic name of cations is based on the charge of the metal ion.

Element	Charge	Roman Numeral	Systematic/Stock Name
Iron	Fe^{2+}	Iron(II)	Iron(II) ion
	Fe^{3+}	Iron(III)	Iron(III) ion
Copper	Cu^{+}	Copper(I)	Copper(I) ion
	Cu^{2+}	Copper(II)	Copper(II) ion
Tin	Sn^{2+}	Tin(II)	Tin(II) ion
	Sn^{4+}	Tin(IV)	Tin(IV) ion
Lead	Pb^{2+}	Lead(II)	Lead(II) ion
	Pb^{4+}	Lead(IV)	Lead(IV) ion
Chromium	Cr^{2+}	Chromium(II)	Chromium(II) ion
	Cr^{3+}	Chromium(III)	Chromium(III) ion
Gold	Au^{+}	Gold(I)	Gold(I) ion
	Au^{3+}	Gold(III)	Gold(III) 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 3.5.2 lists the names of some common monatomic ions.

Table 3.5.2: 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 polyatomic ions have their own characteristic names, as discussed earlier.

Example 3.5.1

Name each ion.

a. Ca^{2+}
b. SO_3^{2-}

a. Solution

Step 1: Identify the element

- **Ca** stands for calcium, which is an alkaline earth metal in Group 2 of the periodic table.

Step 2: Determine the charge

- **Ca^{2+}** indicates that the calcium ion has a +2 charge. Calcium typically loses two electrons to form this cation.

Step 3: Name the ion

- Metal ions like calcium are named by simply using the name of the element (Calcium) followed by the word "ion" to indicate the charged form.

Result: The name of Ca^{2+} is calcium ion.

b. Solution

Step 1: Identify the group of atoms

- SO_3^{2-} is a polyatomic ion, made of sulfur (S) and oxygen (O) atoms.

Step 2: Determine the charge

- The ion has a charge of -2, as indicated by the 2- sign.

Step 3: Name the ion

- The ion SO_3^{2-} is named Sulfite. It's important to note that sulfite is different from sulfate (SO_4^{2-}), which has one more oxygen atom. The suffix "-ite" indicates fewer oxygen atoms, while "-ate" would indicate a higher oxygen count.

Result: The name of SO_3^{2-} is Sulfite ion.

Naming Ionic Compounds

When naming ionic compounds, the basic rule is to first identify and name the cation (the positively charged ion), followed by the anion (the negatively charged ion). Both parts of the name are written without the word "ion" at the end. The order of naming is always cation first, then anion. This is a standard naming convention used for both simple and polyatomic ionic compounds.

Figure 3.5.1 depicts how the name of an ionic compound formed between Ba^{2+} and NO_3^- is deduced.

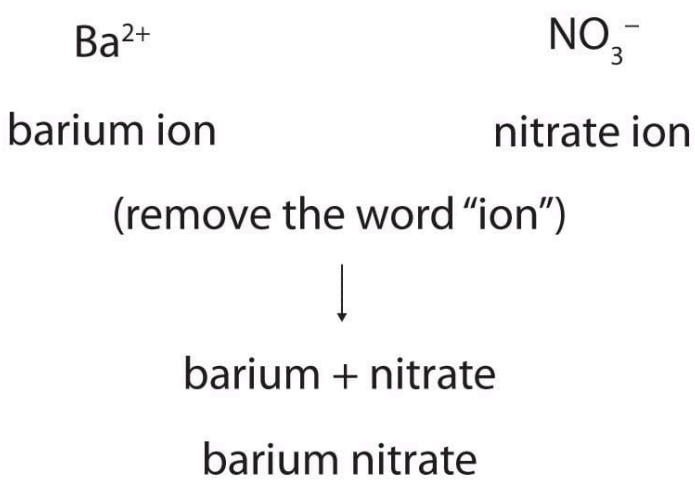


Figure 3.5.1: Barium nitrate ($Ba(NO_3)_2$) is named by identifying the cation barium (Ba^{2+}) with a fixed +2 charge and the anion nitrate (NO_3^-), resulting in the name barium nitrate. (CC BY-NC-SA 3.0; anonymous via source). Detailed Description of Figure 3.5.1.

The compound's name does not indicate that there are two nitrate ions for every barium ion. You must determine the relative numbers of ions by balancing the positive and negative charges.

Writing Formulas for Ionic Compounds

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. If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively.

Figure 3.5.2 indicates the synopsis of how to name simple ionic compounds, providing a clear and concise guide for identifying and naming cations and anions in ionic compounds. The process involves identifying the cation and anion first. If the cation can have more than one possible charge, the Stock system is used, where Roman numerals in parentheses indicate the charge of the cation. For example, $FeCl_2$ would be named iron(II) chloride (or ferrous chloride), and $CuSO_4$ would be named copper(II) sulfate (or cupric sulfate). If the cation has only one possible charge, the name of the cation is followed by the name of the anion, such as KBr (potassium bromide) or $NaNO_3$ (sodium nitrate). This method ensures precise and systematic naming of ionic compounds,

especially for elements with multiple oxidation states.

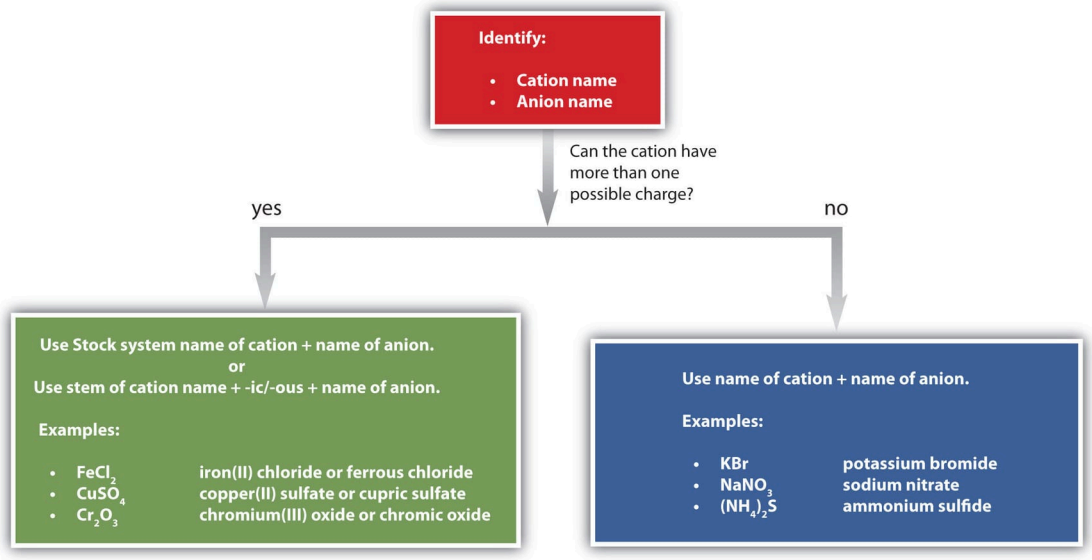


Figure 3.5.2: A Guide to Naming Simple Ionic Compounds. Follow these steps to name a simple ionic compound. Details in Text. (CC BY-NC-SA 3.0; anonymous via source). Detailed Description of Figure 3.5.2.

Example 3.5.2

Name the ionic compound formed by each pair of ions.

a. Na^+ and Br^-
b. Mg^{2+} and Br^-

a. Solution

Step 1: Identify the ions

- Na^+ is a sodium ion with a charge of +1.
- Br^- is a bromide ion with a charge of -1.

Step 2: Combine the ions

- To form a neutral ionic compound, the total positive charge must balance the total negative charge.
- Since Na^+ has a charge of +1 and Br^- has a charge of -1, they combine in a 1:1 ratio to form a neutral compound.

Step 3: Write the formula

- The resulting ionic compound is NaBr (sodium bromide).

b. Solution

Step 1: Identify the ions

- Mg^{2+} is a magnesium ion with a charge of +2.
- Br^- is a bromide ion with a charge of -1.

Step 2: Combine the ions

- To balance the charges, we need two Br^- ions for each Mg^{2+} ion, because two negative charges will balance the two positive charges.
- Therefore, the ions combine in a 1:2 ratio.

Step 3: Write the formula

- The resulting ionic compound is $MgBr_2$ (magnesium bromide).

Example 3.5.3

Name the ionic compound formed by each pair of ions.

a. Cr^{3+} and NO_3^-
b. Fe^{2+} and PO_4^{3-}

a. Solution

Step 1: Identify the ions

- Cr^{3+} is a chromium ion with a charge of +3.
- NO_3^- is a nitrate ion with a charge of -1.

Step 2: Combine the ions

- To form a neutral ionic compound, the total positive charge must balance the total negative charge.
- The chromium ion has a charge of +3, and the nitrate ion has a charge of -1.
- To balance the charges, we need three nitrate ions (NO_3^-) for every chromium ion (Cr^{3+}) because three negative charges will balance the three positive charges.

Step 3: Write the formula

- The resulting ionic compound is $Cr(NO_3)_3$ (chromium(III) nitrate).

b. Solution

Step 1: Identify the ions

- Fe^{2+} is an iron(II) ion with a charge of +2.
- PO_4^{3-} is a phosphate ion with a charge of -3.

Step 2: Combine the ions

- To balance the charges, we need a combination of Fe^{2+} and PO_4^{3-} such that the total positive charge equals the total negative charge.
- The iron(II) ion has a charge of +2, and the phosphate ion has a charge of -3.
- To balance the charges, we need 3 Fe^{2+} ions to balance the 2 PO_4^{3-} ions because 3(2) = 6 (positive charge) and 2(3) = 6 (negative charge).

Step 3: Write the formula

- The resulting ionic compound is $Fe_3(PO_4)_2$ (iron(II) phosphate).

Key Takeaways

- Name the cation first, then the anion (drop “ion”).
- Use Roman numerals for cations with multiple charges.

- Add “-ide” to simple anions; use common names for polyatomic ions.

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3.6: Characteristics of Ionic Compounds

Physical Properties of Ionic Compounds

Naturally occurring ionic crystals often display remarkable color and brilliance. Their distinct geometric shapes result from the highly regular and orderly arrangement of ions within their **crystal lattice**. This structured pattern enables the crystals to grow in a range of angular, symmetrical forms, enhancing their visual appeal. The vibrant colors seen in many of these crystals are largely due to the presence of transition metal ions. These metals interact with light in specific ways, absorbing and reflecting certain wavelengths based on electronic transitions within the ions. Figure 3.6.1 highlights several examples of these colorful ionic crystals, showcasing the relationship between their internal structure and striking external appearance.

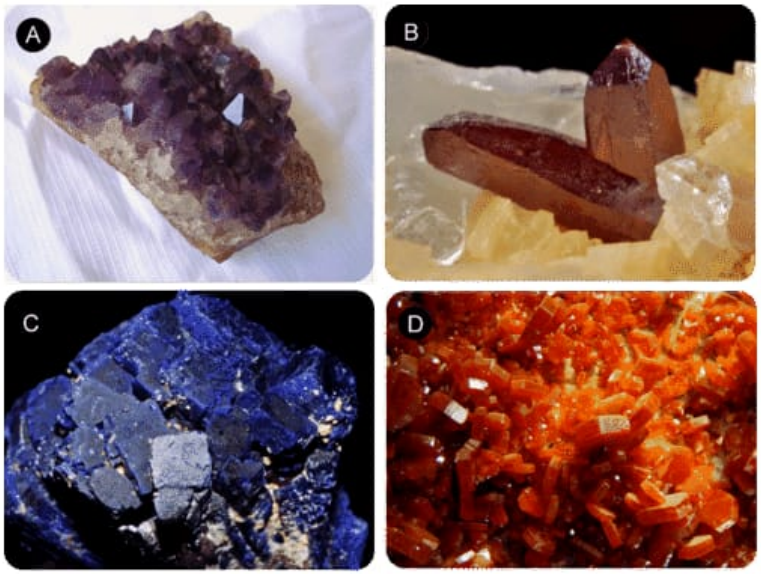


Figure 3.6.1: In nature, the ordered arrangement of ionic solids gives rise to beautiful crystals, (A)Amethyst (B) Cinnabar (C) Azurite (D) Vanadinite(CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 3.6.1.

Ionic compounds have high melting and boiling points due to strong ionic bonds, are hard but brittle, and do not conduct electricity in solid form. They can conduct electricity when melted or dissolved in water. Many are soluble in water and tend to form crystalline structures, with ions arranged in a repeating pattern. These properties stem from the strong electrostatic forces between ions.

Melting Points

Because of the many simultaneous attractions between cations and anions that occur, ionic crystal lattices are very strong. The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal. For example, sodium chloride has a melting temperature of about 800°C. As a comparison, the molecular compound water melts at 0 °C.

Example 3.6.1

Which substance has a higher melting point: calcium oxide (CaO) or ethanol (C_2H_5OH)? Explain your reasoning.

Solution

Calcium oxide (CaO) has a higher melting point than ethanol because CaO is an ionic compound composed of calcium ions (Ca^{2+}) and oxide ions (O^{2-}) held together by strong electrostatic forces in a crystal lattice. These strong ionic bonds require a lot of energy to break, resulting in a high melting point. Ethanol is a molecular compound held together by weaker intermolecular forces such as hydrogen bonds, so it has a much lower melting point compared to CaO.

Shattering

Ionic compounds are generally hard, but brittle. Why? It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbor. However, when that happens, it brings ions of the same charge next to each other (see Figure 3.6.2 below). The repulsive forces between like-charged ions cause the crystal to shatter. When an ionic crystal breaks, it tends to do so along smooth planes because of the regular arrangement of the ions.

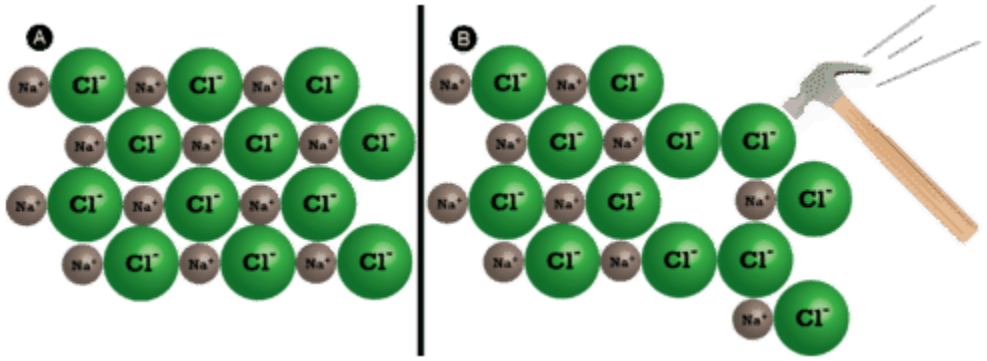


Figure 3.6.2: (A) The sodium chloride crystal is shown in two dimensions. (B) When struck by a hammer, the negatively-charged chloride ions are forced near each other and the repulsive force causes the crystal to shatter. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 3.6.2.

Conductivity

The electrical conductivity of a solution depends on whether the dissolved substance produces ions. Ionic compounds like sodium nitrate ($NaNO_3$) dissolve in water by fully separating into sodium and nitrate ions. These free ions move easily and carry electric current, so the solution conducts electricity very well. On the other hand, weak acids such as carbonic acid (H_2CO_3) only partially ionize in water. Most of the acid remains as whole molecules, with only a small portion forming hydrogen and bicarbonate ions. This limited ion production results in low conductivity. Lastly, molecular compounds like sugar (sucrose, $C_{12}H_{22}O_{11}$) do not produce ions at all when dissolved. They remain as neutral molecules, so the solution does not conduct electricity. These examples show how the type of substance and the degree of ion formation directly affect the **electrical conductivity** of a solution.

Figure 3.6.3 indicates the conductivity comparison of three different solutions. The figure shows three experiments in which two electrodes, connected to a light bulb, are placed in beakers containing three different substances: ethanol (a molecular compound), potassium chloride (KCl, an ionic compound), and acetic acid (weak acid) solution. The brightness of the light bulb indicates the level of conductivity in each solution.

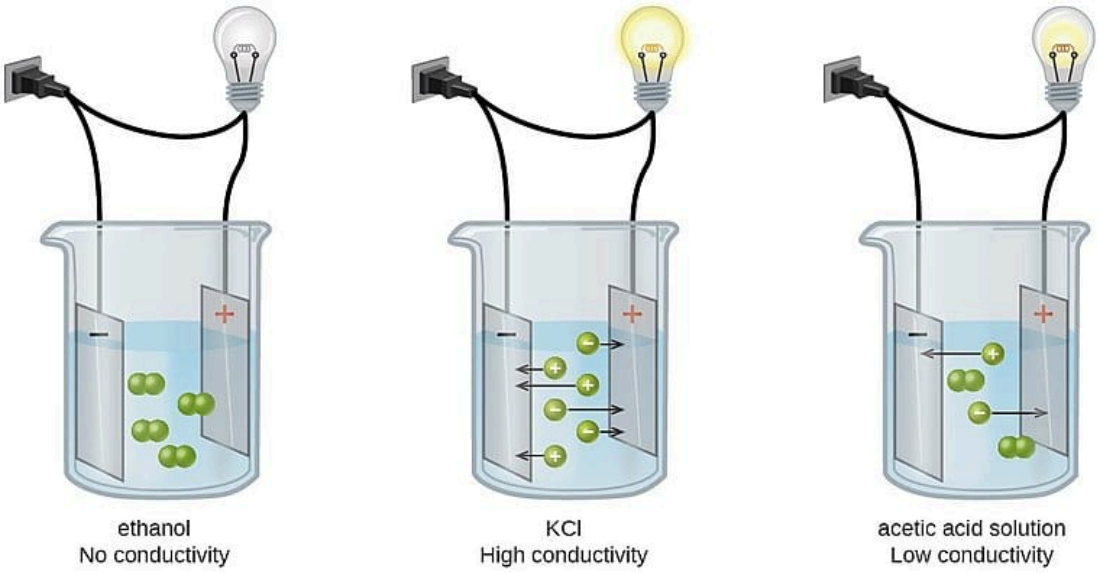


Figure 3.6.3: The conductivity of ethanol (no conductivity), potassium chloride (KCl) with high conductivity and acetic acid solution with low conductivity on the right. (CC BY-SA 4.0; Karishma50 via Wikimedia). Detailed Description of Figure 3.6.3.

Example 3.6.2

Compare the melting point and electrical conductivity of magnesium bromide ($MgBr_2$) and methanol (CH_3OH). Explain the reasons for the differences between the two substances.

Solution

Magnesium bromide ($MgBr_2$) is an ionic compound composed of magnesium ions and bromide ions held together by strong electrostatic forces in a crystal lattice. Because of these strong ionic bonds, $MgBr_2$ has a high melting point. When dissolved in water, $MgBr_2$ dissociates completely into ions, allowing the solution to conduct electricity well.

Methanol (CH_3OH) is a molecular compound made of neutral molecules held together by relatively weaker intermolecular forces, mainly hydrogen bonding. These forces require less energy to overcome, so methanol has a much lower melting point compared to $MgBr_2$. Additionally, methanol does not ionize in water, so it does not conduct electricity when dissolved.

Key Takeaways

- Ionic compounds have high melting points.
- Ionic compounds are hard and brittle.
- Ionic compounds dissociate into ions when dissolved in water.
- Solutions of ionic compounds and melted ionic compounds conduct electricity, but solid materials do not.
- An ionic compound can be identified by its chemical formula: metal + nonmetal or polyatomic ions.

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3.7: Summary

Ionic compounds are formed when atoms transfer electrons, resulting in the formation of positively charged ions (cations) and negatively charged ions (anions). This transfer usually occurs between metals and nonmetals. Metals, which have a tendency to lose electrons, become positively charged ions, while nonmetals gain electrons and become negatively charged ions. The electrostatic force of attraction between these oppositely charged ions holds the compound together in a stable structure.

Ionic compounds typically have high melting and boiling points due to the strong ionic bonds between the ions. They are usually crystalline solids at room temperature and tend to dissolve in water, where the ions separate and move freely. This ability to dissociate in water gives ionic compounds the property of electrical conductivity when dissolved or melted, as the ions are free to move.

Examples of ionic compounds include sodium chloride (NaCl), where sodium (Na) donates an electron to chlorine (Cl), forming $Na^{\diamondsuit}$ and $Cl^{\diamondsuit}$ ions, which are attracted to each other and form a stable lattice structure. Ionic compounds are also generally brittle, and their structure tends to shatter when subjected to stress.

In summary, ionic compounds are essential in many chemical processes and everyday materials, characterized by their high melting points, electrical conductivity in molten or aqueous form, and crystal lattice structures.

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3.8: Key Terms

- Chemical bond: The force of attraction between atoms that allows the formation of molecules and compounds, typically resulting from the interaction of electrons in the outer shells of atoms.
- Octet rule: A chemical principle that atoms tend to form bonds in a way that results in eight electrons in their outermost electron shell, which provides greater stability.
- Ions: Charged particles that are formed when atoms or molecules lose or gain one or more electrons, leading to an imbalance between the number of protons and electrons, resulting in a positive charge (cation) or negative charge (anion).
- Ionic bond: A type of chemical bond formed when one atom transfers electrons to another, resulting in the formation of oppositely charged ions that attract each other.
- Valence Electrons: The electrons in the outermost shell of an atom that are involved in chemical bonding.
- Electron Transfer: The process by which electrons are transferred from one atom to another during the formation of an ionic bond, creating cations and anions.
- Cation: A positively charged ion formed when an atom loses one or more electrons. This occurs because the atom now has more protons than electrons, resulting in a net positive charge. Most metals form cations.
- Isoelectronic: Describes two atoms or ions that have the same electron configuration. For example, the sodium ion (Na^+) is isoelectronic with the neon atom because both have the same number of electrons in their outermost shell.
- Anion: A negatively charged ion formed when an atom gains one or more electrons, resulting in more electrons than protons. Most nonmetals form anions.
- Lewis Electron Dot Diagram: A visual representation of an atom's valence electrons, using dots placed around the element symbol to show possible bonding and lone pairs.
- Main group elements: Elements in Groups 1, 2, and 13–18; they have predictable valence electrons and include metals, nonmetals, and metalloids.
- Charge balance: The principle that the total positive and negative charges in a compound or reaction must be equal, resulting in an overall neutral system.
- Net zero charge: A condition where the total positive and negative charges in a compound or molecule cancel out, resulting in an overall charge of zero.
- Chemical Formula: Chemical formula: A chemical formula represents the elements in a compound and their relative proportions using symbols and subscripts. For example, H_2O represents two hydrogen atoms and one oxygen atom in a water molecule.
- Formula unit: The simplest ratio of ions in an ionic compound that represents the smallest whole number ratio of the ions involved. It is used to describe ionic compounds in their solid form, as opposed to molecules, which apply to covalent compounds.
- Empirical Formula: The simplest whole-number ratio of elements in a compound.
- Lowest-whole number ratio: The simplest ratio of atoms or ions in a compound that reflects its empirical formula, showing the smallest possible integers that maintain charge balance.
- Polyatomic ions: Ions that consist of two or more atoms covalently bonded together, carrying a charge as a group. These ions behave as a single entity in chemical reactions. Examples include the sulfate ion (SO_4^{2-}) and nitrate ion (NO_3^-).
- Formula mass: The sum of the atomic masses of all the atoms in a formula unit of a compound. It is calculated by adding up the atomic masses (in atomic mass units, amu) of each element in the compound, based on the number of atoms of each element present in the formula.
- Nomenclature: The system of naming chemical compounds according to specific rules to ensure consistency and clarity in identifying substances.
- Crystal Lattice: A three-dimensional arrangement of ions in an ionic solid, where each ion is surrounded by oppositely charged ions in a regular, repeating pattern.
- Electrical conductivity: The ability of ionic compounds to conduct electricity when dissolved in water or melted, due to the movement of ions.

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3.9: Alternative text for Chapter 3

Figure 3.2.1: The Formation of a Sodium Ion

The image is divided into two parts: On the left is a sodium atom (Na) with 11 protons and 11 electrons arranged in three energy levels; 2 in the first, 8 in the second, and 1 valence electron in the third ($3s^1$). This outer electron is easily lost. An arrow points to the right where the sodium (Na^+) is shown. After losing one electron, Na⦰ has 10 electrons with an electron configuration of $1s^2$ $2s^2$ $2p^6$, identical to that of neon, making it isoelectronic with neon. The ion is more stable with a full octet in the second energy level, and the process is represented by the equation: $Na \rightarrow Na^+ + e^-$.

Figure 3.2.2: The Formation of a Chlorine Ion

The image is divided into two parts: A chlorine atom (Cl), shown on the left, has 17 protons and 17 electrons. These electrons are arranged in three energy levels: 2 in the first shell, 8 in the second, and 7 in the third, with the third level containing seven valence electrons ($3s^2$ $3s^5$). Because it is one electron short of a full octet, chlorine readily gains one electron to become more stable. This process forms a chloride ion (Cl^-) shown on the right, which now has 18 electrons. The additional electron completes the third energy level, giving the chloride ion a full outer shell ($3s^2$ $3p^6$). The resulting electron configuration is $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$, which is the same as that of the noble gas argon, making the chloride ion isoelectronic with argon. This gain of an electron gives the ion a negative charge (-1). The overall process is represented by the equation: $Cl + e^- \rightarrow Cl^-$.

Figure 3.2.3: Relationship Between Periodic Table Groups and Ionic Charges

The image shows a simplified chart of common ionic charges arranged by periodic table groups. On the left side are Groups 1A and 2A. Group 1A contains positive one ions listed vertically from top to bottom as hydrogen ion H-plus, lithium ion Li-plus, sodium ion Na-plus, potassium ion K-plus, and rubidium ion Rb-plus. Group 2A is immediately to the right and contains positive two ions listed as magnesium ion Mg-two-plus, calcium ion Ca-two-plus, and strontium ion Sr-two-plus.

On the right side of the image is a block representing Groups 3A through 8A. In Group 3A, aluminum ion Al-three-plus is shown. Group 5A lists nitrogen ion N-three-minus and phosphorus ion P-three-minus. Group 6A includes oxygen ion O-two-minus, sulfur ion S-two-minus, and selenium ion Se-two-minus. Group 7A lists halide ions fluorine F-minus, chlorine Cl-minus, bromine Br-minus, and iodine I-minus arranged from top to bottom. Group 8A is labeled but contains no ions, indicating noble gases typically do not form common ions.

Figure 3.2.4: Lewis Electron Dot Diagrams Represent the Valence Electrons of an Element

Lewis electron dot diagrams represent the valence electrons of an element as dots placed around the element's chemical symbol. For the elements lithium (Li) through neon (Ne), which include the second period of the periodic table, these diagrams show the number of valence electrons corresponding to each element's group. Lithium has one dot, representing its single valence electron, while neon has eight dots arranged around its symbol to represent a full valence shell. These diagrams help visualize how atoms bond by sharing or transferring valence electrons.

Figure 3.2.5: Electron Transfer Process During the Formation of Sodium Chloride (NaCl)

The image illustrates the electron transfer process during the formation of sodium chloride (NaCl). An arrow shows the single valence electron from the sodium (Na) atom being transferred to the chlorine (Cl) atom. This transfer results in the formation of a positively charged sodium ion (Na⁺) and a negatively charged chloride ion (Cl⁻). Another arrow indicates the electrostatic attraction between these oppositely charged ions, leading to the formation of a neutral ionic compound, NaCl.

Figure 3.2.6: Electron Transfer Process During the Formation of Lithium Bromide (LiBr)

The image illustrates the electron transfer process in the formation of lithium bromide (LiBr). An arrow shows the single valence electron from the lithium (Li) atom transferring to the bromine (Br) atom. This results in a positively charged lithium ion (Li⁺) and a negatively charged bromide ion (Br⁻). Another arrow indicates the electrostatic attraction between these oppositely charged ions, which combines them into a neutral ionic compound, LiBr.

Figure 3.2.7: Electron Transfer Process During the Formation of Magnesium Bromide ($MgBr_2$)

The image depicts the electron transfer process during the formation of magnesium bromide ($MgBr_2$). Two arrows show the two valence electrons from the magnesium (Mg) atom being transferred—one to each bromine (Br) atom. This results in a magnesium ion with a +2 charge (Mg^{2+}) and two bromide ions each with a –1 charge (Br^-). Another set of arrows indicates the electrostatic attractions between the positively charged magnesium ion and the two negatively charged bromide ions, forming the neutral ionic compound $MgBr_2$.

Figure 3.2.8: Common and Variable Charges of Monatomic Ions in the Periodic Table

This chart summarizes predictable ionic charges based on periodic table groups. Elements in Group 1A typically form positive one ions, Group 2A form positive two ions, and aluminum in Group 3A forms a positive three ion. On the nonmetal side, Group 5A elements shown form negative three ions, Group 6A form negative two ions, and Group 7A form negative one ions. Group 8A noble gases generally do not form common ions. The image helps students recognize the pattern that metals tend to lose electrons and form positive ions, while nonmetals tend to gain electrons and form negative ions.

Figure 3.3.1: The Crystal Lattice Structure of Sodium Chloride (NaCl)

The image depicts the crystal lattice structure of sodium chloride (NaCl), showing a repeating three-dimensional arrangement of ions. In this lattice, positively charged sodium ions (Na⁺) and negatively charged chloride ions (Cl⁻) alternate in all directions, creating a stable and balanced structure through strong electrostatic attractions. Each sodium ion is surrounded by several chloride ions, and each chloride ion is similarly surrounded by sodium ions. This arrangement forms a continuous network without discrete, individual NaCl molecules. Instead, the compound exists as an extended ionic solid where the ratio of sodium to chloride ions is consistently 1:1. The image helps visualize how ionic compounds form large, rigid crystals rather than separate molecules.

Figure 3.3.2: Crossing Charges Method for Writing Ionic Formulas

Diagram illustrating the crossing charges method for writing ionic formulas. The metal ion M has a charge of n^+, and the nonmetal ion X has a charge of m^-. Arrows show the charge number n from M pointing downward to become the subscript next to X, and the charge number mmm from X pointing downward to become the subscript next to M. This swapping of charge magnitudes results in the formula M_mX_n, representing the balanced ratio of ions in the compound.

Figure 3.3.3: Writing the Chemical Formula for Aluminum Oxide Using the Crossing Charges Method

Diagram showing the charges on aluminum (Al) and oxygen (O) ions used to write the chemical formula. Aluminum has a charge of 3^+ and oxygen has a charge of 2^-. Two arrows illustrate the crossing charges method: the 3 from aluminum becomes the subscript for oxygen, and the 2 from oxygen becomes the subscript for aluminum. This results in the chemical formula Al_2O_3, representing the balanced ionic compound formed by aluminum and oxygen.

Figure 3.3.4: Writing the Chemical Formula for Using the Crossing Charges Method

The image presents a two-part question asking for the chemical formulas of compounds formed by (a) barium and nitrate, and (b) ammonium and phosphate ions. In part (a), barium has a charge of 2^+ and nitrate (NO_3) has a charge of 1^-. Using the crossing charges method, the charge of barium (2) becomes the subscript for nitrate, resulting in the formula $Ba(NO_3)_2$. In part (b), ammonium (NH_4) has a charge of 1^+ and phosphate (PO_4) has a charge of 3^-. Crossing these charges gives the formula $(NH_4)_3PO_4$, showing the balanced ionic compound.

Figure 3.4.1: Calculating the Formula Mass of Sodium Chloride (NaCl)

The image has two sections. On the left, a table calculates the formula mass of sodium chloride (NaCl). The first column lists the elements: sodium (Na) and chlorine (Cl). The second column shows the quantity of each element in one formula unit: 1 for Na and 1 for Cl. A multiplication sign appears between the quantity and the average atomic mass in the fourth column: 22.99 amu for Na and 35.45 amu for Cl. The fifth column shows the subtotal for each element's mass: 22.99 amu for Na and 35.45 amu for Cl. At the bottom, the total formula mass is shown as 58.44 amu. On the right side of the image, a 3D crystal structure of NaCl is displayed, showing a cubic lattice of alternating sodium and chloride ions, each surrounded by oppositely charged ions.

Figure 3.4.2: Calculating the Formula Mass of Aluminum Sulfate, $Al_2(SO_4)_3$

The image is divided into two sections. On the left, a table shows the calculation of the formula mass for aluminum sulfate, $Al_2(SO_4)_3$. The first column lists the components: Al, S, and O. The second column shows the quantities: 2 aluminum atoms, 3 sulfur atoms, and 12 oxygen atoms (from 3 sulfate groups, each containing 4 oxygen atoms). A multiplication sign separates the quantity from the average atomic mass listed in the fourth column: 26.98 amu for Al, 32.06 amu for S, and 16.00 amu for O. The fifth column provides the subtotals: 53.96 amu from Al, 96.18 amu from S, and 192.00 amu from O. The final row shows the total formula mass as 342.14 amu. On the right side of the image, a 3D crystal lattice structure illustrates alternating aluminum and sulfate ions arranged in a regular pattern, representing the compound's solid-state structure.

Figure 3.5.1: Naming an Ionic Compound with a Polyatomic Ion

The image shows the process for naming the ionic compound containing barium and nitrate. On the left, the cation is identified as barium (Ba^{2+}), a Group 2 metal with a fixed +2 charge. On the right, the anion is identified as nitrate (NO_3^-), a polyatomic ion with a -1 charge. Two nitrate ions are required to balance the +2 charge of one barium ion, resulting in the formula $Ba(NO_3)_2$. Since barium has only one possible charge, no Roman numeral is needed in the name. The compound is named simply barium nitrate.

Figure 3.5.2: Flowchart or Structured Guide for Naming Simple Ionic Compounds

The image provides a flowchart or structured guide for naming simple ionic compounds. It begins with identifying the cation and anion in the chemical formula. If the cation has more than one possible charge (commonly transition metals), the Stock system

is used: the cation's name is followed by a Roman numeral in parentheses to indicate its charge—for example, $FeCl_2$ is named iron(II) chloride, and $FeCl_3$ is named iron(III) chloride. The chart may also show the alternate common system names: ferrous chloride and ferric chloride, respectively. If the cation has only one possible charge (such as alkali or alkaline earth metals), the Roman numeral is omitted, and the name of the cation is directly followed by the anion name—for instance, KBr is potassium bromide and $NaNO_3$ is sodium nitrate. The visual helps users distinguish when Roman numerals are required and reinforces systematic naming rules for ionic compounds. Labels include:

Synopsis of how to name simple ionic compounds

Top box:
Identify:

- Cation name
- Anion name

Decision question:
Can the cation have more than one possible charge?

Two branches follow:

Left branch: yes

Instruction box:
Use Stock system name of cation + name of anion.
or
Use stem of cation name + -ic/-ous + name of anion.

Examples:

- FeCl2 — iron(II) chloride or ferrous chloride
- CuSO4 — copper(II) sulfate or cupric sulfate
- Cr2O3 — chromium(III) oxide or chromic oxide

Right branch: no

Instruction box:
Use name of cation + name of anion.

Examples:

- KBr — potassium bromide
- NaNO3 — sodium nitrate
- (NH4)2S — ammonium sulfide

Figure 3.6.1: Ionic Solids Gives Rise to Beautiful Crystals

The image is divided into four parts, each showing a different mineral. In the top left is Amethyst, a purple variety of quartz (SiO_2) whose color comes from iron ions within its crystal structure. The top right shows Cinnabar, a bright red mineral made of mercury(II) sulfide (HgS) and the primary ore of mercury. The bottom left features Azurite, a deep blue copper mineral with the formula $Cu_3(CO_3)_2(OH)_2$, known for its vibrant color and unique crystal forms. Finally, the bottom right displays Vanadinite, a reddish-orange mineral composed of lead chloro vanadate ($Pb_5(VO_4)_3Cl$), which is the main ore of vanadium and forms distinctive hexagonal crystals.

Figure 3.6.2: Structure and Fracture of a Sodium Chloride Crystal

The image is divided into two parts. On the left (A), a two-dimensional model of a sodium chloride (NaCl) crystal is shown, illustrating the regular arrangement of positively charged sodium ions and negatively charged chloride ions in a repeating, grid-like structure. On the right (B), the same crystal is shown after being struck by a hammer. The force causes the layers to shift,

bringing like-charged chloride ions close together. Their repulsion causes the structure to break apart, demonstrating why ionic crystals like NaCl shatter rather than bend under pressure.

Figure 3.6.3: The Conductivity of Three Solutions

The figure is divided into three sections, each showing the conductivity of a different liquid solution. On the left, a container filled with ethanol is shown. There are no glowing bulbs or visible electrical activity, indicating no conductivity, as ethanol is a non-electrolyte and does not produce ions in solution. In the middle, a container holds a potassium chloride (KCl) solution. Electrodes are connected to a circuit, and a bright light bulb is glowing, demonstrating high conductivity. This is because KCl is a strong electrolyte that fully dissociates into potassium and chloride ions in water. On the right, the container contains a solution of acetic acid, a weak acid. The light bulb glows dimly, indicating low conductivity. This occurs because acetic acid is a weak electrolyte that only partially ionizes in solution, producing fewer free ions. The image illustrates the concept that ionic and molecular compounds differ significantly in their ability to conduct electricity in aqueous solutions.

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4: COVALENT COMPOUNDS AND BONDING

Introduction

Atoms form chemical bonds to achieve greater stability, usually by completing their outer electron shells, similar to noble gases. There are two main types of bonds: Ionic bonds, where electrons are transferred between atoms forming charged ions, and covalent bonds, where atoms share electron pairs. Covalent bonds typically form between non-metal atoms with similar electron affinities, allowing them to achieve stable configurations by sharing electrons. For example, two hydrogen atoms share electrons to form H_2.

This chapter focuses on covalent bonding, which is essential for creating a variety of compounds, from simple molecules like hydrogen to complex biological molecules like cholesterol. Cholesterol's structure ($C_{27}H_{46}O$) is a complex arrangement of carbon, hydrogen, and oxygen atoms forming a multi-ringed molecule with both rigid and flexible regions, as shown in Figure 4.1. This unique structure allows cholesterol to play important roles in maintaining cell membrane stability and serving as a precursor for certain hormones.

However, excessive cholesterol intake or poor metabolism can cause cholesterol buildup (plaque) in arteries, leading to heart attacks, strokes, and circulatory problems. To reduce these risks, doctors often recommend diets low in cholesterol, blood tests to monitor levels, and medications to lower cholesterol.

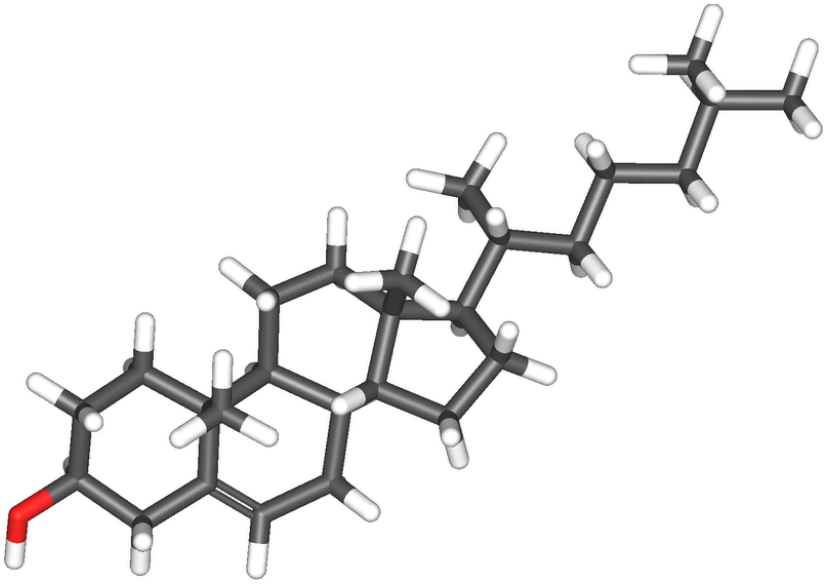


Figure 4.1: Stick Model of Cholesterol. (CC BY-NC-SA 3.0; Sbrools via Wikimedia). Detailed Description of Figure 4.1.

Learning Objectives

- Understand covalent bond formation through the sharing of electrons between nonmetals.
- Identify and differentiate between polar and nonpolar covalent bonds.
- Draw Lewis structures to represent covalent bonds and electron sharing.
- Predict molecular shapes using the VSEPR theory.
- Understand the properties of covalent compounds, such as low melting points and poor electrical conductivity.
- Name and write formulas for covalent compounds using systematic naming conventions.
- Compare and contrast covalent and ionic bonds in terms of their formation, properties, and behavior.
- Understanding covalent bonding in simple organic compounds.

4.1: Formation of a Covalent Bond

4.2: Covalent versus Ionic Compounds
4.3: Drawing Lewis Structures
4.4: Characteristics of Covalent Bonds
4.5: Understanding Molecular Shape and Bonding
4.6: Introduction to Organic Chemistry
4.7: Summary
4.8: Key Terms
4.9: Alternative text for Chapter 4

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4.1: Formation of a Covalent Bond

Introduction

Covalent bonding is a fundamental concept in chemistry where two atoms share electrons in order to achieve a more stable electron configuration, usually resembling the nearest noble gas configuration. This type of bond typically occurs between nonmetals, as they have similar tendencies to attract electrons.

The formation of a **covalent bond** happens when atoms, particularly those of nonmetals, come close together. Instead of transferring electrons (as in ionic bonding), the atoms share one or more pairs of electrons. This **electron sharing** allows each atom to fill its outer electron shell, leading to a more stable arrangement.

Covalent bonds

Covalent bonds are formed when two atoms share electrons to achieve a more stable electron configuration. Atoms can share different numbers of electrons, leading to single, double, or triple bonds, depending on how many pairs of electrons are shared between the atoms.

Formation of Single Bonds

A **single bond** is formed when two atoms share one pair of electrons (two electrons in total). Single bonds are the simplest and most common type of covalent bond, primarily formed between nonmetal atoms. The shared electrons allow each atom to achieve a more stable electron configuration, typically resembling that of the nearest noble gas.

Figure 4.1.1 illustrates the formation of single covalent bonds in all four molecules with one (HCl), two (H_2O), three (NH_3) and four (CH_4) single bonds respectively. When a nonmetal atom combines with another nonmetal atom, they usually make a covalent bond. A covalent bond is a pair of shared electrons, called a **bond pair** of electrons, where each bonded atom contributes one electron. For example, chlorine has seven valence electrons and needs one more to complete its octet. Hydrogen has one valence electron and requires one more to acquire the electron configuration of helium, i.e., **duet** instead of the octet. Hydrogen and chlorine combine by sharing one electron to make the compound HCl. Similarly, oxygen, nitrogen, and carbon make 2, 3, and 4 covalent bonds with hydrogen to complete their octet and make compounds H_2O, NH_3, and CH_4, respectively, as illustrated below.

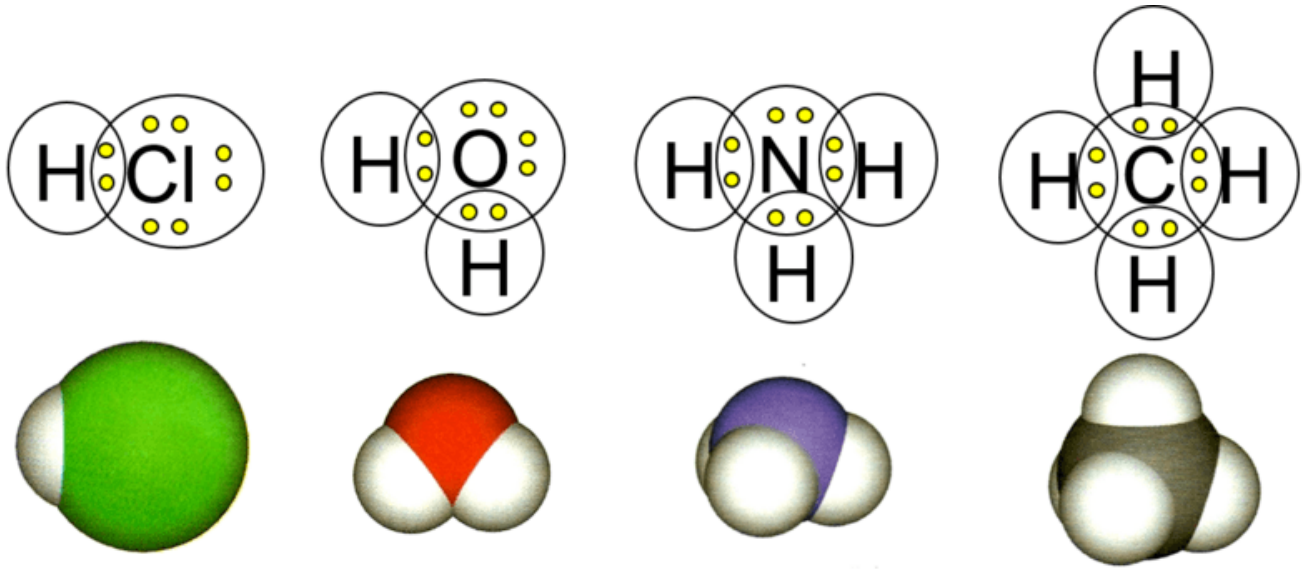


Figure 4.1.1: Examples of covalent bonds where Cl, O, N, and C acquire octet by sharing 1, 2, 3, and 4 electrons, respectively. Hydrogen shares one electron and acquires a duet, i.e., the electron configuration of helium. 2nd row is the space-filling model of the molecule where the white sphere is hydrogen, and the colored is the other atom. (CC BY-SA 4.0; Drago Karlo via Wikimedia). Detailed Description of Figure 4.1.1.

Formation of a Double Bond

A **double bond** occurs when two atoms share two pairs of electrons (a total of 4 electrons). Double bonds are typically formed between nonmetals and are stronger than single bonds but weaker than triple bonds. This type of bond is commonly seen in molecules where atoms need to share more electrons to achieve a stable electron configuration, typically an octet.

In the formation of an O_2 molecule, two oxygen atoms, each with 6 valence electrons, come together to share electrons and achieve a stable electron configuration. Each oxygen atom needs 2 more electrons to complete its octet, so they share two pairs of electrons, forming a double bond. This allows both oxygen atoms to have 8 electrons in their outer shell—6 from their own electrons and 2 from the shared bonding electrons. The resulting structure is a stable O_2 molecule, with each oxygen atom having a complete octet of electrons in its valence shell, represented by a double bond between the two oxygen atoms.

Figure 4.1.2 illustrates the formation of a double bond between two oxygen atoms. O_2 has a double bond (O=O), where each line between the atoms represent one covalent bond for a total of two covalent bonds in O_2. A valence electron pair that is not involved in bonding is called a **lone pair** (nonbonding pair). Two nonbonding pair and two bonding pairs complete the octet of each oxygen atom in the O_2 molecule.

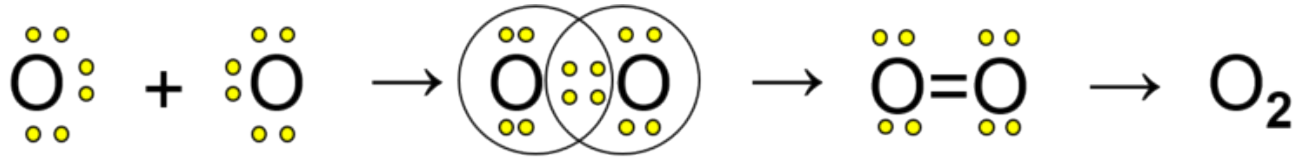


Figure 4.1.2: Illustration of two covalent bond formations between two oxygen atoms. (CC BY-SA 4.0; Drago Karlo via Wikimedia). Detailed Description of Figure 4.1.2.

Formation of a Triple Bond

A **triple bond** is formed when two atoms share three pairs of electrons (a total of 6 electrons). Triple bonds are the strongest and shortest type of covalent bond, and they are typically formed between atoms that need to share a large number of electrons to complete their outer electron shells and achieve stability.

In the formation of an N_2 molecule, two nitrogen atoms, each with 5 valence electrons, come together to share electrons and achieve a stable electron configuration. Each nitrogen atom needs 3 more electrons to complete its octet, so they share three pairs of electrons, forming a triple bond. This allows both nitrogen atoms to have 8 electrons in their outer shell—5 from their own electrons and 3 from the shared bonding electrons. The resulting structure is a stable N_2 molecule, with each nitrogen atom having a complete octet of electrons in its valence shell, represented by a triple bond between the two nitrogen atoms.

Figure 4.1.3 illustrates the formation of a triple bond between two nitrogen atoms. N_2 has a triple bond (N≡N), where each line between the atoms represent one covalent bond for a total of three covalent bonds in N_2. One nonbonding pair and three bonding pairs complete the octet of each nitrogen atom in the N_2 molecule, as shown below.

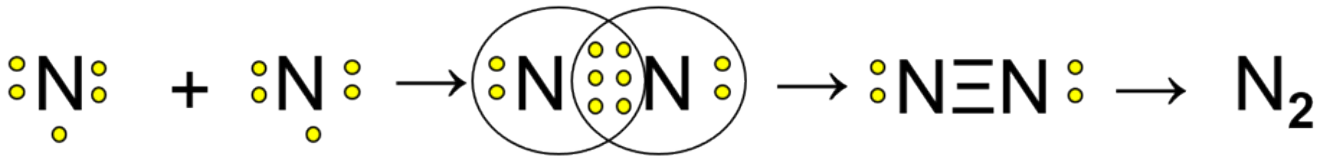


Figure 4.1.3: Illustration of three covalent bond formations between two nitrogen atoms. Details in Text (CC BY-SA 4.0; Drago Karlo via Wikimedia). Detailed Description of Figure 4.1.3.

Note

The bonding pair of electrons counts towards the total valence electrons of each bonded atom., i.e., in O_2 each oxygen atom has eight valence electrons; four in the nonbonding pair and four in two bonding pairs. In N_2 each nitrogen has eight valence electrons; two in the nonbonding pair and six in three bonding pairs.

Example 4.1.1

Determine how many bonding and nonbonding electron pairs are present on each atom in the HCl.

Solution

Hydrogen has 1 valence electron, while chlorine has 7 valence electrons**.** In forming the HCl molecule, hydrogen shares its one electron with chlorine to form a single covalent bond. This bond consists of 1 bonding pair of electrons (2 electrons total) shared between the two atoms.

Now, looking at each atom individually: Hydrogen participates in one bonding pair and has no nonbonding pairs. This is because hydrogen only needs two electrons to complete its outer shell (following the duet rule), and it gets these two electrons from sharing with chlorine. On the other hand, chlorine shares one of its seven valence electrons with hydrogen to form the bond. The remaining six valence electrons on chlorine are not involved in bonding and exist as three nonbonding pairs (also known as lone pairs).

Therefore, in the HCl molecule, hydrogen has one bonding pair and no nonbonding pairs, while chlorine has one bonding pair and three nonbonding pairs. This bonding arrangement satisfies the stability requirements for both atoms — hydrogen achieves a full outer shell with two electrons, and chlorine reaches an octet with eight electrons total (two in the bond and six in lone pairs).

Example 4.1.2

Determine how many bonding and nonbonding electron pairs are present on each atom in the Cl_2.

Solution

The Cl_2 molecule consists of two chlorine atoms bonded together. Each chlorine atom has 7 valence electrons. To achieve a stable octet, each chlorine shares one electron with the other, forming a single covalent bond. This shared pair of electrons constitutes one bonding pair, which holds the two chlorine atoms together.

After forming the bond, each chlorine atom is left with 6 nonbonding electrons, or 3 lone pairs. These electrons are not involved in bonding and remain localized around each chlorine atom. So, for each chlorine atom in Cl_2, there is 1 bonding pairs (shared between the two atoms) and 3 nonbonding pairs (localized on that atom).

Therefore, in the Cl_2 molecule, each chlorine atom has one bonding pair and three nonbonding pairs. The molecule as a whole has 1 bonding pair (shared between the atoms) and 6 nonbonding pairs in total (3 on each atom), satisfying the octet rule for both atoms.

Key Takeaways

- In each example, the bonding pairs of electrons count toward the total valence electrons of each atom.
- For H_2, each hydrogen atom ends up with 2 valence electrons (a duet).
- For O_2, each oxygen atom ends up with 8 valence electrons (4 nonbonding electrons + 4 bonding electrons from the double bond).
- For N_2, each nitrogen atom ends up with 8 valence electrons (2 nonbonding electrons + 6 bonding electrons from the triple bond).

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4.2: Covalent versus Ionic Compounds

Covalent Compounds

Covalent compounds are formed when two or more non-metal atoms share electrons to achieve a more stable electron configuration, typically resembling that of noble gases. This sharing of electrons creates a bond called a **covalent bond**, which is characterized by the attraction between the shared electrons and the positively charged nuclei of the atoms involved.

What elements make covalent bonds? Covalent bonds form when two or more nonmetals combine. For example, both hydrogen and oxygen are nonmetals, and when they combine to make water, they do so by forming covalent bonds. 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.

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). Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals), silicon dioxide (SiO_2) will be a covalently bound molecule (a semi-metal and a non-metal) and $MgCl_2$ will be ionic (a metal and a non-metal).

Characteristics of Covalent (Molecular) Compounds

Compounds that contain covalent bonds (also called molecular compounds) exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds. For example, water (molecular compound) boils at 100 °C while sodium chloride (ionic compound) boils at 1413 °C. In fact, many covalent compounds are liquids or gases at room temperature, and, in their solid states, they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good conductors of electricity when dissolved in water, most covalent compounds, being electrically neutral, are poor conductors of electricity in any state.

Example 4.2.1

Is each compound formed from ionic bonds, covalent bonds, or both?

a. Na_2O
b. $LiBr$
c. N_2O_4

a. Solution

- Na_2O is made up of Na^+ (sodium ion) and O^{2-} (oxide ion).
- Sodium (Na) is a metal, and oxygen (O) is a non-metal.
- Metals like sodium typically lose electrons to form cations (Na^+), while non-metals like oxygen gain electrons to form anions (O^{2-}).
- The opposite charges between Na^+ and O^{2-} attract each other, forming an ionic bond.
- Conclusion: Ionic bond.

b. Solution

- LiBr consists of Li^+ (sodium ions) and Br^- (bromine ions).
- Lithium (Li) is a metal and forms a cation (Li^+), while bromine is a nonmetal and forms (Br^-).
- The attraction between the sodium cations and bromine anions is an ionic bond.
- Conclusion: Ionic bond.

c. Solution

- N_2O_4 consists of two nitrogen atoms (N) and four oxygen atoms (O).
- Both nitrogen and oxygen are non-metals.
- In non-metallic compounds, atoms share electrons to form covalent bonds.
- The electrons are shared between nitrogen and oxygen atoms to create the molecule N_2O_4.
- Conclusion: Covalent bond.

Example 4.2.2

Is each compound are formed from ionic bonds, covalent bonds, or both?

a. F_2
b. PCl_3

a. Solution

- F_2 consists of two fluorine atoms (F).
- Fluorine atoms are non-metals, and since they have similar electronegativities, they share a pair of electrons to form a covalent bond.
- Conclusion: Covalent bond.

b. Solution

- PCl_3 consists of phosphorus (P) and chlorine (Cl) atoms, both of which are non-metals.
- In this case, phosphorus and chlorine share electrons to form covalent bonds.
- Conclusion: Covalent bond.

Molecular Formulas

The chemical formulas for covalent compounds are referred to as **molecular formula** because these compounds exist as separate, discrete molecules. Typically, a molecular formula begins with the nonmetal that is closest to the lower left corner of the periodic table, except that hydrogen is almost never written first (H_2O is the prominent exception). Then the other nonmetal symbols are listed. Numerical subscripts are used if there is more than one of a particular atom. In covalent molecules, subscripts represent the number of atoms of each element involved in the molecule. These subscripts are crucial for understanding the molecule's structure and chemical properties. The subscripts reflect the stoichiometry of the atoms, ensuring the correct balance of elements to form a stable compound.

Figure 4.2.1 illustrates the molecular formula of ammonia (NH_3) and emphasizes the importance of subscripts in a covalent molecule. In the molecular formula for ammonia, the "3" next to hydrogen (H) indicates that three hydrogen atoms are bonded to one nitrogen atom. The formula does not need to show a subscript for nitrogen (N) because, by convention, if an element is present only once in the molecule, the subscript "1" is not written. Therefore, in ammonia, "NH_3" implies one nitrogen atom and three hydrogen atoms, but the "1" next to nitrogen is not explicitly shown.

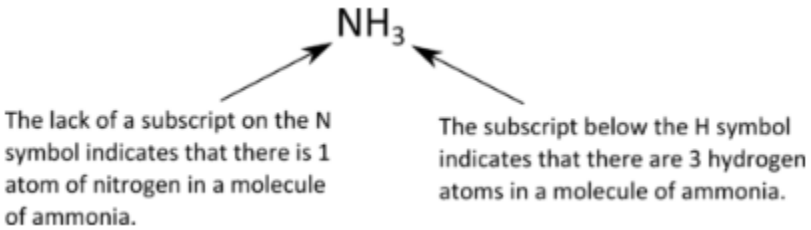


Figure 4.2.1: An arrow points to the N and says, "The lack of a subscript on the N symbol indicates that there is 1 atom of nitrogen in a molecule of ammonia". An arrow points to H and says, "The subscript below the H symbol indicates that there are 3 hydrogen atoms in a molecule of ammonia". (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.2.1.

Naming Covalent Compounds

Naming *binary* (two-element) covalent 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 4.2.1 lists these numerical prefixes. Normally, no prefix is added to the first element's name if there is only one atom of the first element in a molecule. If the second element is oxygen, the trailing vowel is usually omitted from the end of a polysyllabic prefix but not a monosyllabic one (that is, we would say "monoxide" rather than "monooxide" and "trioxide" rather than "troxide").

Table 4.2.1: 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.**	

Let us practice by naming the compound whose molecular formula is CCl_4. The name begins with the name of the first element—carbon. The second element, chlor*ine*, becomes chlor*ide*, and we attach the correct numerical prefix ("tetra-") to indicate that the molecule contains four chlorine atoms. Putting these pieces together gives the name *carbon tetrachloride* for this compound.

Example 4.2.3

Write the molecular formula for each compound.

a. chlorine trifluoride
b. phosphorus pentachloride
c. sulfur dioxide
d. dinitrogen pentoxide

Solution

If there is no numerical prefix on the first element's name, we can assume that there is only one atom of that element in a molecule.

a. ClF_3
b. PCl_5
c. SO_2
d. N_2O_5 (The *di-* prefix on nitrogen indicates that two nitrogen atoms are present.)

Example 4.2.4

Write the name for each compound.

a. BrF_5
b. S_2F_2
c. CO

Solution

a. bromine pentafluoride
b. disulfur difluoride
c. carbon monoxide

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

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system that we will introduce in in a separate section.

Key Takeaways

- The chemical formula of a simple covalent compound can be determined from its name.
- The name of a simple covalent compound can be determined from its chemical formula.

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4.3: Drawing Lewis Structures

Introduction

A **Lewis structure**, also known as a Lewis dot structure, is a way to represent the bonding between atoms in a molecule and the lone pairs of electrons that may exist around atoms. The main purpose of Lewis structures is to visually depict the arrangement of atoms, bonding, and lone pairs in a molecule. This helps us understand how molecules interact, their shapes, and their chemical behavior.

An important concept in constructing Lewis structures is the octet rule, which states that atoms tend to form bonds in a way that allows them to achieve a stable electron configuration, typically by having 8 electrons in their valence shell (for most atoms). This is similar to the electron configuration of noble gases, which are stable. However, there are exceptions to this rule, such as hydrogen, which follows the duet rule (achieving 2 electrons in its valence shell).

By following the octet rule (or duet rule in the case of hydrogen), Lewis structures help predict how atoms will bond, the number of bonds they will form, and the lone pairs of electrons that will remain on the atoms. This structure provides a valuable framework for understanding the molecular geometry and reactivity of compounds.

Drawing Lewis Structures

For very simple molecules and molecular ions, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms.

Figure $4.3.1$ illustrates the Lewis structure of three simple molecules like HBr (hydrogen bromide) H_2S (hydrogen sulfide) and N_2 (nitrogen). The Lewis structures can be drawn by pairing up the unpaired electrons from the constituent atoms to form bonds and satisfy their valence electron requirements. In HBr, hydrogen has one valence electron, and bromine has seven. To form a stable molecule, the hydrogen atom shares its single electron with bromine, creating a single covalent bond between them. Bromine, which has seven electrons, uses one to bond with hydrogen, leaving it with three lone pairs of electrons. The Lewis structure of HBr can be represented as H - Br, with bromine having three lone pairs. In H_2S, sulfur has six valence electrons, while each hydrogen atom has one. To form a stable structure, sulfur pairs one electron from each hydrogen, creating two single covalent bonds. This bonding satisfies sulfur's need for two electrons from hydrogen atoms, and the remaining four electrons on sulfur are placed as two lone pairs. The Lewis structure of H_2S is H - S - H, with sulfur having two lone pairs. Lastly, in the case of N_2 (nitrogen molecule), each nitrogen atom has five valence electrons. Since nitrogen needs three more electrons to complete its octet, two nitrogen atoms will share three pairs of electrons, forming a triple bond. The remaining two electrons on each nitrogen atom are placed as lone pairs. The Lewis structure of N_2 is represented as :N ≡ N:, where the triple bond between the two nitrogen atoms is shown, and each nitrogen atom has one lone pair of electrons. These structures are drawn by pairing unpaired electrons to form bonds, ensuring each atom achieves a stable electron configuration, following the principles of valency and the octet rule.

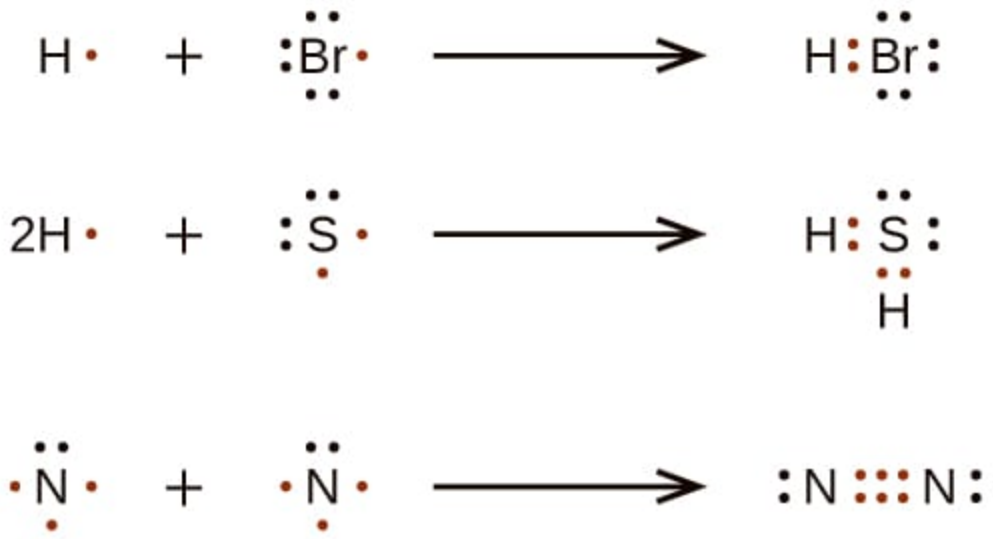


Figure 4.3.1: The Lewis structure of three simple molecules like HBr (hydrogen bromide) H_2S (hydrogen sulfide) and N_2 (nitrogen). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.1.

For more complicated molecules and molecular ions, it is helpful to follow the step-by-step procedure outlined here:

1. Determine the total number of valence (outer shell) electrons among all the atoms. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.
2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom. (Generally, the least electronegative element should be placed in the center.) Connect each atom to the central atom with a single bond (one electron pair).
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen), completing an octet around each atom.
4. Place all remaining electrons on the central atom.
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

> In a Lewis structure, the total number of electrons must equal the sum of the valence electrons of all atoms, adjusted for any ion charge.

Let us determine the Lewis structures of OF_2 and HCN as examples in the following procedure:

Determine the total number of valence (outer shell) electrons in the molecule or ion. For a molecule, we add the number of valence electrons (use the main group number) on each atom in the molecule. This is the total number of electrons that must be used in the Lewis structure.

$$O + 2\,(F) = OF_2$$

$$6e^- + (2 \times 7e^-) = 20e^-$$

$$H + C + N = HCN$$

$$1e^- + 4e^- + 5e^- = 10e^-$$

Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom and connecting each atom to the central atom with a single (one electron pair) bond. See Figure 4.3.2. Note that H and F can only form one bond, and are always on the periphery rather than the central atom.



Figure 4.3.2: The molecule OF_2 (oxygen difluoride) is shown with oxygen as the central atom bonded to two fluorine atoms, each forming a single bond and positioned on the periphery. In HCN (hydrogen cyanide), carbon is the central atom, bonded to hydrogen on one side and nitrogen on the other. Hydrogen and fluorine, which form only one bond, are shown correctly placed at the outer edges of the molecules. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.2.

Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons, as seen in Figure 4.3.3.

- In OF_2, six electrons are placed on each F.
- In HCN, six electrons placed on N

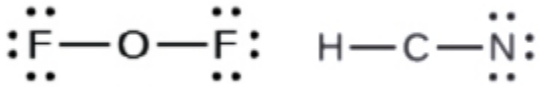


Figure 4.3.3: Distribution of Lone Pairs to Complete Octets on Terminal Atoms. In OF_2 (oxygen difluoride), each fluorine atom receives three lone pairs of electrons to complete its octet. In HCN (hydrogen cyanide), the terminal nitrogen atom is assigned three lone pairs to fulfill its octet. Hydrogen, which only needs two electrons, does not receive any lone pairs. CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.3.

Place all remaining electrons on the central atom, as seen in Figure 4.3.4.

- In OF_2, 4 electrons are placed on O.
- In HCN: no electrons remain (the total valence of 10 e^- is reached) so nothing changes.



Figure 4.3.4: Placement of Remaining Electrons on Central Atoms. In OF_2, the central oxygen atom already has a complete octet after forming two bonds with fluorine and receiving two lone pairs. In HCN, the central carbon atom has not yet reached an octet; additional bonding is required to complete its valence shell. CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.4.

Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

- In OF_2, each atom has an octet as drawn, so nothing changes, as seen in Figure 4.3.5.

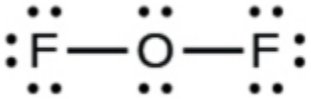


Figure 4.3.5: Electron Rearrangement in OF_2 to Form Multiple Bonds with the Central Oxygen Atom.
In OF_2, the outer fluorine atoms already have complete octets with three lone pairs each, so there is no need to rearrange electrons to form multiple bonds with the central oxygen atom. The oxygen atom maintains two single bonds and two lone pairs to satisfy its octet. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.5.

In HCN, form two more C–N bonds, as seen in Figure 4.3.6.

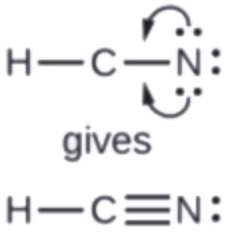


Figure 4.3.6: Electron Rearrangement in HCN to Form Multiple Bonds with the Central Carbon Atom. In HCN, the outer nitrogen atom has a complete octet with three lone pairs. To satisfy the octet rule for the central carbon atom, electrons are rearranged to form a triple bond between carbon and nitrogen, resulting in a stable structure where carbon, nitrogen, and hydrogen all have complete valence shells. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.6.

Finally, check to see if the total number of valence electrons are present in the Lewis structure. And then, inspect if the H atom has 2 electrons surrounding it and if each of the main group atoms is surrounded by 8 electrons.

Multiple Bonds

In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares only two electrons. Review HCN in Step 5 above. Another example is carbon dioxide (CO_2). CO_2 has a total valence of $4e^- + (2 \times 6e^-) = 16e^-$. Following steps 1 to 4, we draw as illustrated in Figure 4.3.7.

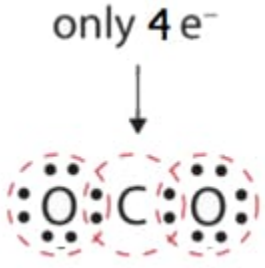


Figure 4.3.7: Initial Distribution of All Valence Electrons In Carbon Dioxide. Carbon dioxide (CO_2) has a total of 16 valence electrons: 4 from the central carbon atom and 6 from each oxygen atom. In the initial structure, carbon is placed at the center with one oxygen on each side. Single bonds are shown as pairs of dots between atoms to represent shared electron pairs. Each oxygen atom receives three lone pairs to complete its octet. All 16 electrons are used in this arrangement, but the central carbon is left with only four electrons, demonstrating the need to form multiple bonds to. satisfy the octet rule for all atoms. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.7.

This does not give the carbon atom a complete octet; only four electrons are in its valence shell. This arrangement of shared electrons is far from satisfactory. In this case, more than one pair of electrons must be shared between two atoms for both atoms to have an octet, as seen in Figure 4.3.8.



Figure 4.3.8: Electron Rearrangement in CO_2 to Satisfy the Octet Rule. To complete the octet of the central carbon atom in carbon dioxide (CO_2), lone pair electrons from each oxygen atom are shifted toward the carbon. This is illustrated by two arrows, each pointing from a lone pair on an oxygen atom to the bonding region between that oxygen and the carbon atom. CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.8.

A second electron pair from each oxygen atom must be shared with the central carbon atom shown by the arrows above. A lone

pair from each O must be converted into a bonding pair of electrons, as seen in Figure 4.3.9.

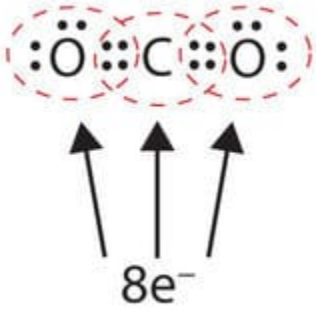


Figure 4.3.9: Final Lewis Structure of CO_2 with Bonds Shown as Dots.
In the final structure of carbon dioxide (CO_2), lone pairs from each oxygen atom are shared with the central carbon to form two double bonds. These shared electron pairs are shown as dots for clarity. Each atom—carbon and both oxygens—has eight electrons around it, satisfying the octet rule and resulting in a stable, linear molecule. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.9.

In this arrangement, the carbon atom shares four electrons (two pairs) with the oxygen atom on the left and four electrons with the oxygen atom on the right. There are now eight electrons around each atom. Two pairs of electrons shared between two atoms make a double bond between the atoms, which is represented by a double dash, as seen in Figure 4.3.10.

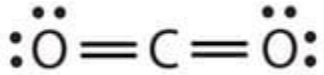


Figure 4.3.10: Final Lewis Structure of CO_2 with Bonds Shown as Lines. This representation of carbon dioxide (CO_2) shows the final Lewis structure using lines to indicate bonding pairs of electrons. Each double bond between carbon and oxygen is represented by two lines. All atoms—carbon and both oxygens—have complete octets, with eight electrons surrounding each, demonstrating a stable molecular structure. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.10.

Some molecules contain triple bonds (like HCN, shown above). Triple bonds are covalent bonds in which three pairs of electrons are shared by two atoms. Another compound that has a triple bond is acetylene (C_2H_2), whose Lewis diagram is as follows, see Figure 4.3.11.

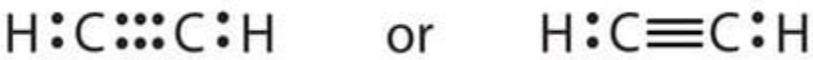


Figure 4.3.11: Lewis structure of acetylene (C_2H_2) in two linear formats. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.11.

Example 4.3.1

Draw the Lewis diagram for each molecule.

1. N_2
2. CH_2O (The carbon atom is the central atom.) One application of CH_2O, also called formaldehyde, is the preservation of biological specimens. Aqueous solutions of CH_2O are called formalin and have a sharp, characteristic (pungent) odor.

Solution

1. The total number of electrons is 2 x 5 = 10 electrons. The bond between the two nitrogen atoms is a triple bond. The Lewis diagram for N_2 is seen in Figure 4.3.12.



Figure 4.3.12: Lewis structure of nitrogen gas (N_2) in two linear formats. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.12.

2. The total number of electrons is 4 x 2(1) + 6 = 12 electrons. In CH_2O, the central atom is surrounded by two different types of atoms. The Lewis diagram that fills each atom's valence electron shell is as follows, see Figure 4.3.13.

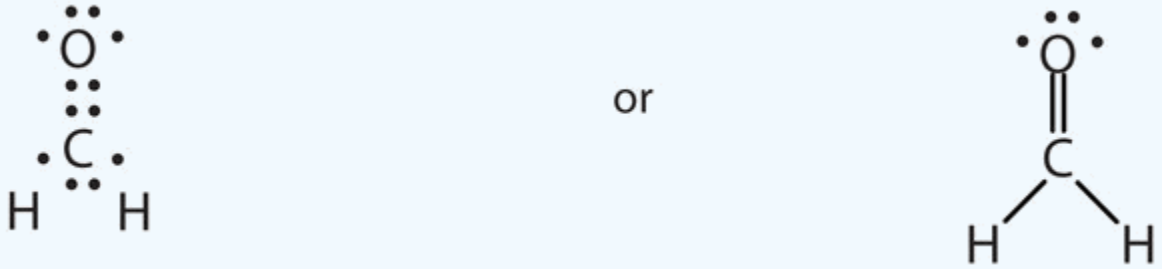


Figure 4.3.13: Lewis structure of formaldehyde (CH_2O) in two linear formats. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.13.

Example 4.3.2

Which is the correct Lewis structure for N_2H_2? See Figure 4.3.14.

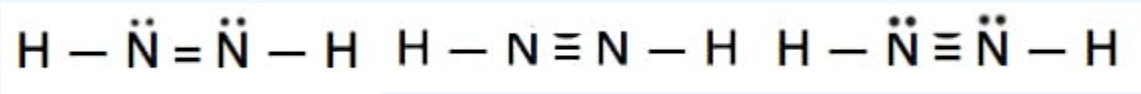


Figure 4.3.14: Which of the possible Lewis structures for diazene (N_2H_2) best satisfies the octet rule and represents the most stable form? (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.3.14.

Solution

Lewis structure on the left is the correct answer. It has a total of (2 x $5e^-$) + (2 x $1e^-$) = $12e^-$. Each of the N atoms satisfy the octet requirement and the H atoms follow the duet rule.

Lewis structure on the middle is electron deficient. It has only $10e^-$ instead of 12.

Lewis structure on the right has 14 (2 extra) electrons. The N atoms do not satisfy the octet.

Key Takeaways

- A Lewis structure shows the bonding and non-bonding electrons around individual atoms in a molecule.
- Some molecules must have multiple covalent bonds between atoms to satisfy the octet rule.
- A double bond contains four electrons and a triple bond contains six electrons.

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4.4: Characteristics of Covalent Bonds

Bond Length

Covalent bonds are one of the most common types of chemical bonds that hold atoms together to form molecules. These bonds are formed when two non-metal atoms share electrons in order to achieve a more stable electron configuration, often resembling that of the nearest noble gas. Understanding the characteristics of covalent bonds is crucial for explaining the behavior, properties, and reactivity of many compounds in chemistry.

Bond length is the average distance between the nuclei of two bonded atoms in a molecule and is a key factor in determining the physical and chemical properties of molecular compounds. It varies depending on the types of atoms involved and the number of shared electron pairs—shorter bonds are generally stronger and more stable. Bond length affects molecular geometry, which in turn influences how molecules interact, react, and fit together in biological and chemical systems. Accurate knowledge of bond lengths helps chemists predict reactivity, design new compounds, and understand molecular behavior.

For example, the covalent bond in the hydrogen molecule (H_2) has a certain length (about 7.4×10^{-11} m). Other covalent bonds also have known bond lengths, which are dependent on both the identities of the atoms in the bond and whether the bonds are single, double, or triple bonds. Table $4.4.1$ lists the approximate bond lengths for some single covalent bonds. The exact bond length may vary depending on the identity of the molecule but will be close to the value given in the table.

Table $4.4.1$: Approximate Bond Lengths of Some Single Bonds

Bond	Length ($\times 10^{-12}\ m$)
H–H	74
H–C	110
H–N	100
H–O	97
H–I	161
C–C	154
C–N	147
C–O	143
N–N	145
O–O	145

Table $4.4.2$ compares the lengths of single covalent bonds with those of double and triple bonds between the same atoms. Without exception, as the number of covalent bonds between two atoms increases, the bond length decreases. With more electrons between the two nuclei, the nuclei can get closer together before the internuclear repulsion is strong enough to balance the attraction.

Table $4.4.2$: Comparison of Bond Lengths for Single and Multiple Bonds

Bond	Length ($\times 10^{-12}\ m$)
C–C	154
C=C	134
C≡C	120
C–N	147
C=N	128

Bond	Length ($\times 10^{-12}\ m$)
C≡N	116
C–O	143
C=O	120
C≡O	113
N–N	145
N=N	123
N≡N	110
O–O	145
O=O	121

Electronegativity and Bond Type

The two idealized extremes of chemical bonding: (1) Ionic bonding, in which one or more electrons are transferred completely from one atom to another, and the resulting ions are held together by purely electrostatic forces and (2) Covalent bonding, in which electrons are shared equally between two atoms. Most compounds, however, have **polar covalent bonds**, which means that electrons are shared unequally between the bonded atoms. **Electronegativity** determines how the shared electrons are distributed between the two atoms in a polar covalent bond (Figure 4.4.1). The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a polar covalent bond are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarized the electron distribution and the larger the partial charges of the atoms. Recall that a lowercase Greek delta (δ) is used to indicate that a bonded atom possesses a partial positive charge, indicated by δ^+, or a partial negative charge, indicated by δ^-, and a bond between two atoms that possess partial charges is a polar bond.

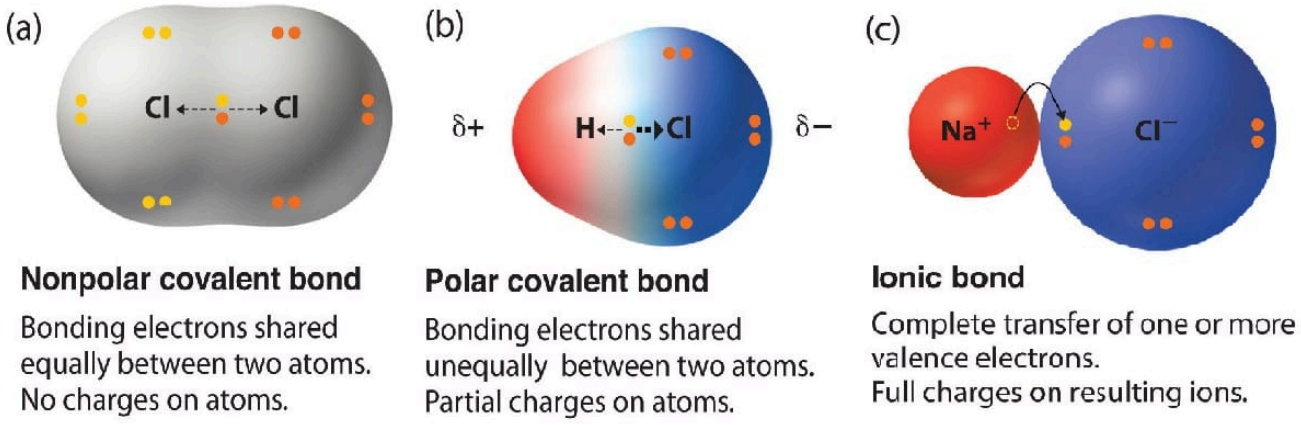


Figure 4.4.1: The Electron Distribution in a Nonpolar Covalent Bond, a Polar Covalent Bond, and an Ionic Electron-rich (negatively charged) regions are shown in blue; electron-poor (positively charged). (CC BY-SA-NC; Anonymous via LibreTexts). Detailed Description of Figure 4.4.1.

Any covalent bond between atoms of different elements is a polar bond, but the degree of polarity varies widely. Some bonds between different elements are only minimally polar, while others are strongly polar. Ionic bonds can be considered the ultimate in polarity, with electrons being transferred rather than shared. To judge the relative polarity of a covalent bond, chemists use electronegativity, which is a relative measure of how strongly an atom attracts electrons when it forms a covalent bond. There are various numerical scales for rating electronegativity. Figure 4.4.2 shows one of the most popular—the Pauling scale.

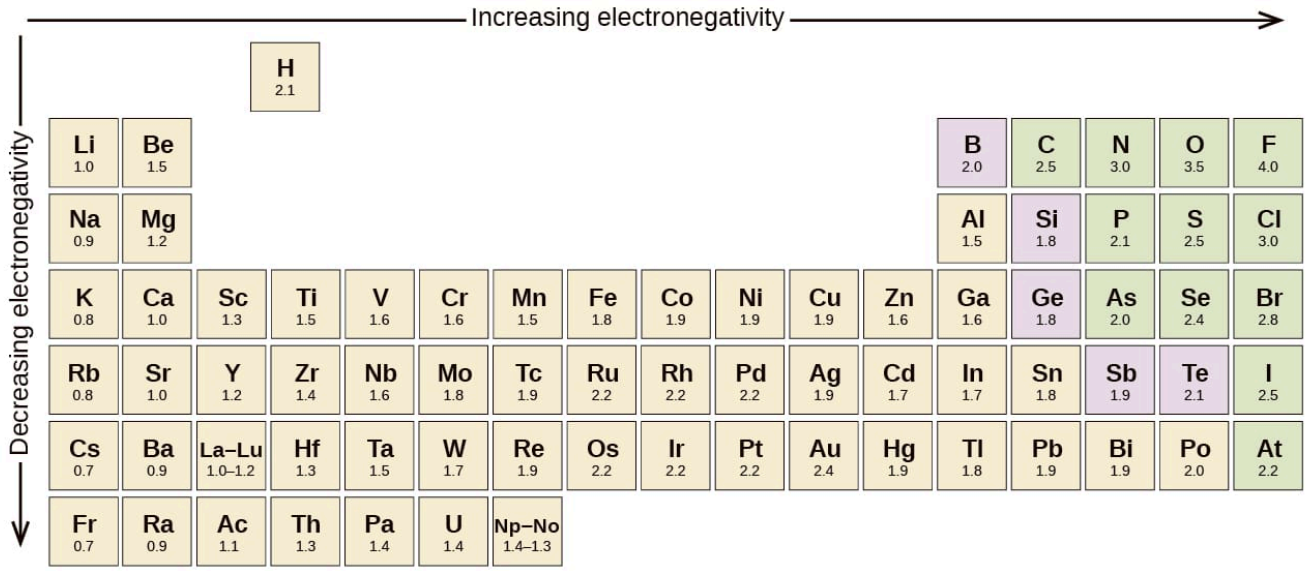


Figure 4.4.2: The electronegativity values derived by Pauling follow predictable periodic trends with the higher electronegativities toward the upper right of the periodic table. Fluorine has the highest value (4.0). (CC BY 4.0; via OpenStax). Detailed Description of Figure 4.4.2.

Looking Closer: Linus Pauling

Arguably the most influential chemist of the 20th century, Linus Pauling (1901–94) is the only person to have won two individual (that is, unshared) Nobel Prizes (see Figure 4.4.3). In the 1930s, Pauling used new mathematical theories to enunciate some fundamental principles of the chemical bond. His 1939 book The Nature of the Chemical Bond is one of the most significant books ever published in chemistry.

By 1935, Pauling's interest turned to biological molecules, and he was awarded the 1954 Nobel Prize in Chemistry for his work on protein structure. (He was very close to discovering the double helix structure of DNA when James Watson and James Crick announced their own discovery of its structure in 1953.) He was later awarded the 1962 Nobel Peace Prize for his efforts to ban the testing of nuclear weapons.

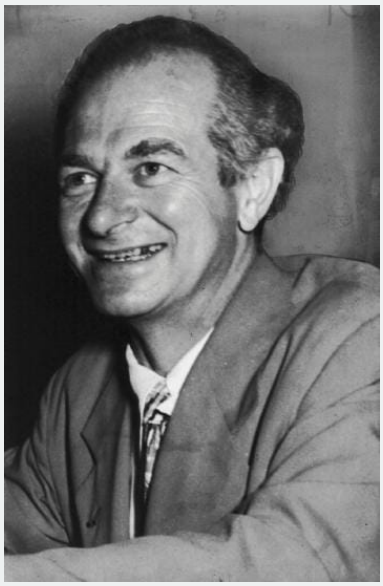
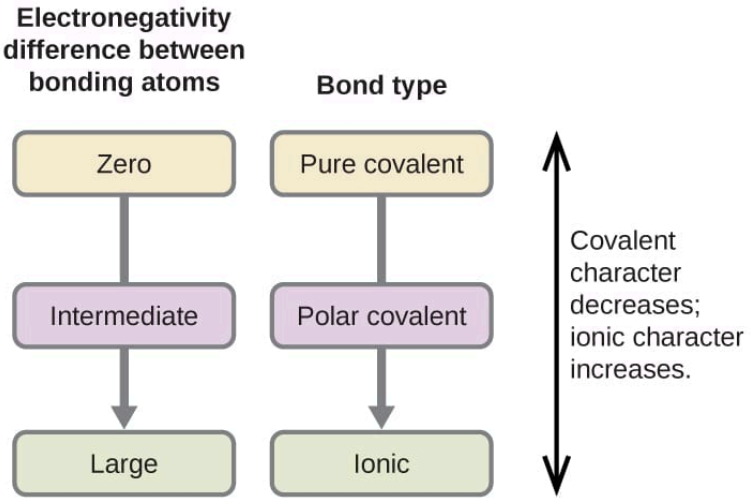


Figure 4.4.3: Linus Pauling was one of the most influential chemists of the 20th century. (CC BY-SA-NC; Anonymous via LibreTexts). Detailed Description of Figure 4.4.3.

In his later years, Pauling became convinced that large doses of vitamin C would prevent disease, including the common cold. Most clinical research failed to show a connection, but Pauling continued to take large doses daily. He died in 1994, having spent a lifetime establishing a scientific legacy that few will ever equal.

The polarity of a covalent bond can be judged by determining the difference in the electronegativities of the two atoms making the bond. The greater the difference in electronegativities, the greater the imbalance of electron sharing in the bond. Although there are no hard and fast rules, the general rule is if the difference in electronegativities is less than about 0.4, the bond is considered nonpolar; if the difference is greater than 0.4, the bond is considered polar. If the difference in electronegativities is large enough (generally greater than about 1.8), the resulting compound is considered ionic rather than covalent. An electronegativity difference of zero, of course, indicates a **nonpolar covalent bond**.



Bond Type	Electronegativity Difference
pure covalent	< 0.4
polar covalent	between 0.4 and 1.8
ionic	> 1.8

Figure 4.4.4: As the electronegativity difference increases between two atoms, the bond becomes more ionic. (CC BY 4.0; OpenStax). Detailed Description of Figure 4.4.4.

Example 4.4.1

Describe the electronegativity difference between each pair of atoms and the resulting polarity (or bond type).

a. C and H
b. H and H
c. Na and Cl
d. O and H

Solution

a. Carbon has an electronegativity of 2.5, while the value for hydrogen is 2.1. The difference is 0.4, which is rather small. The C–H bond is therefore considered nonpolar.
b. Both hydrogen atoms have the same electronegativity value—2.1. The difference is zero, so the bond is nonpolar.
c. Sodium's electronegativity is 0.9, while chlorine's is 3.0. The difference is 2.1, which is rather high, and so sodium and chlorine form an ionic compound.
d. With 2.1 for hydrogen and 3.5 for oxygen, the electronegativity difference is 1.4. We would expect a very polar bond. The sharing of electrons between O and H is unequal with the electrons more strongly drawn towards O.

Example 4.4.2

Describe the electronegativity (EN) difference between each pair of atoms and the resulting polarity (or bond type).

a. C and O
b. K and Br

c. N and N
d. Cs and F

Solution

a. The EN difference is 1.0 , hence polar. The sharing of electrons between C and O is unequal with the electrons more strongly drawn towards O.
b. The EN difference is greater than 1.8, hence ionic.
c. Identical atoms have zero EN difference, hence nonpolar.
d. The EN difference is greater than 1.8, hence ionic.

Key Takeaways

- Covalent bonds between different atoms have different bond lengths.
- Covalent bonds can be polar or nonpolar, depending on the electronegativity difference between the atoms involved.

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4.5: Understanding Molecular Shape and Bonding

Introduction

In this section, we will explore Lewis structures, which are a way to represent the bonding between atoms in a molecule, as well as the lone pairs of electrons that may exist. By using Lewis structures, you will learn how to visualize the arrangement of atoms and electrons, helping to predict the chemical behavior of molecules. These structures provide important insights into the bonding, molecular shape, and reactivity of compounds. Understanding how to draw and interpret Lewis structures is essential for advancing in chemistry, as it lays the foundation for more complex topics like **molecular geometry** and reaction mechanisms. Let's dive into this fundamental concept to see how the electron configuration shapes the world of molecules.

Molecular Shape: VSEPR Theory

Unlike ionic compounds, with their extended crystal lattices, covalent molecules are discrete units with specific three-dimensional shapes. The shape of a molecule is determined by the fact that covalent bonds, which are composed of negatively charged electrons, tend to repel one another. This concept is called the **VSEPR (Valence Shell Electron Pair Repulsion) theory.** According to VSEPR theory, the shape of a molecule is determined by the repulsion between electron pairs in the valence shell of atoms, which arrange themselves to minimize repulsion and maximize distance between them.

In $BeCl_2$ (beryllium chloride), the bond angle can be explained using VSEPR (Valence Shell Electron Pair Repulsion) theory (Figure 4.5.1). According to VSEPR theory, the electron pairs around a central atom will arrange themselves in a way that minimizes the repulsion between them, leading to a specific molecular geometry. In the case of $BeCl_2$, beryllium is the central atom, and it is bonded to two chlorine atoms.

Beryllium (Be) has two valence electrons, and each chlorine (Cl) has seven valence electrons. Since beryllium forms two single bonds with chlorine atoms, there are two bonding pairs of electrons around the central beryllium atom, and no lone pairs of electrons on the beryllium atom itself. With only two bonding pairs and no lone pairs on the central atom, the electron pairs in $BeCl_2$ will adopt a linear geometry to maximize the distance between them. According to VSEPR theory, the ideal bond angle between the two bonding pairs in a linear geometry is 180°.

Thus, the bond angle in$BeCl_2$ is 180°, and the molecule adopts a linear shape.

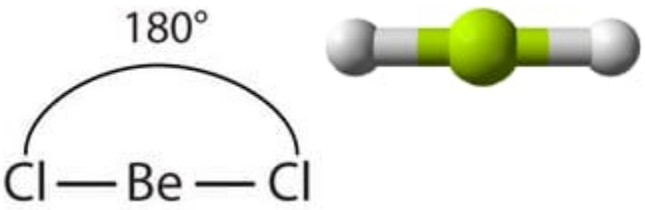


Figure 4.5.1: The bond angle in $BeCl_2$, determined by VSEPR theory, is 180° due to the linear arrangement of two bonding pairs of electrons around the central beryllium atom and no lone pairs. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.1.

Again, in BF_3 (boron trifluoride), the bond angle can be explained using VSEPR theory (see Figure 4.5.2). Boron (B) has three valence electrons, and each fluorine (F) has seven valence electrons. Since boron forms three single bonds with fluorine atoms, there are three bonding pairs of electrons around the central boron atom, and no lone pairs of electrons on the boron atom itself. With three bonding pairs and no lone pairs on the central atom, the electron pairs in BF_3 will adopt a trigonal planar geometry to maximize the distance between them. According to VSEPR theory, the ideal bond angle between the three bonding pairs in a trigonal planar geometry is 120°.

Thus, the bond angle in BF_3 is 120°, and the molecule adopts a trigonal planar shape.

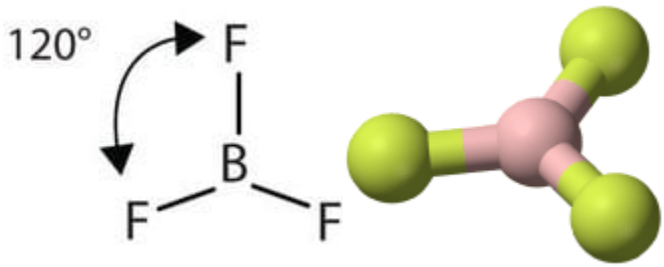


Figure 4.5.2: The bond angle in BF_3, determined by VSEPR theory, is 120° due to the trigonal planar arrangement of three bonding pairs of electrons around the central boron atom and no lone pairs. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.2.

The molecules $BeCl_2$ and BF_3 actually violates octet rule; however, such exceptions are rare and will not be discussed in this text.

> Try sticking three toothpicks into a marshmallow or a gumdrop and see if you can find different positions where your "bonds" are farther apart than the planar 120° orientation.

Determining Molecular Geometry by Atom Count

Molecular geometry refers to the three-dimensional arrangement of atoms within a molecule, and it can be predicted using VSEPR theory. According to VSEPR theory, the geometry of a molecule is determined by the number of bonding pairs of electrons around the central atom. The number of atoms directly bonded to the central atom (bonding pairs) plays a key role in determining the overall shape.

Basic Concept

- Central Atom: This is the atom that is surrounded by other atoms in the molecule.
- Bonding Pairs: These are pairs of electrons involved in covalent bonds between the central atom and the surrounding atoms.
- Electron Pairs: These include bonding pairs (shared between atoms) and lone pairs (non-bonding electrons on the central atom).

The number of atoms (or bonding pairs of electrons) around the central atom helps determine the electron geometry, which in turn dictates the molecular geometry. Lone pairs of electrons on the central atom also influence the shape by pushing bonding pairs further apart, though lone pairs themselves are not directly counted as atoms. See the Table 4.5.1 below.

Table 4.5.1: Describes how to use the number of atoms around the central atom to determine the geometry of a molecule, as well as examples.

Atoms Around Central Atom	Molecular Geometry	Example
2 AB_2	Linear	$BeCl_2$
3 AB_3	Trigonal Planar	BF_3
4 AB_4	Tetrahedral	CCl_4

Example 4.5.1

What is the molecular geometry and bond angle of the following compounds?

a. BCl_3 (boron trichloride)
b. CCl_4 (carbon tetrachloride)
c. HCl (hydrochloric acid)

Solution

a. BCl_3 (boron trichloride)

- Bonding electrons: 3 bonding pairs
- Geometry: Trigonal Planar
- Bond angle: 120°

b. CCl_4 (carbon tetrachloride)

- Bonding electrons: 4 bonding pairs
- Geometry: Tetrahedral
- Bond angle: 109.5°

c. HCl (hydrochloric acid)

- Bonding electrons: 1 bonding pair
- Geometry: Linear
- Bond angle: 180°

In determining the shapes of molecules, it is useful to first determine the Lewis diagram for a molecule. The shapes of molecules with multiple bonds are determined by treating the multiple bonds as a single bond. For example, Figure 4.5.3 shows the Lewis diagram of formaldehyde (CH_2O).

Since formaldehyde has three bonding pairs of electrons around the central carbon atom, it adopts a shape similar to BF_3 (boron trifluoride). Both molecules have a trigonal planar geometry, with bond angles of 120°.

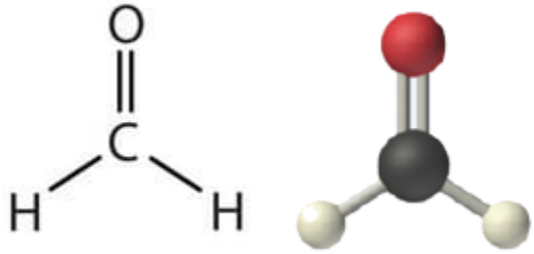


Figure 4.5.3: The Lewis diagram of formaldehyde (CH_2O) illustrates a trigonal planar geometry with three bonding pairs around the central carbon atom. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.3.

Molecules With Lone Pairs Around Central Atom

Molecules with lone electron pairs around the central atom have a shape based on the position of the atoms, not the electron pairs. For example, NH_3 has **one lone electron pair** and **three bonded electron pairs** as shown in Figure 4.5.4. These four electron pairs repel each other and adopt a tetrahedral arrangement. However, the shape of the molecule is described in terms of the positions of the atoms, not the lone electron pairs. Thus, NH_3 is said to have a **trigonal pyramidal** shape, not a tetrahedral one.

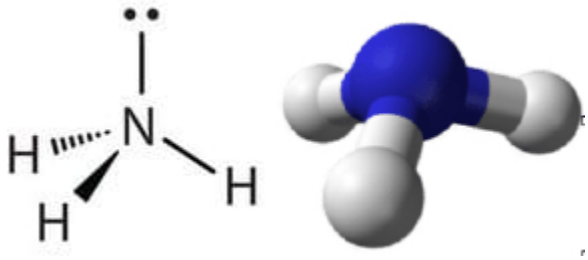


Figure 4.5.4: The Lewis diagram of formaldehyde (NH_3) illustrates a trigonal planar geometry with three bonding pairs around the central carbon atom. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.4.

Similarly, H_2O has two lone pairs of electrons around the central oxygen atom and two bonded electron pairs as shown in Figure 4.5.5. Although the four electron pairs adopt a tetrahedral arrangement, the shape of the molecule is described by the positions of the atoms only. The shape of H_2O is bent with an approximate 109.5° angle.

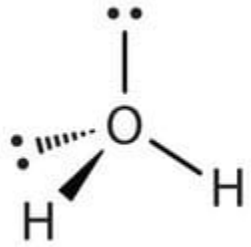


Figure 4.5.5: The Lewis diagram of formaldehyde (H_2O) illustrates a bent geometry with three bonding pairs around the central carbon atom. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.5.

To determine the electronic and molecular geometry, follow the steps below in relation to the table. This applies to both bonding only and with lone pairs:

1. Draw the Lewis structure: Determine the arrangement of atoms and the distribution of electrons.
2. Count the electron regions: Include both bonding pairs and lone pairs around the central atom.
3. Refer to the table: Identify the electron geometry based on the total number of electron regions (bonding + lone pairs).
4. Determine the molecular geometry: Based on the positions of the atoms (bonding pairs only) and ignore lone pairs.
5. Predict bond angles: Use the molecular geometry to estimate bond angles as shown in the table.

By following these steps, you can determine the molecular shape and bond angles for molecules, whether considering only bonding or including lone pairs.

> For example, If a molecule has 2 bonds and 0 lone pairs, it is linear. If a molecule has 3 bonds and 0 lone pairs, it is trigonal planar. If a molecule has 3 bonds and 1 lone pair, it is bent or angular. If a molecule has 4 bonds and 0 lone pairs, it is tetrahedral. If a molecule has 4 bonds and 1 lone pair, it is trigonal pyramid. If a molecule has 4 bonds and 2 lone pairs, it is bent or angular.

To determine the geometry of a molecule, you can use the number of atoms and lone pairs around the central atom, based on VSEPR theory. See the Figure 4.5.6 below.

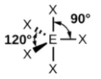
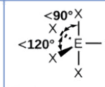
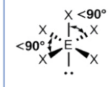
Number of electron regions	Electron region geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square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

Figure 4.5.6: Determining Molecular Geometry Based on Electron Regions and Lone Pairs Around the Central Atom. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.6.

Example 4.5.2

What is the geometry of the ammonium ion, NH_4^+? Its Lewis structure is shown below. How is this different from ammonia, NH_3? See Figure 4.5.7.

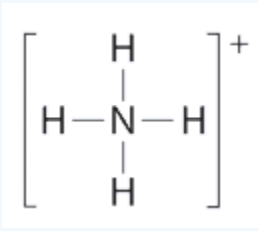


Figure 4.5.7: The Lewis Structure of NH_4^+. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.7.

Solution

In ammonium ion, the central atom N has 4 bonds and no lone pair. It is equivalent to the below in Figure 4.5.8. Hence, this is *tetrahedral*.

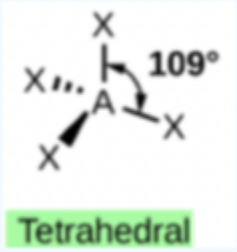


Figure 4.5.8:The ammonium ion (NH_4◈) has a central nitrogen atom labeled as A, bonded to four hydrogen atoms labeled as X, with no lone pairs on nitrogen. This arrangement corresponds to a tetrahedral molecular geometry. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.8.

In ammonia (NH_3), shown below, N has 3 bonds and one lone pair. See Figure 4.5.9.

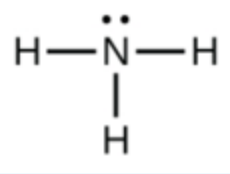


Figure 4.5.9: The Lewis Structure of NH_3. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.9.

It is equivalent to the structure in Figure 4.5.10. Hence, the shape of this molecule is trigonal pyramid.



Figure 4.5.10: The ammonia molecule (NH_3) has a central nitrogen atom labeled as A, bonded to three hydrogen atoms labeled as X, with one lone pair on nitrogen. This arrangement results in a trigonal pyramidal molecular geometry. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.10.

Example 4.5.3

What is the molecular shape of nitrosyl chloride (NOCl), a highly corrosive, reddish-orange gas? Its Lewis structure is shown below in Figure 4.5.11.

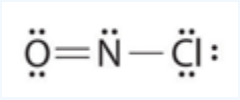


Figure 4.5.11: The Lewis Structure of nitrosyl chloride (NOCl). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.11.

Solution

Focus on the central atom, N. It has a double bond to O, count this as one bond. It also has a single bond to Cl. Thus, N has 2 bonds and one lone pair. These 3 electron pairs will spread out 120 degrees from each other. But, since the shape is defined by the arrangement of the atoms only, the shape is bent or angular. If you consult Figure 4.5.12, this molecule is equivalent to the structure below. Hence, two bonds and one lone pair has a bent or angular shape.

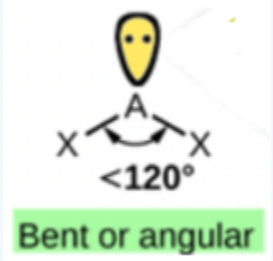


Figure 4.5.12: The nitrosyl chloride (NOCl) molecule has a central nitrogen atom labeled as A, forming two bonds—one double bond to an oxygen atom and one single bond to a chlorine atom—and has one lone pair of electrons. The bonded atoms are labeled as X. With two bonding regions and one lone pair around the central atom, the molecular geometry is bent (angular). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.12.

Molecular Polarity

In general, a molecule is nonpolar if all its bonds are nonpolar. Examples are I_2, O_2, H_2, CH_4, C_2H_6 and C_3H_8.

In general, a molecule is polar if it contains polar bonds EXCEPT when the bond polarities cancel each other. For example, in Figure 4.5.13, the shape of the CO_2 molecule (linear) orients the two C=O polar bonds directly opposite each other, thus cancelling each other's effect. Carbon dioxide (CO_2) is a nonpolar molecule.

On the other hand, water is a bent molecule because of the two lone pairs on the central oxygen atom. Because of the bent shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the individual H-O polar bonds represented by the two red arrows are not directly opposite each other. These two dipoles don't cancel each other out. In fact, the net dipole (blue arrow) points upward. There is a resultant partial positive charge at one end (between the two H atoms) and a partial negative charge on the other end (where O is located). The uneven distribution of charge or the overall dipole is shown by the blue arrow below. Hence, water is polar (has + and - poles) while carbon dioxide is nonpolar.

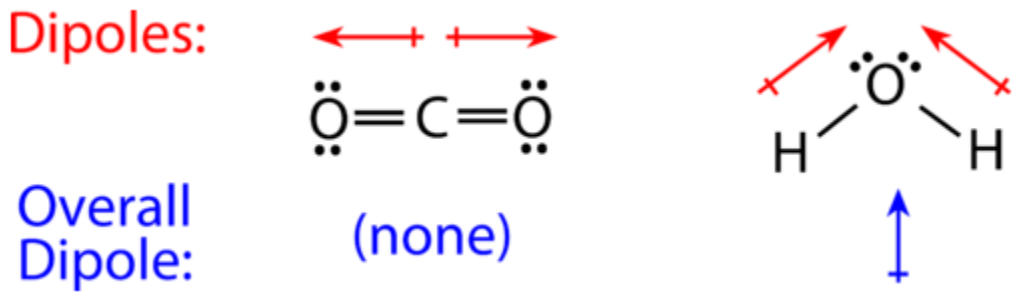


Figure 4.5.13: Bond dipoles can cancel or combine depending on molecular shape. In carbon dioxide, CO2, the bond dipoles cancel, so there is no overall dipole. In water, H2O, the bond dipoles do not cancel, so the molecule has an overall dipole. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.13.

Similarly, in BF_3 (trigonal planar), the effect of a B-F bond is cancelled by the sum of the other two B-F bonds (see video). Hence, a trigonal planar molecule (BF_3) is nonpolar because the bond polarities cancel each other, but a trigonal pyramidal molecule (NH_3) is polar. See Figure 4.5.14.

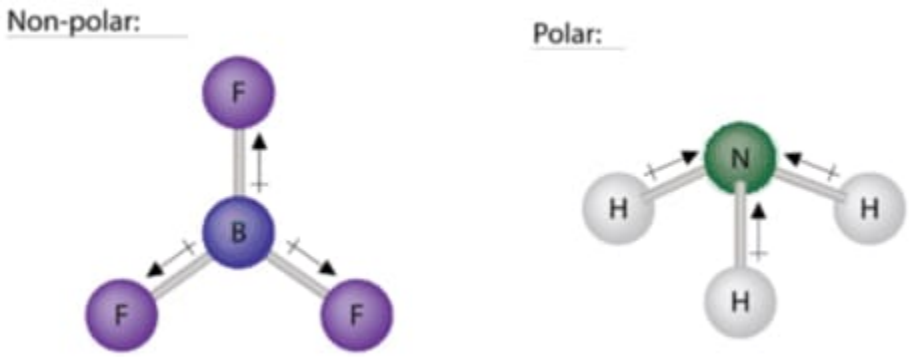


Figure 4.5.14: Molecular geometry affects overall polarity. Boron trifluoride, BF3, is nonpolar because its bond dipoles cancel, while ammonia, NH3, is polar because its bond dipoles do not cancel. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.14.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as CCl_4 is ***nonpolar***. However, if the peripheral atoms are not of the same electronegativity, the bond polarities don't cancel and the molecule becomes ***polar***, as in CH_3Cl. Figure 4.5.15.

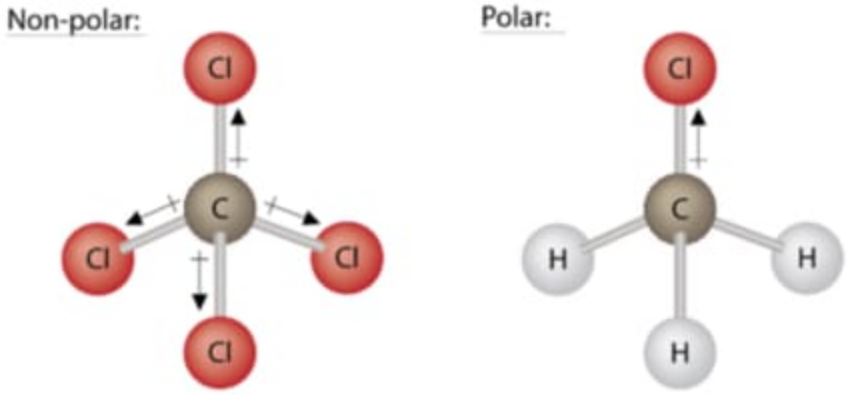


Figure 4.5.15: Molecular shape affects overall polarity. Carbon tetrachloride, CCl4, is nonpolar because its bond dipoles cancel, while chloroform, CHCl3, is polar because its bond dipoles do not cancel. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.15.

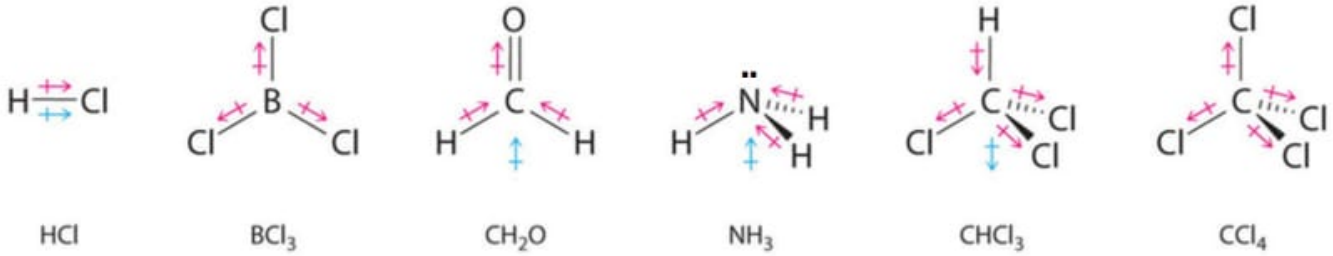


Figure 4.5.16: Examples of bond dipoles and overall molecular polarity for HCl, BCl3, CH2O, NH3, CHCl3, and CCl4. The diagram shows how molecular shape affects whether bond dipoles combine to produce a net dipole. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.16.

Physical Properties of Molecular Compounds

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. Molecular compounds sometimes called covalent compounds, display a wide range of physical properties due to the different types of intermolecular attractions such as different kinds of polar interactions. The melting and boiling points of molecular compounds are generally quite low compared to those of ionic compounds. This is because the energy required to disrupt the intermolecular between molecules is far less than the energy required to break the ionic bonds in a crystalline ionic compound. Since molecular compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. Ionic compounds do not conduct electricity in the solid state because of their rigid structure, but conduct well when either molten or dissolved into a solution. The water solubility of molecular compounds is variable and depends primarily on the type of intermolecular forces involved. Substances that exhibit hydrogen bonding or dipole-dipole forces are generally water soluble, whereas those that exhibit only London dispersion forces are generally insoluble. Most, but not all, ionic compounds are quite soluble in water. Table 4.5.2 summarizes some of the differences between ionic and molecular compounds.

Table 4.5.2: Comparison of Ionic and Molecular Compounds.

Property	Ionic Compounds	Molecular Compounds
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - transfer of electron(s) between atoms	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable

Property	Ionic Compounds	Molecular Compounds
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor

In summary, covalent compounds are softer, have lower boiling and melting points, are more flammable, are less soluble in water and do not conduct electricity compared to ionic compounds. The individual melting and boiling points, solubility and other physical properties of molecular compounds depend on molecular polarity.

Example 4.5.4

Describe the shape of each molecule. Is it polar or nonpolar?

a. PCl_3
b. CO_2

Solution

a. The Lewis diagram for PCl_3 is as follows:

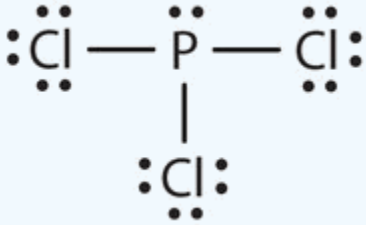


Figure 4.5.17: Lewis structure of phosphorus trichloride (PCl_3), showing three P–Cl single bonds and lone pairs on phosphorus and each chlorine. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.17.

Focus on the central atom, P that has 3 bonds and one lone pair in Figure 4.5.17. The four electron pairs arrange themselves tetrahedrally, but the lone electron pair is not considered in describing the molecular shape. Like NH_3, this molecule is pyramidal. The 3 P-Cl bonds don't cancel each other. This is polar.

b. The Lewis diagram for CO_2 is as follows:

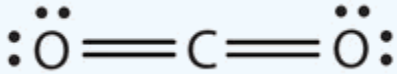


Figure 4.5.18: The Lewis Structure of CO_2. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.18.

Focus on the central atom, C in Figure 4.5.18. The multiple bonds are treated as one group, hence C has 2 bonds and zero lone pair. CO_2 has only two groups of electrons that repel each other. They will direct themselves 180° apart from each other, so CO_2 molecules are linear. This is highly symmetrical, with the two opposite dipoles cancelling each other. The CO_2 molecule is nonpolar.

Bond Polarity and Molecular Polarity

If there is only one bond in the molecule, the bond polarity determines the molecular polarity. Any diatomic molecule in which the two atoms are the same element must be a nonpolar molecule. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. A polar molecule is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. 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. Hence, a molecule with two poles is called a dipole. A simplified way to depict polar molecules like HF is pictured below (Figure 4.5.19.

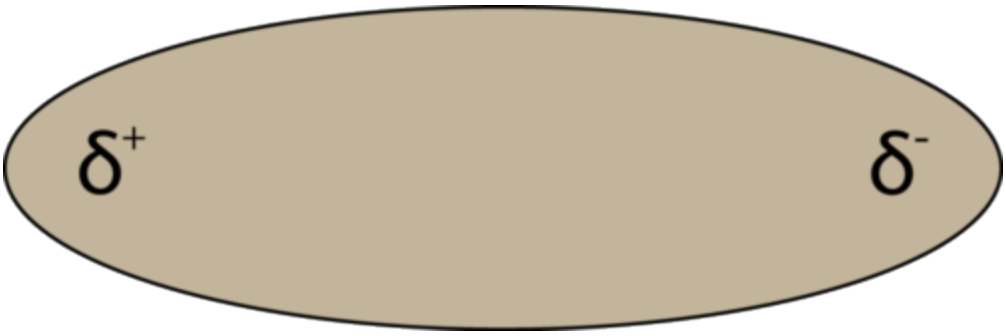


Figure 4.5.19: A molecular dipole results from the unequal distribution of electron density throughout a molecule. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.19.

When placed between oppositely charged plates, polar molecules orient themselves so that their positive ends are closer to the negative plate and their negative ends are closer to the positive plate (Figure 4.5.20).

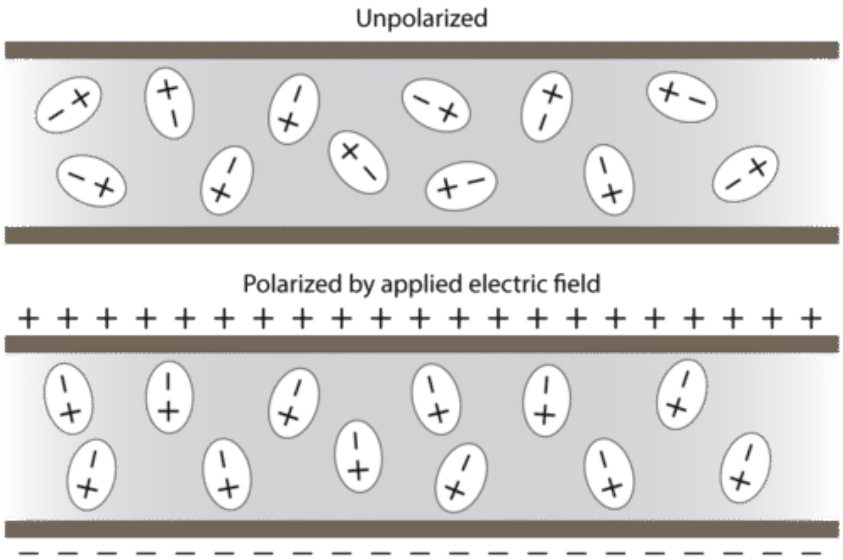


Figure 4.5.20: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves to maximize the attraction between opposite charges (bottom). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.20.

Experimental techniques involving electric fields can be used to determine if a certain substance is composed of polar molecules and to measure the degree of polarity.

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. (Figure 4.5.21) is a comparison between carbon dioxide and water. Carbon dioxide (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 directly opposite each other, they cancel each other out, and the overall molecular polarity of CO_2 is zero. CO_2 is a nonpolar molecule. In H_2O, the orientation of the two O–H bonds is bent: one end of the molecule has a partial positive charge, and the other end has a partial negative charge. In short, the molecule itself is polar. The polarity of water has an enormous impact on its physical and chemical properties. For example, the boiling point of water (100°C) is high for such a small molecule due to the fact that polar molecules attract each other strongly. On the other hand, the nonpolar carbon dioxide becomes a gas at –77°C, almost 200° lower than the temperature at which water boils. More details will be presented in the next section.



Figure 4.5.21: Physical Properties and Polarity. The physical properties of water and carbon dioxide are affected by their polarities. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.5.21.

Key Takeaways

- Molecule shape is determined by electron pair repulsion to minimize distance between them.
- Electronegativity differences create polar bonds with unequal electron sharing.
- A molecule is polar if it has an uneven electron distribution and a net dipole moment.
- Properties like boiling point and solubility depend on intermolecular forces.

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4.6: Introduction to Organic Chemistry

Introduction

Organic chemistry is the study of carbon-based compounds, and its significance stretches across many areas of science and industry. Carbon's ability to form a wide range of structures, from simple hydrocarbons to complex macromolecules, is central to its versatility. Functional groups modify the behavior and reactivity of organic compounds, making them essential for understanding processes from biological systems to chemical manufacturing. Organic chemistry shapes the world around us, from the creation of medicines and materials to understanding the complex chemistry of life itself.

Alkanes

The simplest organic compounds are the **hydrocarbons**, compounds composed of carbon and hydrogen atoms only. Some hydrocarbons have only single bonds and appear as a chain (which can be a straight chain or can have branches) of carbon atoms also bonded to hydrogen atoms. These hydrocarbons are called **alkanes** (saturated hydrocarbons). Each alkane has a characteristic, systematic name depending on the number of carbon atoms in the molecule. These names consist of a stem that indicates the number of carbon atoms in the chain plus the ending -ane. The stem meth- means one carbon atom, so methane is an alkane with one carbon atom. Similarly, the stem eth- means two carbon atoms; ethane is an alkane with two carbon atoms. Continuing, the stem prop- means three carbon atoms, so propane is an alkane with three carbon atoms. The stem but- means four carbon atoms; butane is an alkane with four carbon atoms. Figure 4.6.1 gives the Lewis structures, condensed structural formulas and molecular formulas of the four simplest alkanes. In the **condensed structural formula**, the covalent bonds are understood to exist between each carbon and the hydrogens associated with it, as well as between carbon atoms.

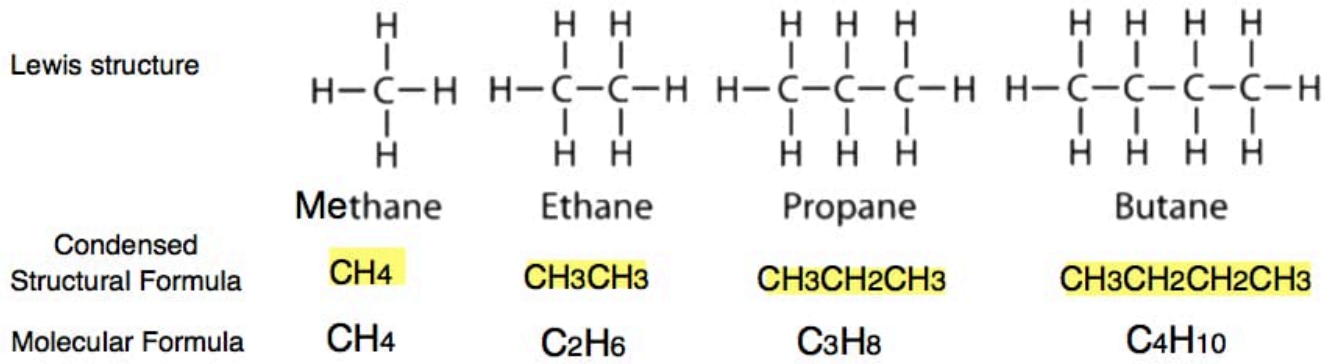


Figure 4.6.1: Comparison of methane, ethane, propane, and butane showing Lewis structures, condensed structural formulas, and molecular formulas. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.1.

Alkenes

Some hydrocarbons have one or more carbon–carbon double bonds (denoted C=C). These hydrocarbons are called **alkenes**. Figure 4.6.2 shows the formulas and the molecular models of the two simplest alkenes. Note that the names of alkenes have the same stem as the alkane with the same number of carbon atoms in its chain but have the ending *-ene*. Thus, ethene is an alkene with two carbon atoms per molecule, and propene is a compound with three carbon atoms and one double bond.

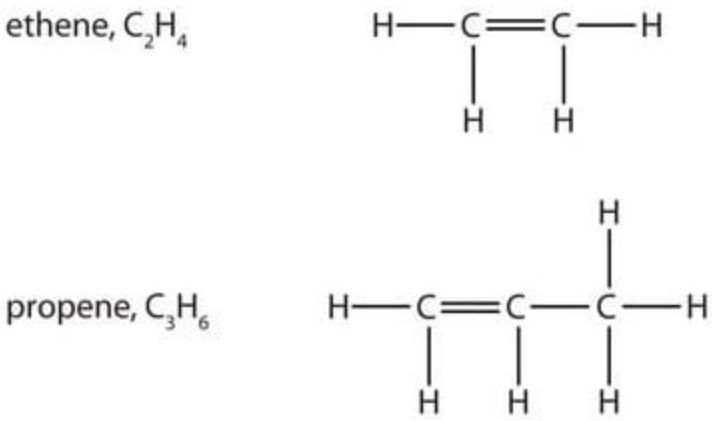


Figure 4.6.2: Comparison of the Lewis structures of ethene and propene. Both molecules contain a carbon-carbon double bond; ethene has two carbon atoms, and propene has three carbon atoms. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.2.

Alkynes

Alkynes are hydrocarbons with a carbon–carbon triple bond (denoted C≡C) as part of their carbon skeleton. Below is the formula and the molecular model of the simplest alkyne and its systematic name is ethyne. Its common name is acetylene (see Figure 4.6.3) below. Its chemical formula is C_2H_2.

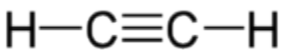


Figure 4.6.3: The Lewis Structure of acetylene (ethyne). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.3.

The names for **alkynes** have the same stems as for alkanes but with the ending *-yne*.

To Your Health: Saturated and Unsaturated Fats

Hydrocarbons are not the only compounds that can have carbon–carbon double bonds. A group of compounds called fats can have them as well, and their presence or absence in the human diet is becoming increasingly correlated with health issues.

Fats are combinations of long-chain organic compounds (fatty acids) and glycerol ($C_3H_8O_3$). (For more information on fats, see Chapter 17) The long carbon chains can have either all single bonds, in which case the fat is classified as *saturated*, or one or more double bonds, in which case it is a *monounsaturated* or a *polyunsaturated* fat, respectively. Saturated fats are typically solids at room temperature; beef fat (tallow) is one example. Mono- or polyunsaturated fats are likely to be liquids at room temperature and are often called oils. Olive oil, flaxseed oil, and many fish oils are mono- or polyunsaturated fats.

Studies have linked higher amounts of saturated fats in people's diets with a greater likelihood of developing heart disease, high cholesterol, and other diet-related diseases. In contrast, increases in unsaturated fats (either mono- or polyunsaturated) have been linked to a lower incidence of certain diseases. Thus, there have been an increasing number of recommendations by government bodies and health associations to decrease the proportion of saturated fat and increase the proportion of unsaturated fat in the diet. Most of these organizations also recommend decreasing the total amount of fat in the diet.

Recently, certain fats called trans fats have been implicated in the presence of heart disease. These are fats from animal sources and are also produced when liquid oils are exposed to partial hydrogenation, an industrial process that increases their saturation. Trans fats are used in many prepared and fried foods. Because they bring with them the health risks that naturally occurring saturated fats do, there has been some effort to better quantify the presence of trans fats in food products. US law now requires that food labels list the amount of trans fat in each serving.

Since the early 1900's, the US Department of Agriculture has been providing science-based dietary guidelines for the public. The most current version is the MyPlate illustration that gives a simple, visual picture of how much of what kind of foods make up a good, balanced diet (see Figure 4.6.4. It recommends minimizing daily intake of sugars, the "bad fats", trans and

saturated fat, and sodium. "Good fats", unsaturated fats or oils, are not considered a food group but do contain essential nutrients and therefore are included as part of a healthy eating pattern. The difference as simple as the difference between a single and double carbon–carbon bond, good and bad fats, can have a significant impact on health.

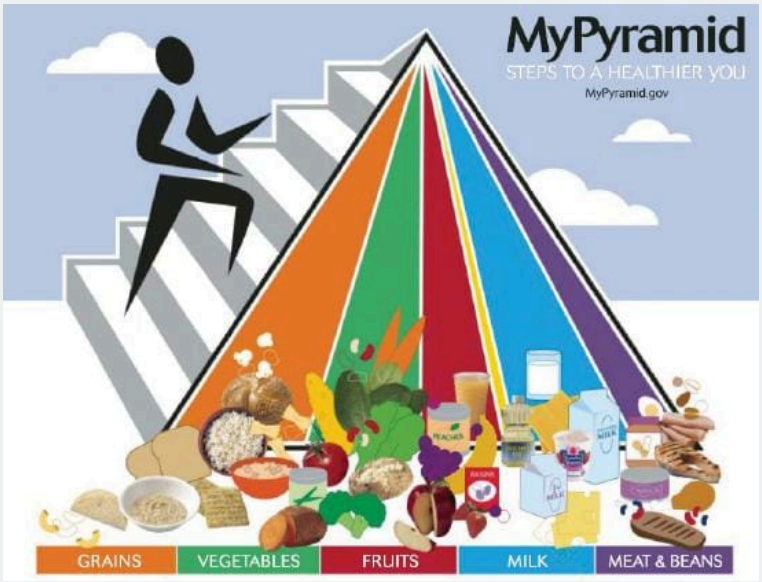


Figure 4.6.4: MyPlate is a reminder that everything you eat and drink matters. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.4.

Functional Groups

The carbon–carbon double and triple bonds are examples of **functional groups** in organic chemistry. A functional group is a specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to a molecule. Alkanes have no functional group. A carbon–carbon double bond is considered a functional group because carbon–carbon double bonds chemically react in specific ways that differ from reactions of alkanes (for example, under certain circumstances, alkenes react with water); a carbon–carbon triple bond also undergoes certain specific chemical reactions. In the remainder of this section, we introduce two other common functional groups.

If an OH group (also called a hydroxyl group) is substituted for a hydrogen atom in a hydrocarbon molecule, the compound is an alcohol. Alcohols are named using the parent hydrocarbon name but with the final *-e* dropped and the suffix *-ol* attached. The two simplest alcohols are methanol and ethanol. Figure 4.6.5 shows their formulas along with a molecular model of each.

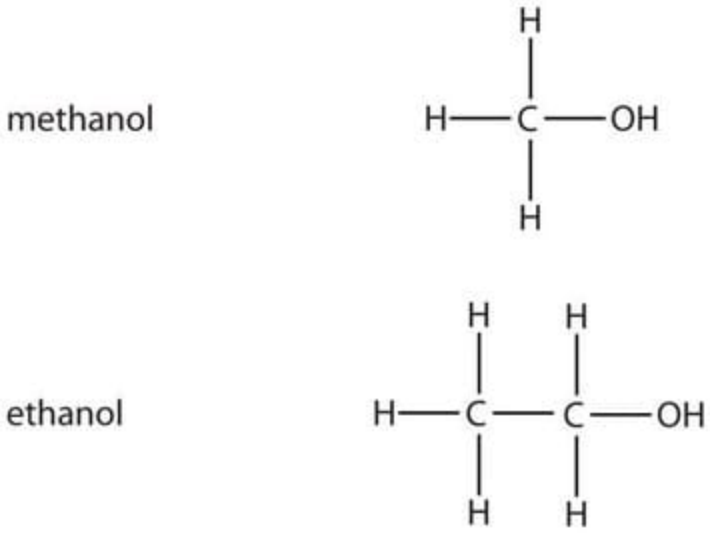


Figure 4.6.5: The Two Simplest Organic Alcohol Compounds. Alcohols have an OH functional group in the molecule. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.5.

Cholesterol, described in the chapter-opening essay, has an alcohol functional group, as its name implies.

Alcohol

Ethanol (also called ethyl alcohol) is the alcohol in alcoholic beverages. Other alcohols include methanol (or methyl alcohol), which is used as a solvent and a cleaner, and isopropyl alcohol (or rubbing alcohol), which is used as a medicinal disinfectant. Neither methanol nor isopropyl alcohol should be ingested, as they are toxic even in small quantities.

Another important family of organic compounds has a carboxyl group, in which a carbon atom is double-bonded to an oxygen atom and to an OH group. Compounds with a carboxyl functional group are called carboxylic acids, and their names end in *-oic acid*. Figure 4.6.6 shows the formulas and the molecular models of the two simplest carboxylic acids, perhaps best known by the common names formic acid and acetic acid. The carboxyl group is sometimes written in molecules as COOH.

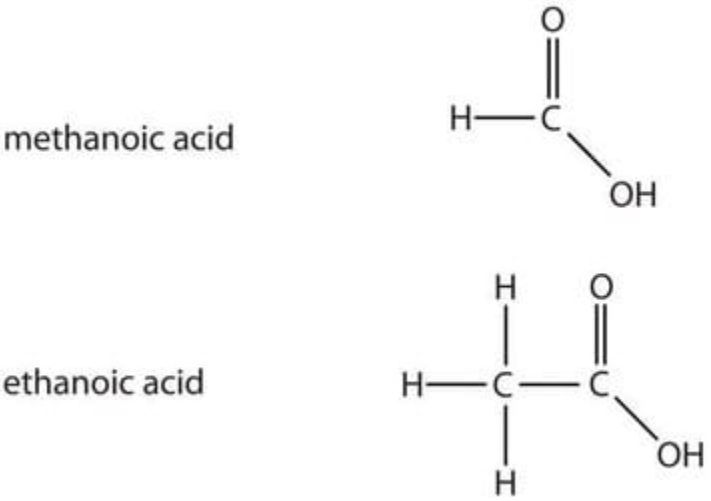


Figure 4.6.6: The two smallest carboxylic acids are formic acid (found in the stingers of ants) and acetic acid (found in vinegar). (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.6.

The condensed structures of methanoic acid and ethanoic acid are HCOOH and CH_3COOH, respectively.

Many organic compounds are considerably more complex than the examples described here. Many compounds, such as cholesterol discussed in the chapter-opening essay, contain more than one functional group. The formal names can also be quite complex.

Example 4.6.1

Identify the functional group(s) in each molecule as a double bond, a triple bond, an alcohol, or a carboxyl. Assuming the structures are arranged left to right as a to c.

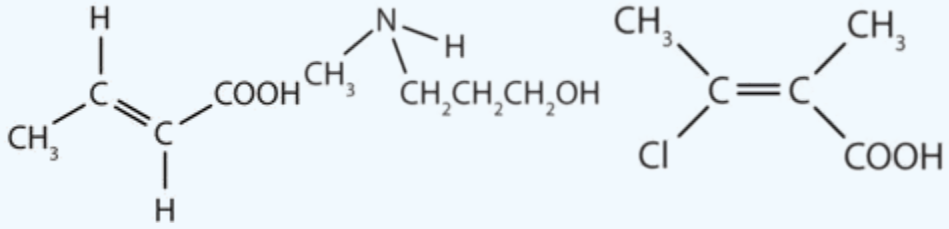


Figure 4.6.7: Identify the functional group(s) in each molecule as a double bond. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.7.

Solution

Answer available at the end of this section.

Example 4.6.2

Identify the functional group(s) in each molecule as a double bond, a triple bond, an alcohol, or a carboxyl. Assuming the structures are arranged left to right as a to d.

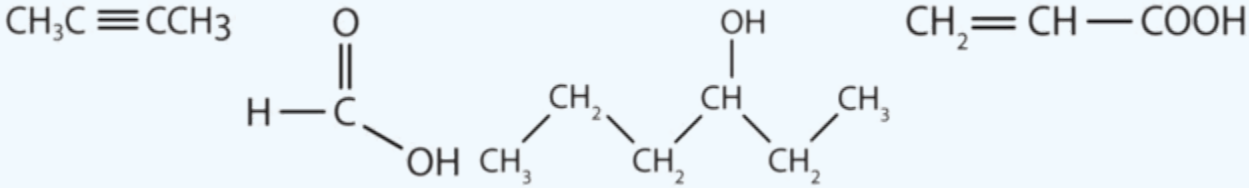


Figure 4.6.8: Identify the functional group(s) in each molecule as a double bond, a triple bond, an alcohol, or a carboxyl. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 4.6.8.

Solution

a. triple bond (alkyne)
b. carboxyl group
c. alcohol group
d. double bond (alkene) and carboxyl group

Career Focus: Forensic Chemist

The main job of a forensic chemist is to identify unknown materials and their origins. Although forensic chemists are most closely associated in the public mind with crime labs, they are employed in pursuits as diverse as tracing evolutionary patterns in living organisms, identifying environmental contaminants, and determining the origin of manufactured chemicals.

In a crime lab, the forensic chemist has the job of identifying the evidence so that a crime can be solved. The unknown samples may consist of almost anything—for example, paint chips, blood, glass, cloth fibers, drugs, or human remains. The forensic chemist subjects them to a variety of chemical and instrumental tests to discover what the samples are. Sometimes these samples are extremely small, but sophisticated forensic labs have state-of-the-art equipment capable of identifying the smallest amount of unknown sample.

Another aspect of a forensic chemist's job is testifying in court. Judges and juries need to be informed about the results of forensic analyses, and it is the forensic chemist's job to explain those results. Good public-speaking skills, along with a broad background in chemistry, are necessary to be a successful forensic chemist.

Key Takeaways

- Organic chemistry is the study of carbon-containing compounds and their chemical reactions.
- Organic molecules are classified based on their bonding patterns, the types of atoms they contain, and the presence of functional groups—distinct groups of atoms, such as alcohols, carboxylic acids, and alkenes, that determine a molecule's reactivity.
- Functional groups are the chemically active sites in molecules and play a key role in predicting and understanding chemical behavior.
- Lewis structures depict the bonding between carbon and other atoms in organic molecules, helping to visualize functional groups and predict molecular shape and reactivity.

Answers

Example 4.6.1

a. This molecule has a double bond and a carboxyl functional group.
b. This molecule has an alcohol functional group.
c. This molecule has a carbon-carbon double bond and a carboxyl functional group.

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4.7: Summary

Covalent compounds are formed when two or more nonmetal atoms share electrons to achieve a stable electron configuration. In this type of bonding, atoms form molecules that are held together by the shared electrons, and the resulting compound can be either polar or nonpolar, depending on the distribution of electrons. If the atoms involved have similar electronegativities, the electrons are shared equally, and the molecule is nonpolar. If there is a significant difference in electronegativity between the atoms, the electrons are shared unevenly, leading to a polar molecule with partial positive and negative charges on opposite ends.

The naming of covalent compounds follows specific rules based on the number of atoms involved. For example, the prefix system is used to indicate the number of each type of atom in the compound, such as "mono-" for one, "di-" for two, and "tri-" for three. The first element in the compound is named as the element itself, while the second element is modified to end in "-ide." For example, CO_2 is carbon dioxide, while SO_3 is sulfur trioxide.

Lewis structures are a useful way to represent covalent compounds, showing how atoms are connected and where the electrons are shared. In a Lewis structure, dots represent valence electrons, and lines are used to indicate shared electron pairs. These structures help visualize the bonding and predict the molecule's shape and properties.

The shape of a molecule is determined by the arrangement of atoms around the central atom, which is predicted using the VSEPR (Valence Shell Electron Pair Repulsion) theory. This theory states that electron pairs around a central atom will arrange themselves to minimize repulsion, leading to specific shapes such as linear, bent, trigonal planar, or tetrahedral, depending on the number of bonding and lone pairs of electrons. The geometry of a molecule affects its polarity and overall behavior, as molecules with certain shapes may exhibit dipole moments if they have an uneven distribution of charge.

In addition, the basics of organic chemistry focus on carbon-containing compounds, which form the backbone of countless molecules including biological molecules and fuels. Organic molecules often involve covalent bonding and are characterized by the presence of functional groups—specific groupings of atoms that determine the chemical reactivity and properties of the compound. Understanding covalent bonding, molecular shape, and functional groups is essential for studying the wide variety of organic compounds found in nature and industry.

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4.8: Key Terms

- Covalent Bond: A bond formed when two nonmetal atoms share electrons to achieve a stable electron configuration.
- Electron Sharing: The process by which atoms in a covalent bond share electrons to form a stable bond.
- Single Bond: A covalent bond where two atoms share one pair of electrons.
- Bond Pair: A pair of electrons shared between two atoms in a covalent bond.
- Duet: The duet states that hydrogen and helium atoms are stable with two electrons in their outermost shell, rather than eight, as in the octet rule. This rule applies to small atoms that only require two electrons to achieve a stable configuration.
- Double Bond: A covalent bond where two atoms share two pairs of electrons.
- Lone Pair: A pair of electrons that are not shared and belong to a single atom in a molecule.
- Triple Bond: A covalent bond where two atoms share three pairs of electrons.
- Molecular formula**:** Molecular formula shows the exact number of each type of atom in a molecule. It does not show how the atoms are arranged or bonded.
- Lewis Structure: A Lewis structure is a diagram that represents the bonding between atoms in a molecule and shows the lone pairs of electrons around atoms. It uses dots to represent electrons and lines to represent bonds, helping to visualize how atoms are connected and how electrons are distributed.
- Bond Length: The average distance between the nuclei of two bonded atoms in a molecule. It is typically shorter for stronger bonds (e.g., triple bonds are shorter than single bonds).
- Electronegativity: A measure of an atom's ability to attract and hold electrons in a covalent bond. The greater the electronegativity difference between atoms, the more polar the bond.
- Polar Covalent Bond: Polar covalent bond occurs when atoms with different electronegativities share electrons unequally. The more electronegative atom attracts the electrons more strongly, creating partial charges in the molecule.
- Nonpolar Covalent Bond: A type of chemical bond where two atoms share electrons equally, typically occurring between two identical nonmetals, such as in molecules like O_2 or H_2.
- Structural Formula: A formula that shows the bonding arrangement of atoms in a molecule, often represented by lines between atoms to indicate bonds.
- Molecular Shape: The three-dimensional arrangement of atoms in a molecule, determined by the number of bonding and lone electron pairs around the central atom.
- VSEPR (Valence Shell Electron Pair Repulsion) theory: states that the shape of a molecule is determined by the repulsion between electron pairs in the valence shell of atoms, which arrange themselves to minimize repulsion and maximize distance between them.
- Molecular Geometry: The three-dimensional arrangement of atoms in a molecule, determined by the VSEPR theory, which minimizes electron pair repulsion.
- Intermolecular Forces: Forces of attraction or repulsion between molecules, including hydrogen bonds, dipole-dipole interactions, and London dispersion forces.
- Dipole Moment: A measure of the polarity of a molecule, arising from the separation of positive and negative charges due to uneven electron distribution.
- Resonance Structures: Different Lewis structures that represent the same molecule, used when no single Lewis structure can accurately describe the bonding in a molecule.
- Hydrogen Bond: A strong intermolecular force between a hydrogen atom, which is covalently bonded to a highly electronegative atom (e.g., oxygen, nitrogen, or fluorine), and another electronegative atom.
- Hydrocarbons: Organic compounds made only of carbon and hydrogen. They include alkanes (single bonds), alkenes (double bonds), alkynes (triple bonds), and aromatics (ring structures).
- Alkanes: Saturated hydrocarbons containing only single bonds between carbon atoms. They follow the general formula C_nH_{2n+2} and include compounds like methane, ethane, and propane.
- Condensed Structural Formula: Simplified way of representing a molecule that shows the arrangement of atoms without drawing all the bonds explicitly. For example, propane can be written as $CH_3CH_2CH_3$.
- Alkenes: A hydrocarbon that contains at least one carbon–carbon double bond (C=C). General formula C_nH_{2n} for molecules with one double bond and no rings.
- Alkynes: A hydrocarbon that contains at least one carbon–carbon triple bond (C≡C). General formula C_nH_{2n-2} for molecules with one triple bond and no rings.
- Functional Groups: A functional group is a specific arrangement of atoms within a molecule that determines the molecule's characteristic chemical reactions. These groups are the reactive sites in organic compounds, such as –OH in alcohols or –COOH in carboxylic acids. Identifying functional groups helps predict how a compound will behave in chemical processes.

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4.9: Alternative text for Chapter 4

Figure 4.1: Molecular Structure of Cholesterol Highlighting Covalent Bonds

The image displays the molecular structure of cholesterol ($C_{27}H_{46}O$), a large organic compound composed of carbon, hydrogen, and oxygen atoms. The structure includes four fused carbon rings forming a rigid core and a hydrocarbon tail extending from one side. A single hydroxyl group (–OH), represented in red to highlight the oxygen and hydrogen atoms, is attached to one of the rings. This –OH group is the molecule's only polar feature, while the rest of the molecule is nonpolar. The complex, covalently bonded structure illustrates how covalent bonding allows for the formation of large and biologically important molecules.

Figure 4.1.1: Examples of Covalent Molecules Demonstrating Electron Sharing

The image presents several examples of covalent molecules, each demonstrating how different atoms achieve stable electron configurations by sharing electrons. In all cases, Hydrogen, shown as a white sphere, forms one covalent bond by sharing one electron, achieving a duet or two electrons total, like the noble gas Helium. The second row elements Chlorine (Cl), Oxygen (O), Nitrogen (N), and Carbon (C) are shown bonding in ways that allow each of them to reach a stable octet, or eight electrons in their outer shell. In the first example, Hydrogen bonds with Chlorine (Cl), which shares one electron to complete its octet. In the second, a water molecule (H_2O) is shown, where an Oxygen atom shares two electrons by forming single bonds with two Hydrogen atoms. The third example shows ammonia (NH_3), where a Nitrogen atom shares three electrons by forming three bonds with Hydrogen atoms. In the fourth, methane (CH_4) is shown, where a Carbon atom shares four electrons, one with each Hydrogen atom, completing its octet.

Figure 4.1.2: Formation of a Diatomic Oxygen Molecule (O_2)

The image illustrates the formation of a diatomic oxygen molecule (O_2) using Lewis structures, with yellow dots representing valence electrons. At the beginning, two separate oxygen atoms are shown, each with six yellow dots around them, indicating their valence electrons. A plus sign between the atoms shows they are unbonded and approaching one another. In the next step, the two oxygen atoms move closer, and two yellow dots from each atom begin to overlap, signaling the start of covalent bonding. In the following stage, a double bond is represented by two dashed lines between the oxygen atoms, indicating that each atom is sharing two electrons, forming a total of four shared electrons in a double covalent bond. Each oxygen still has two lone pairs (four yellow dots) on the outer sides. In the final step, the completed molecule is shown as O_2 representing a stable diatomic oxygen molecule held together by a double bond.

Figure 4.1.3: Formation of a Diatomic Nitrogen Molecule (N_2)

The image illustrates the formation of a diatomic nitrogen molecule (N_2) using Lewis structures, with yellow dots representing valence electrons. Initially, two separate nitrogen atoms are shown, each with five yellow dots around them, representing their valence electrons. A plus sign between the atoms indicates that they are unbonded and about to interact. In the next step, the atoms move closer together, and three yellow dots from each nitrogen atom begin to overlap in the space between them. This overlapping area indicates the formation of covalent bonds. In the following stage, a triple bond is represented by three dashed lines between the nitrogen atoms, showing that each atom is sharing three electrons, for a total of six shared electrons. Each nitrogen atom retains one lone pair (two yellow dots) outside the bonding region. In the final step, the molecule is written as "N_2," representing a stable nitrogen molecule held together by a strong triple covalent bond.

Figure 4.2.1: The Molecular Formula of Ammonia

The figure shows the molecular formula of ammonia, written as NH_3. The chemical formula NH_3 uses a subscript "3" next to hydrogen to indicate that three hydrogen atoms are bonded to one nitrogen atom. There is no subscript shown next to nitrogen,

which means only one nitrogen atom is present. In chemical notation, it is a standard convention not to include the number "1" as a subscript; its absence automatically implies a single atom of that element. This figure helps explain how subscripts in molecular formulas communicate the number of atoms of each element in a covalent compound, reinforcing the idea that ammonia consists of four atoms in total: one nitrogen and three hydrogen atoms.

Figure 4.3.1: Lewis structure of a simple covalent molecules

The image is divided into three rows, each showing the Lewis structure of a simple covalent molecule. In the first row, the structure of hydrogen bromide (HBr) is shown. A hydrogen atom is bonded to a bromine atom by a single covalent bond, represented by a single line. Bromine has three lone pairs of electrons surrounding it, completing its octet, while hydrogen has no lone pairs and satisfies the duet rule. In the second row, the Lewis structure of hydrogen sulfide (H_2S) is displayed. A sulfur atom is at the center, bonded to two hydrogen atoms on either side with single covalent bonds. Sulfur also has two lone pairs of electrons, shown above and below the atom, allowing it to achieve a full octet. In the third row, the structure of nitrogen gas (N_2) is shown, with two nitrogen atoms connected by a triple bond, represented by three parallel lines. Each nitrogen atom also has one lone pair of electrons. These Lewis structures illustrate how atoms share electrons in covalent bonds to achieve stable electron configurations, following the duet rule for hydrogen and the octet rule for second-row elements.

Figure 4.3.2: The Lewis Structure of Three Molecules

The image contains two simple molecular skeleton structures side by side. On the left is oxygen difluoride (OF_2). The central atom is oxygen (O), with two fluorine (F) atoms bonded to it. Each fluorine atom is connected to the oxygen atom by a single bond (one line), representing a pair of shared electrons. The fluorine atoms are positioned on either side of the oxygen atom, and no other bonds are shown, reflecting that fluorine forms only one bond and is located on the periphery. On the right is hydrogen cyanide (HCN). The central atom is carbon (C), bonded to a hydrogen (H) atom on one side and a nitrogen (N) atom on the other. Each bond is represented by a single line. Hydrogen, which forms only one bond, is placed at one end of the molecule, while nitrogen is at the opposite end, both connected to carbon.

Figure 4.3.3: Initial Electron Distribution in OF_2 and HCN Lewis Structures

The image is divided into two sections.: the left section shows the Lewis structure of OF_2, with a central oxygen atom bonded to two fluorine atoms, each fluorine having three lone pairs. The right section shows the Lewis structure of HCN, with hydrogen single-bonded to carbon, and carbon single-bonded to nitrogen. The nitrogen atom has three lone pairs, and and hydrogen has none.

Figure 4.3.4: Lewis Structures Showing Lone Pairs and Incomplete Octet in HCN

The image is divided into two sections. The left section shows the Lewis structure of OF_2, with oxygen in the center single-bonded to two fluorine atoms, each fluorine having three lone pairs. The oxygen atom has two lone pairs, completing its octet. The right section shows the Lewis structure of HCN, with hydrogen single-bonded to carbon, and carbon single-bonded to nitrogen. The nitrogen atom has three lone pairs, but the carbon atom does not yet have a full octet.

Figure 4.3.5: No Multiple Bonds Needed in OF_2: Octet Completion with Single Bonds

In oxygen difluoride (OF_2), electron rearrangement to form multiple bonds is not necessary. Each fluorine atom is a terminal atom and already has a complete complete octet by forming a single bond with oxygen and holding three lone pairs. The central oxygen atom also reaches an octet by forming two single bonds (one with each fluorine) and keeping two lone pairs of its own. Since both the central and outer atoms achieve full valence shells, there is no need to create double or triple bonds in this structure. The bonding remains with two single bonds, and the molecule is stable as is.

Figure 4.3.6: Formation of the Carbon-Nitrogen Triple Bond in HCN

In hydrogen cyanide (HCN), the terminal nitrogen atom has a complete octet with three lone pairs. The central carbon atom initially forms single bonds with hydrogen and nitrogen and does not yet have a full octet. To satisfy the octet rule, electrons are rearranged to form multiple bonds between carbon and nitrogen. This process is shown with two arrows below the initial Lewis structure, illustrating how electron pairs from nitrogen's lone pairs move toward the bonding region. Below the arrows, the final Lewis structure shows a triple bond between carbon and nitrogen, completing the valence shells of all atoms.

Figure 4.3.7: Initial Distribution of All Valence Electrons in Carbon Dioxide

Carbon dioxide (CO_2) has a total of 16 valence electrons—4 from the central carbon atom and 6 from each oxygen atom. In the initial Lewis structure, carbon is placed in the center with one oxygen atom on each side. Single bonds between carbon and oxygen are represented by pairs of dots, indicating shared electron pairs. Each oxygen atom receives three lone pairs to complete its octet. All 16 electrons are accounted for in this arrangement; however, the central carbon atom is left with only four electrons, showing that multiple bonds are necessary to satisfy the octet rule for carbon.

Figure 4.3.8: Electron Movement and Bond Formation in Carbon Dioxide

To satisfy the octet rule for the central carbon atom in carbon dioxide (CO_2), lone pair electrons from each oxygen atom are shifted toward the carbon. This electron movement forms additional bonds, helping carbon complete its octet. The process is illustrated by two arrows, each pointing from a lone pair on an oxygen atom toward the bonding region between that oxygen and the carbon atom.

Figure 4.3.9: Final Lewis Structure of CO_2 with Bonds Shown as Dots

In the final Lewis structure of carbon dioxide (CO_2), lone pairs from each oxygen atom are shared with the central carbon atom, forming two double bonds. These shared electron pairs are depicted as dots to clearly show bonding. Each atom—carbon and both oxygens—has a complete octet of eight electrons, satisfying the octet rule and resulting in a stable, linear molecule.

Figure 4.3.10: Final Lewis Structure of Carbon Dioxide (CO_2) Using Lines to Represent Bonds

The image shows the final Lewis structure of carbon dioxide (CO_2) using lines to represent bonds. A central carbon atom is double bonded to two oxygen atoms, one on each side. Each double bond is represented by two solid lines. Both oxygen atoms also have two lone pairs of electrons shown as dots, while the central carbon atom has no lone pairs. All atoms have complete octets, and the molecule is linear in shape.

Figure 4.3.11: Complete Lewis Structure of Acetylene (C_2H_2) in Two Linear Formats

The image displays the complete Lewis structure of acetylene (C_2H_2) in two linear formats. On the left, the structure is drawn using only electron dots. Two carbon atoms are shown in the center, each sharing three pairs of dots to form a triple bond. Each carbon also shares a single pair of dots with one hydrogen atom. On the right, the same structure is shown using lines to represent bonds: a triple bond (three parallel lines) connects the two carbon atoms, and each carbon is single-bonded (one line) to a hydrogen atom. No lone pairs are present, and all atoms have a full valence shell.

Figure 4.3.12: Lewis Structure of Nitrogen Gas (N_2) in Dot-and-Line Representations

In N_2 (nitrogen gas), the Lewis structure can be described in two equivalent ways. On the left (dot-only structure), each nitrogen atom contributes five valence electrons, for a total of 10. The two nitrogen atoms share three pairs of electrons, shown as six dots between them, representing a triple bond. Each nitrogen atom also has one lone pair of electrons, shown as two dots on each atom. This representation displays all valence electrons explicitly, including both bonding and nonbonding pairs.

On the right (line-and-dot structure), the three shared electron pairs are represented as three lines, forming a triple bond (N≡N) between the two nitrogen atoms. Each nitrogen atom still has one lone pair, shown as a pair of dots next to each atom. This version emphasizes the bonding arrangement while still indicating the presence of lone pairs, and both nitrogen atoms satisfy the octet rule.

Figure 4.3.13: Lewis Structure of CH_2O. Dot-Only (Left) and Line–Bond with Lone Pairs (Right)

In CH_2O (formaldehyde), the Lewis structure can be described in two ways. On the left (dot-only structure), all 12 valence electrons are shown explicitly as pairs of dots. Carbon is in the center with two hydrogen atoms attached on either side, each bond represented by a pair of dots showing shared electrons. Carbon also forms a double bond with oxygen, shown as two shared electron pairs (between C and O). Oxygen carries two lone pairs of electrons, shown as four additional dots grouped in pairs on the oxygen atom. This full dot representation makes every valence electron visible, whether it is bonding or nonbonding.

On the right (line-and-dot structure), the bonding electrons are simplified into lines while only lone pairs remain as dots. Carbon is shown in the center with single lines connecting it to each hydrogen atom (H—C—H), and a double line between carbon and oxygen (C=O). Oxygen still has two lone pairs of electrons, drawn as pairs of dots next to the atom. This version emphasizes the molecular framework and bonding pattern more clearly, while still indicating that oxygen has nonbonding electron pairs and that all atoms satisfy their octet (or duet for hydrogen).

Figure 4.3.14: Lewis Structure of N_2. Dot-Only (Left) and Line–Bond with Lone Pairs (Right)

In diazene (N_2H_2), the Lewis structure can be described in two equivalent ways. On the left, the structure is drawn with dots representing lone pairs and lines representing bonds: a double bond between the two nitrogen atoms (N=N), with each nitrogen bonded to one hydrogen atom and carrying one lone pair. On the right, the same structure is shown in an alternative dot-and-line format, again illustrating the N=N double bond, N–H single bonds, and one lone pair on each nitrogen atom.

Figure 4.4.1: Examples of Nonpolar Covalent, Polar Covalent, and Ionic Bonds

The image is divided into three labeled parts. On the left (A) is Cl_2, representing a nonpolar covalent bond where two identical chlorine atoms share electrons equally, resulting in no charge separation. In the middle (B) is HCl, showing a polar covalent bond between hydrogen and chlorine; the electrons are more attracted to chlorine, leading to partial negative and positive charges on chlorine and hydrogen, respectively. On the right (C) is NaCl, illustrating an ionic bond formed by the complete transfer of an electron from sodium to chlorine, creating oppositely charged ions (Na^+ and Cl^-) held together by electrostatic attraction.

Figure 4.4.2: Electronegativity Trends on the Periodic Table

In the figure, a periodic table is displayed with arrows showing electronegativity trends. At the top of the table, a horizontal arrow points to the right and is labeled ‘increasing electronegativity.’ On the left side of the table, a vertical arrow points downward and is labeled ‘decreasing electronegativity.’ These arrows indicate that electronegativity increases toward the upper right corner of the periodic table, where fluorine has the highest value.

Figure 4.4.3: Portrait of Linus Pauling

The figure shows a portrait of Linus Pauling, a prominent 20th-century chemist, depicted in a formal setting. He appears as an older man with neatly combed hair, wearing a suit and tie, and looking directly toward the camera with a composed expression. The image emphasizes his professional stature and historical significance, often associated with his groundbreaking contributions to chemical bonding theory, molecular biology, and peace activism, including his two individual Nobel Prizes in Chemistry and Peace.

Figure 4.4.4: Relationship Between Electronegativity Difference and Bond Type

The image is divided into four vertical columns. The first column shows the difference in electronegativity between bonding atoms: at the bottom, the difference is zero; in the middle, it is intermediate; and at the top, the difference is large. The second column aligns these differences with bond types: pure covalent bond for zero difference, polar covalent bond for intermediate difference, and ionic bond for a large difference in electronegativity. The third column displays a vertical arrow pointing upward, indicating that covalent character decreases upward, while ionic character increases upward as electronegativity difference increases. The fourth column contains a table summarizing all three bond types along with their typical electronegativity differences and their placement on the electronegativity scale. This layout visually explains how the type of bond correlates with the difference in electronegativity between atoms.

Figure 4.5.1: The Bond Angle in $BeCl_2$

The image is divided into two parts. On the left, a Lewis structure of beryllium chloride ($BeCl_2$) shows a central beryllium atom with single bonds extending to one chlorine atom on each side. No lone pairs are shown on beryllium. The bond angle between the two chlorine atoms is labeled as 180°, indicating a linear geometry. On the right, a ball-and-stick model visually represents the same structure: a small central ball for beryllium is connected in a straight line to two larger balls for chlorine atoms, illustrating the molecule's linear shape.

Figure 4.5.2: The Bond Angle in BF_3

The image is divided into two parts. On the left, the Lewis structure of boron trifluoride (BF_3) shows a central boron atom bonded to three fluorine atoms arranged evenly around it. Each bond is a single line, and there are no lone pairs on the boron atom. The fluorine atoms are positioned at 120° angles, forming a flat, trigonal planar shape. On the right, a ball-and-stick model depicts the same structure: a central ball representing boron is connected to three surrounding balls for fluorine atoms, all lying in the same plane and spaced equally apart, illustrating the trigonal planar geometry predicted by VSEPR theory.

Figure 4.5.3: The Bond Angle in CH_2O

The image is divided into two parts. On the left, the Lewis structure of formaldehyde (CH_2O) shows a central carbon atom double bonded to an oxygen atom and single bonded to two hydrogen atoms. The atoms are arranged in a trigonal planar geometry with bond angles close to 120°. The structure follows typical VSEPR predictions, although lone pairs on the oxygen atom are not shown in this depiction. On the right, a ball-and-stick model shows a central ball for carbon connected to two smaller hydrogen atoms and one larger oxygen atom. All atoms lie in the same plane, representing the flat, trigonal planar shape of the molecule.

Figure 4.5.4: The Bond Angle in NH_3

The image is divided into two parts. On the left, the Lewis structure of ammonia (NH_3) shows a central nitrogen atom single bonded to three hydrogen atoms. A lone pair of electrons is clearly shown on the nitrogen atom. The electron geometry is tetrahedral, but due to the lone pair, the molecular shape is trigonal pyramidal, with bond angles slightly less than 109.5°. On the right, a ball-and-stick model represents the same molecule: the nitrogen atom is at the center, connected to three hydrogen atoms arranged in a pyramid-like shape. The lone pair on nitrogen is indicated, emphasizing how it influences the molecule's three-dimensional geometry.

Figure 4.5.5: The Bond Angle in H_2O

The image shows the Lewis structure of a water (H_2O) molecule. The central oxygen atom is single bonded to two hydrogen atoms. Two lone pairs of electrons are shown on the oxygen, occupying positions above the atom. This arrangement gives the molecule a bent geometry, with a bond angle of approximately 104.5°, as predicted by VSEPR theory. The structure visually reflects how lone pairs on the oxygen repel bonding pairs, resulting in the characteristic V shape (bent) of the water molecule.

Figure 4.5.6: Molecular Geometry Determined by Electron Regions and Lone Pairs

The figure is organized as a table with six columns, where each row represents a different number of electron regions (ranging from 2 to 6) around a central atom labeled X. The first column lists the number of electron regions: 2, 3, 4, 5, and 6. The remaining five columns display corresponding Lewis structures for molecular shapes with 0 to 4 lone pairs on the central atom E (another symbol used for the central element). Each Lewis structure shows bonding regions as lines connecting to outer atoms and lone pairs as dots placed on the central atom. Below each Lewis structure, the name of the molecular geometry is written, reflecting how the presence of lone pairs alters the shape predicted by VSEPR theory

Figure 4.5.7: Lewis Structure of the Ammonium Ion (NH_4^+)

The figure shows a two-dimensional Lewis structure of the ammonium ion (NH_4^+) drawn using straight lines for bonds. A central nitrogen (N) atom is placed in the middle, with four hydrogen (H) atoms positioned around it—one above, one below, one to the left, and one to the right—forming a cross-like arrangement in the plane. Each hydrogen is connected to the nitrogen by a single straight line representing a covalent bond (a shared pair of electrons). No lone pairs are shown on the nitrogen atom. The entire structure is enclosed within square brackets, and a superscript plus sign (+) is written outside the brackets to indicate the overall positive charge of the ion. The diagram emphasizes the four N–H single bonds and the ionic charge, using a simple flat (2D) representation rather than a $+$-dimensional geometry

Figure 4.5.8: Tetrahedral Representation of the Ammonium Ion (NH_4^+)

The figure shows a three-dimensional representation of a generic tetrahedral structure labeled AX_4^+. The central atom ‘A’ is positioned in the middle, with four surrounding atoms labeled ‘X.’ The structure uses wedge-and-dash notation to indicate three-dimensional geometry. Two X atoms are drawn in the plane of the page, connected to A by straight lines. One X atom is connected to A by a solid wedge, indicating a bond projecting out of the page toward the viewer. Another X atom is connected by a dashed wedge, indicating a bond extending behind the plane of the page. All four A–X bonds are single covalent bonds, and no lone pairs are shown on the central atom. The entire structure is enclosed in square brackets, with a superscript positive charge (+) shown outside the brackets, indicating an overall +1 charge. The diagram illustrates a tetrahedral arrangement around the central atom using a 3D perspective

Figure 4.5.9: Lewis Structure of Ammonia (NH_3)

The figure shows a two-dimensional Lewis structure of ammonia (NH_3). A central nitrogen (N) atom is connected to three hydrogen (H) atoms using single straight lines representing covalent bonds. The hydrogen atoms are arranged around the nitrogen in a roughly trigonal planar layout in the plane of the page. The nitrogen atom has one lone pair of electrons shown as two dots above it. The structure is enclosed in a simple flat 2D representation, emphasizing bonding pairs and the lone pair on nitrogen.

Figure 4.5.10: Trigonal Pyramidal Representation of Ammonia (NH_3)

The figure shows a three-dimensional representation of ammonia (NH_3), labeled generically as AX_3. The central atom ‘A’ represents nitrogen, and it is bonded to three surrounding atoms labeled ‘X,’ representing hydrogen atoms. The geometry is trigonal pyramidal. One A–X bond is shown as a solid wedge projecting out of the plane toward the viewer, one is shown as a dashed wedge projecting behind the plane, and one is shown as a straight line in the plane. A lone pair of electrons is shown on atom A as two dots. This illustrates the three-dimensional structure of ammonia and the effect of the lone pair on molecular shape.

Figure 4.5.11: Lewis Structure of Nitrosyl Chloride (NOCl)

The figure shows the Lewis structure of nitrosyl chloride (NOCl). A nitrogen (N) atom is positioned in the center, bonded to an oxygen (O) atom and a chlorine (Cl) atom. The N–O bond is shown as a double bond, and the N–Cl bond is shown as a single bond. Lone pairs are displayed on the oxygen and chlorine atoms. The nitrogen atom also has one lone pair of electrons. The structure is arranged in a bent (angular) layout, reflecting the electron geometry around nitrogen in NOCl.

Figure 4.5.12: Bent Representation of Nitrosyl Chloride (NOCl)

The figure shows a three-dimensional representation of nitrosyl chloride (NOCl) using a generic AX_2 notation. The central atom 'A' represents nitrogen, and it is bonded to two surrounding atoms labeled 'X,' representing oxygen and chlorine. Both A–X bonds are shown as straight lines. The two X atoms are arranged in a bent geometry rather than linear, with the central atom A at the vertex. A lone pair of electrons is shown on atom A as two dots. The diagram illustrates the bent molecular shape of NOCl due to lone pair repulsion around the central atom.

Figure 4.5.13: Lewis Structures and Dipole Representations Showing Polarity in CO_2 and H_2O

The image shows a side-by-side comparison of the Lewis structures of carbon dioxide (CO_2) on the left and water (H_2O) on the right, each illustrating molecular polarity through dipole arrows. In the left image, the CO_2 molecule is shown as a linear structure, with a central carbon atom double bonded to two oxygen atoms, one on each side. Two red arrows, representing bond dipoles, point from the carbon atom toward each oxygen atom, indicating that electrons are pulled toward the more electronegative oxygen atoms. The arrows are equal in length and point in exactly opposite directions, so their effects cancel out. There is no blue arrow indicating a net dipole because the molecule is nonpolar. In the right image, the H_2O molecule is shown with a bent geometry, where the central oxygen atom is single bonded to two hydrogen atoms and also has two lone pairs (noted on the oxygen). Two red arrows point from each hydrogen atom toward the oxygen atom, showing the polar O–H bonds. Because of the bent shape, these red arrows are not directly opposite each other, and do not cancel. A blue arrow pointing upward from between the hydrogen atoms toward the oxygen shows the net molecular dipole. This indicates that water is a polar molecule, with a partial positive charge near the hydrogen side and a partial negative charge on the oxygen side.

Figure 4.5.14: Geometry Determines Polarity: Nonpolar BF_3 and Polar NH_3

The figure compares two molecular structures: BF_3 on the left and NH_3 on the right, to illustrate the relationship between molecular geometry and polarity. On the left, the BF_3 (boron trifluoride) molecule is shown in a trigonal planar geometry. A central boron atom is bonded to three fluorine atoms arranged symmetrically at 120° angles in a flat triangle. Three black arrows represent bond dipoles, each pointing from the boron atom toward a fluorine atom. Due to the symmetrical shape of the molecule, these bond dipoles are equal in magnitude and arranged evenly around the central atom, so they cancel out, resulting in no overall dipole moment. Thus, BF_3 is nonpolar. On the right, the NH_3 (ammonia) molecule is shown in a trigonal pyramidal geometry. A central nitrogen atom is bonded to three hydrogen atoms (the lone pair on nitrogen is not shown). Three black arrows represent the individual N–H bond dipoles, each pointing from hydrogen toward nitrogen. Because of the molecule's asymmetric, nonplanar shape, these dipoles do not cancel out. A fourth black arrow, larger and pointing upward, represents the net dipole of the molecule, indicating that NH_3 is polar, with a partial negative charge near the nitrogen atom and a partial positive region near the hydrogen atoms.

Figure 4.5.15: Effect of Molecular Symmetry on Polarity in Tetrahedral Molecules — CCl_4 vs. CH_3Cl

The image compares the molecular polarity of CCl_4 (carbon tetrachloride) on the left and CH_3Cl (chloromethane) on the right, both shown in tetrahedral geometries. In both structures, the central carbon atom is colored green. On the left, CCl_4 has four chlorine atoms (colored red) symmetrically arranged around the central carbon. Each C–Cl bond is shown with a black arrow pointing outward from the green carbon to each red chlorine atom, representing bond dipoles. Due to the molecule's symmetry, these dipoles cancel out, resulting in no net dipole moment. The molecule is nonpolar. On the right, CH_3Cl has one chlorine atom (red) and three hydrogen atoms (gray) bonded to the central green carbon atom. A single black arrow points from carbon toward the

red chlorine atom, representing the only significant polar bond. No arrows are shown between hydrogen and carbon, since these bonds are considered nonpolar in this context. Because the structure is asymmetrical and the C–Cl bond is polar, the dipoles do not cancel, making CH_3Cl a polar molecule with a net dipole toward the red chlorine atom.

Figure 4.5.16: Comparison of Molecular Geometry and Polarity

The image shows six molecular structures displayed from left to right, each illustrating molecular geometry and polarity using red arrows for individual bond dipoles and a blue arrow to indicate the overall dipole moment, if present.

1. HCl (hydrogen chloride): A simple linear molecule with a red arrow pointing from hydrogen to chlorine, indicating electron density shifts toward chlorine. A blue arrow overlays the red one, showing a net dipole from H to Cl. This molecule is polar.
2. BCl_3 (boron trichloride): A trigonal planar molecule with three red arrows pointing from boron to each chlorine, arranged symmetrically. Due to this symmetry, the bond dipoles cancel. No blue arrow is shown. BCl_3 is nonpolar.
3. CH_2O (formaldehyde): A trigonal planar molecule. Two red arrows point from hydrogen atoms to carbon, and a stronger red arrow points from carbon to oxygen. A blue arrow points from the carbon toward the oxygen, indicating a net dipole in that direction. CH_2O is polar.
4. NH_3 (ammonia): A trigonal pyramidal structure with three red arrows pointing from hydrogen atoms toward the central nitrogen atom. These arrows do not cancel due to the shape. A blue arrow points upward from the molecule, indicating the overall dipole. NH_3 is polar.
5. $CHCl_3$ (chloroform): A tetrahedral molecule with three red arrows pointing from carbon to each chlorine atom and one hydrogen atom. Due to uneven distribution of chlorine atoms, the bond dipoles do not cancel. A blue arrow points toward the side with more chlorine atoms. $CHCl_3$ is polar.
6. CCl_4 (carbon tetrachloride): A tetrahedral molecule with four red arrows pointing symmetrically from carbon to each chlorine atom. These bond dipoles cancel due to perfect symmetry. No blue arrow is shown. CCl_4 is nonpolar.

Figure 4.5.17: Lewis Structure of Phosphorus Trichloride (PCl_3)

The figure shows the Lewis structure of phosphorus trichloride (PCl_3). At the center is a phosphorus (P) atom, represented by its chemical symbol. Surrounding it are three chlorine (Cl) atoms positioned around the phosphorus, typically arranged in a trigonal pyramidal shape. Each chlorine atom is connected to the phosphorus atom by a single covalent bond, represented by a pair of shared electrons (a line). The phosphorus atom has one lone pair of electrons, shown as two dots placed on the phosphorus, giving it a total of four electron domains. Each chlorine atom has three lone pairs of electrons (six dots) arranged around it, in addition to the bonding pair shared with phosphorus. The structure illustrates that phosphorus forms three single bonds with chlorine and retains one lone pair, resulting in a trigonal pyramidal molecular geometry. The diagram emphasizes electron pairs, bonding connections, and the spatial arrangement of atoms in the molecule.

Figure 4.5.18: Lewis Structure of Carbon Dioxide (CO_2)

The figure shows the Lewis structure of carbon dioxide (CO_2). A central carbon (C) atom is positioned between two oxygen (O) atoms in a straight horizontal line, forming a linear arrangement. The carbon atom is connected to each oxygen atom by a double covalent bond, represented by two pairs of shared electrons (two lines) on each side. Each oxygen atom has two lone pairs of electrons (four dots) placed around it, while the carbon atom has no lone pairs. The structure shows that carbon forms a total of four bonds (two double bonds), satisfying its octet, and each oxygen also achieves a complete octet through its double bond and lone pairs. The overall geometry is linear, with a bond angle of 180 degrees, illustrating equal electron distribution on both sides of the central carbon atom.

Figure 4.5.19: A Molecule With an Unequal Distribution of Electron Density

The image shows an oval-shaped electron cloud representing a molecule with an unequal distribution of electron density. One end of the cloud is labeled with a partial negative charge (δ–), where the electron density is higher, and the opposite end is labeled with a partial positive charge (δ+), where the electron density is lower. This illustrates a molecular dipole, with an arrow pointing from the δ+ to the δ– end to indicate the direction of the dipole moment.

Figure 4.5.20: Orientation of Nonpolar and Polar Molecules in an Electric Field

The image depicts multiple identical polar molecules in two scenarios. At the top, in the absence of an electric field, the molecules are randomly oriented despite having permanent dipoles. At the bottom, in the presence of an electric field, the same type of polar molecules align uniformly—positive ends toward the negative field side and negative ends toward the positive field side—demonstrating molecular polarization due to the external electric field.

Figure 4.5.21: Influence of Molecular Geometry and Polarity on Physical Properties

In the image on the left, water (H_2O) is shown with a bent molecular geometry. This shape, combined with the large difference in electronegativity between oxygen and hydrogen, creates a polar molecule with a strong net dipole moment. The polarity of water results in significant intermolecular forces—specifically, hydrogen bonding—which are much stronger than the van der Waals forces found in nonpolar molecules. These strong interactions give water its relatively high boiling point of 100°C and make it a liquid at room temperature. Additionally, water's polarity makes it an excellent solvent for other polar substances and ionic compounds, contributing to its nickname as the "universal solvent." In contrast, the image on the right shows carbon dioxide (CO_2), a molecule with a linear geometry. Although each C=O bond is polar due to oxygen's higher electronegativity, the symmetry of the molecule causes the dipoles to cancel out, resulting in a nonpolar molecule. Because CO_2 is nonpolar, it only exhibits weak London dispersion forces between its molecules. These weak forces mean that CO_2 has a much lower boiling point (–77°C) and exists as a gas under standard conditions. Unlike water, CO_2 does not dissolve well in polar solvents and has limited solubility in water. The stark contrast in physical properties between these two small molecules highlights how crucial molecular polarity is in determining behavior such as boiling point, state of matter, and solubility.

Figure 4.6.1: Lewis Structures and Formulas of the First Four Alkanes

The image presents a comparison of four hydrocarbons—methane, ethane, propane, and butane—organized in three rows. The top row shows their complete Lewis structures, where each atom and bond is explicitly represented. Methane consists of a single carbon atom bonded to four hydrogen atoms. Ethane has two carbon atoms bonded together, each attached to three hydrogen atoms. Propane features a straight chain of three carbon atoms with the appropriate number of hydrogen atoms completing the octet rule. Butane extends this pattern to a four-carbon chain with ten hydrogen atoms total. The middle row displays the condensed structural formulas of these hydrocarbons: CH_4 for methane, CH_3CH_3 for ethane, $CH_3CH_2CH_3$ for propane, and $CH_3CH_2CH_2CH_3$ for butane. These notations simplify the structure by grouping atoms bonded to each carbon. The bottom row shows the molecular formulas, summarizing the number of atoms in each compound: CH_4, C_2H_6, C_3H_8, and C_4H_{10}, respectively. This layout provides a visual and symbolic overview of how the structure and composition of alkanes grow with increasing carbon chain length.

Figure 4.6.2: Lewis Structures and Molecular Formulas of Ethene and Propene

The image is divided into two main columns. On the left, the names and molecular formulas of the two simplest alkenes are listed: Ethene (C_2H_4) at the top and Propene (C_3H_6) at the bottom. On the right, each compound's corresponding Lewis structure is shown. Ethene's structure features a double bond between the two carbon atoms, with each carbon also bonded to two hydrogen atoms. Propene's Lewis structure shows a double bond between the first two carbon atoms, with the third carbon single-bonded to the second and carrying three hydrogen atoms. All atoms follow the octet rule.

Figure 4.6.3: Lewis Structure of Ethyne Highlighting the Triple Bond

The Lewis structure of ethyne (C_2H_2) shows two carbon atoms connected by a triple bond, with each carbon atom also bonded to one hydrogen atom. The triple bond is clearly emphasized with three lines between the carbons, illustrating the strong covalent bonding typical of alkynes. The structure is linear, with all atoms having complete valence shells.

Figure 4.6.4: USDA MyPyramid Dietary Guidelines Illustration

The figure shows the USDA MyPyramid dietary guide. The pyramid is divided into several colored vertical sections that widen toward the bottom and narrow toward the top, illustrating recommended proportions of different food groups. From left to right, the sections represent grains, vegetables, fruits, oils, milk, and meat & beans. Each section is a different color and contains a simple icon representing its food group. A small figure of a person climbing steps on the side of the pyramid emphasizes the importance of physical activity. The base of the pyramid is the widest, showing grains as the largest recommended portion of daily intake, while the top is the narrowest, indicating that oils and fats should be consumed in smaller amounts. The overall design visually communicates balance, variety, moderation, and the idea that healthier eating patterns include larger proportions of plant-based foods and limited intake of fats and sugars.

Figure 4.6.5: Methanol and Ethanol Highlighting the Alcohol Functional Group

The image shows two alcohols arranged vertically. On the left, the names “Methanol” (top) and “Ethanol” (bottom) are displayed. To the right of each name are their respective Lewis structures. Methanol’s structure features a single carbon atom bonded to three hydrogens and one hydroxyl group (–OH), highlighting the alcohol functional group. Ethanol’s structure, below it, consists of two carbon atoms in a chain with the terminal carbon bonded to the hydroxyl group. Both structures use lines to represent bonds and show lone pairs on the oxygen atom, emphasizing the presence of the alcohol group that characterizes these compounds.

Figure 4.6.6: Methanoic Acid and Ethanoic Acid Highlighting the Carboxylic Acid Functional Group

The image displays two organic acids arranged vertically. On the left, the names “Methanoic Acid” (top) and “Ethanoic Acid” (bottom) are shown. To the right of each name are their corresponding Lewis structures. Methanoic acid’s structure shows a single carbon atom bonded to a hydroxyl group (–OH) and double-bonded to an oxygen atom, forming the characteristic carboxyl functional group. Ethanoic acid’s structure, below it, has two carbon atoms in a chain, with the terminal carbon bonded to the same carboxyl group. Both structures clearly depict bonds with lines and include lone pairs on the oxygen atoms. This emphasizes the carboxyl group, which defines these acids’ chemical properties.

Figure 4.6.7: Structures of Three Organic Compounds (Left to Right)

Figure showing the chemical structures of 2-Butenoic acid, 3-(methylamino)propan-1-ol, and 3-Chloro-2-methylbut-2-enoic acid. The left structure shows a four-carbon chain with a carbon–carbon double bond between the second and third carbons and a terminal carboxylic acid group. The middle structure shows a three-carbon chain with a hydroxyl group attached to carbon 1 and a methylamino group attached to carbon 3. The right structure shows a four-carbon carboxylic acid chain with a double bond between carbons 2 and 3, a methyl group attached to carbon 2, and a chlorine atom attached to carbon 3.

Figure 4.6.8: Structures of Four Organic Compounds (Left to Right)

A figure showing four organic compounds arranged horizontally. The first is 2-butyne, a four-carbon alkyne with a triple bond between carbons 2 and 3. The second is methanoic acid (formic acid), the simplest carboxylic acid with a single carbon bonded to a carboxyl group. The third is 3-hexanol, a six-carbon chain alcohol with the hydroxyl group on carbon 3. The fourth is 2-butenoic acid (crotonic acid), a four-carbon carboxylic acid with a double bond between carbons 2 and 3.

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5: INTRODUCTION TO CHEMICAL REACTIONS

Introduction

Chemical change is a central concept in chemistry. The goal of chemists is to know how and why a substance changes in the presence of another substance or even by itself. Because there are tens of millions of known substances, there are a huge number of possible chemical reactions. In this chapter, we will find that many of these reactions can be classified into a small number of categories according to certain shared characteristics.

Learning Objectives

1. Define a chemical reaction.
2. Explain the roles of subscripts and coefficients in chemical equations.
3. Use a balanced chemical equation to represent a chemical reaction.
4. Balance chemical reactions to ensure conservation of mass.
5. Assess the relationship between reactants and products in a chemical reaction.
6. Calculate the amount of one substance that will react with or be produced from a given amount of another substance.

5.1: Prelude to Introduction to Chemical Reactions

5.2: The Law of Conservation of Matter

5.3: Chemical Equations

5.4: How to Balance Chemical Equations

5.5: Quantitative Relationships Based on Chemical Equations

5.6: Summary

5.7: Key Terms

5.8: Alternative text for Chapter 5

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5.1: Prelude to Introduction to Chemical Reactions

The Essential Role of Yeast in Food

Although yeast has been used for thousands of years, its true nature has been known only for the last two centuries. Yeasts are single-celled fungi. About 1,000 species are recognized, but the most common species is *Saccharomyces cerevisiae*, which is used in bread making. Other species are used for the fermentation of alcoholic beverages. Some species can cause infections in humans.

Yeasts live primarily on sugars, such as glucose ($C_6H_{12}O_6$). They convert glucose into carbon dioxide (CO_2) and ethanol (C_2H_5OH) in a chemical transformation that is represented as follows:

$$C_6H_{12}O_{6(s)} \rightarrow 2\ CO_{2(g)} + 2\ C_2H_5OH_{(\ell)}$$

Bread making depends on the production of carbon dioxide. The gas, which is produced in tiny pockets in bread dough, acts as a leavening agent: it expands during baking and makes the bread rise. Leavened bread is softer, lighter, and easier to eat and chew than unleavened bread. The other major use of yeast, fermentation, depends on the production of ethanol, which results from the same chemical transformation. Some alcoholic beverages, such as champagne, can also be carbonated using the carbon dioxide produced by the yeast.

Yeast is among the simplest life forms on Earth, yet it is absolutely necessary for at least two major food industries. Without yeast to turn dough into bread and juice into wine, these foods and food industries would not exist today.



Figure 5.1.1: The bread-making industry depends on a chemical reaction performed by yeast. Several round, rustic loaves of bread with a golden-brown crust, lightly dusted with flour. The loaves have deep cracks on the surface, indicating a crisp crust and soft interior. (CC BY-SA 4.0; Tomas Castelazo via Wikimedia Commons).

Key Takeaway

Yeast helps bread rise by turning sugar into carbon dioxide and ethanol through a chemical reaction.

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5.2: The Law of Conservation of Matter

Understanding the Law of Conservation of Matter

In science, a **law** is a general statement that explains a large number of observations. Before being accepted, a law must be verified many times under many conditions. Laws are therefore considered the highest form of scientific knowledge and are generally thought to be inviolable. Scientific laws form the core of scientific knowledge.

One scientific law that provides the foundation for understanding in chemistry is the **law of conservation of matter**. It states that in any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant. A concise way of expressing this law is to say that the amount of matter in a system is *conserved*.

The law of conservation of matter, also referred to as the "law of conservation of mass," states that "In a chemical reaction, matter is neither created nor destroyed." This principle indicates that during any physical or chemical change, the total mass of the products is equal to the total mass of the reactants.

What does this mean for chemistry? In any chemical change, one or more initial substances change into a different substance or substances. Both the initial and final substances are composed of atoms because all matter is composed of atoms. According to the law of conservation of matter, matter is neither created nor destroyed, so we must have the same number and kind of atoms after the chemical change as were present before the chemical change.

$$\underbrace{\overbrace{\mathrm{HgO\,(s)}}^{\text{Mercuric oxide}}}_{100\text{ g}} \rightarrow \underbrace{\overbrace{\mathrm{Hg\,(l)}}^{\text{Mercury}}}_{92.6\text{ g}} + \underbrace{\overbrace{\mathrm{O_2\,(g)}}^{\text{Oxygen}}}_{7.4\text{ g}}$$

Example 5.2.1

If heating 10 grams of $\mathrm{CaCO_3}$ produces 4.4 g of $\mathrm{CO_2}$ and 5.6 g of CaO, show that these observations are in agreement with the law of conservation of mass.



Figure 5.2.1: A sample of calcium carbonate ($CaCO_3$). (Public domain; Walkerma via Wikimedia Commons).

Solution

- Mass of the reactants, $\mathrm{CaCO_3}$: $10\ g$
- Mass of the products, $\mathrm{CO_2}$ and CaO: $4.4\ g + 5.6\ g = 10\ g$.

Because the mass of the reactants = the mass of the products, the observations are in agreement with the law of conservation of mass

Key Takeaway

The amount of matter in a closed system is conserved.

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5.3: Chemical Equations

Conservation of Atoms in Chemical Reactions

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. In Figure 5.3.1 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.

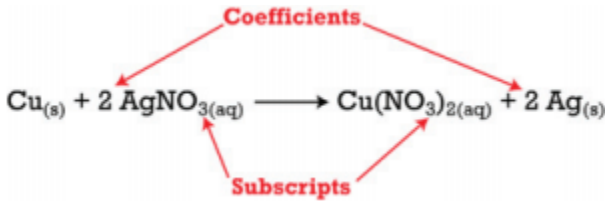


Figure 5.3.1: Chemical equation of solid copper and aqueous silver nitrate forming aqueous copper (II) nitrate and solid silver (CC BY-NC; via CK-12). Detailed Description of Figure 5.3.1.

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.

Water (H_2O) is composed of hydrogen and oxygen. Suppose we imagine a process in which we take some elemental hydrogen (H_2) and elemental oxygen (O_2) and let them react to make water. The statement

"hydrogen and oxygen react to make water"

is one way to represent that process, which is called a **chemical reaction**. Figure 5.3.2 shows a rather dramatic example of this very reaction.

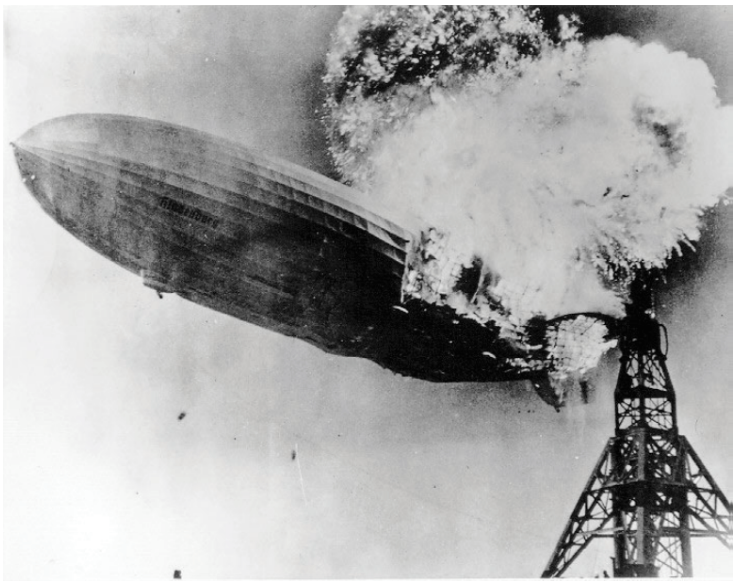


Figure 5.3.2: Zeppelin the Hindenburg on fire at the mooring mast of Lakehurst. The front of the airship is tilted downward while flames and smoke billow from the center. Debris falls from the explosion, and parts of the skeletal structure are visible through the fire. In this reaction, hydrogen gas in the zeppelin reacts with oxygen in the air to form water, while ballast water is released downward. (Public domain; Sam Shere (1905–1982) via Wikimedia Commons).

To simplify the writing of reactions, we use formulas instead of names when we describe a reaction. We can also use symbols to represent other words in the reaction. A plus sign connects the initial substances (and final substances, if there is more than one), and an arrow ($\rightarrow$) represents the chemical change, as shown in reaction 5.3.1:

$$\mathrm{H_2 + O_2 \ \rightarrow H_2O} \tag{5.3.1}$$

This statement is one example of a **chemical equation**, an abbreviated way of using symbols to represent a chemical change. The substances on the left side of the arrow are called **reactants**, and the substances on the right side of the arrow are called **products**. It is not uncommon to include a phase label with each formula—(s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for a substance dissolved in water, also known as an *aqueous solution*. If we included phase labels for the reactants and products, under normal environmental conditions, the reaction for 5.3.2 would be:

$$\mathrm{H_2\,(g) + O_2\,(g) \ \rightarrow H_2O\ (\ell)} \tag{5.3.2}$$

This equation is still not complete because *it does not satisfy the* law of conservation of matter. Count the number of atoms of each element on each side of the arrow. On the reactant side, there are two H atoms and two O atoms; on the product side, there are two H atoms and only one oxygen atom. Figure 5.3.2 shows that the equation is not balanced because the number of oxygen atoms on each side is not the same .

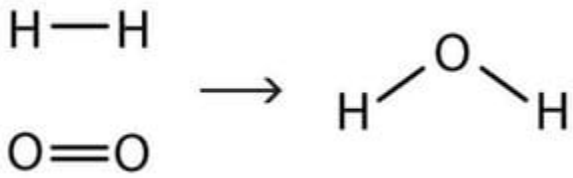


Figure 5.3.3: By counting the atoms of each element, we can see that the reaction is not balanced as written. (CC BY-NC-SA 3.0; Ball via Introduction to Chemistry: General, Organic, and Biological).

To make this chemical equation conform to the law of conservation of matter, we must revise the amounts of the reactants and the products as necessary to get the same number of atoms of a given element on each side. Because every substance has a characteristic chemical formula, we cannot change the chemical formulas of the individual substances. For example, we cannot change the formula for elemental oxygen to O. However, we can assume that different numbers of reactant molecules or product molecules may be involved. For instance, as shown in Figure 5.3.3, perhaps two water molecules are produced, not just one:

$$\mathrm{H_2\,(g) + O_2\,(g) \ \rightarrow 2\,H_2O\ (\ell)} \tag{5.3.3}$$

The 2 preceding the formula for water is called a coefficient. It implies that two water molecules are formed. There are now two oxygen atoms on each side of the equation.

By inserting the coefficient 2 in front of the formula for water, we have also changed the number of hydrogen atoms on the product side as well. As a result, we no longer have the same number of hydrogen atoms on each side. In reaction 5.3.4, this can be easily fixed by putting a coefficient of 2 in front of the diatomic hydrogen reactant:

$$2\,H_2\,(g) + O_2\,(g) \;\rightarrow 2\,H_2O\;(\ell) \tag{5.3.4}$$

Now we have four hydrogen atoms and two oxygen atoms on each side of the equation. The law of conservation of matter is satisfied because we now have the same number of atoms of each element in the reactants and in the products. We say that the reaction is **balanced** (Figure 5.3.3). The diatomic oxygen has a coefficient of 1, which typically is not written but assumed in balanced chemical equations.

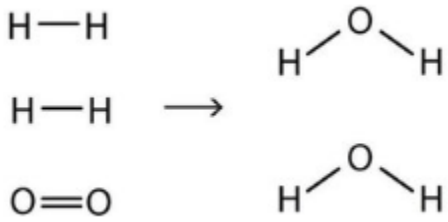


Figure 5.3.4: By counting the atoms of each element, we can see that the reaction is now balanced. (CC BY-NC-SA 3.0; Ball via Introduction to Chemistry: General, Organic, and Biological).

Proper chemical equations should be balanced. Writing balanced reactions is a chemist's way of acknowledging the law of conservation of matter.

Key Takeaways

- Chemical reactions are represented by chemical equations that list reactants and products.
- Proper chemical equations are balanced; the same number of each element's atoms appears on each side of the equation.

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5.4: How to Balance Chemical Equations

Guidelines for Writing and Balancing Chemical Equations

Balancing chemical equations is essential because it reflects the law of conservation of mass, ensuring that the number of atoms of each element is the same on both sides of the equation. This helps accurately represent how substances react and form new products. A balanced chemical equation shows the precise amounts of reactants needed to produce a specific amount of product. The process of balancing chemical equations involves a systematic approach: first, count the number of atoms of one element on one side of the equation, then check the number of atoms of that same element on the other side. If they do not match, adjust the coefficient as needed. Next, examine another element and repeat this back-and-forth process until all elements have the same number of atoms on both sides of the arrow. When balancing chemical equations, it is generally advisable to balance hydrogen (H) and oxygen (O) last. While the order of balancing elements does not usually affect the outcome, balancing H and O at the end is often more effective. This is because these elements are commonly found in multiple molecules on both sides of the equation, which allows for easier adjustments to the coefficients for the other elements first. Once those are balanced, you can finalize the equation by adjusting H and O.

Below are guidelines for writing and balancing chemical equations.

1. Determine the correct chemical formulas for each reactant and product. Write the skeleton equation.
2. Count the number of atoms of each element that appears as a reactant and as a product. If a polyatomic ion is unchanged on both sides of the equation, count it as a unit.
3. Balance each element one at a time by placing coefficients in front of the formulas. No coefficient is written for a 1. It is best to begin by balancing elements that only appear in one chemical formula on each side of the equation. NEVER change the subscripts in a chemical formula - you can only balance equations by using coefficients.
4. Check each atom or polyatomic ion to be sure that they are equal on both sides of the equation.
5. Make sure that all coefficients are in the lowest possible ratio. If necessary, reduce to the lowest ratio.

For example, to balance the equation

Step 1

Write the skeleton equation with the correct formulas for Reaction 5.4.1:

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + HCl} \tag{5.4.1}$$

Step 2

Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 C atom	1 C atom
4 H atoms	1 H atoms
2 Cl atom	5 Cl atoms

Step 3

In reaction 5.4.2, we find that both sides are already balanced with one carbon atom. So we proceed to balance the hydrogen atoms. We find that the reactant side has four hydrogen atoms, so the product side must also have four hydrogen atoms. This is balanced by putting a 4 in front of the HCl:

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{5.4.2}$$

Reactants	Products
1 C atom	1 C atom
4 H atoms	4 H atoms
2 Cl atom	8 Cl atoms

Now, in reaction 5.4.3, each side has four hydrogen atoms. The product side has a total of eight chlorine atoms (four from the CCl_4 and four from the four molecules of HCl), so we need eight chlorine atoms as reactants. Because elemental chlorine is a diatomic molecule, we need four chlorine molecules to get a total of eight chlorine atoms. We add another 4 in front of the Cl_2 reactant:

$$\mathrm{CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{5.4.3}$$

Reactants	Products
1 C atom	1 C atom
4 H atoms	4 H atoms
8 Cl atom	8 Cl atoms

Step 4

Now, we check: each side has one carbon atom, four hydrogen atoms, and eight chlorine atoms. The chemical equation is balanced. And, the coefficients are in the lowest possible ratio.

Example 5.4.1

Is each chemical equation balanced?

a. $2Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
b. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$
c. $AgNO_3(aq) + 2KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

Solution

a. By counting, we find two sodium atoms and two oxygen atoms in the reactants and four sodium atoms and two oxygen atoms in the products. This equation is not balanced.
b. The reactants have one carbon atom, four hydrogen atoms, and four oxygen atoms. The products have one carbon atom, four hydrogen atoms, and four oxygen atoms. This equation is balanced.
c. The reactants have one silver atom, one nitrogen atom, three oxygen atoms, two potassium atoms, and two chlorine atoms. The products have one silver atom, one chlorine atom, one potassium atom, one nitrogen atom, and three oxygen atoms. Because there are different numbers of chlorine and potassium atoms, this equation is not balanced.

Building on the previous example, the next one further explains this concept in detail.

Example 5.4.2

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

Step 1: Write the skeleton equation with the correct formulas.

$$\mathrm{Pb(NO_3)_2\,(aq) + NaCl\,(aq) \rightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$$

Step 2: Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 Pb atom	1 Pb atom
2 NO_3^- ions	1 NO_3^- ions
1 Na atom	1 Na atom
1 Cl atom	2 Cl atoms

Step 3: Solve. The nitrate ions and the chlorine atoms are unbalanced. Start by placing a 2 in front of the $NaCl$. This increases the reactant counts to 2 Na atoms and 2 Cl atoms. Then place a 2 in front of the $NaNO_3$. The result is:

$$Pb(NO_3)_2\,(aq) + 2NaCl\,(aq) \rightarrow 2NaNO_3\,(aq) + PbCl_2\,(s)$$

Step 4: The new count for each atom and polyatomic ion becomes:

Reactants	Products
1 Pb atom	1 Pb atom
2 NO_3^- ions	2 NO_3^- ions
2 Na atoms	2 Na atoms
2 Cl atoms	2 Cl atoms

Step 5: Think about the result.

The equation is now balanced since there are equal numbers of atoms of each element on both sides of the equation. And, the coefficients are in the lowest possible ratio

Key Takeaways

- Coefficients are used to balance chemical equations, but chemical formulas must not be altered.
- Begin balancing with elements that appear in only one compound on both sides of the equation.
- Elements in their pure form (elemental state) should be balanced last.
- Always verify the balanced equation by counting each type of atom on both the reactant and product sides.

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5.5: Quantitative Relationships Based on Chemical Equations

Understanding Stoichiometry

Analyzing the numerical relationships in chemical equations allows us to predict the precise amounts of reactants consumed and products formed in a chemical reaction. A balanced chemical equation not only describes some of the chemical properties of substances—by showing us what substances react with what other substances to make what products—but also shows numerical relationships between the reactants and the products. The study of these numerical relationships is called **stoichiometry**. The stoichiometry of chemical equations revolves around the coefficients in the balanced chemical equation because these coefficients determine the molecular ratio in which reactants react and products are made.

One way to understand chemical reactions is to think of them as "recipes" (Figure $5.5.1$). Just like recipes, a chemical equation specifies all the ingredients and their relative amounts, as well as the quantity of the product produced. Consider a stoichiometry analogy from the kitchen. A recipe for 1 dozen biscuits requires 2 cups of flour, 1 egg, 4 tablespoons of shortening, 1 teaspoon of salt, 1 teaspoon of baking soda, and 1 cup of milk. If this were expressed as a chemical equation, it would be written as:

$$\text{2 c flour + 1 egg + 4 tbsp shortening + 1 tsp salt + 1 tsp baking soda + 1 c milk} \rightarrow \text{12 biscuits}$$

This equation gives ratios of how much of what reactants are needed to make how much of what product. Two cups of flour, when combined with the proper amounts of the other ingredients, will yield 12 biscuits. One teaspoon of baking soda (when also combined with the right amounts of the other ingredients) will make 12 biscuits. One egg must be combined with 1 cup of milk to yield the product food. Other relationships can also be expressed.

Ratios derived from the equation can be used for predictive purposes. For example, if 4 cups of flour are available and there is enough of the other ingredients, how many biscuits can be made? In this case, a double recipe of 24 biscuits can be made.

To determine this answer mathematically, a conversion factor can be used, similar to the method introduced in Chapter 1. Since 2 cups of flour produce 12 biscuits, an equivalency ratio can be established:

$$\frac{12 \text{ biscuits}}{2 \text{ c flour}}$$

This ratio can then be used to formally convert flour to biscuits:

$$4 \text{ c flour} \times \frac{12 \text{ biscuits}}{2 \text{ c flour}} = 24 \text{ biscuits}$$

In the same way, by constructing similar ratios, it is possible to determine how many biscuits can be made from any given amount of an ingredient. Doubling or halving a recipe is a practical example of this type of calculation. Applying these principles to chemical reactions follows a similar process.



Figure 5.5.1: The Science of Proportions. Just like a recipe shows how much of each ingredient is needed for the proper reaction to take place, a balanced chemical equation specifies the exact proportions of reactants required to produce the desired products in a chemical reaction. (CC-BY-SA-3.0; Peng via Wikimedia Commons). Detailed Description of Figure 5.5.1.

Consider the following balanced chemical equation:

$$2\,C_2H_2 + 5\,O_2 \rightarrow 4\,CO_2 + 2\,H_2O \tag{5.5.1}$$

The coefficients on the chemical formulas give the ratios in which the reactants combine and the products form. Table 5.5.1 takes a statement from the balanced chemical reaction to generate the following statements and construct the corresponding ratios:

Table 5.5.1: Ratios and inverse ratios from the balanced chemical reaction.

Statement from the Balanced Chemical Reaction	Ratio	Inverse Ratio
two C_2H_2 molecules react with five O_2 molecules	$\frac{2\ C_2H_2\ \text{molecules}}{5\ O_2\ \text{molecules}}$	$\frac{5\ O_2\ \text{molecules}}{2\ C_2H_2\ \text{molecules}}$
two C_2H_2 molecules react to make four CO_2 molecules	$\frac{2\ C_2H_2\ \text{molecules}}{4\ CO_2\ \text{molecules}}$	$\frac{4\ CO_2\ \text{molecules}}{2\ C_2H_2\ \text{molecules}}$
five O_2 molecules react to make two H_2O molecules	$\frac{5\ O_2\ \text{molecules}}{2\ H_2O\ \text{molecules}}$	$\frac{2\ H_2O\ \text{molecules}}{5\ O_2\ \text{molecules}}$
four CO_2 molecules are made at the same time as two H_2O molecules	$\frac{2\ H_2O\ \text{molecules}}{4\ CO_2\ \text{molecules}}$	$\frac{4\ CO_2\ \text{molecules}}{2\ H_2O\ \text{molecules}}$

Other relationships are possible; in fact, 12 different conversion factors can be constructed from this balanced chemical equation. In each ratio, the unit is assumed to be molecules because that is how we are interpreting the chemical equation.

Any of these fractions can be used as a conversion factor to relate an amount of one substance to an amount of another substance. For example, suppose we want to know how many CO_2 molecules are formed when 26 molecules of C_2H_2 are reacted. As usual with a conversion problem, we start with the amount we are given—26C_2H_2—and multiply it by a conversion factor that cancels out our original unit and introduces the unit we are converting to—in this case, CO_2. That conversion factor is $\frac{4CO_2}{2C_2H_2}$, which is composed of terms that come directly from the balanced chemical equation. Thus, we have

$$26\ C_2H_2\ \text{molecules} \times \frac{4\ CO_2\ \text{molecules}}{2\ C_2H_2\ \text{molecules}}$$

The molecules of C_2H_2 cancel, and we are left with molecules of CO_2. Multiplying through, we get

$$26\ C_2H_2\ \text{molecules} \times \frac{4\ CO_2\ \text{molecules}}{2\ C_2H_2\ \text{molecules}} = 52\ CO_2\ \text{molecules}$$

Thus, 52 molecules of CO_2 are formed.

This application of stoichiometry is extremely powerful in its predictive ability, as long as we begin with a balanced chemical equation. Without a balanced chemical equation, the predictions made by simple stoichiometric calculations will be incorrect.

Example 5.5.1

Start with this balanced chemical equation.

$$KMnO_4 + 8\,HCl + 5\,FeCl_2 \longrightarrow 5\,FeCl_3 + MnCl_2 + 4\,H_2O + KCl$$

1. Verify that the equation is indeed balanced.
2. Give 2 ratios that give the relationship between HCl and $FeCl_3$.

Solution

1. Each side has 1 K atom and 1 Mn atom. The 8 molecules of HCl yield 8 H atoms, and the 4 molecules of H_2O also yield 8 H atoms, so the H atoms are balanced. The Fe atoms are balanced, as we count 5 Fe atoms from 5 $FeCl_2$ reactants and 5 $FeCl_3$ products. As for Cl, on the reactant side, there are 8 Cl atoms from HCl and 10 Cl atoms from the 5 $FeCl_2$ formula units, for a total of 18 Cl atoms. On the product side, there are 15 Cl atoms from the 5 $FeCl_3$ formula units, 2 from the $MnCl_2$ formula unit, and 1 from the KCl formula unit. This is a total of 18 Cl atoms in the products, so the Cl atoms are balanced. All the elements are balanced, so the entire chemical equation is balanced.
2. Because the balanced chemical equation tells us that 8 HCl molecules react to make 5 $FeCl_3$ formula units, we have the following 2 ratios: $\frac{8\text{ HCl molecules}}{5\text{ FeCl}_3\text{ formula units}}$ and $\frac{5\text{ FeCl}_3\text{ formula units}}{8\text{ HCl molecules}}$. There are a total of 42 possible ratios. Can you find the other 40 relationships?

Key Takeaway

A balanced chemical equation gives the ratios in which molecules of substances react and are produced in a chemical reaction.

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5.6: Summary

Scientific laws are general statements that apply to a wide variety of circumstances. One important law in chemistry is the law of conservation of matter, which states that in any closed system, the amount of matter stays constant.

Chemical equations are used to represent chemical reactions. Reactants change chemically into products. The law of conservation of matter requires that a proper chemical equation be balanced. Coefficients are used to show the relative numbers of reactant and product molecules.

In stoichiometry, quantities of reactants and/or products can be related to each other using the balanced chemical equation. The coefficients in a balanced chemical reaction are used to devise the proper ratios that relate the number of molecules of one substance to the number of molecules of another substance.

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5.7: Key Terms

- **Law:** A general statement that explains a large number of observations.
- **Law of Conservation of Matter:** In any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant.
- **Chemical reaction**: A representation of a chemical change.
- **Chemical equation**: An abbreviated way of using symbols to represent a chemical change.
- **Reactants**: A substance on the left side of the arrow in a chemical equation.
- **Products**: A substance on the right side of the arrow in a chemical equation.
- **Subscript:** A subscript is a small number written slightly below and to the right of a chemical symbol in a formula. It indicates the number of atoms of that element in a molecule.
- **Coefficient**: A number that gives the number of molecules of a substance in a balanced chemical equation.
- **Balanced**: A property of a chemical equation when there are the same number of atoms of each element in the reactants and products.
- **Stoichiometry**: The study of the numerical relationships between the reactants and the products in a balanced chemical equation.

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5.8: Alternative text for Chapter 5

Figure 5.3.1: Coefficients and subscripts.

Coefficients are the numbers before a chemical formula (e.g., "2" in "$2AgNO_3$") represents the coefficient. Coefficients indicate the number of molecules or formula units involved in the reaction. The arrows point to the coefficients in $2AgNO_3$ and 2Ag. Subscripts are the small numbers within a chemical formula (e.g., "3" in "NO_3" and "2" in "$Cu(NO_3)_2$") represent subscripts. Subscripts indicate the number of atoms of each element within a single molecule or formula unit. The arrows highlight the subscripts in NO_3 and $Cu(NO_3)_2$. Coefficients tell us how many molecules or units participate in the reaction. Subscripts indicate the number of atoms in each molecule.

Figure 5.5.1: Chemical Reactions and Recipes - The Science of Proportions.

The image depicts a kitchen counter set up for baking, featuring various ingredients and tools. On the counter, there is a mixing bowl with a liquid mixture, a measuring cup, a box of baking soda or powder, a bottle of liquid (possibly vanilla extract or cooking oil), a bag of flour, and a carton of eggs. An open cookbook with illustrated recipes outlining ingredient quantities is also visible, suggesting preparation for a baking activity.

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6: CLASSIFICATION OF CHEMICAL REACTIONS

Introduction

While there are millions of possible chemical reactions, most can be grouped into a few fundamental reaction types. Classifying reactions serves two key purposes: identifying similarities between different reactions and predicting reaction products based on known patterns. Some reactions may fit into multiple categories, highlighting the interconnected nature of chemical processes. Understanding these classifications provides a foundation for analyzing and predicting chemical behavior.

Learning Objectives

1. Use solubility rules to predict whether a precipitate will form and write and balance chemical equations for precipitation reactions.
2. Write and balance chemical equations in molecular, total ionic, and net ionic formats.
3. Identify an acid-base neutralization reaction and predict its products.
4. Assign oxidation numbers to atoms in simple compounds and recognize a reaction as an oxidation-reduction reaction.
5. To identify oxidation-reduction reactions with organic compounds.

6.1: Introduction to Classifying Chemical Reactions

6.2: Precipitation Reactions and Solubility Guidelines

6.3: Acids, Bases, and Neutralization Reactions

6.4: Oxidation-Reduction (Redox) Reactions

6.5: Recognizing Redox Reactions

6.6: Redox Reactions in Organic Chemistry and Biochemistry

6.7: Summary

6.8: Key Terms

6.9: Alternative text for Chapter 6

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6.1: Introduction to Classifying Chemical Reactions

Classifying Chemical Reactions Based on Reaction Processes

Chemical reactions are categorized into different types to help scientists better understand their behavior and predict the products they will yield. These reactions can be grouped into three main types based on the specific chemical processes involved: precipitation reactions, acid-base neutralization reactions, and redox reactions.

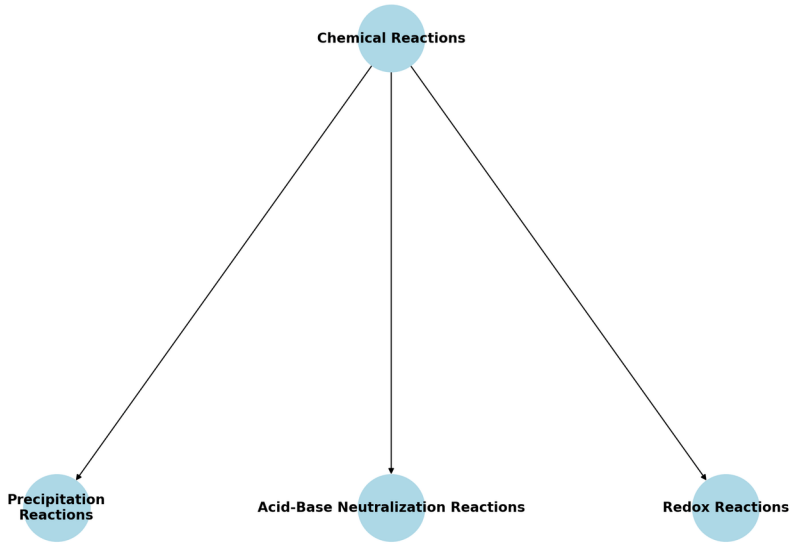


Figure 6.1.1: Classification of Chemical Reaction. This flowchart was generated using ChatGPT tools based on user input and is not subject to copyright restrictions. Free to use and distribute with attribution to ChatGPT. (OpenAI. (2024) ChatGTP (GPT-4, Jan 8 version). Detailed Description of Figure 6.1.1.

Understanding Chemical Reactions Through Atomic Behavior

Another way to classify chemical reactions is by examining the reaction patterns, such as synthesis (or combination), decomposition, single displacement, and double displacement. In this classification, the letters (A, B, C, D) represent atoms or groups of atoms.

Table 6.1.1: Classification of Chemical Reactions through atomic behaviour

Type of Reaction	Generic Equation
synthesis (or combination)	A+B → AB
decomposition	AB → A+B
single displacement	A+BC → AC + B
double displacement	AB+CD → AD + CB

Synthesis (Combination) Reactions

Synthesis (combination) reaction is a chemical reaction that makes a single substance from two or more reactants. In the balanced chemical equation, reactants combine to give only one product (Equation 6.1.1).

$$4\,\mathrm{Fe(s)} + 3\,\mathrm{O_2(g)} \rightarrow 2\,\mathrm{Fe_2O_3(s)} \tag{6.1.1}$$

Decomposition Reactions

A **decomposition reaction** is the reverse of a combination reaction. In a decomposition reaction, The reactant is broken down into two or more smaller substances. There is only one substance initially (Equation 6.1.2).

$$2\,\mathrm{NH_3\,(g)} \rightarrow \mathrm{N_2\,(g)} + 3\,\mathrm{H_2\,(g)} \tag{6.1.2}$$

Single-Displacement Reactions

A **single displacement reaction** is a reaction in which one element replaces another element in a compound (Equation 6.1.3).

$$\mathrm{Br_2\,(g)} + 2\,\mathrm{KI\,(aq)} \rightarrow \mathrm{I_2\,(g)} + 2\,\mathrm{KBr\,(aq)} \tag{6.1.3}$$

Double-Displacement Reactions

A **double displacement reaction** is a reaction in which the first elements of two compounds are exchanged to produce two different compounds (Equation 6.1.4, Equation 6.1.5). NOTE: Precipitation or acid-base neutralization reactions are Double displacement reactions.

$$2\,\mathrm{AgNO_3\,(aq)} + \mathrm{BaCl_2\,(aq)} \rightarrow 2\,\mathrm{AgCl\,(s)} + \mathrm{Ba(NO_3)_2\,(aq)} \tag{6.1.4}$$

$$\mathrm{NaOH\,(aq)} + \mathrm{HNO_3\,(aq)} \rightarrow \mathrm{H_2O(\ell)} + \mathrm{NaNO_3\,(aq)} \tag{6.1.5}$$

Key Takeaways

- Chemical reactions are classified into three main types: precipitation, acid-base neutralization, and redox reactions.
- Reactions can also be categorized by atomic behaviour, including synthesis, decomposition, single displacement, and double displacement.

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6.2: Precipitation Reactions and Solubility Guidelines

Overview

There are many types of chemical reactions that you will learn about in chemistry. In this chapter, we will focus on various types of chemical reactions, **precipitation reactions** (a type of double-displacement reaction), **acid-base neutralization reactions**, and **oxidation-reduction reactions**.

The practice of barter (trading one thing for another) has been in existence since the beginning of time. In the past, for example, items like chickens were bartered for newspapers. Person A had something that person B wanted, and vice versa. So, when person A and person B traded items, they each had something new. Some chemical reactions are like that—compounds swap parts, and the products are new materials.

A double-displacement 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-displacement (also called double-replacement) reaction is shown below (Reaction $6.2.1$)):

$$\mathrm{AB} + \mathrm{CD} \rightarrow \mathrm{AD} + \mathrm{CB} \tag{6.2.1}$$

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 (Reaction 6.2.2):

$$2\mathrm{KI}\,(aq) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow 2\mathrm{KNO_3}\,(aq) + \mathrm{PbI_2}\,(s) \tag{6.2.2}$$

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 $6.2.1$). The other product of the reaction, potassium nitrate, remains soluble.



Figure $6.2.1$: Formation of lead iodide precipitate. (CC BY 2.0; Paige Powers via Flickr). Detailed Description of Figure $6.2.1$.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble, or insoluble). Table 6.2.1 lists some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules. If a compound is soluble, we use the (*aq*) label with it, indicating that it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

Table 6.2.1: Solubility Rules: Soluble Compounds and Their Exceptions

These compounds generally dissolve in water (are soluble):	**Exceptions:**
All compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+	None
All compounds of NO_3^- and $C_2H_3O_2^-$	None
Compounds of Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds of SO_4^{2-}	Hg_2^{2+}, Pb^{2+}, Sr^{2+}, Ba^{2+}

Table 6.2.2: Solubility Rules: Insoluble Compounds and Their Exceptions

These compounds generally do not dissolve in water (are insoluble):	**Exceptions:**
Compounds of CO_3^{2-} and PO_4^{3-}	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+
Compounds of OH^-	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, Sr^{2+}, and Ba^{2+}

For example, consider the possible double-replacement reaction between Na_2SO_4 and $SrCl_2$. The solubility rules say that all ionic sodium compounds are soluble and all ionic chloride compounds are soluble, except for Ag^+, Hg_2^{2+}, and Pb^{2+}, which are not being considered here. Therefore, Na_2SO_4 and $SrCl_2$ are both soluble. The possible double-replacement reaction products are NaCl and $SrSO_4$. Are these soluble? NaCl is (by the same rule we just quoted), but what about $SrSO_4$? Compounds of the sulfate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:

$$\mathrm{Na_2SO_4\,(aq) + SrCl_2\,(aq) \longrightarrow 2\,NaCl\,(aq) + SrSO_4\,(s)} \tag{6.2.3}$$

You would expect to see a visual change corresponding to $SrSO_4$ precipitating out of solution (Reaction 6.2.3).

Equations for Ionic Reactions

Many chemical reactions involving ionic compounds occur in aqueous solutions.

When ions participate in reactions, chemical equations can be expressed at different levels of detail, including molecular, total ionic, and net ionic equations, to better represent the reaction process.

Molecular equations don't explicitly represent the ionic species that are present in the solution (6.2.4):

$$\mathrm{2\,AgNO_3\,(aq) + CaCl_2\,(aq) \rightarrow 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)} \tag{6.2.4}$$

When ionic compounds dissolve in water, they may dissociate into their individual ions, rather than remaining as neutral compounds. Representing these dissociated ions explicitly results in a **total ionic equation**, which more accurately reflects the reaction in solution (6.2.5):

$$2\,\mathrm{Ag^+}_{(aq)} + 2\,\mathrm{NO_3^-}_{(aq)} + \mathrm{Ca^{2+}}_{(aq)} \rightarrow \mathrm{Ca^{2+}}_{(aq)} + 2\,\mathrm{NO_3^-}_{(aq)} + 2\,\mathrm{AgCl(s)} \quad (6.2.5)$$

Some ions appear unchanged on both sides of the equation—these are called **spectator ions**. Their role is to maintain charge neutrality, but they do not chemically or physically change during the reaction. Elimination of the spectator ions results in a **net ionic equation** (6.2.6):

$$\mathrm{Ag^+}_{(aq)} + \mathrm{Cl^-}_{(aq)} \rightarrow \mathrm{AgCl(s)} \quad (6.2.6)$$

Example 6.2.1:

Will a precipitation reaction occur? If so, identify the products.

1. $Ca(NO_3)_2$ + KBr → ?
2. NaOH + $FeCl_2$ → ?

Solution

1. According to the solubility rules, both $Ca(NO_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)—namely, $CaBr_2$ and KNO_3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and $FeCl_2$ are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $Fe(OH)_2$. NaCl is soluble, but, according to the solubility rules, $Fe(OH)_2$ is not. Therefore, a reaction would occur, and $Fe(OH)_2$(s) would precipitate out of solution. The balanced chemical equation is

$$2\,\mathrm{NaOH\,(aq)} + \mathrm{FeCl_2\,(aq)} \longrightarrow 2\,\mathrm{NaCl\,(aq)} + \mathrm{Fe(OH)_2\,(s)}$$

Key Takeaways

- Precipitation reactions occur when two aqueous ionic compounds react to form an insoluble solid (precipitate), which can be predicted using solubility rules.
- Double-replacement reactions involve the exchange of ions between two reactants, forming new products, one of which may be a precipitate, gas, or molecular compound.
- Ionic equations can be written at different levels of detail, including molecular equations, total ionic equations, and net ionic equations, to better represent the reaction.
- Spectator ions do not participate in the chemical reaction and remain unchanged on both sides of the equation; removing them results in the net ionic equation.

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6.3: Acids, Bases, and Neutralization Reactions

Understanding Acids, Bases, and Neutralization Reactions

Acid is any compound that produces hydrogen ions (H^+) in an aqueous solution, and the chemical opposite of an acid is a base, which is a compound that produces hydroxide ions (OH^-) in an aqueous solution. These original definitions were proposed by Arrhenius (the same person who proposed ion dissociation) in 1884, so they are referred to as the **Arrhenius definition** of an acid and a base, respectively.

You may recognize that, based on the description of a hydrogen atom, an H^+ ion is a hydrogen atom that has lost its lone electron; that is, H^+ is simply a proton. Do we really have bare protons moving about in aqueous solution? No. What is more likely is that the H^+ ion has attached itself to one (or more) water molecule(s). To represent this chemically, we define the **hydronium ion** as H_3O^+, which represents an additional proton attached to a water molecule. We use the hydronium ion as the more logical way that a hydrogen ion appears in an aqueous solution, although in many chemical reactions H^+ and H_3O^+ are treated equivalently.

The reaction of an acid and a base is called a **neutralization reaction**. Although acids and bases have their own unique chemistries, the acid and base "cancel" each other's chemistry to produce a rather innocuous substance, water. In fact, the general reaction between an acid and a base is as shown in Reaction $6.3.1$:

$$\text{acid} + \text{base} \longrightarrow \text{water} + \text{salt} \tag{6.3.1}$$

where the term **salt** is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. (In chemistry, the word *salt* refers to more than just table salt.) For example, the balanced chemical equation for the Reaction $6.3.2$, between HCl(aq) and KOH(aq) is:

$$\mathrm{HCl\,(aq)} + \mathrm{KOH\,(aq)} \longrightarrow \mathrm{H_2O(\ell)} + \mathrm{KCl\,(aq)} \tag{6.3.2}$$

where the salt is KCl. By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between HCl(aq) and $Mg(OH)_2$(aq), additional molecules of HCl and H_2O are required to balance the chemical equation for the Reaction $6.3.3$:

$$2\,\mathrm{HCl\,(aq)} + \mathrm{Mg(OH)_2\,(aq)} \longrightarrow 2\,\mathrm{H_2O(\ell)} + \mathrm{MgCl_2\,(aq)} \tag{6.3.3}$$

Here, the salt is $MgCl_2$. (This is one of several reactions that take place when a type of antacid, a base, is used to treat stomach acid.)

Example $6.3.1$:

Write the neutralization reactions between each acid and base.

a. HNO_3(aq) and $Ba(OH)_2$(aq)
b. H_3PO_4(aq) and $Ca(OH)_2$(aq)

Solution

First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation.

a. The expected products are water and barium nitrate, so the initial chemical reaction is

$$\mathrm{HNO_3\,(aq)} + \mathrm{Ba(OH)_2\,(aq)} \longrightarrow \mathrm{H_2O(\ell)} + \mathrm{Ba(NO_3)_2\,(aq)}$$

To balance the equation, we need to realize that there will be two H_2O molecules, so two HNO_3 molecules are required:

$$2\,HNO_3\,(aq) + Ba(OH)_2\,(aq) \longrightarrow 2\,H_2O(\ell) + Ba(NO_3)_2\,(aq)$$

This chemical equation is now balanced.

b. The expected products are water and calcium phosphate, so the initial chemical equation is

$$H_3PO_4\,(aq) + Ca(OH)_2\,(aq) \longrightarrow H_2O(\ell) + Ca_3(PO_4)_2\,(s)$$

According to the solubility rules, $Ca_3(PO_4)_2$ is insoluble, so it has an (s) phase label. To balance this equation, we need two phosphate ions and three calcium ions. We end up with six water molecules to balance the equation:

$$2\,H_3PO_4\,(aq) + 3\,Ca(OH)_2\,(aq) \longrightarrow 6\,H_2O(\ell) + Ca_3(PO_4)_2\,(s)$$

This chemical equation is now balanced.

Neutralization reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical Reaction $6.3.5$, between HCl(aq) and $Fe(OH)_3$(s) still proceeds according to the equation:

$$3\,HCl\,(aq) + Fe(OH)_3\,(s) \longrightarrow 3\,H_2O(\ell) + FeCl_3\,(aq) \tag{6.3.4}$$

even though $Fe(OH)_3$ is not soluble. When one realizes that $Fe(OH)_3$(s) is a component of rust, this explains why some cleaning solutions for rust stains contain acids; the neutralization reaction produces products that are soluble and wash away. (Washing with acids like HCl is one way to remove rust and rust stains, but HCl must be used with caution!)

Key Takeaways

- Acids produce hydrogen ions in aqueous solutions, while bases produce hydroxide ions, according to the Arrhenius definition.
- A neutralization reaction occurs when an acid reacts with a base, producing water and a salt.
- Salts formed in acid-base reactions can be soluble or insoluble and are commonly found in everyday applications, such as antacids treating stomach acid.

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6.4: Oxidation-Reduction (Redox) Reactions

Understanding Oxidation-Reduction (Redox) Reactions

When zinc metal is submerged into a quantity of aqueous HCl, the following reaction occurs (Figure $6.4.1$):

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow H_2\,(g) + ZnCl_2\,(aq)} \tag{6.4.1}$$

This is one example of what is sometimes called a *single displacement reaction* because Zn replaces H in combination with Cl.

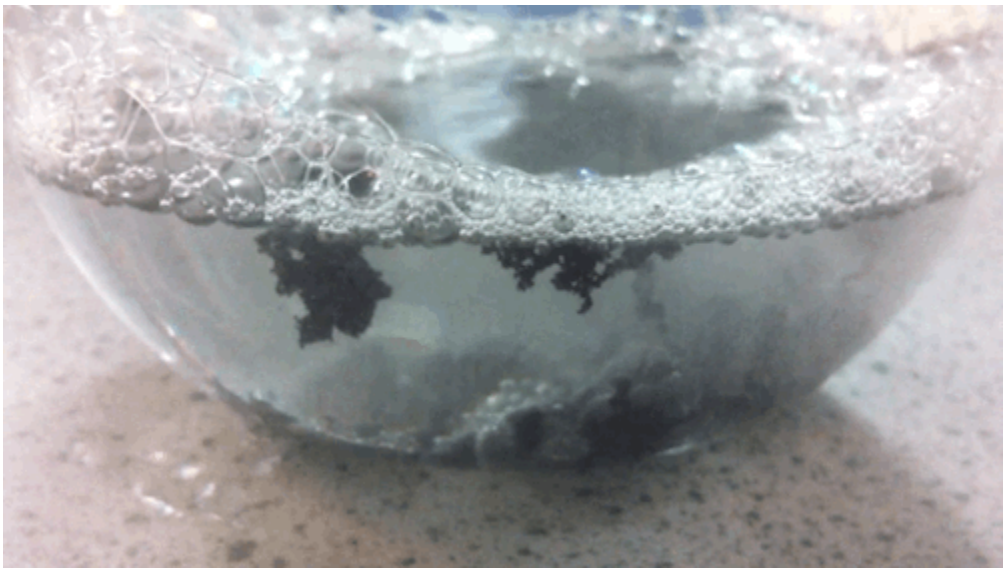


Figure $6.4.1$: Reaction between zinc metal and aqueous hydrochloric acid, demonstrating the release of hydrogen gas as visible bubbles. (CC BY-NC-SA 3.0; Lizeth Guadalupe Cervantes via Introductory, Conceptual, and GOB Chemistry). Detailed Description of Figure $6.4.1$.

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow H_2\,(g) + Zn^{2+}\,(aq) + 2\,Cl^-\,(aq)} \tag{6.4.2}$$

Viewed this way, the net reaction seems to be a charge transfer between zinc and hydrogen atoms. (There is no net change experienced by the chloride ion.) In fact, electrons are being transferred from the zinc atoms to the hydrogen atoms (which ultimately make a molecule of diatomic hydrogen), changing the charges on both elements.

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called **oxidation**. The gain of electrons is called **reduction**. Because any loss of electrons by one substance must be accompanied by a gain in electrons by something else, oxidation and reduction always occur together. As such, electron-transfer reactions are also called **oxidation-reduction reactions** or simply **redox reactions**. The atom that loses electrons is oxidized, and the atom that gains electrons is reduced. Also, because we can think of the species being oxidized as causing the reduction, the species being oxidized is called the **reducing agent**, and the species being reduced is called the **oxidizing agent**.

Although the two reactions occur together, it can be helpful to write the oxidation and reduction reactions separately as **half reactions**. In half reactions, we include only the reactant being oxidized or reduced, the corresponding product species, any other species needed to balance the half reaction, and the electrons being transferred. Electrons that are lost are written as products; electrons that are gained are written as reactants. For example, consider the following equation, written without the chloride ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)} \tag{6.4.3}$$

zinc atoms are oxidized to $\mathrm{Zn^{2+}}$. The half reaction for the oxidation reaction, omitting phase labels, is as follows:

$$\mathrm{Zn} \longrightarrow \mathrm{Zn}^{2+} + 2\,\mathrm{e}^- \tag{6.4.4}$$

This half reaction is balanced in terms of the number of zinc atoms, and it also shows the two electrons that are needed as products to account for the zinc atom losing two negative charges to become a 2+ ion. With half reactions, there is one more item to balance: the overall charge on each side of the reaction. In this case, both sides of the reaction have a net charge of zero.

Hydrogen is reduced in the reaction. The balanced reduction half reaction is as follows:

$$2\,\mathrm{H}^+ + 2\,\mathrm{e}^- \longrightarrow \mathrm{H}_2 \tag{6.4.5}$$

There are two hydrogen atoms on each side, and the two electrons written as reactants serve to neutralize the 2+ charge on the reactant hydrogen ions. Again, the overall charge on both sides is zero.

The overall reaction, as presented in section 6.4.6, is obtained by combining the two half-reactions through addition:

$$\begin{aligned} \mathrm{Zn} &\longrightarrow \mathrm{Zn}^{2+} + \cancel{2e^-} \\ \underline{2H^+ + \cancel{2e^-} \rightarrow H_2} \quad & \\ \mathrm{Zn} + 2\mathrm{H}^+ &\rightarrow \mathrm{Zn}^{2+} + \mathrm{H2} \end{aligned} \tag{6.4.6}$$

Because we have two electrons on each side of the equation, they can be canceled. This is the key criterion for a balanced redox reaction: the electrons have to cancel exactly. If we check the charge on both sides of the equation, we see they are the same—2+. (In reality, this positive charge is balanced by the negative charges of the chloride ions, which are not included in this reaction because chlorine does not participate in the charge transfer.)

Redox reactions are often balanced by balancing each individual half reaction and then combining the two balanced half reactions. Sometimes a half reaction must have all of its coefficients multiplied by some integer for all the electrons to cancel. The following example demonstrates this process.

Example 6.4.1: Reducing Silver Ions

Write and balance the redox reaction that has silver ions and aluminum metal as reactants and silver metal and aluminum ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Solution

We start by using symbols of the elements and ions to represent the reaction:

$$\mathrm{Ag}^+ + \mathrm{Al} \longrightarrow \mathrm{Ag} + \mathrm{Al}^{3+}$$

The equation looks balanced as it is written. However, when we compare the overall charges on each side of the equation, we find a charge of +1 on the left but a charge of +3 on the right. This equation is not properly balanced. To balance it, let us write the two half reactions. Silver ions are reduced, and it takes one electron to change Ag^+ to Ag:

Reduction half-reaction:

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

Aluminum is oxidized, losing three electrons to change from Al to Al^{3+}:

Oxidation half-reaction:

$$\mathrm{Al} \longrightarrow \mathrm{Al}^{3+} + 3\,\mathrm{e}^-$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{aligned} 3\,(\mathrm{Ag}^+ + \mathrm{e}^-) &\longrightarrow 3\,\mathrm{Ag} \\ \mathrm{Al} &\longrightarrow \mathrm{Al}^{3+} + 3\,\mathrm{e}^- \\ 3\,\mathrm{Ag}^+ + \mathrm{Al} &\longrightarrow 3\,\mathrm{Ag} + \mathrm{Al}^{3+} \end{aligned} \tag{6.4.7}$$

Now the equation is balanced, not only in terms of elements but also in terms of charge.

- The substance oxidized is the reactant that had undergone oxidation: Al
- The substance reduced is the reactant that had undergone reduction: Ag^+
- The reducing agent is the same as the substance oxidized: Al
- The oxidizing agent is the same as the substance reduced: Ag^+

Link to Learning: Redox Reactions and Pacemaker Batteries

All batteries use redox reactions to supply electricity because electricity is basically a stream of electrons being transferred from one substance to another. Pacemakers—surgically implanted devices for regulating a person's heartbeat—are powered by tiny batteries, so the proper operation of a pacemaker depends on a redox reaction.

Pacemakers used to be powered by NiCad batteries, in which nickel and cadmium (hence the name of the battery) react with water according to this redox reaction:

$$\mathrm{Cd\,(s) + 2\,NiOOH\,(s) + 2\,H_2O(\ell) \longrightarrow Cd(OH)_2\,(s) + 2\,Ni(OH)\ 2\,(s)}$$

The cadmium is oxidized, while the nickel atoms in NiOOH are reduced. Except for the water, all the substances in this reaction are solids, allowing NiCad batteries to be recharged hundreds of times before they stop operating. Unfortunately, NiCad batteries are fairly heavy batteries to be carrying around in a pacemaker. Today, the lighter lithium/iodine battery is used instead. The iodine is dissolved in a solid polymer support, and the overall redox reaction is as follows:

$$\mathrm{2\,Li\,(s) + I_2\,(s) \longrightarrow 2\,LiI\,(s)}$$

Lithium is oxidized, and iodine is reduced. Although the lithium/iodine battery cannot be recharged, one of its advantages is that it lasts up to 10 years. Thus, a person with a pacemaker does not have to worry about periodic recharging; about once per decade a person requires minor surgery to replace the pacemaker/battery unit. Lithium/iodine batteries are also used to power calculators and watches.



Figure 6.4.2: Small button cell batteries like these are commonly used to power compact electronic devices such as watches, pacemakers, and calculators due to their long life and compact size. (CC BY-SA; Gerhard H Wrodnigg via Wikipedia).

Oxidation and reduction can also be defined in terms of changes in composition. The original meaning of oxidation was "adding oxygen," so when oxygen is added to a molecule, the molecule is being oxidized. The reverse is true for reduction: if a molecule loses oxygen atoms, the molecule is being reduced. For example, the acetaldehyde (CH_3CHO) molecule takes on an oxygen atom to become acetic acid (CH_3COOH).

$$\mathrm{2\,CH_3CHO + O_2 \longrightarrow 2\,CH_3COOH}$$

Thus, acetaldehyde is being oxidized.

Similarly, reduction and oxidation can be defined in terms of the gain or loss of hydrogen atoms. If a molecule adds hydrogen atoms, it is being reduced. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of acetaldehyde into ethanol (CH_3CH_2OH), hydrogen atoms are added to acetaldehyde, so the acetaldehyde is being reduced:

$$CH_3CHO + H_2 \longrightarrow CH_3CH_2OH$$

Table 6.4.1 Summary of How Oxidation and Reduction Affect Oxygen and Hydrogen Levels in Some Reactions.

Process	Change in oxygen (some reactions)	Change in hydrogen (some reactions)
Oxidation	gain	lose
Reduction	lose	gain

Example 6.4.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $N_2 \rightarrow NH_3$
b. $CH_3CH_2OHCH_3 \rightarrow CH_3COCH_3$
c. HCHO → HCOOH

Solution

a. Hydrogen is being added to the original reactant molecule, so reduction is occurring.
b. Hydrogen is being removed from the original reactant molecule, so oxidation is occurring.
c. Oxygen is being added to the original reactant molecule, so oxidation is occurring.

Key Takeaways

- Oxidation-reduction (redox) reactions involve the transfer of electrons, where one substance loses electrons (oxidation) and another gains electrons (reduction).
- Half-reactions are used to separate oxidation and reduction processes, helping balance redox reactions by ensuring both mass and charge are conserved.
- The substance that loses electrons is the reducing agent, while the substance that gains electrons is the oxidizing agent.
- Balancing redox reactions requires ensuring both element and charge balance, often using half-reactions to simplify the process.

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6.5: Recognizing Redox Reactions

Assigning Oxidation Numbers

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use **oxidation numbers** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on four rules. Oxidation numbers are not necessarily equal to the charge on the atom (although sometimes they can be); we must keep the concepts of charge and oxidation numbers separate.

The rules for assigning oxidation numbers to atoms are as follows:

1. Atoms in their elemental state are assigned an oxidation number of 0.
2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge. Oxidation numbers are usually written with the sign first, then the magnitude, to differentiate them from charges.
3. In compounds, fluorine is assigned a –1 oxidation number; oxygen is usually assigned a –2 oxidation number (except in peroxide compounds [where it is –1] and in binary compounds with fluorine [where it is positive]); and hydrogen is usually assigned a +1 oxidation number (except when it exists as the hydride ion [H^-], in which case rule 2 prevails).
4. In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).

Here are some examples for practice. In H_2, both H atoms have an oxidation number of 0 by rule 1. In $MgCl_2$, magnesium has an oxidation number of +2, while chlorine has an oxidation number of –1 by rule 2. In H_2O, the H atoms each have an oxidation number of +1, while the O atom has an oxidation number of –2, even though hydrogen and oxygen do not exist as ions in this compound (rule 3). By contrast, by rule 3, each H atom in hydrogen peroxide (H_2O_2) has an oxidation number of +1, while each O atom has an oxidation number of –1. We can use rule 4 to determine oxidation numbers for the atoms in SO_2. Each O atom has an oxidation number of –2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the S atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it means only that the S atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Example 6.5.1

Assign oxidation numbers to the atoms in each substance.

1. Cl_2
2. GeO_2
3. $Ca(NO_3)_2$

Solution

1. Cl_2 is the elemental form of chlorine. Rule 1 states each atom has an oxidation number of 0.
2. By rule 3, oxygen is normally assigned an oxidation number of –2. For the sum of the oxidation numbers to equal the charge on the species (zero), the Ge atom is assigned an oxidation number of +4.
3. $Ca(NO_3)_2$ can be separated into two parts: the Ca^{2+} ion and the NO_3^- ion. Considering these separately, the Ca^{2+} ion has an oxidation number of +2 by rule 2. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of –2, and there are three of them. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation $x + 3(-2) = -1$

   where x is the oxidation number of the N atom and the –1 represents the charge on the species. Evaluating for x,

   $x + (-6) = -1x = +5$

Thus the oxidation number on the N atom in the NO_3^- ion is +5.

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction can also be defined in terms of increasing or decreasing oxidation numbers, respectively.

Example 6.5.2:

Identify what is being oxidized and reduced in this redox reaction.

$$2\,\mathrm{Na} + \mathrm{Br_2} \longrightarrow 2\,\mathrm{NaBr}$$

Solution

Both reactants are the elemental forms of their atoms, so the Na and Br atoms have oxidation numbers of 0. In the ionic product, the Na^+ ions have an oxidation number of +1, while the Br^- ions have an oxidation number of –1.

$$\underset{0}{2Na} + \underset{0}{Br_2} \rightarrow \underset{+1-1}{2NaBr}$$

Sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to –1, so it is being reduced:

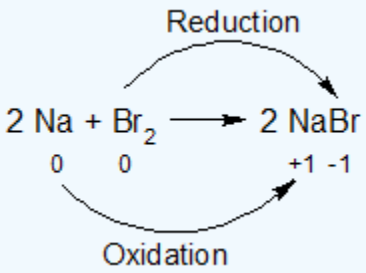


Figure 6.5.1: Illustration of a redox reaction where sodium (Na) is oxidized from 0 to +1 and bromine (Br_2) is reduced from 0 to -1, forming sodium bromide (NaBr). The diagram highlights both oxidation and reduction processes with curved arrows. (CC BY 4.0; Gorkem Ermut via www.acdlabs.com [ACD]). Detailed Description of Figure 6.5.1.

Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by sodium (two, one lost from each Na atom) is gained by bromine (two, one gained for each Br atom).

Oxidation reactions can become quite complex, as attested by the following Redox Reaction 6.5.1:

$$6H^+(aq) + 2\underset{+7}{MnO_4^-}(aq) + 5\underset{-1}{H_2O_2}(l) \rightarrow 2\underset{+2}{Mn^{2+}}(aq) + 5\underset{0}{O_2}(g) + 8H_2O(l) \tag{6.5.1}$$

To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced?

Chemistry in Everyday Life: Fortifying Food with Iron

Iron is an essential mineral in our diet; iron-containing compounds like the heme protein in hemoglobin could not function without it. Most biological iron has the form of the Fe^{2+} ion; iron with other oxidation numbers is almost inconsequential in human biology (although the body does contain an enzyme to reduce Fe^{3+} to Fe^{2+}, so Fe^{3+} must have some biological significance, albeit minor). To ensure that we ingest enough iron, many foods are enriched with iron. Although Fe^{2+}

compounds are the most logical substances to use, some foods use "reduced iron" as an ingredient (bread and breakfast cereals are the most well-known examples). Reduced iron is simply iron metal; iron is added as a fine metallic powder. The metallic iron is oxidized to Fe^{2+} in the digestive system and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use Fe^{2+} salts as an additive?

Although it is difficult to establish conclusive reasons, a search of scientific and medical literature suggests a few reasons. One reason is that fine iron filings do not affect the taste of the product. The size of the iron powder (several dozen micrometers) is not noticeable when chewing iron-supplemented foods, and the tongue does not detect any changes in flavor that can be detected when using Fe^{2+} salts. Fe^{2+} compounds can affect other properties of foodstuffs during preparation and cooking, like dough pliability, yeast growth, and color. Finally, of the common iron substances that might be used, metallic iron is the least expensive. These factors appear to be among the reasons why metallic iron is the supplement of choice in some foods.

and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use Fe^{2+} salts as an additive?

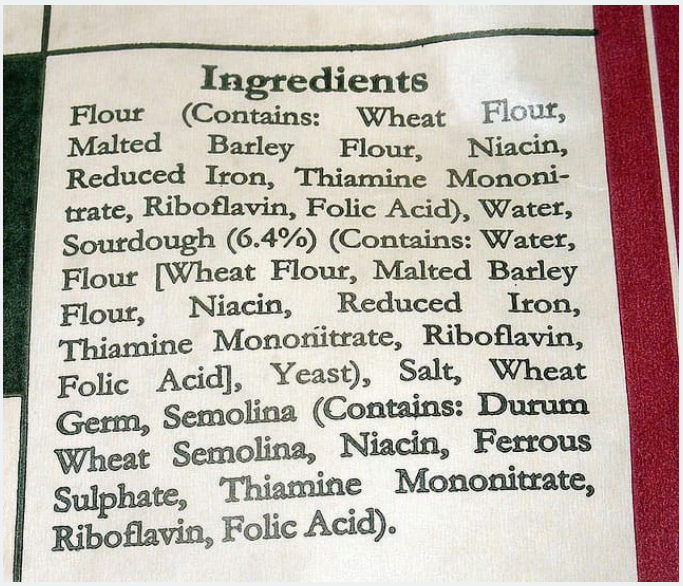


Figure 6.5.2: Many prepared foods list reduced iron in their ingredients list. (CC BY-NC-SA 3.0; Ball via Beginning Chemistry). Detailed Description of Figure 6.5.2.

Key Takeaways

- Oxidation numbers help track electron transfer in redox reactions and are assigned based on a set of rules.
- Elements in their natural state have an oxidation number of zero, while common ions and compounds follow predictable oxidation number assignments (e.g., oxygen is usually -2, hydrogen is usually +1).
- Redox reactions involve a simultaneous change in oxidation numbers, where one element increases its oxidation number (oxidation) and another decreases its oxidation number (reduction).
- Balanced redox reactions ensure that the total number of electrons lost equals the total number of electrons gained, maintaining charge balance.

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6.6: Redox Reactions in Organic Chemistry and Biochemistry

Combustion Reactions

Oxidation-reduction reactions are of central importance in organic chemistry and biochemistry. These reactions involve elements across the entire periodic table and are integral to countless processes in nature, biological systems, and industrial applications.

The burning of fuels that provide the energy to maintain our civilization and the metabolism of food that furnishes the energy to keep us alive both involve redox reactions.



Figure 6.6.1: The burning of natural gas is not only a combustion reaction but also a redox reaction. Similar reactions include the burning of gasoline and coal. These are also redox reactions. (CC BY-NC-SA 3.0; Lizeth Guadalupe Cervantes via Introductory, Conceptual, and GOB Chemistry).

All **combustion reactions** are also redox reactions. A typical combustion reaction is the burning of methane, the principal component of natural gas (Figure 6.6.1).

$$\mathrm{CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O} \tag{6.6.1}$$

Respiration

In **respiration**, the biochemical process by which the oxygen we inhale in air is used to oxidize foodstuffs to carbon dioxide and water, redox reactions provide energy to living cells. A typical respiratory reaction is the oxidation of glucose ($\mathrm{C_6H_{12}O_6}$), the simple sugar we encountered in the chapter-opening essay that makes up the diet of yeast:

$$\mathrm{C_6H_{12}O_6 + 6\,O_2 \rightarrow 6\,CO_2 + 6\,H_2O} \tag{6.6.2}$$

Organic chemists use a variety of redox reactions. For example, potassium dichromate ($\mathrm{K_2Cr_2O_7}$) is a common oxidizing agent that can be used to oxidize alcohols (symbolized by the general formula ROH). The product of the reaction depends on the location of the OH functional group in the alcohol molecule, the relative proportions of alcohol and the dichromate ion, and reaction conditions such as temperature. If the OH group is attached to a terminal carbon atom and the product is distilled off as it forms, the product is an aldehyde, which has a terminal carbonyl group (C=O) and is often written as RCHO. One example is the reaction used by the Breathalyzer to detect ethyl alcohol ($\mathrm{CH_3CH_2OH}$) in a person's breath:

$$\mathrm{3\,CH_3CH_2OH + Cr_2O_7^{2-} + 8\,H^+ \rightarrow 3\,CH_3CHO + 2\,Cr^{3+} + 7\,H_2O} \tag{6.6.3}$$

If the product acetaldehyde ($\mathrm{CH_3CHO}$) is not removed as it forms, it is further oxidized to acetic acid ($\mathrm{CH_3COOH}$). In this case, the overall reaction is as follows:

$$\mathrm{3\,CH_3CH_2OH + 2\,Cr_2O_7^{2-} + 16\,H^+ \rightarrow 3\,CH_3COOH + 4\,Cr^{3+} + 11\,H_2O} \tag{6.6.4}$$

In this reaction, the chromium atom is reduced because it loses oxygen atoms from $Cr_2O_7^{2-}$ to Cr^{3+}. On the other hand, the

carbon atom in ethanol is oxidized. In the **oxidation** of ethyl alcohol (CH_3CH_2OH, a.k.a. ethanol) to form acetaldehyde (CH_3CHO, a.k.a. ethanal), the number of bonds to oxygen has increased, and the number of hydrogen atoms has decreased from six to four. Either or both of these indicate that an oxidation has occurred.

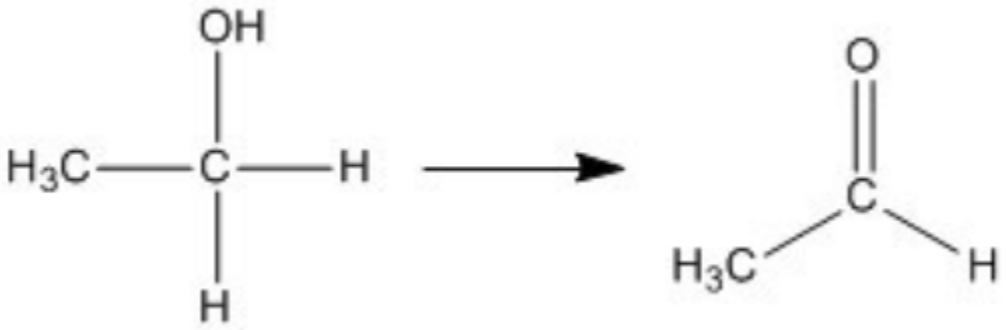


Figure 6.6.2: Oxidation of ethanol (CH_3CH_2OH) to form acetaldehyde (CH_3CHO). (CC BY-NC-SA 3.0; Lizeth Guadalupe Cervantes via Introductory, Conceptual, and GOB Chemistry).

In the oxidation of acetaldehyde to acetic acid (a.k.a. ethanoic acid), the carbon atom that gained an additional oxygen is the element oxidized.

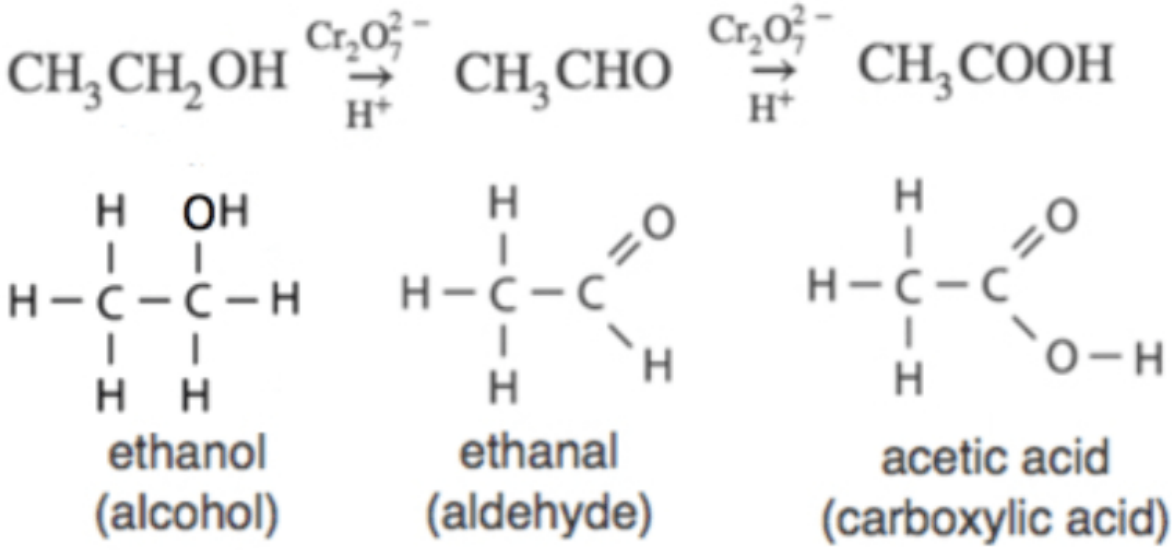


Figure 6.6.3: Ethanol (CH_3CH_2OH) is oxidized to ethanal (CH_3CHO), an aldehyde, and further oxidized to acetic acid (CH_3COOH), a carboxylic acid, using acidified dichromate ($Cr_2O_7^{2-}$/H^+) as the oxidizing agent. The structures of each compound are shown with hydrogen, carbon, and oxygen atoms. (CC BY-NC-SA 3.0; Lizeth Guadalupe Cervantes via Introductory, Conceptual, and GOB Chemistry). Detailed Description of Figure 6.6.3.

When the OH group of the alcohol is bonded to an interior carbon atom, the oxidation of an alcohol will produce a ketone (the formulas of ketones are often written as RCOR, and the carbon–oxygen bond is a double bond). The simplest ketone is derived from 2-propanol ($CH_3CHOHCH_3$). It is the common solvent acetone [$(CH_3)_2CO$], which is used in varnishes, lacquers, rubber cement, and nail polish remover. Acetone can be formed by the following redox reaction:

$$3\,CH_3CHOHCH_3 + Cr_2O_7^{2-} + 8\,H^+ \rightarrow 3\,(CH_3)_2CO + 2\,Cr^{3+} + 7\,H_2O \tag{6.6.5}$$

As we have just seen, aldehydes and ketones can be formed by the oxidation of alcohols. Conversely, aldehydes and ketones can be reduced to alcohols. Reduction of the carbonyl group is important in living organisms. For example, in **anaerobic metabolism**, in which biochemical processes take place in the absence of oxygen, pyruvic acid ($CH_3COCOOH$) is reduced to lactic acid (

$CH_3CHOHCOOH$) in the muscles.

$$CH_3COCOOH \rightarrow CH_3CHOHCOOH \tag{6.6.6}$$

(Pyruvic acid is both a carboxylic acid and a ketone; only the ketone group is reduced.) The buildup of lactic acid during vigorous exercise is responsible in large part for the fatigue that we experience.

In food chemistry, the substances known as **antioxidants** are reducing agents. Ascorbic acid (vitamin C; $C_6H_8O_6$) is thought to retard potentially damaging oxidation of living cells. In the process, it is oxidized to dehydroascorbic acid ($C_6H_6O_6$). In the stomach, ascorbic acid reduces the nitrite ion (NO_2^-) to nitric oxide (NO):

$$C_6H_8O_6 + 2\,H^+ + 2\,NO_2^- \rightarrow C_6H_6O_6 + 2\,H_2O + 2\,NO \tag{6.6.7}$$

If a reaction in Equation 6.6.7 did not occur, nitrite ions from foods would oxidize the iron in hemoglobin, destroying its ability to carry oxygen.

Tocopherol (vitamin E) is also an antioxidant. In the body, vitamin E is thought to act by scavenging harmful by-products of metabolism, such as the highly reactive molecular fragments called free radicals. In foods, vitamin E acts to prevent fats from being oxidized and thus becoming rancid. Vitamin C is also a good antioxidant (Figure 6.6.2).



Figure 6.6.4: Citrus Fruits. Citrus fruits, such as oranges, lemons, and limes, are good sources of vitamin C, which is an antioxidant. (Public domain via Wikipedia).

Finally, and of greatest importance, green plants carry out the redox reaction that makes possible almost all life on Earth. They do this through a process called **photosynthesis**, in which carbon dioxide and water are converted to glucose ($C_6H_{12}O_6$). The synthesis of glucose requires a variety of proteins called enzymes and a green pigment called chlorophyll that converts sunlight into chemical energy (Figure 6.6.3). The overall change that occurs is as follows:

$$6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2 \tag{6.6.8}$$

In this reaction, carbon dioxide is reduced to glucose, and water is oxidized to oxygen gas. Other reactions convert the glucose to more complex carbohydrates, plant proteins, and oils.



Figure 6.6.3: Life on Earth. Photosynthesis is the fundamental process by which plants use sunlight to convert carbon dioxide and water into glucose and oxygen. Then plants make more complex carbohydrates. It is the ultimate source of all food on Earth, and it is a redox reaction. (Public domain; Wikipedia).

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6.7: Summary

Chemical reactions can be classified by type. Combination reactions (also called composition reactions) make a substance from other substances. Decomposition reactions break one substance down into multiple substances. Combustion reactions combine molecular oxygen with the atoms of another reactant.

Oxidation reactions are reactions in which an atom loses an electron. Reduction reactions are reactions in which an atom gains an electron. These two processes always occur together, so they are collectively referred to as oxidation-reduction (or redox) reactions. The species being oxidized it called the reducing agent, while the species being reduced is the oxidizing agent. Alternate definitions of oxidation and reduction focus on the gain or loss of oxygen atoms, or the loss or gain of hydrogen atoms. Redox reactions are easily balanced if the overall reaction is first separated into half reactions, which are individually balanced.

Oxidation-reduction reactions are common in organic and biological chemistry. Respiration, the process by which we inhale and metabolize oxygen, is a series of redox reactions. In the absence of oxygen, redox reactions still occur in a process called anaerobic metabolism. Antioxidants such as ascorbic acid also play a part in the human diet, acting as reducing agents in various biochemical reactions. Photosynthesis, the process by which plants convert water and carbon dioxide to glucose, is also based on redox reactions.

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6.8: Key Terms

- **Synthesis (Combination) Reaction:** A chemical reaction that makes a single substance from two or more reactants.
- **Decomposition Reaction:** A chemical reaction in which a single substance is converted into two or more products.
- **Single Displacement Reaction:** A single displacement reaction (also called a single replacement reaction) is a type of chemical reaction in which one element replaces another element in a compound.
- **Double Displacement Reaction**: A double displacement reaction (also called a double replacement reaction) is a type of chemical reaction in which the positive and negative ions of two compounds swap places, forming two new compounds.
- **Precipitation Reaction:** A precipitation reaction is a type of double displacement reaction in which two soluble reactants in aqueous solution react to form an insoluble solid called a precipitate.
- **Precipitate:** A precipitate is an insoluble solid that forms when two liquid or aqueous solutions react in a precipitation reaction. The solid separates from the solution and may appear as cloudiness, sediment, or crystals.
- **Solubility Rules:** Solubility rules are a set of guidelines that help predict whether a given ionic compound will dissolve in water (soluble) or form a precipitate (insoluble) in an aqueous solution. These rules are based on experimental observations and are commonly used in precipitation reactions.
- **Molecular equation:** A molecular equation for a precipitation reaction is a balanced chemical equation that shows all reactants and products as neutral compounds without breaking them into their ionic forms. This equation does not explicitly show dissociation in aqueous solutions.
- **Total Ionic Equation:** A total ionic equation is a balanced chemical equation that shows all soluble ionic compounds as their dissociated ions in aqueous solution while keeping insoluble substances, gases, and liquids in their molecular form.
- **Net Ionic Equation:** A net ionic equation is a simplified chemical equation that shows only the ions and molecules directly involved in the chemical reaction, excluding spectator ions (ions that do not change during the reaction).
- **Spectator Ions:** Spectator ions are ions that do not participate in a chemical reaction. They remain unchanged in both the reactant and product sides of a total ionic equation and do not affect the overall reaction.
- **Arrhenius Definition**: An Arrhenius acid increases H^+ (hydrogen ions) in water, and an Arrhenius base increases OH^- (hydroxide ions) in water.
- **Hydronium ion**: A hydronium ion (H_3O^+) forms when a hydrogen ion (H^+) from an acid joins with a water molecule (H_2O). It's the form that hydrogen ions take in water and it's what makes a solution acidic.
- **Neutralization Reaction:** A neutralization reaction is a type of chemical reaction in which an acid and a base react to form water and a salt. This reaction typically occurs in aqueous solution and results in the neutralization of the acidic and basic properties.
- **Salt**: A salt or ionic compound is a chemical compound consisting of an assembly of positively charged ions (cations) and negatively charged ions (anions).
- **Oxidation:** The loss of electrons by an atom.
- **Reduction:** The gain of electrons by an atom.
- **Oxidation-Reduction (Redox) Reactions:** A chemical reaction in which electrons are transferred from one atom to another. Also, called a redox reaction.
- **Reducing Agent:** A species that causes reduction.
- **Oxidizing Agent:** A species that causes oxidation, which is itself reduced.
- **Half Reactions:** A chemical reaction that shows only oxidation or reduction.
- **Oxidation Numbers:** An oxidation number (or oxidation state) is a numerical value assigned to an atom in a chemical compound that represents the number of electrons lost, gained, or shared in a chemical bond. It helps determine how electrons are transferred in oxidation-reduction (redox) reactions.
- **Combustion Reaction:** A chemical reaction accompanied by the production of light and/or heat, typically a vigorous reaction because of combination with oxygen.
- **Respiration:** The biochemical process by which the oxygen we inhale oxidizes foodstuffs to carbon dioxide and water.
- **Anaerobic Metabolism:** A biochemical process that takes place in the absence of oxygen.
- **Antioxidants:** A substance in foods that acts as a reducing agent.
- **Photosynthesis:** The process by which plants use solar energy to convert carbon dioxide and water to glucose.

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6.9: Alternative text for Chapter 6

Figure 6.1.1: Classification of Chemical Reactions

This flowchart represents the hierarchical classification of chemical reactions. At the top, a central node labeled "Chemical Reactions" branches out into three categories: "Precipitation Reactions" on the left. "Acid-Base Neutralization Reactions" in the center. "Redox Reactions" on the right.

Figure 6.2.1: Formation of Lead Iodide Precipitate

An image of a chemical reaction in a laboratory flask, likely showing the formation of a bright yellow precipitate. The substance appears to be lead(II) iodide (PbI_2), a product of the double displacement reaction between a lead(II) salt and potassium iodide in aqueous solution. The vibrant yellow color of lead(II) iodide is a classic visual in chemistry experiments, often used to demonstrate precipitation reactions.

Figure 6.4.1: Reaction of Zinc with Hydrochloric Acid

Zinc metal reacts with aqueous hydrochloric acid, producing bubbles as a result of hydrogen gas being released during the reaction.

Figure 6.5.1: Redox Reaction of Sodium and Bromine

Chemical equation showing a redox reaction: 2 Na + Br_2 → 2 NaBr. Oxidation states are labeled as 0 for both Na and Br_2 on the left, and +1 for Na and -1 for Br in NaBr on the right. Curved arrows indicate oxidation of sodium (Na) and reduction of bromine (Br_2), with labels "Oxidation" and "Reduction" respectively.

Figure 6.5.2: Ingredients List for Bread

The ingredients list contains Flour (Contains: wheat flour, malted barley flour, niacin, reduced iron, thiamine mononitrate, riboflavin, folic acid), water, sourdough (6.4%) (contains: water, flour [wheat flour, malted barley flour, niacin, reduced iron, thiamine mononitrate, riboflavin, folic acid], yeast), salt, wheat germ, semolina (contains: durum wheat semolina, niacin, ferrous sulphate, thiamine mononitrate, riboflavin, folic acid).

Figure 6.6.3: Oxidation of Ethanol to Acetic Acid

A chemical reaction sequence showing the stepwise oxidation of ethanol (CH_3CH_2OH) to acetic acid (CH_3COOH) using acidified dichromate ($Cr_2O_7^{2-}$/H^+) as the oxidizing agent. The first step shows ethanol, a primary alcohol, being oxidized to ethanal (CH_3CHO), an aldehyde. The second step shows ethanal being further oxidized to acetic acid, also known as ethanoic acid, which is a carboxylic acid. Each compound is accompanied by its structural formula: ethanol with a hydroxyl group (-OH) bonded to a carbon atom; ethanal with a carbonyl group (C=O) bonded to a hydrogen (aldehyde group); and acetic acid with a carboxyl group (-COOH). The names and functional groups of each compound are labeled.

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7: QUANTITIES IN CHEMICAL REACTIONS

Introduction

We often describe chemical reactions in terms of individual atoms and molecules, but the reactions around us involve far larger quantities of matter. Even a tiny sample of a substance contains billions upon billions of particles, far too many to count directly. Chemistry provides tools to relate these enormous numbers to quantities we can measure in the laboratory. This chapter introduces stoichiometry, the study of the numerical relationships between reactants and products in balanced chemical equations. By applying stoichiometry, we connect the microscopic world of atoms and molecules with the measurable, macroscopic quantities used in chemical reactions.

Learning Objectives

- Define the mole as a unit for the amount of substance and relate it to Avogadro's number.
- Calculate formula masses and molar masses of covalent and ionic compounds.
- Express and interpret the masses of moles of atoms, molecules, and compounds.
- Convert between mass and moles of a substance.
- Use balanced chemical equations to determine molar relationships between reactants and products.
- Perform mass–mole–mass conversions between different substances in a chemical reaction.

7.1: Prelude to Quantities in Chemical Reactions

7.2: The Mole

7.3: Atomic and Molar Masses

7.4: Mole to Mass Conversions

7.5: Mole to Mole Conversions

7.6: Mass to Mass Conversions

7.7: Reaction Yields

7.8: Summary

7.9: Key Terms

7.10: Alternative Text for Chapter 7

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7.1: Prelude to Quantities in Chemical Reactions

Amounts Matter: From Lavoisier to Everyday Life

"When the disengaged gasses are carefully examined, they are found to weigh **113.7 grs.**; these are of two kinds, viz. **144 cubical inches** of carbonic acid gas, weighing **100 grs.** and **380 cubical inches** of a very light gas, weighing only 13.7 grs and, when the water which has passed over into the bottle [labeled] H is carefully examined, it is found to have lost **85.7 grs.** of its weight. Thus, in this experiment, **85.7 grs.** of water, joined to **28 grs.** of charcoal, have combined in such a way as to form **100 grs.** of carbonic acid, and **13.7 grs.** of a particular gas capable of being burnt. (Bold emphasis added.)"

In the paragraph above from the Elements of Chemistry, Antoine Lavoisier (1743–94) is explaining an experiment in which he was trying to demonstrate that water is not an element but instead is composed of hydrogen (the gas "capable of being burnt") and oxygen (Figure 7.1.1). This is a historical account of a groundbreaking experiment and illustrates the importance of amounts in chemistry. Lavoisier was pointing out that the initial total mass of water and charcoal, 85.7 g plus 28 g, equals the final total mass of carbonic acid and the particular gas, 100 g plus 13.7 g. In this way, he was illustrating the law of conservation of matter. It is another way of saying that amounts matter.

Amounts do matter and in a variety of circumstances. The chapter-opening essay in Chapter 1 tells the story of a nurse who mistakenly read "2–3 mg" as "23 mg" and administered the higher and potentially fatal dose of morphine to a child. Food scientists who work in test kitchens must keep track of specific amounts of ingredients as they develop new products for us to eat. Quality control technicians measure amounts of substances in manufactured products to ensure that the products meet company or government standards. Supermarkets routinely weigh meat and produce and charge consumers by the ounce or the pound.

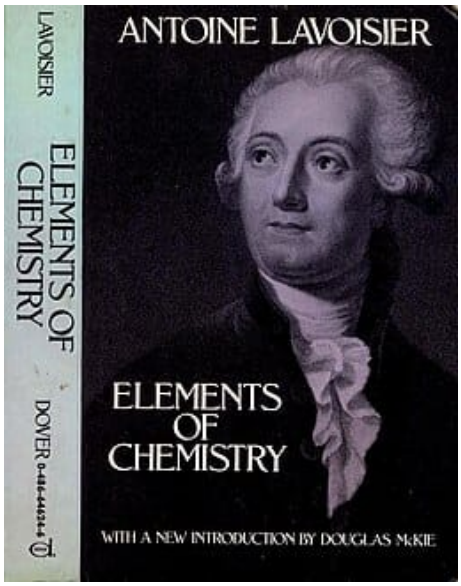


Figure 7.1.1: Cover of Antoine Lavoisier's *Elements of Chemistry*, featuring a portrait of Lavoisier, often regarded as the "Father of Modern Chemistry. (Public Domain; anonymous via Wikimedia).

Key Takeaways

- Lavoisier proved water is a compound and demonstrated the law of conservation of matter.
- Careful measurement of amounts is essential in chemistry.
- Accurate quantities matter in both science and everyday life.

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7.2: The Mole

Counting Particles in Chemistry

In everyday life, the measurement of quantities is carried out either by counting individual items or by weighing them in bulk, with the choice determined by practicality. For example, it is far more efficient to weigh a kilogram of rice than to count each grain, whereas it is easier to count a dozen eggs or pencils than to determine their mass. The same principle applies in the laboratory, where substances must be accurately measured to prepare solutions or conduct reactions. However, the actual participants in chemical processes are atoms, ions, molecules, and formula units—entities far too small to be counted directly. For example, a teaspoon of water (about 5 mL) contains 2 x 10^{23} water molecules (Figure 7.2.1).



Figure 7.2.1: A teaspoon of water (5 mL) holds about 2×10^{23} molecules —That's more molecules than stars in the universe! (OpenAI. (2025) Image generated using ChatGPT (GPT-5). This image was generated using ChatGPT tools based on user input and is not subject to copyright restrictions. Free to use and distribute with attribution to ChatGPT)

To address this challenge, chemists introduced the concept of the **mole**, a unit that enables the enumeration of chemical particles and represents a large number of atoms or molecules. The mole is a unit of amount similar to familiar units like pair, dozen, and gross. It provides a specific measure of *the number* of atoms or molecules in a sample of matter. Just as a dozen implies 12 items, a mole (abbreviated as mol) represents 6.022 × 10^{23} things. The number 6.022 × 10^{23}, called **Avogadro's number** (abbreviated as N_A) after the 19th-century chemist Amedeo Avogadro, is the number we use in chemistry to represent macroscopic amounts of atoms and molecules (Figure 7.2.2).



Figure 7.2.2: Each sample contains 6.02 × 10^{23} molecules or formula units—1.00 mol of the compound or element. Clockwise from the upper left: 1-octanol ($C_8H_{17}OH$), mercury(II) iodide (HgI_2), methanol (CH_3OH), and powdered sulfur (S_8). (CC BY 4.0; Sahar Atwa via OpenStax). Detailed Description of Figure 7.2.2.

Thus, if we have 6.022 × 10^{23} sodium atoms, we say we have 1 mol of sodium atoms. If we have 2 mol of sodium atoms, we have 2 × (6.022 × 10^{23}) sodium atoms, or 1.2044 × 10^{24} sodium atoms. Similarly, if we have 0.5 mol of benzene (C_6H_6) molecules, we have 0.5 × (6.022 × 10^{23}) C_6H_6 molecules, or 3.011 × 10^{23} C_6H_6 molecules.

> A mole represents a very large number! 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.

Notice that we are applying the mole unit to different types of chemical entities. The word “mole” represents a number of things—6.022 × 10^{23} of them—but does not, by itself, specify what “they” are. The chemical entities can be atoms, molecules, formula units, and ions. This specific information needs to be specified accurately.

Most elements are made up of individual atoms, such as helium. However, some elements consist of molecules, such as the diatomic elements, nitrogen, hydrogen, oxygen, fluorine, chlorine, bromine, and iodine. One mole of helium (He) consists of 6.022 × 10^{23} helium atoms, but one mole of nitrogen contains 6.022 × 10^{23} nitrogen (N_2) molecules. The basic units of molecular (covalent) compounds are molecules as well. The molecules of "compounds" consist of different kinds of atoms, while the molecules of "elements" consist of only one type of atom. For example, the molecules of ammonia (NH_3) consist of nitrogen and hydrogen atoms, while nitrogen (N_2) molecules have nitrogen (N) atoms only. Compounds that are ionic, like sodium chloride (NaCl), are represented by ionic formulas. One mole of NaCl, for example, refers to 6.022 × 10^{23} formula units of NaCl. And, one formula unit of NaCl consists of one sodium ion and one chloride ion.

A mole is a quantity that contains 6.022×10^{23} items. It serves as a conversion factor to relate the number of moles of a substance to the number of atoms or molecules. Avogadro's number can be used in dimensional analysis problems as a conversion factor or ratio. The relationship of moles can be expressed as follows:

$$1 \text{ mol} = 6.022 \times 10^{23} \text{atoms}$$

We can convert this relationship into two possible conversion factors shown below:

$$\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \text{ or } \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}$$

If the number of atoms is given, we can convert it into "moles" by multiplying the given value by the conversion factor on the left. However, if the number of "mol" is given, the appropriate conversion factor to use is the one at the right.

In the context of molecular compounds, the relationship between moles is represented as follows:

$$1\ \text{mol} = 6.022 \times 10^{23}\text{molecules}$$

$$\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{molecules}} \text{ or } \frac{6.022 \times 10^{23}\ \text{molecules}}{1\ \text{mol}}$$

If working with ionic compounds (such as NaCl), the mole relationship is expressed as follows:

$$1\ \text{mol NaCl} = 6.022 \times 10^{23}\ \text{NaCl formula units}$$

$$\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{NaCl formula units}} \text{ or } \frac{6.022 \times 10^{23}\ \text{NaCl formula units}}{1\ \text{mol}}$$

Example 7.2.1

How many atoms are present in 2.76 mol of Li?

Solution

The mole is a fundamental unit of measurement in chemistry, serving as a conversion factor for quantifying amounts of substances. It provides a standardized method for converting between the number of particles and moles of a substance.

$$1\ \text{mol} = 6.022 \times 10^{23}\text{atoms Li}$$

By using Avogadro's number as a conversion factor, we can calculate the total number of atoms.

$$2.76\ \cancel{\text{mol Li}} \times \frac{6.022 \times 10^{23}\text{atoms Li}}{1\ \cancel{\text{mol Li}}} = 1.66 \times 10^{24}\text{atoms Li}$$

Conversion Between Moles and Atoms, Molecules and Ions

We can understand numerical relationships in chemical formulas with a simple analogy: Asking how many wheels are on a given number of tricycles is similar to asking how many hydrogen atoms are in a given amount of water (H_2O). For example, suppose we want to know the number of wheels on 5 tricycles. The conversion factor comes from our everyday knowledge—each tricycle has 3 wheels (Figure 7.2.3).



Figure 7.2.3: Tricycle-to-Wheel Ratio. Each tricycle has 3 wheels, and we can express this relationship as the ratio 1 tricycle : 3 wheels. (OpenAI. (2025) Image generated using ChatGPT (GPT-5). This image was generated using ChatGPT tools based on user input and is not subject to copyright restrictions. Free to use and distribute with attribution to ChatGPT)

We can express that relationship as a ratio between wheels and tricycles:

$$3 \text{ wheels} : 1 \text{ tricycle}$$

With this ratio, we can write a conversion factor to determine the number of wheels in 5 tricycles:

$$5 \text{ tricycles} \times \frac{3 \text{ wheels}}{1 \text{ tricycle}} = 15 \text{ wheels}$$

Similarly, a chemical formula gives us ratios between elements and molecules. For example, the formula for water (H_2O) indicates that there are two hydrogen atoms per one water molecule. We can write this as:

$$2 \text{ H atoms} : 1 \text{ } H_2O \text{ molecule}$$

Just as we can scale up tricycles to dozens, we can scale molecules to moles. In terms of moles, we write:

$$2 \text{ mol H} : 1 \text{ mol } H_2O$$

With conversion factors such as these, which come directly from chemical formulas, we can determine the amounts of the constituent elements present in a given amount of a compound.

Using our unit conversion techniques learned in Chapter 1, we can use the mole relationship and the chemical formula to convert back and forth between the moles and the number of chemical entities (atoms, molecules or ions).

Because 1 N_2 molecule contains 2 N atoms, 1 mol of N_2 molecules (6.022×10^{23} molecules) has 2 mol of N atoms. 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 ethanol (C_2H_6O), we can construct the following relationships (Table 7.2.1):

Table 7.2.1: Molecular Relationships

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular 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 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular Relationships
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{1\text{ mol O atoms}}$

The following example illustrates how we can use these relationships as conversion factors.

Example 7.2.2

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms, hydrogen atoms, and oxygen atoms does it have?

Solution

Using the relationships in Table 7.2.1, we apply the appropriate conversion factor for each element:

$$2.5\ \cancel{\text{mol }C_2H_6O\text{ molecules}} \times \frac{2\text{ mol C atoms}}{1\ \cancel{\text{mol }C_2H_6O\text{ molecules}}} = 5\text{ mol C atoms}$$

Note how the unit *mol C_2H_6O molecule*s cancels algebraically. Similar equations can be constructed for determining the number of H and O atoms:

$$2.5\ \cancel{\text{mol }C_2H_6O\text{ molecules}} \times \frac{6\text{ mol H atoms}}{1\ \cancel{\text{mol }C_2H_6O\text{ molecules}}} = 15\text{ mol H atoms}$$

$$2.5\ \cancel{\text{mol }C_2H_6O\text{ molecules}} \times \frac{1\text{ mol O atoms}}{1\ \cancel{\text{mol }C_2H_6O\text{ molecules}}} = 2.5\text{ mol O atoms}$$

Example 7.2.3

How many formula units are present in 2.34 mol of NaCl? How many ions are in 2.34 mol of NaCl?

Solution

Typically, in a problem like this, we start with what we are given and apply the appropriate conversion factor. Here, we are given a quantity of 2.34 mol of NaCl, to which we can apply the definition of a mole as a conversion factor:

$$2.34\ \cancel{\text{mol NaCl}} \times \frac{6.022 \times 10^{23}\text{ formula units NaCl}}{1\ \cancel{\text{mol NaCl}}} = 1.41 \times 10^{24}\text{ formula units NaCl}$$

Since each formula unit contains two ions, there are

$$1.41 \times 10^{24}\ \cancel{\text{formula units NaCl}} \times \frac{2\text{ ions}}{1\ \cancel{\text{formula unit NaCl}}} = 2.82 \times 10^{24}\text{ ions}$$

in the sample.

Example 7.2.4

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76 \ \cancel{\text{mol H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{molecules H}_2\text{O}}{\cancel{\text{mol H}_2\text{O}}} = 1.66 \times 10^{24} \text{molecules H}_2\text{O}$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24} \ \cancel{\text{molecules H}_2\text{O}} \times \frac{3 \text{ atoms}}{1 \ \cancel{\text{molecule H}_2\text{O}}} = 4.99 \times 10^{24} \text{ atoms}$$

Key Takeaways:

- One mole (mol) is defined as 6.022×10^{23} particles (atoms, molecules, ions, or formula units).
- This constant is known as Avogadro's number (N_A).
- The mole serves as a bridge between the microscopic scale of atoms and molecules and the macroscopic scale of grams and liters.
- Chemical formulas indicate the mole ratios of atoms within compounds (e.g., H_2O contains 2 mol of H atoms and 1 mol of O atoms per 1 mol of H_2O).
- Conversion factors based on Avogadro's number allow conversion between moles, number of particles, and mass.

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7.3: Atomic and Molar Masses

Connecting the Mole to Atomic Weight and amu

Now that we have introduced the mole and practiced using it as a conversion factor, we ask the obvious question: why is the mole *that particular* number of things? Why is it 6.022×10^{23} and not 1×10^{23} or even 1×10^{20}?

In the previous section, we learned that chemistry uses a unit called the mole. The mole (mol) is a counting term similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. Since it is impossible to hand-count the number of molecules correctly, you must weigh them instead. This is a common method for dealing with all kinds of small objects: Nails, nuts, and grains of rice are all weighed rather than counted.

A mole is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure ^{12}C weighing exactly 12 g. One Latin connotation for the word “mole” is “large mass” or “bulk,” which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and a fundamental property, number of atoms, molecules, and so forth.

The number in a mole, Avogadro’s number, is related to the relative sizes of the **atomic mass unit** and **gram mass units**. Whereas one hydrogen atom has a mass of approximately 1 amu, 1 mol of H atoms has a mass of approximately 1 gram. And whereas one sodium atom has an approximate mass of 23 amu, 1 mol of Na atoms has an approximate mass of 23 grams.

The atomic weight of an element is the weighted average mass of the naturally occurring isotopes of that element, expressed relative to one-twelfth the mass of a carbon-12 atom. Atomic weights allow chemists to determine the mass of one mole of an element in grams, thereby linking the microscopic scale of atoms to measurable laboratory quantities. A single carbon-12 atom has a mass of exactly 12 atomic mass units (amu) by definition. An oxygen atom (^{16}O) has an average atomic weight of about 16.00 amu, and a hydrogen atom (^{1}H), the lightest element, has an average atomic mass of about 1.008 amu (Figure 7.3.1).

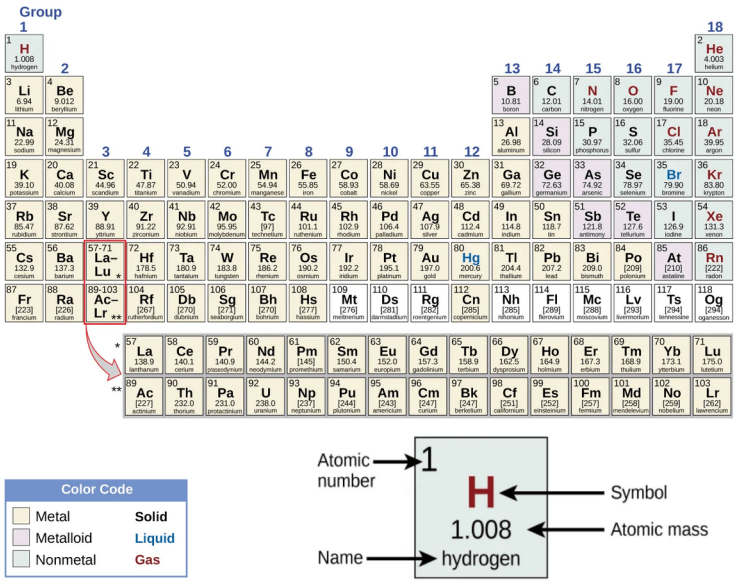


Figure 7.3.1: Relationship Between Atomic Weight, Atomic Mass, and Molar Mass. Hydrogen's atomic weight is 1.008 amu. Hydrogen's molar mass is 1.008 amu. One mole of hydrogen atoms weighs 1.008 g. (CC BY 4.0; anonymous OpenStax). Detailed Description of Figure 7.3.1.

One mole of a substance has the same mass in grams as one atom or molecule has in atomic mass units. The numbers in the periodic table that we identified as the atomic masses of the atoms not only tell us the mass of one atom in amu (atomic mass unit) but also tell us the mass of 1 mol of atoms in grams. The **molar mass** of any substance is numerically equivalent to its atomic weight in amu. According to the atomic mass unit (amu) definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). A mole of ^{12}C weighs 12 g (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses

are measured relative to that of the amu-reference substance, ^{12}C. Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu.

> One mole of a substance has the same mass in grams as one atom has in atomic mass units (amu).

But it also goes beyond carbon. Previously, we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the mass in grams of 1 mole of atoms of that element, while the molecular mass of a compound is the mass in grams of 1 mole of molecules of that compound. Sometimes these masses are called molar masses to emphasize the fact that they are the mass for 1 mol of things.

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (Figure 7.3.2).

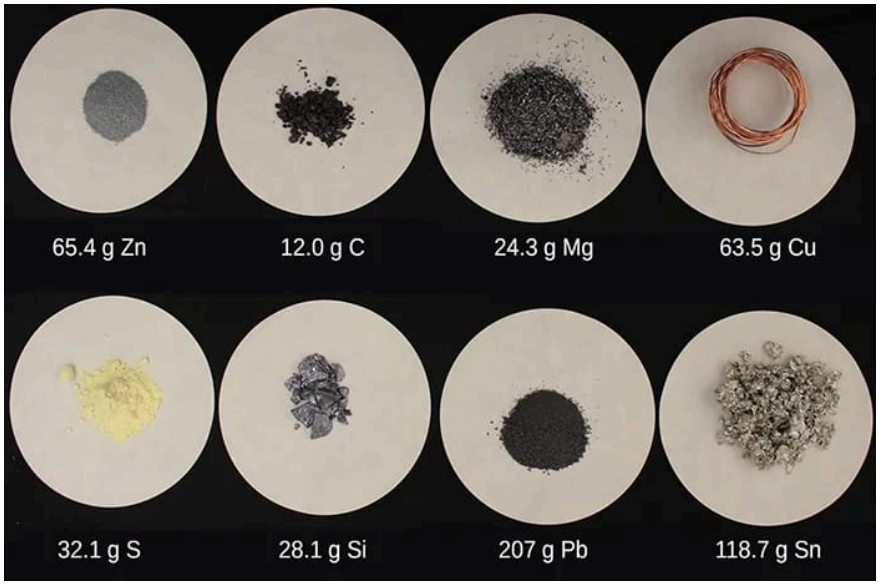


Figure 7.3.2: Each sample contains 6.022×10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g of zinc, 12.0 g of carbon, 24.3 g of magnesium, and 63.5 g of copper. From left to right (bottom row): 32.1 g of sulfur, 28.1 g of silicon, 207 g of lead, and 118.7 g of tin. (CC BY 4.0; Mark Ott via OpenStax). Detailed Description of Figure 7.3.2.

The molar mass of any substance is numerically equivalent to its atomic weight in amu. While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water after a rainfall. Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on Earth, each person would receive more than 100 billion molecules. (Table 7.3.1)

Table 7.3.1: Average Atomic Mass, Molar Mass, and Number of Atoms per Mole for Selected Elements. The table shows that the numerical value of atomic mass in amu is equal to the molar mass in grams per mole, and that one mole of each element contains Avogadro's number of atoms (6.022×10^{23}).

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
Cl	33.45	35.45	6.022×10^{23}

Remember that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022 × 10^{23} atoms:

$$12 \text{ g C} = 6.022 \times 10^{23} \text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50 × 10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25} \cancel{\text{ atoms C}} \times \frac{12.0000 \text{ g C}}{6.022 \times 10^{23} \cancel{\text{ atoms C}}} = 299 \text{ g C}$$

Defining Molecular Weight and Molar Mass

To determine how many molecules of a given substance are in a certain mass, it is helpful to refine a quantity called molecular weight. Molecular weight (MW) is the average mass of a molecule, calculated by summing the atomic weights of all the atoms it contains.

For example, the molecular weight of water (H_2O) is determined by summing the contributions of its atoms: two hydrogen atoms (2 × 1.008 amu) plus one oxygen atom (16.00 amu), giving a total of 18.02 amu:

$$\begin{aligned} \text{Atomic weight of 2 H} &= 2 \times 1.008 \text{ amu} = 2.016 \text{ amu} \\ \text{Atomic weight of 1 O} &= 1 \times 16.00 \text{ amu} = 16.00 \text{ amu} \\ \text{Molecular weight of } H_2O & = 18.016 \text{ amu} \end{aligned}$$

The molar mass of a substance is defined as the mass of one mole of its molecules. It represents the total mass of all the atoms present in a single mole of that molecule. Molar mass is expressed in units of grams per mole (g/mol).

For water, H_2O, Molar mass is:

$$\begin{aligned} \text{Molar mass of 2 H} &= 2 \times 1.008 \text{ g} = 2.016 \text{ g} \\ \text{Molar mass of 1 O} &= 1 \times 16.00 \text{ g} = 16.00 \text{ g} \\ \text{Molar mass of } H_2O &= 18.016 \text{ g} \end{aligned}$$

Example 7.3.1

Chloroform, ($CHCl_3$), is a molecular compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the “anti-stick” polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. What is the molecular mass (amu) for this compound?

Solution

The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. Figure 7.3.3 outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37

Figure 7.3.3: The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.3.3.

Example 7.3.2

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $C_2H_5O_2N$. What is the molar mass for this compound?

Solution

The molar mass of glycine is computed in the same fashion as its molecular mass. One mole of glycine, $C_2H_5O_2N$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen (Figure 7.3.4):

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
				Molar mass (g/mol compound)	75.07

Figure 7.3.4: The average mass of a glycine molecule, $C_2H_5O_2N$, is 75.07 grams, which is the sum of the molar masses of each of its constituent atoms. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.3.4.

Formula Mass of Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The

formula mass for an ionic compound is calculated in the same way as the molecular mass for covalent compounds: by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it cannot correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu.

$$\begin{aligned} \text{Atomic weight of 1 Na} &= 1 \times 22.99\,\text{amu} &&= 22.99\,\text{amu} \\ \text{Atomic weight of 1 Cl} &= 1 \times 35.45\,\text{amu} &&= 35.45\,\text{amu} \\ \text{Molecular weight of NaCl} & &&= 58.44\,\text{amu} \end{aligned}$$

The molar mass of an ionic compound is defined as the mass of one mole of its formula units. It represents the total mass of all the atoms present in a single mole of the compound's formula units. Molar mass is expressed in units of grams per mole (g/mol).

Example 7.3.3

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. In the formula $Al_2(SO_4)_3$, the subscript "3" outside the parentheses indicates that there are three sulfate groups present. Each sulfate group, SO_4^{2-}, consists of one sulfur atom and four oxygen atoms. When multiplied by three, this gives a total of three sulfur atoms and twelve oxygen atoms. Combined with the two aluminum atoms indicated by the subscript "2," the complete formula unit of aluminum sulfate contains two aluminum atoms, three sulfur atoms, and twelve oxygen atoms.

Following the approach outlined above, the formula mass for this compound is calculated as follows:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Formula mass	342.14

Figure 7.3.5: The average mass of an aluminum sulfate, $Al_2(SO_4)_3$, is 342.14 grams, which is the sum of the molar masses of each of its constituent atoms. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.3.5.

Key Takeaways

- The mole connects atomic masses in amu to measurable masses in grams.
- Molecular weight is in amu per molecule, while molar mass is in grams per mole.
- The mass of moles of atoms and molecules is expressed in units of grams.
- Formula mass is the sum of atomic masses in an ionic compound's formula unit, expressed in amu or g/mol.

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7.4: Mole to Mass Conversions

Using Molar Mass as a Conversion Factor

One of the main reasons the mole is such a useful unit in laboratory work is that it provides a direct link between mass and the number of entities in a sample. If you know the number of moles of a substance, you can calculate its mass or the number of particles it contains. Conversely, if the mass or number of entities is known, you can determine the number of moles.

As explained in the previous section, the molar mass of a substance is defined as the mass in grams of one mole—that is, Avogadro's number of molecules or formula units. Because molar mass directly relates the number of moles of a substance to its corresponding mass in grams, it serves as a useful conversion factor in chemical calculations. The most straightforward application of molar mass as a conversion factor is in mole–gram conversions, or in the reverse process of converting grams to moles (Figure 7.4.1).

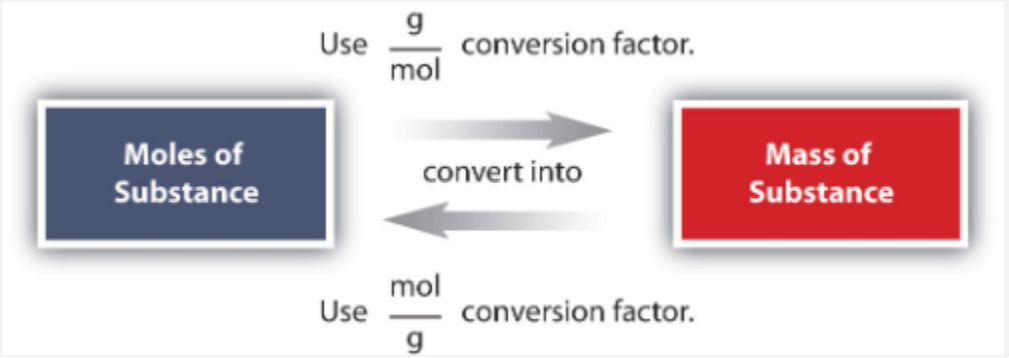


Figure 7.4.1: A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles. (CC BY-NC-SA 3.0; anonymous via The Basics of General, Organic, and Biological Chemistry v. 1.0). Detailed Description of Figure 7.4.1.

For example, 1 mol of Al has a mass of 26.98 g. Stated mathematically,

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

We can divide both sides of this expression by either side to get one of two possible conversion factors:

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \quad \text{and} \quad \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$$

The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do "by eye."

Deriving Grams from Moles

To convert a given amount in moles to the corresponding mass in grams, the number of moles is multiplied by the molar mass. This method provides a straightforward way to determine the mass of an element from a known amount in moles. The general relationship is expressed as:

$$\text{number of moles} \times \frac{\text{molar mass (g)}}{1 \text{ mol}} = \text{mass (g)}$$

Example 7.4.1

What is the mass of 3.987 mol of Al?

Solution

The first step in a conversion problem is to decide what conversion factor to use. Because we are starting with mole units, we want a conversion factor that will cancel the mole unit and introduce the unit for mass in the numerator. Therefore, we should use the $\frac{26.98\text{ g Al}}{1\text{ mol Al}}$ conversion factor. We start with the given quantity and multiply by the conversion factor (Figure 7.4.2):

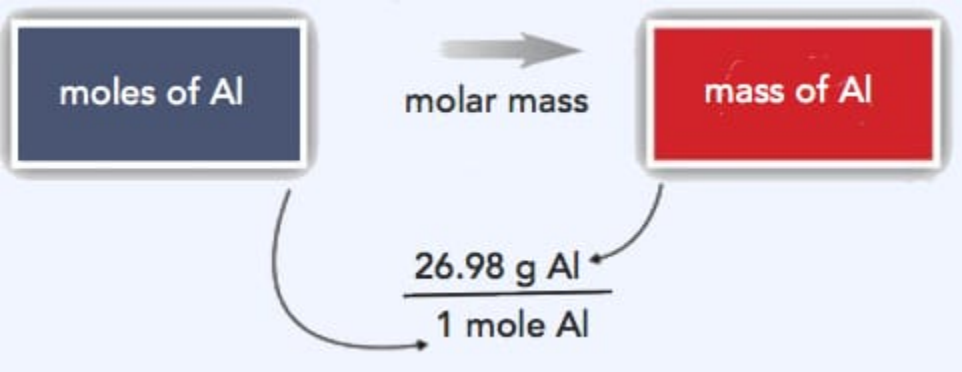


Figure 7.4.2: The molar mass of aluminum (26.98 g/mol) acts as the conversion factor, allowing conversion from moles of Al to grams of Al and vice versa. (CC BY-NC-SA 4.0; anonymous via CHEM1306: Health Chemistry I). Detailed Description of Figure 7.4.2.

$$3.987\text{ mol Al} \times \frac{26.98\text{ g Al}}{1\text{ mol Al}}$$

Note that the mol units cancel algebraically. (The quantity 3.987 mol is understood to be in the numerator of a fraction that has 1 in the unwritten denominator.) Canceling and solving gives

$$3.987\text{ mol Al} \times \frac{26.98\text{ g Al}}{1\text{ mol Al}} = 107.6\text{ g Al}$$

Our final answer is expressed to four significant figures

Example 7.4.2

A biochemist needs 0.00655 mol of bilirubin ($C_{33}H_{36}N_4O_6$) for an experiment. How many grams of bilirubin will that be?

Solution

To convert from moles to mass, we need the molar mass of bilirubin, which we can determine from its chemical formula:

$$\begin{aligned}
\text{Molar mass of 33 C} &= 33 \times 12.01\text{ g} &&= 396.33\text{ g} \\
\text{Molar mass of 36 H} &= 36 \times 1.008\text{ g} &&= 36.29\text{ g} \\
\text{Molar mass of 4 N} &= 4 \times 14.01\text{ g} &&= 56.04\text{ g} \\
\text{Molar mass of 6 O} &= 6 \times 16.00\text{ g} &&= 96.00\text{ g} \\
\text{Molar mass of } C_{33}H_{36}N_4O_6 &= &&584.66\text{ g/mol}
\end{aligned}$$

The molar mass of bilirubin is 584.73 g. Using the relationship

1 mol bilirubin = 584.66 g bilirubin

We can construct the appropriate conversion factor for determining how many grams there are in 0.00655 mol. Following the steps from Figure 7.4.2:

$$0.00655\text{ mol bilirubin} \times \frac{584.66\text{ g bilirubin}}{\text{mol bilirubin}} = 3.83\text{ g bilirubin}$$

The mol bilirubin unit cancels. The biochemist needs 3.83 g of bilirubin

Deriving Moles from Grams

To convert a given mass in grams to the amount in moles, the mass is divided by the molar mass (that is, multiplied by 1/molar mass). This method provides a straightforward way to determine the number of moles of an element from a measurable sample. The general relationship is expressed as:

$$\text{mass (g)} \times \frac{1\,\text{mol}}{\text{molar mass (g)}} = \text{number of moles}$$

Example 7.4.3

How many moles are present in 100.0 g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Solution

$$100.0\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} = 3.706\ \text{mol Al}$$

Deriving the Number of Entities (Atoms, Molecules, or Ions) from the Mass of a Substance

To convert a given mass in grams to the corresponding amount in moles, the mass is divided by the molar mass. This relationship allows chemists to determine how many atoms of an element are present in a measurable sample. Because the molar mass of an element is numerically equal to its atomic weight expressed in grams per mole, this conversion provides a direct bridge between the macroscopic scale of grams and the microscopic scale of atoms. For example, dividing 54.94 g of manganese by its molar mass (54.94 g/mol) shows that the sample contains exactly 1.00 mol of Mn atoms, or 6.022×10^{23} atoms.

$$1\ \text{mol Mn} = 54.94\ \text{g} = 6.022\times10^{23}\ \text{atoms}$$

Mass and the number of entities are not directly proportional to one another; instead, both are related through the concept of the mole. Consequently, any conversion between mass and number of entities must first pass through the number of moles as the connecting step. For example, to find the number of atoms in a given mass, the general relationship is expressed as:

$$\text{mass (g)} \times \frac{1\,\text{mol}}{\text{molar mass (g)}} \times \frac{6.022 \times 10^{23}\,\text{entities}}{1\,\text{mol}} = \text{number of entities}$$

Example 7.4.4

Copper is commonly used to fabricate electrical wire (Figure 7.4.3). How many copper atoms are in 5.00 g of copper wire?



Figure 7.4.3: A spool of copper wire, made up of countless copper (Cu) atoms arranged in a metallic structure. (CC BY 4.0; Emilian Robert Vicol via OpenStax).

Solution

The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (N_A) to convert this molar amount to number of Cu atoms:

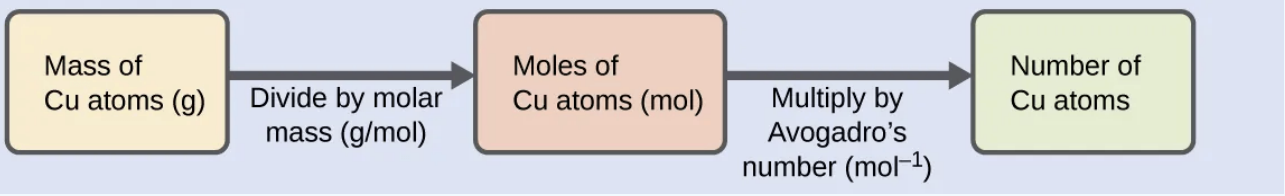


Figure 7.4.4: The mass of copper (Cu) in grams is first converted to moles by dividing by the molar mass. The moles are then multiplied by Avogadro's number to determine the total number of Cu atoms. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.4.4.

Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth N_A, or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$5.00\ \text{g Cu} \times \frac{1\ \text{mol Cu}}{63.55\ \text{g Cu}} \times \frac{6.022 \times 10^{23}\ \text{Cu atoms}}{1\ \text{mol Cu}} = 4.74 \times 10^{22}\ \text{Cu atoms}$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 4.74 x 10^{22} as expected.

Example 7.4.5

A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$). Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution

To determine the number of moles from a given mass in grams, the mass is divided by the compound's molar mass. Conversely, to convert from moles back to mass, the molar mass is again required as the conversion factor. For saccharin ($C_7H_5NO_3S$), this value can be obtained directly from its chemical formula by summing the atomic masses of all constituent atoms:

$$\begin{aligned}
\text{Molar mass of 7 C} &= 7 \times 12.01\,\text{g} &&= 84.07\,\text{g} \\
\text{Molar mass of 5 H} &= 5 \times 1.008\,\text{g} &&= 5.04\,\text{g} \\
\text{Molar mass of 1 N} &= 1 \times 14.01\,\text{g} &&= 14.01\,\text{g} \\
\text{Molar mass of 3 O} &= 3 \times 16.00\,\text{g} &&= 48.00\,\text{g} \\
\text{Molar mass of 1 S} &= 1 \times 32.07\,\text{g} &&= 32.07\,\text{g} \\
\text{Molar mass of } C_7H_5NO_3S &= &&= 183.19\,\text{g/mol}
\end{aligned}$$

The molar mass of bilirubin is 183.19 g. Using the relationship

1 mol saccharin ($C_7H_5NO_3S$) = 183.19g bilirubin

The number of molecules in a given mass of compound is computed by first deriving the number of moles and then multiplying by Avogadro's number:

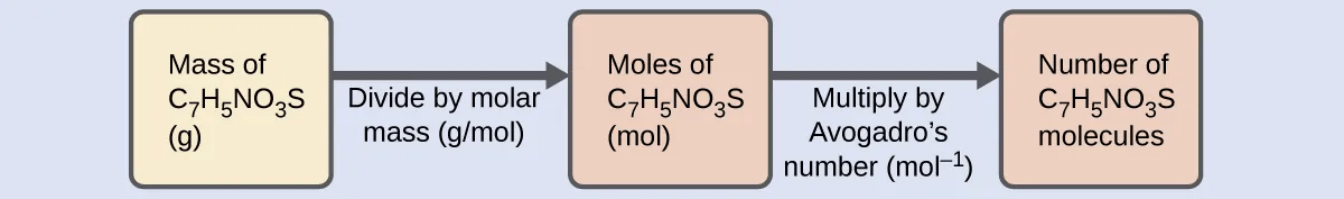


Figure 7.4.5: The mass of saccharin (C7H5NO3S) in grams is first converted to moles by dividing by the molar mass. The moles are then multiplied by Avogadro's number to determine the total number of Cu atoms. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.4.5.

Using the provided mass and molar mass for saccharin yields:

$$\begin{aligned}
&0.0400\ \cancel{\text{g } C_7H_5NO_3S}\left(\frac{\cancel{\text{mol } C_7H_5NO_3S}}{183.19\ \cancel{\text{g } CH_5NO_3S}}\right)\left(\frac{6.022 \times 10^{23}\ {}_7CH_5NO_3S\text{ molecules}}{1\ \cancel{\text{mol } C_7H_5NO_3S}}\right) \\
&= 1.31 \times 10^{20}\ {}_7CH_5NO_3S\text{ molecules}
\end{aligned}$$

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

$$1.31 \times 10^{20}\ {}_7CH_5NO_3S\text{ molecules}\left(\frac{7\,\text{C atoms}}{1\ {}_7CH_5NO_3S\text{ molecule}}\right) = 9.17 \times 10^{20}\,\text{C atoms}$$

Deriving the Mass of a Substance from the Number of Entities (Atoms, Molecules, or Ions)

To determine the number of moles of a compound from a given mass in grams, the mass of the sample is divided by the compound's molar mass. The molar mass of a compound is obtained by summing the atomic masses of all the atoms in its formula. This method allows chemists to connect measurable laboratory quantities to the molecular scale, where reactions actually occur. For example, the molar mass of water (H_2O) is 18.02 g/mol, derived from 2(1.01 g/mol for H) + 16.00 g/mol for O. If a sample contains 36.04 g of H_2O, dividing by 18.02 g/mol shows that the sample corresponds to 2.00 mol of water molecules, or $2.00 \times 6.022 \times 10^{23} = 1.20 \times 10^{24}$ molecules of H_2O.

1 mol H_2O=18.02 g=6.022×10^{23} molecules

The conversion between the number of entities and mass is not direct; it must proceed through the number of moles, which

serves as the essential bridge. For instance, when determining the mass corresponding to a specified number of atoms, the general relationship can be expressed as:

$$\text{number of entities} \times \frac{1\,\text{mol}}{6.022 \times 10^{23}\,\text{entities}} \times \frac{\text{molar mass (g)}}{1\,\text{mol}} = \text{mass (g)}$$

Example 7.4.6

Cortisone, $C_{21}H_{28}O_5$, is a synthetic steroid that is used as an anti-inflammatory drug. Determine the mass in grams of 6.51 x 10^{24} molecules of cortisone.

Solution

To determine the mass in grams from a given number of moles, the mass is divided by the compound's molar mass. For Cortisone, $C_{21}H_{28}O_5$, this value can be obtained directly from its chemical formula by summing the atomic masses of all constituent atoms:

$$\begin{aligned} \text{Molar mass of 21 C} &= 21 \times 12.01\,\text{g} &&= 252.21\,\text{g} \\ \text{Molar mass of 28 H} &= 28 \times 1.008\,\text{g} &&= 28.23\,\text{g} \\ \text{Molar mass of 5 O} &= 5 \times 16.00\,\text{g} &&= 80.00\,\text{g} \\ \text{Molar mass of } C_{21}H_{28}O_5 &= &&360.44\,\text{g/mol} \end{aligned}$$

The molar mass of cortisone, $C_{21}H_{28}O_5$, using the relationship

1 mol cortisone ($C_{21}H_{28}O_5$) = 360.44 g cortisone ($C_{21}H_{28}O_5$)

Using the provided mass and molar mass for cortisone yields:

$$6.51 \times 10^{24}\,C_{21}H_{28}O_5\,\text{molecules} \times \frac{1\text{mol}}{6.022 \times 10^{23}\,\text{molecules}} \times \frac{360.44\,\text{g}}{1\text{mol}} = 3.90 \times 10^{3}\,\text{g}$$

Link to Learning: To Your Health- Minerals

For our bodies to function properly, we need to ingest certain substances from our diets. Among our dietary needs are minerals, the noncarbon elements our body uses for a variety of functions, such developing bone or ensuring proper nerve transmission.

The US Department of Agriculture has established some recommendations for the RDIs of various minerals. The accompanying table lists the RDIs for minerals, both in mass and moles, assuming a 2,000-calorie daily diet.

Table 7.4.1: Essential Minerals and their Composition in Humans

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
Ca	1,000 mg	0.025 mol	1,000 mg	0.025 mol
Cr	35 µg	6.7×10^{-7} mol	25 µg	4.8×10^{-7} mol
Cu	900 µg	1.4×10^{-5} mol	900 µg	1.4×10^{-5} mol
F	4 mg	2.1×10^{-4} mol	3 mg	1.5×10^{-4} mol
I	150 µg	1.2×10^{-6} mol	150 µg	1.2×10^{-6} mol
Fe	8 mg	1.4×10^{-4} mol	18 mg	3.2×10^{-4} mol
K	3,500 mg	9.0×10^{-2} mol	3,500 mg	9.0×10^{-2} mol

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
Mg	400 mg	1.6×10^{-2} mol	310 mg	1.3×10^{-2} mol
Mn	2.3 mg	4.2×10^{-5} mol	1.8 mg	3.3×10^{-5} mol
Mo	45 mg	4.7×10^{-7} mol	45 mg	4.7×10^{-7} mol
Na	2,400 mg	1.0×10^{-1} mol	2,400 mg	1.0×10^{-1} mol
P	700 mg	2.3×10^{-2} mol	700 mg	2.3×10^{-2} mol
Se	55 µg	7.0×10^{-7} mol	55 µg	7.0×10^{-7} mol
Zn	11 mg	1.7×10^{-4} mol	8 mg	1.2×10^{-4} mol

Table 7.4.1 illustrates several things. First, the needs of men and women for some minerals are different. The extreme case is for iron; women need over twice as much as men do. In all other cases where there is a different RDI, men need more than women.

Second, the amounts of the various minerals needed on a daily basis vary widely—both on a mass scale and a molar scale. The average person needs 0.1 mol of Na a day, which is about 2.5 g. On the other hand, a person needs only about 25–35 µg of Cr per day, which is under one millionth of a mole. As small as this amount is, a deficiency of chromium in the diet can lead to diabetes-like symptoms or neurological problems, especially in the extremities (hands and feet). For some minerals, the body does not require much to keep itself operating properly.

Although a properly balanced diet will provide all the necessary minerals, some people take dietary supplements. However, too much of a good thing, even minerals, is not good. Exposure to too much chromium, for example, causes a skin irritation, and certain forms of chromium are known to cause cancer (as presented in the movie *Erin Brockovich*).

Key Takeaways

- Molar mass (g/mol) acts as the conversion factor between mass (grams) and amount (moles).
- To go from mass to moles, divide by molar mass; to go from moles to mass, multiply by molar mass.
- The mole links mass and particles. Use Avogadro's number (6.022×10^{23}) to convert between moles and number of atoms, molecules, or ions.
- Any conversion between mass and number of entities must pass through moles as the bridge.

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7.5: Mole to Mole Conversions

Interpreting Coefficients in Balanced Chemical Equations

Balancing chemical equations involves ensuring that the number of each type of atom is the same on both sides of the reaction. The coefficients placed before chemical formulas indicate the numbers of molecules or formula units of each substance. In this section, the meaning of these coefficients will be extended to a broader chemical context.

Consider the simple chemical equation

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4\,H_2 + 2\,O_2 \longrightarrow 4\,H_2O$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22\,H_2 + 11\,O_2 \longrightarrow 22\,H_2O$$

because 22:11:22 also reduces to 2:1:2. So we can simplify this version of the equation by writing it as

$$2\text{ mol }\ H_2 + 1\ \text{ mol }\ O_2 \longrightarrow 2\text{ mol }\ H_2O$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

We interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? Balanced chemical equations are balanced not only at the molecular level but also in terms of molar amounts of reactants and products. Thus, we can read this reaction as “two moles of hydrogen react with one mole of oxygen to produce two moles of water.”

The ratios we constructed to describe molecules reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\text{ mol }H_2}{1\text{ mol }O_2}\text{ or }\frac{1\text{ mol }O_2}{2\text{ mol }H_2}$$

$$\frac{2\text{ mol }H_2O}{1\text{ mol }O_2}\text{ or }\frac{1\text{ mol }O_2}{2\text{ mol }H_2O}$$

$$\frac{2\text{ mol }H_2}{2\text{ mol }H_2O}\text{ or }\frac{2\text{ mol }H_2O}{2\text{ mol }H_2}$$

We can use these ratios to determine what the amount of a substance, in moles, that will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called **stoichiometry**. The ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products, is called the **stoichiometric factor**.

Example 7.5.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O? The balanced equation is as follows:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

Solution

Since the problem involves quantities of H_2O and O_2, the stoichiometric ratio relating these two substances must be used. Because the given value is in moles of H_2O and the goal is to determine moles of O_2, the conversion factor should be arranged with H_2O in the denominator (to cancel the given unit) and O_2 in the numerator (to introduce the desired unit), as shown in Figure 7.5.1. Thus,

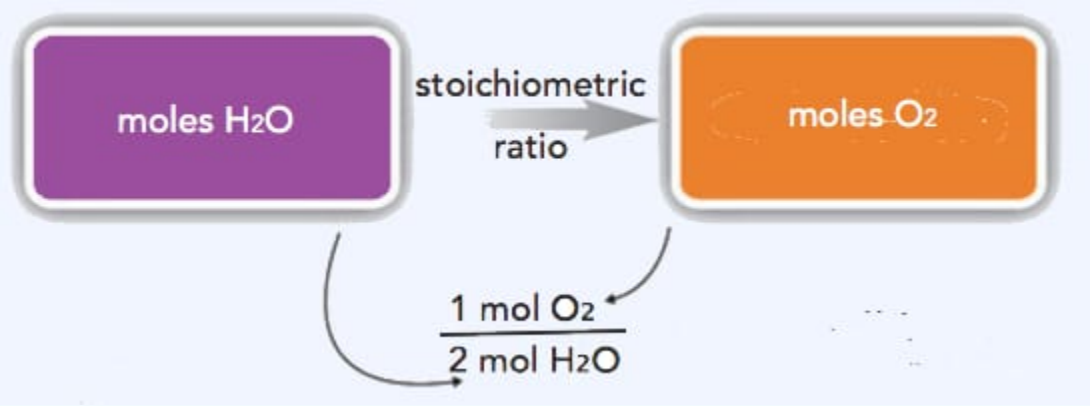


Figure 7.5.1: The stoichiometric ratio from the balanced chemical equation (2 mol H_2O : 1 mol O_2) serves as the conversion factor, allowing conversion from moles of water to moles of oxygen and vice versa. (CC BY-NC-SA 4.0; anonymous via CHEM1306: Health Chemistry I). Detailed Description of Figure 7.5.1.

$$27.6\ \text{mol}\ H_2O \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} = 13.8\ \text{mol}\ O_2$$

To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.

Example 7.5.2

Using $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of hydrogen react with 3.07 mol of oxygen to produce H_2O?

Solution

The balanced equation shows that 2 moles of H_2 react with 1 mole of O_2. This gives a mole ratio of

$$\frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2}$$

Multiplying the given 3.07 mol O_2 by this ratio gives the required moles of hydrogen:

$$3.07\ \text{mol}\ O_2 \times \frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2} = 6.14\ \text{mol}\ H_2$$

Thus, 6.14 moles of hydrogen react with 3.07 moles of oxygen to produce water.

Mole-to-Mole Conversions for Calculating Product Mass in Reactions

A balanced chemical equation is balanced not only in terms of atoms or molecules, but also in terms of moles. Balanced equations can therefore be used to set up mole ratios, which serve as conversion factors for solving stoichiometric problems, such as determining how many moles of substance A react with a given number of moles of substance B. This technique can be extended further by recalling that molar amounts can be related to mass through the molar mass. Using this relationship, stoichiometric questions can also be answered in terms of the masses of substances, in addition to moles. This is accomplished by applying the following sequence (Figure 7.5.2):

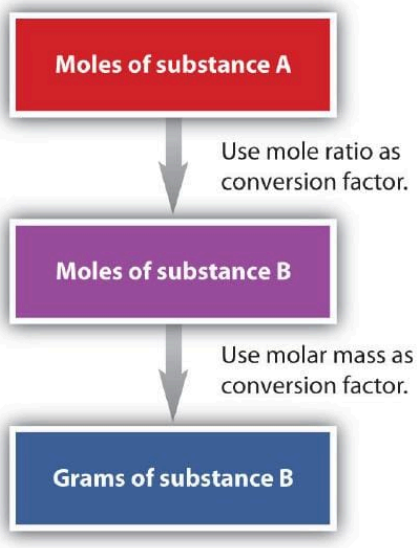


Figure 7.5.2: A flowchart showing the sequence of conversions. Moles of substance A are first related to moles of substance B using the mole ratio from the balanced equation. Then, moles of substance B are converted to grams using its molar mass as the conversion factor. (CC BY-NC-SA 3.0; anonymous via The Basics of General, Organic, and Biological Chemistry v. 1.0). Detailed Description of Figure 7.5.2.

Example 7.5.3

Consider the balanced chemical equation:

$$\mathrm{Fe_2O_3} + 3\,\mathrm{SO_3} \rightarrow \mathrm{Fe_2(SO_4)_3}$$

If 3.59 moles of $\mathrm{Fe_2O_3}$ are available, what mass of $\mathrm{SO_3}$ is reqired to react with it completely?

Solution

Using the mole-mass calculation sequence, we can determine the required mass of $\mathrm{SO_3}$ in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of $\mathrm{SO_3}$ needed. Then, using the molar mass of $\mathrm{SO_3}$ as a conversion factor, we determine the mass that this number of moles of $\mathrm{SO_3}$ has.

As usual, we start with the quantity we were given:

$$3.59\ \cancel{\text{mol Fe}_2\text{O}_3} \times \left(\frac{3\ \text{mol SO}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}}\right) = 10.77\ \text{mol SO}_3$$

The mol $\mathrm{Fe_2O_3}$ units cancel, leaving mol $\mathrm{SO_3}$ unit. Now, we take this answer and convert it to grams of $\mathrm{SO_3}$, using the molar mass of $\mathrm{SO_3}$ as the conversion factor:

$$10.77\ \cancel{\text{mol SO}_3} \times \left(\frac{80.06\ \text{g SO}_3}{1\ \cancel{\text{mol SO}_3}}\right) = 862\ \text{g SO}_3$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of $\mathrm{SO_3}$ will react with 3.59 mol of $\mathrm{Fe_2O_3}$. Many problems of this type can be answered in this manner.

The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:

$$3.59\ \cancel{mol\ Fe_2O_3} \times \underbrace{\left(\frac{3\ \cancel{mol\ SO_3}}{1\ \cancel{mol\ Fe_2O_3}}\right)}_{\text{converts to moles of SO}_3} \times \underbrace{\left(\frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}}\right)}_{\text{converts to grams of SO}_3} = 862\ g\ SO_3$$

We get exactly the same answer when combining all the math steps together.

Mole-to-Mole Conversions for Calculating Product Entities in Reactions

A balanced chemical equation is balanced not only with respect to atoms and molecules, but also in terms of moles. This means that the coefficients in the equation can be used to construct mole ratios, which serve as conversion factors between different substances in the reaction. These mole ratios make it possible to determine how many moles of one substance will react with or produce a specific number of moles of another substance. Because one mole of any substance contains Avogadro's number of representative particles (6.022×10^{23}), this technique can be extended to calculate the actual number of atoms, molecules, or formula units produced or consumed in a reaction. Thus, balanced equations provide a direct link between the macroscopic scale of moles and the microscopic scale of entities, allowing stoichiometric questions to be answered in terms of individual particles as well as moles.

Example 7.5.4

How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$$

Solution

The absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

$$\frac{3\,\mathrm{mol\,CO_2}}{1\,\mathrm{mol\,C_3H_8}}$$

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number (Figure 7.5.3),

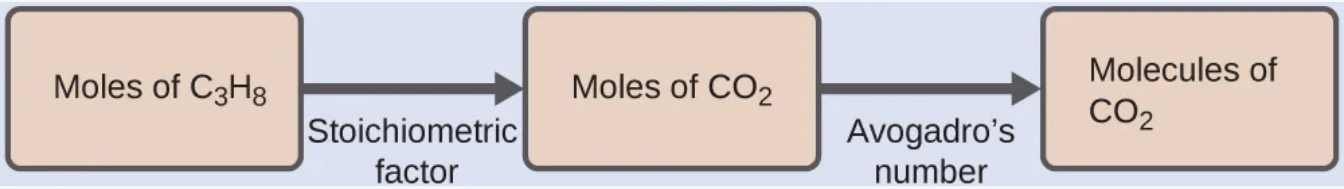


Figure 7.5.3: Calculating the Number of Product Molecules Generated by a Reaction Using Stoichiometry. Stoichiometric ratios convert moles of C3H8 to moles of CO2, and Avogadro's number converts moles of CO2 to molecules. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.5.3.

To determine the number of carbon dioxide molecules produced, we begin with the given moles of propane and use the stoichiometric ratio from the balanced chemical equation. The equation shows that 1 mole of propane reacts to produce 3 moles of carbon dioxide. After converting propane to moles of carbon dioxide, Avogadro's number is applied to convert moles of CO_2 into the actual number of molecules.

$$0.75mol\,{}_3\mathrm{CH_8} \times \frac{3mol\mathrm{CO_2}}{1mol\,\mathrm{C_3H_8}} \times \frac{6.022\times10^{23}\,\mathrm{CO_2\ molecules}}{mol\,\mathrm{CO_2}} = 1.4\times10^{24}\,\mathrm{CO_2\ molecules}$$

Key Takeaways

- The balanced chemical reaction can be used to determine molar relationships between substances.
- The numbers (coefficients) in a balanced equation show how many moles of each substance react or form.
- These numbers can be used like ratios to convert between moles of different substances in a reaction.
- Mole ratios from balanced equations allow conversion from moles of one substance to moles of another.
- Molar mass is then used to convert moles of a substance into grams, making stoichiometric calculations practical in the lab.

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7.6: Mass to Mass Conversions

Converting Mass of One Substance to Mass of Another Using Stoichiometry

In laboratory work, chemicals are almost always measured by mass rather than by directly counting moles. For example, a chemist may need to know how many grams of oxygen are required to completely react with a measured mass of hydrogen, or how many grams of product can be obtained from a known mass of reactant. These types of problems are solved using **mass-to-mass conversions**. The calculation begins by converting the given mass of a substance into moles using its molar mass. The balanced chemical equation then provides the mole ratio to determine the number of moles of another substance. Finally, the moles are converted back into grams using the molar mass. In this way, mass-to-mass conversions allow the theoretical quantities predicted by a balanced chemical equation to be directly applied in practical laboratory settings.

Mass-to-mass stoichiometry calculations involve determining the mass of one substance in a reaction from the known mass of another. The process proceeds through a series of conversions, as shown in the flowchart. First, the given mass of substance A is converted to moles using its molar mass. Next, the balanced chemical equation provides the mole ratio, which is used to calculate the number of moles of substance B. Finally, the moles of substance B are converted back into grams using its molar mass. This sequence ensures that stoichiometric relationships expressed in the balanced equation can be applied to laboratory-scale quantities measured in grams.

Moving from **mole–mass calculations** to mass–mass calculations requires only a small extension. When the starting point is the mass of a substance rather than the number of moles, the mass must first be converted into moles using the substance's molar mass as the conversion factor. Next, the balanced chemical equation provides the mole ratio needed to determine the number of moles of another substance. Finally, those moles are converted back into mass using the molar mass of the second substance. The sequence of steps is therefore (Figure 7.6.1):

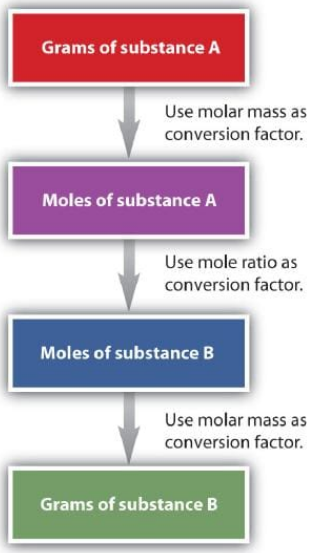


Figure 7.6.1: A flowchart illustrating how to convert from the mass of substance A to the mass of substance B. The process involves four steps: (1) convert grams of substance A to moles using its molar mass, (2) use the mole ratio from the balanced equation to convert to moles of substance B, and (3) convert those moles to grams of substance B using its molar mass. (CC BY-NC-SA 3.0; anonymous via The Basics of General, Organic, and Biological Chemistry v. 1.0). Detailed Description of Figure 7.6.1.

Example 7.6.1

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$\mathrm{MgCl_2\,(aq) + 2\,NaOH\,(aq) \longrightarrow Mg(OH)_2\,(s) + 2\,NaCl\,(aq)}$$

Solution

We must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart (Figure 7.6.2):

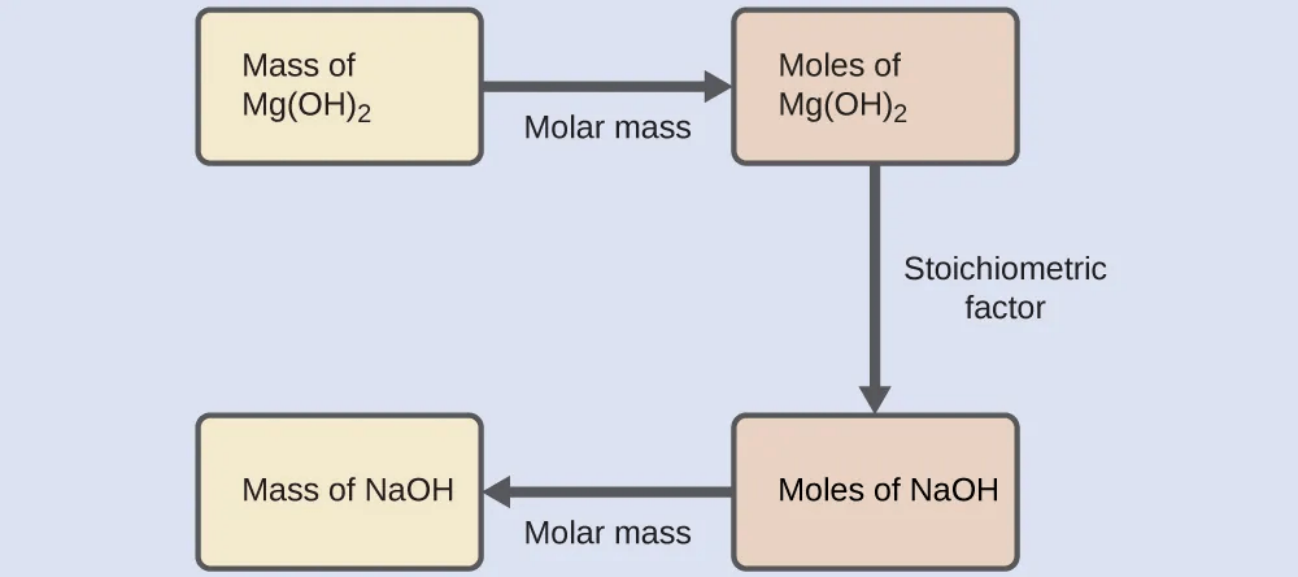


Figure 7.6.2: This flowchart illustrates the mass of Mg(OH)2 is converted to the mass of NaOH by first converting to moles, applying the stoichiometric ratio, and then converting back to mass. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.6.2.

Sixteen grams of $Mg(OH)_2$ is converted to moles, related to NaOH using the 1:2 stoichiometric ratio, and then converted back to grams. The result is 22 g of NaOH.

$$16g\mathrm{Mg(OH)_2} \times \frac{1mol\mathrm{Mg(OH)_2}}{58.3g\mathrm{Mg(OH)_2}} \times \frac{2\mathrm{mol\,NaOH}}{1\,\mathrm{mol\,Mg(OH)_2}} \times \frac{40.0\,\mathrm{g\,NaOH}}{\mathrm{mol\,NaOH}} = 22\,\mathrm{g\,NaOH}$$

Example 7.6.2

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$\mathrm{2\,C_8H_{18} + 25\,O_2 \longrightarrow 16\,CO_2 + 18\,H_2O}$$

Solution

To determine the mass of oxygen gas consumed in the combustion of 702 g of octane (C_8H_{18}), the problem is solved using a sequence of stoichiometric steps based on the given balanced chemical equation. The calculations required are outlined in this flowchart (Figure 7.6.3):

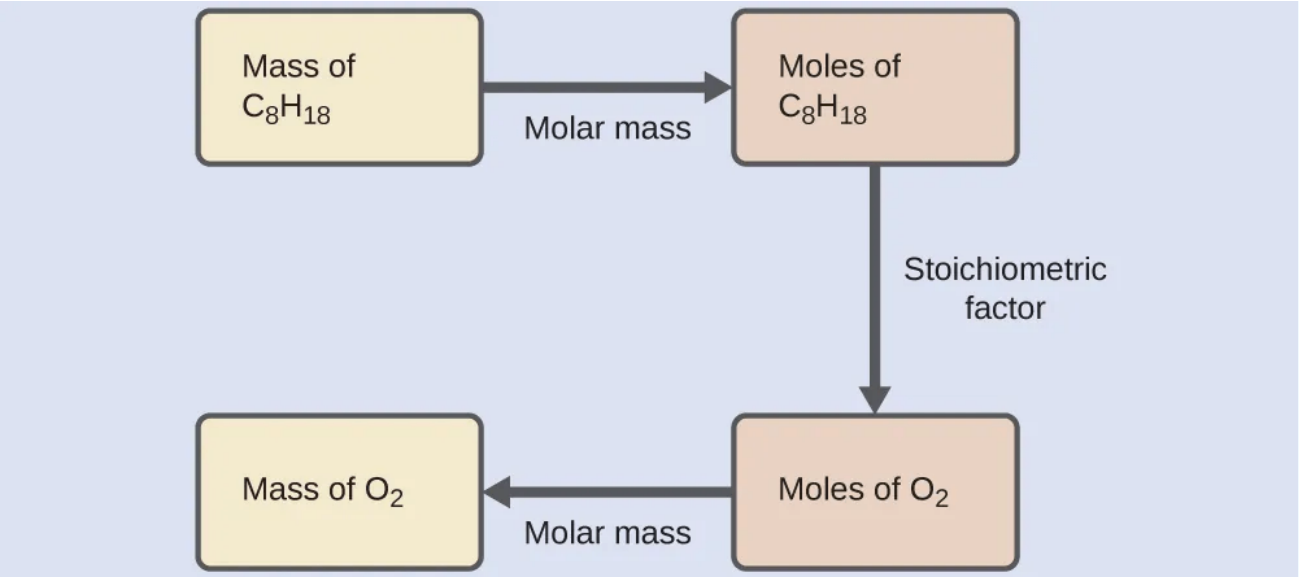


Figure 7.6.3: This diagram shows how to convert between the mass of octane (C_8H_{18}) and the mass of oxygen (O_2) using molar mass and stoichiometric factors, with moles serving as the link between reactants. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.6.3.

To determine the mass of oxygen gas consumed in the combustion of 702 g of octane (C_8H_{18}), the calculation proceeds step by step using stoichiometric relationships from the balanced chemical equation. The result indicates that 702 g of octane requires approximately 2.46×10^3 g of oxygen for complete combustion.

$$702gC_8H_{18} \times \frac{1molC_8H_{18}}{114.23gG_8H_{18}} \times \frac{25molO_2}{2molC_8H_{18}} \times \frac{32.00gO_2}{molO_2} = 2.46 \times 10^3 gO_2$$

Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure 7.6.4 provides a general outline of the various computational steps associated with many reaction **stoichiometry calculations**.

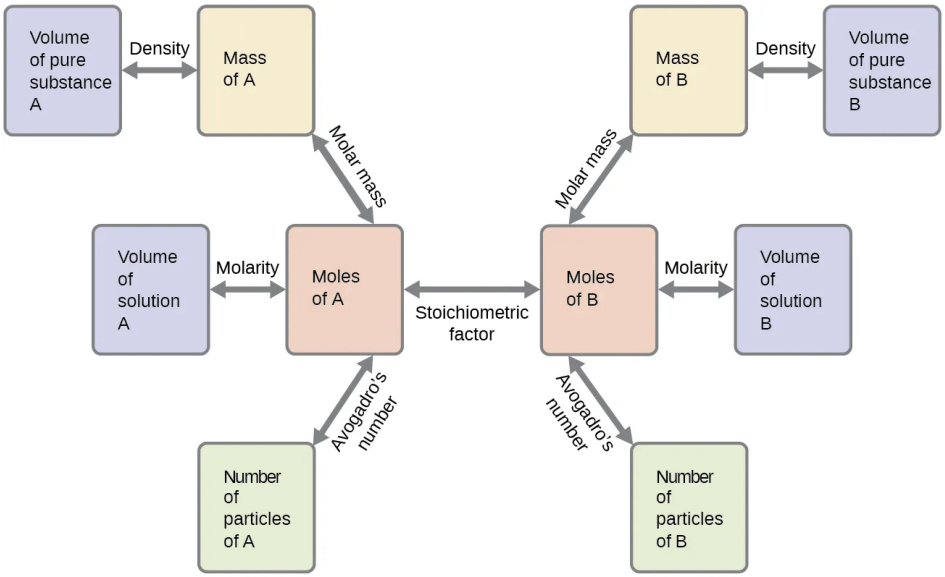


Figure 7.6.4: This flowchart shows how different quantities—mass, moles, volume, and particles—of two reacting substances (A and B) are interconnected. All conversions are linked through the central mole relationship, using molar mass, molarity, density, Avogadro's number, and stoichiometric factors. (CC BY 4.0; anonymous via OpenStax). Detailed Description of Figure 7.6.4.

Link to Learning: Airbags

Airbags (Figure 7.6.5) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2\,\mathrm{NaN_3\,(s)} \longrightarrow 3\,\mathrm{N_2\,(g)} + 2\,\mathrm{Na\,(s)}$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.



Figure 7.6.5: Airbag Inflation Through Chemical Reaction. The image shows deployed airbags inside a car, demonstrating how sodium azide (NaN_3) rapidly decomposes during a collision to produce nitrogen gas, inflating the airbags within fractions of a second to protect passengers. (CC BY 4.0; Jon Seidman) via OpenStax).

Key Takeaways

- Mass-to-mass conversions require first converting the known mass of one substance into moles, then using the mole ratio from the balanced equation to find moles of another substance.
- The final step is converting those moles back into mass using molar mass, allowing stoichiometry to predict how much of a substance is produced or consumed.
- All stoichiometric calculations, whether involving mass, volume, or particles, are connected through the central concept of moles, which act as the bridge between different measurable quantities using molar mass, molarity, density, Avogadro's number, and stoichiometric factors.

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7.7: Reaction Yields

Limiting Reactant

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as **stoichiometric amounts.** All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Consider a food analogy: making grilled cheese sandwiches (Figure 7.7.1):

$$1 \text{ slice of cheese } + 2 \text{ slices of bread } \longrightarrow 1 \text{ sandwich}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in excess.

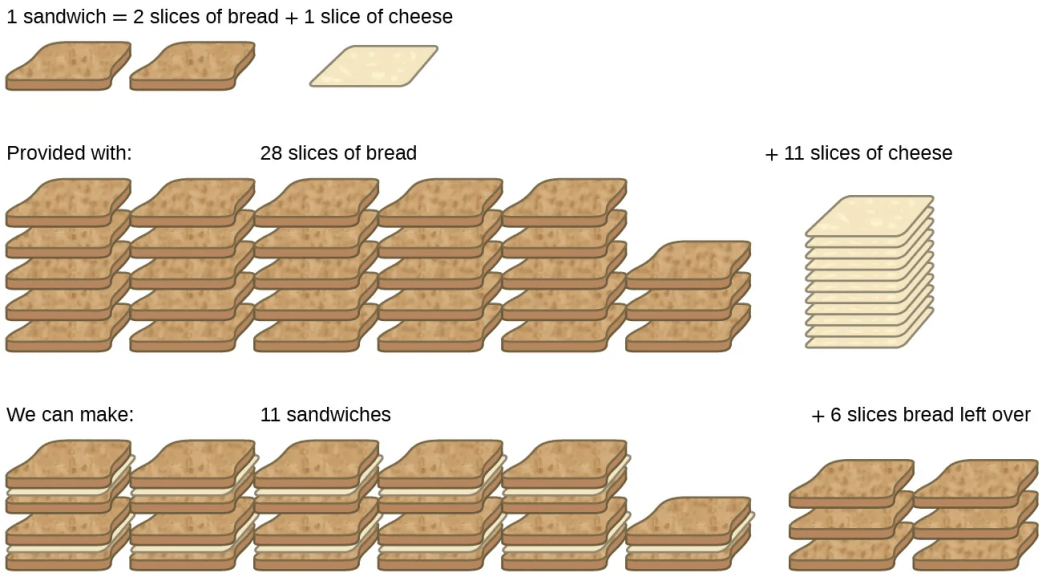


Figure 7.7.1: Limiting Ingredient Analogy with Sandwiches. To make one sandwich, two slices of bread and one slice of cheese are required. Given 28 slices of bread and 11 slices of cheese, only 11 sandwiches can be made, leaving 6 slices of bread unused. This illustrates the concept of a limiting ingredient in determining the maximum product. (CC BY 4.0; anonymous via OpenStax).

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$\mathrm{H_2\,(g) + Cl_2\,(g) \longrightarrow 2\,HCl\,(g)}$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the **limiting reactant**, and the other substance(s) is the **excess reactant**. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen unreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete

reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 3\,\text{mol H}_2 \times \frac{2\,\text{mol HCl}}{1\,\text{mol H}_2} = 6\,\text{mol HCl}$$

Complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 2\,\text{mol Cl}_2 \times \frac{2\,\text{mol HCl}}{1\,\text{mol Cl}_2} = 4\,\text{mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be unreacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure 7.7.2).

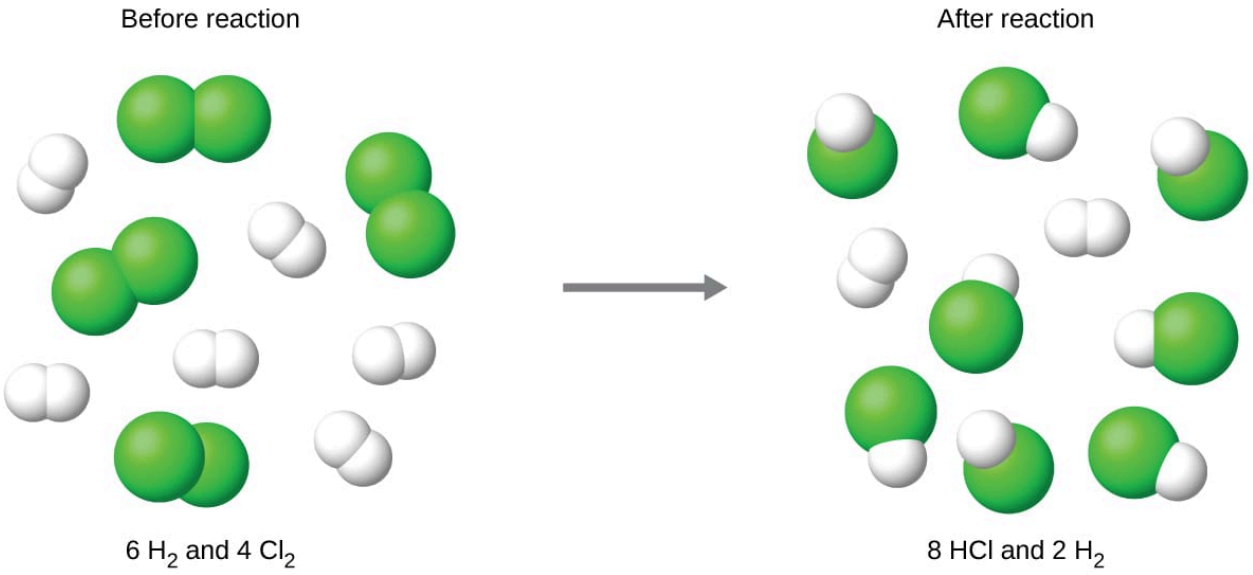


Figure 7.7.2: When hydrogen and chlorine are combined in nonstoichiometric amounts, the limiting reactant determines the maximum amount of HCl produced. In this example, hydrogen is in excess, and chlorine is the limiting reactant. (CC BY 4.0; CFCF via Wikimedia Commons). Detailed Description of Figure 7.7.2.

How to Identify the Limiting Reactant (Limiting Reagent)

There are two ways to determine the limiting reactant. One method is to find and compare the mole ratio of the reactants used in the reaction (Approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the smallest amount of product is the limiting reactant (Approach 2). This section will focus more on the second method.

Approach 1 (The "Reactant Mole Ratio Method"): Find the limiting reactant by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting (excess) reactant.

Approach 2 (The "Product Method"): Find the limiting reactant by calculating and comparing the amount of product that each reactant will produce.

1. Balance the chemical equation for the chemical reaction.

2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reactant.
5. The reactant that produces a larger amount of product is the excess reactant.
6. To find the amount of remaining excess reactant, subtract the mass of excess reactant consumed from the total mass of excess reactant given.

Example 7.7.1

As an example, consider the balanced equation

$$4\,C_2H_3Br_3 + 11\,O_2 \rightarrow 8\,CO_2 + 6\,H_2O + 6\,Br_2$$

What is the limiting reactant if 76.4 grams of $C_2H_3Br_3$ reacted with 49.1 grams of O_2?

Solution

Using Approach 1 (Reactant Mole Ratio Method):

Step 1: Balance the chemical equation.

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br

Step 2: Convert all given information into moles.

$$76.4\ \cancel{\text{g C}_2\text{H}_3\text{Br}_3} \times \frac{1\ \text{mol C}_2\text{H}_3\text{Br}_3}{266.72\ \cancel{\text{g C}_2\text{H}_3\text{B}_3}} = 0.286\ \text{mol C}_2\text{H}_3\text{Br}_3$$

$$49.1\ \cancel{\text{g O}_2} \times \frac{1\ \text{mol O}_2}{32.00\ \cancel{\text{g O}_2}} = 1.53\ \text{mol O}_2$$

Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.

Assuming that all of the oxygen is used up,

$$1.53\ \cancel{\text{mol O}_2} \times \frac{4\ \text{mol C}_2\text{H}_3\text{Br}_3}{11\ \cancel{\text{mol O}_2}} = 0.556\ \text{mol } C_2H_3Br_3 \text{ are required.}$$

Because 0.556 moles of $C_2H_3Br_3$ required > 0.286 moles of $C_2H_3Br_3$ available, $C_2H_3Br_3$ is the limiting reactant.

Using Approach 2 (Product Method):

Step 1: Balance the chemical equation.

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br_2

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$76.4\ \cancel{\text{g C}_2\text{H}_3\text{Br}_3} \times \frac{1\ \cancel{\text{mol C}_2\text{H}_3\text{Br}_3}}{266.72\ \cancel{\text{g C}_2\text{H}_3\text{Br}_3}} \times \frac{8\ \cancel{\text{mol CO}_2}}{4\ \cancel{\text{mol C}_2\text{H}_3\text{Br}_3}} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}} = 25.2\ \text{g CO}_2$$

$$49.1\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{8\ \cancel{\text{mol CO}_2}}{11\ \cancel{\text{mol O}_2}} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}} = 49.1\ \text{g CO}_2$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Therefore, by either method, $C_2H_3Br_3$ is the limiting reactant.

Example 7.7.2

Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$$3Si(s) + 2N_2(g) \longrightarrow Si_3N_4(s)$$

Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

Solution

Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$molSi = 2.00gSi \times \frac{1molSi}{28.09gSi} = 0.0712molSi$$

$$molSi = 2.00gSi \times \frac{1molSi}{28.09gSi} = 0.0712molSi$$

The provided Si:N_2 molar ratio is:

$$\frac{0.0712molSi}{0.0535molN_2} = \frac{1.33molSi}{1molN_2}$$

The stoichiometric Si:N_2 ratio is:

$$\frac{3molSi}{2molN_2} = \frac{1.5molSi}{1molN_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$molSi_3N_4 \text{ produced } = 0.0712molSi \times \frac{1molSi_3N_4}{3molSi^2} = 0.0237molSi_3N_4$$

while the 0.0535 moles of nitrogen would produce

$$\text{mol Si } N_4 \text{ produced } = 0.0535molN \times \frac{1molSi_3N_4}{2molN_2} = 0.0268molSi_3N_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant

Identifying the Limiting Reactant and the Mass of Excess Reactant

When two or more reactants are combined in a chemical reaction, they are not always present in the exact stoichiometric ratio required by the balanced chemical equation. In such cases, one reactant will be completely consumed first, preventing the reaction from continuing even if other reactants remain in excess. This completely used-up reactant is called the limiting reactant, because it limits the amount of product that can form. The other reactant(s), which are present in greater quantity than required, are called excess reactants. Identifying the limiting reactant is essential for accurately predicting the maximum theoretical yield of a reaction. Once the limiting reactant is determined, the amount of the excess reactant that remains unreacted can also be calculated by comparing the actual amount present to the amount required by the stoichiometric ratios. This process is fundamental in laboratory and industrial settings, where controlling reactant usage helps maximize efficiency and minimize

waste.

Example 7.7.3

For example, in the reaction of magnesium metal and oxygen, calculate the mass of magnesium oxide that can be produced if 2.40 g Mg reacts with 10.0 g O_2. Also determine the amount of excess reactant. MgO is the only product in the reaction.

Solution

Following Approach 1:

Step 1: Balance the chemical equation.

$$2\ \mathrm{Mg\ (s)} + \mathrm{O_2\ (g)} \rightarrow 2\ \mathrm{MgO\ (s)}$$

The balanced equation provides the relationship of 2 mol Mg to 1 mol O_2 to 2 mol MgO

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$2.40\ \cancel{\mathrm{g\ Mg}} \times \frac{1\ \cancel{\mathrm{mol\ Mg}}}{24.31\ \cancel{\mathrm{g\ Mg}}} \times \frac{2\ \cancel{\mathrm{mol\ MgO}}}{2\ \cancel{\mathrm{mol\ Mg}}} \times \frac{40.31\ \mathrm{g\ MgO}}{1\ \cancel{\mathrm{mol\ MgO}}} = 3.98\ \mathrm{g\ MgO}$$

$$10.0\ \cancel{\mathrm{g\ O_2}} \times \frac{1\ \cancel{\mathrm{mol\ O_2}}}{32.00\ \cancel{\mathrm{g\ O_2}}} \times \frac{2\ \cancel{\mathrm{mol\ MgO}}}{1\ \cancel{\mathrm{mol\ O_2}}} \times \frac{40.31\ \mathrm{g\ MgO}}{1\ \cancel{\mathrm{mol\ MgO}}} = 25.2\ \mathrm{g\ MgO}$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Mg produces less MgO than does O_2 (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reactant in this reaction.

Step 5: The reactant that produces a larger amount of product is the excess reactant.

O_2 produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore O_2 is the excess reactant in this reaction.

Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reactant consumed from the total mass of excess reactant given.

Mass of excess reactant calculated using the limiting reactant:

$$2.40\ \cancel{\mathrm{g\ Mg}} \times \frac{1\ \cancel{\mathrm{mol\ Mg}}}{24.31\ \cancel{\mathrm{g\ Mg}}} \times \frac{1\ \cancel{\mathrm{mol\ O_2}}}{2\ \cancel{\mathrm{mol\ Mg}}} \times \frac{32.00\ \mathrm{g\ O_2}}{1\ \cancel{\mathrm{mol\ O_2}}} = 1.58\ \mathrm{g\ O_2}$$

or

Mass of excess reactant calculated using the mass of the product:

$$3.98\ \cancel{\mathrm{g\ MgO}} \times \frac{1\ \cancel{\mathrm{mol\ MgO}}}{40.31\ \cancel{\mathrm{g\ MgO}}} \times \frac{1\ \cancel{\mathrm{mol\ O_2}}}{2\ \cancel{\mathrm{mol\ MgO}}} \times \frac{32.0\ \mathrm{g\ O_2}}{1\ \cancel{\mathrm{mol\ O_2}}} = 1.58\ \mathrm{g\ O_2}$$

Mass of total excess reactant given – mass of excess reactant consumed in the reaction:

10.0g O_2 - (available) 1.58g O_2 (used) = 8.42g O_2 (excess)

Therefore, 8.42g O_2 is in excess.

Theoretical Yield, Actual Yield, and Percent Yield

Consider the making of grilled cheese sandwiches as a food analogy. Suppose 3 sandwiches are accidentally burned, leaving only 8 sandwiches. In a chemical reaction analogy, the 8 sandwiches represent the actual yield

In chemical reactions, the concepts of theoretical yield, actual yield, and percent yield are important for understanding efficiency, and these can be illustrated with a grilled cheese sandwich analogy. The **theoretical yield** represents the maximum number of sandwiches that could be made if all ingredients were used perfectly; for example, with 28 slices of bread and 11 slices of cheese, the theoretical yield is 11 sandwiches. The **actual yield** is the number of sandwiches actually produced and edible after preparation; if 3 sandwiches are accidentally burned, only 8 sandwiches remain, making the actual yield 8. In general, yield refers to the amount of product obtained from a process, which can be expressed as either the theoretical or actual yield. The **percent yield** compares the actual yield to the theoretical yield, providing a measure of efficiency. In this example, the percent yield is

$$\text{percent yield } = \frac{8 \text{ sandwiches (actual yield)}}{11 \text{ sandwiches (theoretical yield)}} \times 100\%$$

meaning only about 73% of the possible sandwiches were successfully made. This analogy mirrors how chemists evaluate reaction efficiency in the laboratory (Figure 7.7.3).

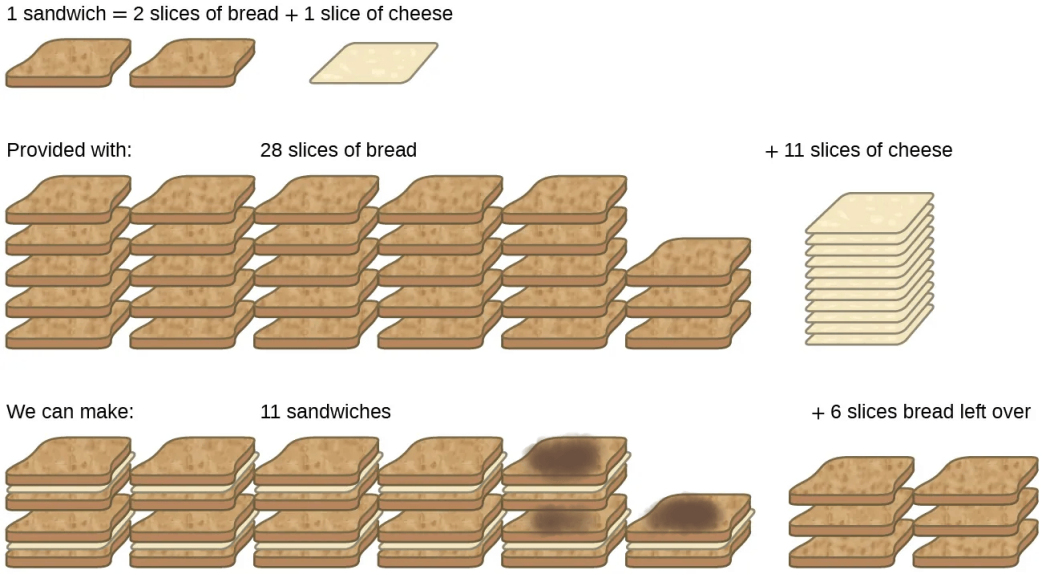


Figure 7.7.3: Illustration of a grilled cheese sandwich analogy: although 11 sandwiches could be made in theory (theoretical yield), 3 are burned, leaving only 8 edible sandwiches (actual yield). (CC BY 4.0; modification of the anonymous work, modified by Gorkem Ermut via OpenStax). Detailed Description of Figure 7.7.3.

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the theoretical yield of the reaction. In practice, the amount of product obtained is called the actual yield, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this chapter). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its percent yield:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

> Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when the percent yield is calculated.

Example 7.7.4

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\mathrm{CuSO_4\,(aq) + Zn\,(s) \longrightarrow Cu\,(s) + ZnSO_4\,(aq)}$$

What is the percent yield?

Solution

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

$$1.274\,\mathrm{g\,CuSO_4} \times \frac{1\,\mathrm{mol\,CuSO_4}}{159.62\,\mathrm{g\,CuSO_4}} \times \frac{1\,\mathrm{mol\,Cu}}{1\,\mathrm{mol\,CuSO_4}} \times \frac{63.55\,\mathrm{g\,Cu}}{1\,\mathrm{mol\,Cu}} = 0.5072\,\mathrm{g\,Cu}$$

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated (via Equation to be

$$\begin{aligned}\text{percent yield} &= \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100\% \\ &= \left(\frac{0.392\,\mathrm{g\,Cu}}{0.5072\,\mathrm{g\,Cu_2}}\right) \times 100\% \\ &= 77.3\%\end{aligned}$$

Example 7.7.5

78.0 grams of Na_2O_2 is reacted with 29.4 grams of H_2O and 43.8 grams of NaOH is produced according to the equation below:

$$\mathrm{Na_2O_2\,(s) + 2\,H_2O\,(l) \longrightarrow 2\,NaOH\,(aq) + H_2O_2\,(l)}$$

Determine the theoretical yield and identify the limiting reactant. What is the percent yield?

Solution

To determine the limiting reactant, the given masses of both reactants are first converted into moles. The molar mass of each reactant is used as the conversion factor, and the stoichiometric coefficients from the balanced chemical equation provide the relationship between reactants and the product, sodium hydroxide (NaOH). By calculating how much NaOH could be formed from each reactant, it is possible to identify which reactant produces the smaller amount of product. That reactant is the limiting reactant because it controls the maximum possible yield of NaOH.

For sodium peroxide (Na_2O_2):

$$78.0\,\mathrm{g\,Na_2O_2} \times \frac{1\,\mathrm{mol\,Na_2O_2}}{77.96\,\mathrm{g\,Na_2O_2}} \times \frac{2\,\mathrm{mol\,NaOH}}{1\,\mathrm{mol\,Na_2O_2}} \times \frac{40\,\mathrm{g\,NaOH}}{1\,\mathrm{mol\,NaOH}} = 80.0\,\mathrm{g\,NaOH}$$

For water (H_2O):

$$29.4\,\mathrm{g\,H_2O} \times \frac{1\,\mathrm{mol\,H_2O}}{18.02\,\mathrm{g\,H_2O}} \times \frac{2\,\mathrm{mol\,NaOH}}{2\,\mathrm{mol\,Na_2O_2}} \times \frac{40\,\mathrm{g\,NaOH}}{1\,\mathrm{mol\,NaOH}} = 65.3\,\mathrm{g\,NaOH}$$

Comparing the two results shows that water produces the smaller amount of NaOH (65.3 g). Therefore, H_2O is the limiting reactant, and the theoretical yield of NaOH is 65.3 grams.

The percent yield compares the actual mass of product obtained to the theoretical mass predicted by stoichiometry. In this case, the actual mass of NaOH collected was 43.8 g. Dividing the actual yield by the theoretical yield and multiplying by 100 gives the percent yield:

$$\text{Percent yield} = \frac{43.8\ \text{g}}{65.3\ \text{g}} \times 100\% = 67.2$$

This indicates that 67.2% of the expected NaOH was recovered, with the remaining percentage representing losses due to side reactions, incomplete reaction, or experimental error.

Key Takeaways

- Limiting Reactant: The reactant that runs out first in a reaction, stopping the process and limiting the amount of product formed.
- Excess Reactant: The reactant(s) left over after the reaction is complete.
- Finding the Limiting Reactant: Compare mole ratios to the balanced equation or calculate how much product each reactant could make—the one that makes less is the limiting reactant.
- Theoretical Yield: The maximum amount of product that should form if everything reacts perfectly.
- Actual Yield: The real amount of product you get in the lab, usually less than the theoretical yield.
- Percent Yield: A measure of reaction efficiency, showing how close the actual yield is to the theoretical yield.

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- 8.6: Limiting Reactant and Theoretical Yield by Henry Agnew, Marisa Alviar-Agnew, Sarick Shah is licensed CC BY-NC-SA 3.0.

7.8: Summary

Chapter Summary

Chemical reactions involve relationships between quantities of reactants and products, which chemists describe using the mole unit. One mole corresponds to 6.022×10^{23} particles (Avogadro's number), whether they are atoms, ions, or molecules. This number is important because it connects the microscopic world of atoms with measurable laboratory quantities. The molar mass of a substance, obtained by summing the atomic masses of its constituent elements from the periodic table, provides the bridge between the mass of a sample (in grams) and the number of moles it contains.

Because one mole of a substance has a definite mass, we can set up mole-mass conversions to convert between moles and grams of a substance. Balanced chemical equations express the relationships between reactants and products in terms of moles, allowing us to perform stoichiometric calculations. Using these mole ratios, we can relate the number of moles of one substance to the moles or mass of another. A mole-mass calculation connects the moles of one substance to the mass of another, while a mass-mass calculation directly relates the masses of two different substances in a reaction.

In real reactions, not all reactants are consumed completely. The limiting reactant is the substance that is used up first and therefore limits the amount of product that can form. The excess reactant is present in greater quantity than needed and remains after the reaction. From the limiting reactant, we calculate the theoretical yield, which is the maximum possible amount of product predicted by stoichiometry. However, experimental conditions often prevent the theoretical yield from being fully achieved. The amount of product actually obtained in the laboratory is the actual yield. To evaluate efficiency, chemists calculate the percent yield as:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Together, these concepts—mole relationships, molar mass, stoichiometric conversions, limiting reactants, and yield calculations—form the foundation for quantitative chemical analysis and are essential for both theoretical predictions and practical laboratory work.

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7.9: Key Terms

- **Mole:** A number of things equal to 6.022×10^{23} items.
- **Avogadro's number:** The value 6.022×10^{23}.
- **Atomic mass unit:** An atomic mass unit (amu) is a unit of mass defined as one-twelfth the mass of a carbon-12 atom, used to express the masses of atoms and subatomic particles.
- **Gram mass unit:** A gram mass unit is the mass of one mole of particles (atoms, molecules, or ions) expressed in grams, numerically equal to the particle's average atomic or molecular mass in amu.
- **Molar mass:** The mass of 1 mol of atoms or molecules.
- **Formula mass:** Formula mass is the sum of the average atomic masses of all atoms represented in the formula of a substance, expressed in atomic mass units (amu).
- **Stoichiometry:** Stoichiometry is the study of the quantitative relationships between reactants and products in a chemical reaction, based on balanced chemical equations.
- **Stoichiometric factor:** A stoichiometric factor is a conversion ratio derived from the coefficients of a balanced chemical equation, used to relate the amounts (in moles) of different reactants or products.
- **Mass-to-mass conversions:** A stoichiometry calculation converting between the mass of one substance and the mass of a different substance in a chemical reaction.
- **Mole-to-mass calculations:** A stoichiometry calculation converting between masses and moles of different substances in a chemical reaction.
- **Stoichiometry calculations:** A stoichiometry calculation converting between masses and moles of different substances in a chemical reaction.
- **Stoichiometric amounts:** Stoichiometric amounts are the exact quantities of reactants needed, in the correct mole ratio from the balanced chemical equation, so that all reactants are completely consumed with no excess.
- **Limiting reactant:** The limiting reactant is the reactant that is completely consumed first in a chemical reaction, thereby limiting the amount of product that can be formed.
- **Excess reactant:** The excess reactant is the substance in a chemical reaction that remains after the limiting reactant is completely used up, because it is present in more than the required stoichiometric amount.
- **Theoretical yield:** Theoretical yield is the maximum amount of product that can be formed from given quantities of reactants, calculated using stoichiometry and assuming complete reaction with no losses.
- **Actual yield:** Actual yield is the measured amount of product obtained from a chemical reaction under real laboratory conditions.
- **Percent yield:** Percent yield is a measure of how efficiently a chemical reaction produces product, expressed as the ratio of the actual yield to the theoretical yield in percentage form.

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7.10: Alternative Text for Chapter 7

Figure 7.2.2: One Mole of Some Familiar Substances

A laboratory setup on a white background showing two shallow ceramic dishes and two clear glass bottles with black screw caps. The left dish contains a pale yellow powder of sulfur, while the right dish holds a bright red-orange powder of mercury(II) iodide. One glass bottle is partially filled with a transparent liquid resembling water, and the other bottle is empty. The colorful powders and clear liquids illustrate common laboratory chemical samples.

Figure 7.3.1: Relationship Between Atomic Weight, Atomic Mass, and Molar Mass. Hydrogen's atomic weight is 1.008 amu

Image of the periodic table of elements arranged by groups (columns) and periods (rows). Each element is displayed with its atomic number, symbol, name, and average atomic mass. Hydrogen (H) is highlighted in the top-left corner with atomic number 1 and atomic mass 1.008. An enlarged callout box at the bottom right explains hydrogen's entry: atomic number at the top, symbol (H) in red in the center, atomic mass (1.008) below the symbol, and the element name (hydrogen) at the bottom. The image includes a color code legend: metals in beige, metalloids in gray, and nonmetals in light blue, with additional shading to indicate physical state at room temperature—solid, liquid, or gas. Two red arrows at the bottom indicate how the lanthanide series (La–Lu) and actinide series (Ac–Lr) are positioned below the main table for compactness.

Figure 7.3.2: One Mole Samples of Different Elements

Image showing eight circular dishes on a black background, each containing a sample of a pure element corresponding to one mole of atoms (6.022×10^{23} atoms). The top row displays, from left to right: 65.4 g of zinc (gray powder), 12.0 g of carbon (black powder), 24.3 g of magnesium (dark metallic shavings), and 63.5 g of copper (coil of reddish-brown wire). The bottom row displays, from left to right: 32.1 g of sulfur (yellow powder), 28.1 g of silicon (shiny gray crystalline pieces), 207 g of lead (dark granular powder), and 118.7 g of tin (cluster of silvery metallic granules). The image illustrates that while each sample contains exactly one mole of atoms, the visible masses differ according to the atomic weights of the elements.

Figure 7.3.3: Calculating the Molecular Mass of $CHCl_3$

A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "C," "H," "C l" and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers "1," "1," and "3" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "12.01," "1.008," and "35.45" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "12.01," "1.008," "106.35," and "119.37." There is a thick black line below the number 106.35. The merged cell under the first five columns reads "Molecular mass."

Figure 7.3.4: Calculating the molar mass of $C_2H_5O_2N$

Table showing the step-by-step calculation of the molecular mass of a compound composed of carbon, hydrogen, oxygen, and nitrogen. The table has four columns: Element, Quantity (mol element per mol compound), Molar Mass (g/mol element), and Subtotal (g/mol compound). For carbon (C), the quantity is 2, molar mass is 12.01, giving a subtotal of 24.02 g/mol. For hydrogen (H), the quantity is 5, molar mass is 1.008, subtotal 5.040 g/mol. For oxygen (O), the quantity is 2, molar mass is 16.00, subtotal 32.00 g/mol. For nitrogen (N), the quantity is 1, molar mass is 14.007, subtotal 14.007 g/mol. At the bottom, the totals are summed to give a molecular mass of 75.07 g/mol compound.

Figure 7.3.5: Calculating the molar mass of $Al_2(SO_4)_3$

Table showing the step-by-step calculation of the formula mass of aluminum sulfate, $Al_2(SO_4)_3$. The table has four columns labeled Element, Quantity, Average atomic mass (amu), and Subtotal (amu). For aluminum (Al), the quantity is 2, the average atomic mass is 26.98 amu, subtotal is 53.96 amu. For sulfur (S), the quantity is 3, the average atomic mass is 32.06 amu, subtotal is 96.18 amu. For oxygen (O), the quantity is 12, the average atomic mass is 16.00 amu, subtotal is 192.00 amu. At the bottom, the formula mass is calculated as 342.14 amu, which is the sum of all subtotals.

Figure 7.4.1: A Simple Flowchart for Converting between Mass and Moles of a Substance

A diagram illustrates the relationship between moles of a substance and mass of a substance using conversion factors. On the left, a box labeled “Moles of Substance” is shown, and on the right, a box labeled “Mass of Substance.” A right-pointing arrow connects the two boxes, labeled “convert into” with the instruction “Use g/mol conversion factor” above it. A left-pointing arrow goes in the opposite direction, from “Mass of Substance” to “Moles of Substance,” labeled “convert into” with the instruction “Use mol/g conversion factor” below it. This diagram demonstrates that multiplying by grams per mole converts moles to mass, and multiplying by moles per gram converts mass to moles.

Figure 7.4.2: Relationship between moles and mass of aluminum

A diagram shows the relationship between moles of aluminum and mass of aluminum. On the left, a box labeled “moles of Al” is connected by a right-pointing arrow labeled “molar mass” to a box on the right labeled “mass of Al.” Below the arrow, a curved arrow pointing from left to right indicates the conversion factor “26.98 g Al / 1 mole Al,” showing that multiplying the number of moles by the molar mass gives the mass in grams.

Figure 7.4.4: Converting Mass to Number of Atoms

A flowchart showing the steps to convert mass of copper atoms into number of atoms. The first box, labeled “Mass of Cu atoms (g),” points to the second box, “Moles of Cu atoms (mol),” with the note “Divide by molar mass (g/mol).” The second box points to the third box, “Number of Cu atoms,” with the note “Multiply by Avogadro’s number (mol^{-1}).”

Figure 7.4.5: Converting Mass to Number of Molecules

A flowchart illustrates the conversion from the mass of a compound to the number of molecules. On the left, a yellow box labeled “Mass of $C_7H_5NO_3S$ (g)” points with an arrow to the middle box, “Moles of $C_7H_5NO_3S$ (mol).” The arrow is labeled “Divide by molar mass (g/mol).” From the middle box, another arrow points right to a peach-colored box labeled “Number of $C_7H_5NO_3S$ molecules.” This arrow is labeled “Multiply by Avogadro’s number (mol^{-1}).”

Figure 7.5.1: Relationship between moles of H_2O and moles of O_2

A flowchart illustrates the stoichiometric relationship between water and oxygen. On the left, a purple box labeled “moles H_2O” points with an arrow labeled “stoichiometric ratio” to an orange box labeled “moles O_2.” Below the arrow, a curved arrow shows the conversion factor “1 mol O_2 / 2 mol H_2O,” indicating that two moles of water produce one mole of oxygen.

Figure 7.5.2: Converting Between Moles and Mass in Stoichiometry

A vertical flowchart illustrates the steps for converting between moles of one substance and grams of another in a chemical reaction. At the top, a red box is labeled “Moles of substance A.” A gray downward arrow points to a purple box labeled “Moles of substance B,” with the annotation on the right reading “Use mole ratio as conversion factor.” Another gray arrow continues downward to a blue box labeled “Grams of substance B,” with the annotation on the right reading “Use molar mass as conversion factor.” The diagram shows that calculations in stoichiometry proceed by first converting moles of one substance to moles of another using the mole ratio from the balanced chemical equation, and then converting moles to grams using molar mass.

Figure 7.5.3: Calculating the Number of Product Molecules Generated by a Reaction Using Stoichiometry

A horizontal flowchart shows the conversion from moles of propane (C_3H_8) to molecules of carbon dioxide (CO_2). On the left, a box labeled “Moles of C_3H_8” is connected by a right-pointing arrow labeled “Stoichiometric factor” to a middle box labeled “Moles of CO_2.” Another right-pointing arrow, labeled “Avogadro’s number,” connects this box to a final box on the right labeled “Molecules of CO_2.” The diagram demonstrates that the stoichiometric ratio from the balanced equation is first used to convert from moles of propane to moles of carbon dioxide, and then Avogadro’s number is applied to convert moles of carbon dioxide into molecules.

Figure 7.6.1: Sequence of Steps in a Mass-to-Mass Stoichiometry Calculation

A vertical flowchart illustrates the steps for performing a mass-to-mass stoichiometric calculation. At the top, a box labeled “Grams of substance A” points downward with a arrow. To the right of the arrow, the text reads “Use molar mass as conversion factor.” The arrow leads to a box labeled “Moles of substance A.” Another arrow continues downward to a box labeled “Moles of substance B,” with the note on the right reading “Use mole ratio as conversion factor.” A third arrow points downward from this box to a box at the bottom labeled “Grams of substance B,” accompanied by the note “Use molar mass as conversion factor.” The diagram shows the complete sequence: starting with grams of a known substance, converting to moles, applying the mole ratio from the balanced chemical equation, and then converting back to grams of the desired substance.

Figure 7.6.2: Mass to Mass

A flowchart illustrates the steps in a mass-to-mass stoichiometric calculation between magnesium hydroxide, $Mg(OH)_2$, and sodium hydroxide, NaOH. The sequence begins with a tan box labeled “Mass of $Mg(OH)_2$.” A right-pointing arrow labeled “Molar mass” leads to another tan box labeled “Moles of $Mg(OH)_2$.” From this box, a downward arrow labeled “Stoichiometric factor” connects to a tan box labeled “Moles of NaOH.” Finally, a left-pointing arrow labeled “Molar mass” connects to the last tan box at the bottom left, labeled “Mass of NaOH.” The diagram demonstrates the process: convert the mass of $Mg(OH)_2$ to moles, use the stoichiometric ratio from the balanced chemical equation to find moles of NaOH, and then convert those moles into the mass of NaOH.

Figure 7.6.3: Stoichiometric Calculations

A flowchart illustrates the steps for solving a mass-to-mass stoichiometry problem involving the combustion of octane (C_8H_{18}) and oxygen gas (O_2). The chart begins with a box labeled "Mass of C_8H_{18}," which connects by a right-pointing arrow labeled "Molar mass" to a box labeled "Moles of C_8H_{18}." From this box, a downward arrow labeled "Stoichiometric factor" points to another box labeled "Moles of O_2." Finally, a left-pointing arrow labeled "Molar mass" connects this box to a box labeled "Mass of O_2." The diagram shows the stepwise sequence: convert the mass of octane to moles using molar mass, use the stoichiometric ratio from the balanced equation to determine moles of oxygen required, and then convert those moles to the mass of oxygen. This highlights how stoichiometric calculations connect masses of reactants in a chemical reaction.

Figure 7.6.4: Stoichiometry Pathways for Substances A and B

A comprehensive stoichiometry flowchart showing the relationships between different quantities of two reacting substances, labeled A and B. At the center, two boxes labeled “Moles of A” and “Moles of B” are connected by a horizontal arrow marked “Stoichiometric factor.” From “Moles of A,” arrows branch out in multiple directions: upward to “Mass of A” (with the arrow labeled “Molar mass”), downward to “Number of particles of A” (arrow labeled “Avogadro’s number”), leftward to “Volume of solution A” (arrow labeled “Molarity”), and diagonally to “Volume of pure substance A” via “Mass of A” (arrow labeled “Density”). Similarly, “Moles of B” connects upward to “Mass of B” (labeled “Molar mass”), downward to “Number of particles of B” (labeled “Avogadro’s number”), rightward to “Volume of solution B” (labeled “Molarity”), and diagonally to “Volume of pure substance B” via “Mass of B” (labeled “Density”). The diagram demonstrates how stoichiometric calculations can move between moles, mass, volume, and number of particles for two substances involved in a chemical reaction, with all conversions linked back to the central mole relationship.

Figure 7.7.2: Limiting Reactant in the Formation of HCl.

The image is a molecular-level illustration showing the limiting reactant concept for the reaction between hydrogen (H_2) and chlorine (Cl_2) to form hydrogen chloride (HCl). On the left, under the label *"Before reaction"*, there are six white diatomic hydrogen molecules (H_2) and four green diatomic chlorine molecules (Cl_2). On the right, under the label *"After reaction"*, the diagram shows that all four Cl_2 molecules have reacted with four H_2 molecules to produce eight hydrogen chloride molecules (HCl), each represented by one green chlorine atom bonded to one white hydrogen atom. Two H_2 molecules remain unreacted, demonstrating that chlorine was the limiting reactant while hydrogen was in excess. Text beneath the diagrams states "6 H_2 and 4 Cl_2" for the reactants and "8 HCl and 2 H_2" for the products. An arrow in the center indicates the progression from reactants to products.

Figure 7.7.3: Grilled Cheese Sandwich

Illustration of a grilled cheese analogy for theoretical vs. actual yield. To make a sandwich, 2 slices of bread and 1 slice of cheese are needed. Provided are 28 slices of bread and 11 slices of cheese, enough to make 11 sandwiches with 6 bread slices left over. The bottom row shows 11 sandwiches made, but 3 of them are visibly burned, leaving only 8 good sandwiches as the actual yield.

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8: ENERGY AND CHEMICAL PROCESSES

Introduction

Energy is a vital component of the world around us. Nearly every physical and chemical process, including all the chemical reactions discussed in previous chapters, occurs with a simultaneous energy change. In this chapter, we will explore the nature of energy and how energy and chemistry are related.

Learning Objectives

- Define energy and heat, and relate calories to nutrition and exercise.
- Relate heat transfer to temperature change, understanding temperature conversions for Celsius, Fahrenheit and Kelvin.
- Determine the heat associated with a phase changes.
- Define bond energy and calculate enthalpy change or heat of reaction and determine if a chemical process is exothermic or endothermic.
- Relate the concept of energy change to chemical reactions that occur in the body.

8.1: Prelude to Energy and Chemical Processes

8.2: Energy, Energy Basics and Its Units

8.3: Heat and Temperature

8.4: Phase Changes

8.5: Bond Energies and Chemical Reactions

8.6: The Energy of Biochemical Reactions

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8.9: Alternative text for Chapter 8

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8.1: Prelude to Energy and Chemical Processes

Overview

The chemistry of the human body, or any living organism, is very complex. Even so, the chemical reactions found in the human body follow the same principles of energy that other chemical reactions follow. Metabolism is the collective term for the chemical reactions that occur in cells and provide energy to keep cells alive.

Some of the energy from metabolism is in the form of heat, and some animals use this heat to regulate their body temperatures. Such warm-blooded animals are called endotherms. Humans are in fact endotherms. In endotherms, problems with metabolism can lead to fluctuations in body temperature. When humans get sick, for instance, our body temperatures can rise higher than normal; we develop a fever. When food is scarce (especially in winter), some endotherms go into a state of controlled decreased metabolism called hibernation. During hibernation, the body temperatures of these endotherms actually decrease. In hot weather or when feverish, endotherms will pant or sweat to rid their bodies of excess heat. Table 8.1.1 shows body temperatures of select endotherms.

Table 8.1.1: Average Body Temperatures of Selected Endotherms

Endotherm	Body Temperature	
	(°F)	(°C)
bird	up to 110	up to 43.5
cat	101.5	38.6
dog	102	38.9
horse	100.0	37.8
human	98.6	37.0
pig	102.5	39.2

Ectotherms, sometimes called cold-blooded animals, do not use the energy of metabolism to regulate body temperature. Instead, they depend on external energy sources, such as sunlight. Fish, for example, will seek out water of different temperatures to regulate body temperature. The amount of energy available is directly related to the metabolic rate of the animal. When energy is scarce, ectotherms may also hibernate.

The connection between metabolism and body temperature is a reminder that energy and chemical reactions are intimately related. A basic understanding of this relationship is especially important when those chemical reactions occur within our own bodies.

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8.2: Energy, Energy Basics and Its Units

Overview

Chemical changes and their accompanying changes in energy are important parts of our everyday world (Figure 8.2.1). The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating, and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.



Figure 8.2.1: Energy in our Daily Lives. The energy involved in chemical changes is important to our daily lives: (a) Food provides the energy you need to get through the rest of the day. (CC BY SA 4.0; via Wikimedia Commons). (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school. (CC BY 2.0; via Wikimedia Commons). and (c) Coal provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily. (CC BY-SA 4.0; European Union via Wikimedia Commons). Detailed Description of Figure 8.2.1.

Over 90% of the energy we use comes originally from the sun. Every day, the sun provides the earth with almost 10,000 times the amount of energy necessary to meet all of the world's energy needs for that day. Our challenge is to find ways to convert and store incoming solar energy so that it can be used in reactions or chemical processes that are both convenient and nonpolluting. Plants and many bacteria capture solar energy through photosynthesis. We release the energy stored in plants when we burn wood or plant products such as ethanol. We also use this energy to fuel our bodies by eating food that comes directly from plants or from animals that got their energy by eating plants. Burning coal and petroleum also releases stored solar energy: These fuels are fossilized plant and animal matter.

Energy

So what really is Energy? **Energy** can be defined as the ability to supply heat or do work. You can understand what this means by thinking about yourself when you feel "energetic." You feel ready to go—to jump up and get something done. When you have a lot of energy, you can perform a lot of work. By contrast, if you do not feel energetic, you have very little desire to do much of anything. This description is not only applicable to you but also to all physical and chemical processes. The quantity of work that can be done is related to the quantity of energy available to do it.

Like matter, energy comes in different types and can generally be classified as follows:

- **Potential energy:** the energy an object has because of its relative position, composition, or condition
- **Kinetic energy:** the energy that an object possesses because of its motion.

Water at the top of a waterfall or dam has potential energy because of its position; when it flows downward through generators, it has kinetic energy that can be used to do work and produce electricity in a hydroelectric plant (Figure 8.2.2). A battery has potential energy because the chemicals within it can produce electricity that can do work.



Figure 8.2.2: Energy Conversion. (a) At the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As the water falls, some of its potential energy is converted into kinetic energy. (CC BY 2.0, Steve Jurvetson via flickr) (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy. (CC BY-SA 4.0; Aleksandar Nikolic via Wikimedia Commons). Detailed Description of Figure 8.2.2.

Energy can be converted from one form into another, but all of the energy present before a change occurs always exists in some form after the change is completed. The combustion engine is one example. Figure 8.2.3 shows a photo of a combustion engine which converts chemical energy into mechanical energy, when gasoline is combusted in the cylinder's of the car's engine. This ultimately results in moving the pistons which make the wheels turn.



Figure 8.2.3: Chemical energy (a type of potential energy) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car's engine, the rapidly expanding gaseous products of this chemical reaction generate mechanical energy (a type of kinetic energy) when they move the cylinders' pistons. (CC BY-SA 4.0; Xlory via Wikimedia Commons). Detailed Description of Figure 8.2.3.

This observation is expressed in the law of conservation of energy:

> During a chemical or physical change, energy can be neither created nor destroyed, it can only be changed from one form to another.

This is also one version of the first law of thermodynamics. Energy can be transferred from one object to another if the objects have different temperatures. The transfer of energy due to temperature differences is called **heat.** Heat is usually released or absorbed, but sometimes the conversion involves light, electrical energy, or some other form of energy. For example, if you hold an ice cube in your hand, the ice cube slowly melts as energy in the form of heat is transferred from your hand to the ice. As your hand loses energy, it starts to feel cold.

Units of Energy

Because of their interrelationships, energy, work, and heat have the same units and are measured in one of two common units: the calorie and the joule.

- **The Joule:** The SI unit of energy, work, and heat is **the joule (J)**. A joule is a small amount of energy. For example, it takes about 4 J to warm 1 mL of H_2O by 1°C. Many processes occur with energy changes in thousands of joules, so the **kilojoule (kJ)** is also common.
- **The Calorie:** Another unit of energy, used widely in the health professions and everyday life, is **the calorie (cal)**. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A calorie is the quantity of heat required to raise the temperature of 1 gram of water by $1^{\circ}\mathrm{C}$. For example, raising the temperature of $100\ \mathrm{g}$ of water from $20^{\circ}\mathrm{C}$ to $22^{\circ}\mathrm{C}$ would require $100 \times 2 = 200\ \mathrm{cal}$.

Calories contained within food are actually **kilocalories (kcal)**. In other words, if a certain snack contains 85 food calories, it actually contains $85\ \mathrm{kcal}$ or $85,000\ \mathrm{cal}$. In order to make the distinction, the dietary calorie is written with a capital C.

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories}$$

To say that the snack "contains" 85 Calories means that $85\ \mathrm{kcal}$ of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. We can use this relationship to convert quantities of energy, work, or heat from one unit to another.The conversion between a joule and a calorie is shown below.

$$1\ \mathrm{J} = 0.2390\ \mathrm{cal}\ \text{or}\ 1\ \mathrm{cal} = 4.184\ \mathrm{J}$$

Although the joule is the proper SI unit for energy, we will focus on the use of the calorie or the kilocalorie (or Calorie) in this chapter because they are widely used by health professionals

Example 8.2.1: Energy Converions

The energy content of a single serving of bread is 70.0 Cal. What is the energy content in calories? In joules?

Solution

This is a simple conversion-factor problem. Using the relationship 1 Cal = 1,000 cal, we can answer the first question with a one-step conversion:

$$70.0\ \mathrm{Cal} \times \frac{1,000\ \mathrm{cal}}{1\ \mathrm{Cal}} = 70,000\ \mathrm{cal}$$

Then we convert calories into joules

$$70,000\ \mathrm{cal} \times \frac{4.184\ \mathrm{J}}{1\ \mathrm{cal}} = 293,000\ \mathrm{J}$$

and then kilojoules

$$293,000\ \text{J} \times \frac{1\ \text{kJ}}{1,000\ \text{J}} = 293\ \text{kJ}$$

The energy content of bread comes mostly from carbohydrates.

Link to Learning: Food Labels

The calorie is used in nutrition to express the energy content of foods. However, because a calorie is a rather small quantity, nutritional energies are usually expressed in **kilocalories (kcal)**, also called **Calories (capitalized; Cal)**. For example, a candy bar may provide 120 Cal (nutritional calories) of energy, which is equal to 120,000 cal. Figure 8.2.4 shows an example of a nutrition label.



Figure 8.2.4: Nutritional Energy. A sample nutrition facts label, with instructions from the U.S. Food and Drug Administration. (CC BY-SA 4.0; via Wikimedia Commons). Detailed Description of Figure 8.2.4.

The Caloric Content of Food

The caloric content of foods is determined by analyzing the food for protein, carbohydrate, fat, water, and "minerals" and then calculating the caloric content using the average values for each component that produces energy (**9 Cal/g for fats, 4 Cal/g for carbohydrates and proteins, and 0 Cal/g for water and minerals**). An example of this approach is shown in Table 8.2.1 for a slice of roast beef.

Table 8.2.1: Approximate Composition and Fuel Value of an 8 oz Slice of Roast Beef

Composition	Calories
97.5 g of water	× 0 Cal/g = 0
58.7 g of protein	× 4 Cal/g = 235
69.3 g of fat	× 9 Cal/g = 624
0 g of carbohydrates	× 4 Cal/g = 0
1.5 g of minerals	× 0 Cal/g = 0
Total mass: 227.0 g	**Total calories: about 900 Cal**

The compositions and caloric contents of some common foods are given in Table 8.2.2.

Table 8.2.2: Approximate Compositions and Fuel Values of Some Common Foods

Food (quantity)	Approximate Composition (%)				Food Value (Cal/g)	Calories
	Water	Carbohydrate	Protein	Fat		
beer (12 oz)	92	3.6	0.3	0	0.4	150
coffee (6 oz)	100	~0	~0	~0	~0	~0
milk (1 cup)	88	4.5	3.3	3.3	0.6	150
egg (1 large)	75	2	12	12	1.6	80
butter (1 tbsp)	16	~0	~0	79	7.1	100
apple (8 oz)	84	15	~0	0.5	0.6	125
bread, white (2 slices)	37	48	8	4	2.6	130
brownie (40 g)	10	55	5	30	4.8	190
hamburger (4 oz)	54	0	24	21	2.9	326
fried chicken (1 drumstick)	53	8.3	22	15	2.7	195
carrots (1 cup)	87	10	1.3	~0	0.4	70

Because the Calorie represents such a large amount of energy, a few of them go a long way.

> An average 73 kg (160 lb) person needs about 67 Cal/h (1600 Cal/day) to fuel the basic biochemical processes that keep that person alive.

This energy is required to maintain body temperature, keep the heart beating, power the muscles used for breathing, carry out chemical reactions in cells, and send the nerve impulses that control those automatic functions. Physical activity increases the amount of energy required but not by as much as many of us hope. A moderately active individual requires about 2500–3000 Cal/day; athletes or others engaged in strenuous activity can burn 4000 Cal/day. Any excess caloric intake is stored by the body for future use, usually in the form of fat, which is the most compact way to store energy. When more energy is needed than the diet supplies, stored fuels are mobilized and oxidized. We usually exhaust the supply of stored carbohydrates before turning to fats, which accounts in part for the popularity of low-carbohydrate diets.

Link to Learning: Energy Expenditure

Most health professionals agree that exercise is a valuable component of a healthy lifestyle. Exercise not only strengthens the body and develops muscle tone but also expends energy. After obtaining energy from the foods we eat, we need to expend that energy somehow, or our bodies will store it in unhealthy ways (e.g., fat). Like the energy content in food, the energy

expenditures of exercise are also reported in kilocalories, usually kilocalories per hour of exercise. These expenditures vary widely, from about 440 kcal/h for walking at a speed of 4 mph to 1,870 kcal/h for mountain biking at 20 mph. Table 8.2.3 lists the energy expenditure for a variety of exercises.

Table 8.2.3: Energy Expenditure of a 180-Pound Person during Selected Exercises

Exercise	Energy Expended (kcal/h)
aerobics, low-level	325
basketball	940
bike riding, 20 mph	830
golfing, with cart	220
golfing, carrying clubs	425
jogging, 7.5 mph	950
racquetball	740
skiing, downhill	520
soccer	680
walking upstairs	1,200
yoga	280

Because some forms of exercise use more energy than others, anyone considering a specific exercise regimen should consult with his or her physician first.

Key Takeaways

- Energy is the ability to do work.
- Heat is the transfer of energy due to temperature differences.
- Energy and heat are expressed in units of joules.

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8.3: Heat and Temperature

Thermal Energy, Temperature, and Heat

The concept of temperature may seem familiar, but many people confuse temperature with heat.

- **Temperature (T)** is a measure of how hot or cold an object is relative to another object (its thermal energy content)
- **Heat (q)** is the flow of thermal energy between objects with different temperatures.

When we touch a hot object, energy flows from the hot object into our fingers, and we perceive that incoming energy as the object being "hot." Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as "cold." In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

Thermal energy is kinetic energy associated with the random motion of atoms and molecules. When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE), and we say that the object is "hot." When the atoms and molecules are moving slowly, they have lower average KE, and we say that the object is "cold" as shown in (Figure 8.3.1).

Temperature is a quantitative measure of "hot" or "cold." Assuming that no chemical reaction or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. And, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

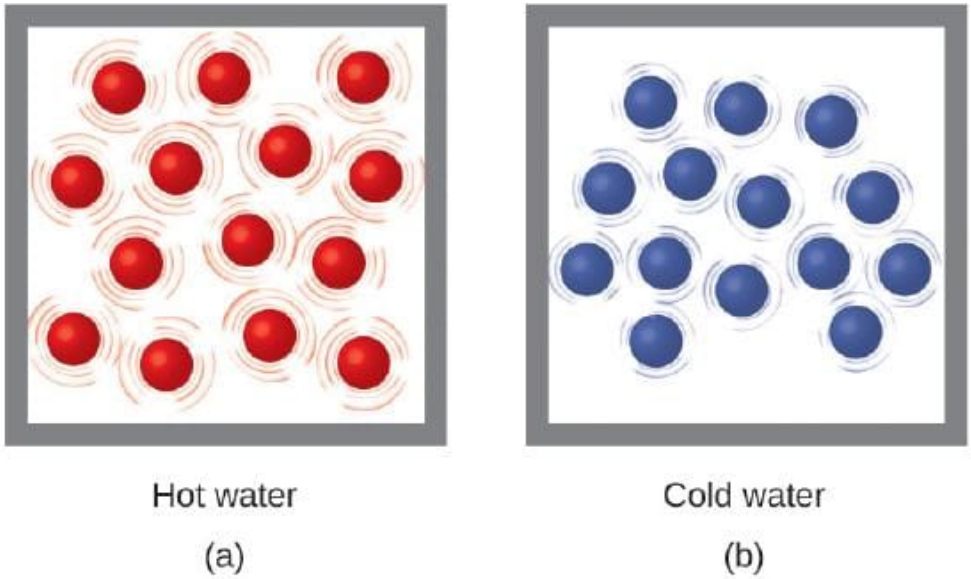


Figure 8.3.1: (a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water. (CC BY 4.0; via OpenStax). Detailed Description of Figure 8.3.1.

Link to Learning: The Effects of Temperature on Molecular Motion

Click on this interactive simulation to view the effects of temperature on molecular motion. Click on the "states" icon to study motion of molecules at each state and how they respond to changes in temperature, and click on the "phase changes" icon to visualize what happens to molecules during a phase change, and how they respond to changes in temperature, pressure and amount based on their phase diagrams.

Heat flow increases the thermal energy of one body and decreases the thermal energy of the other. Suppose we initially have

a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature (Figure 8.3.2).

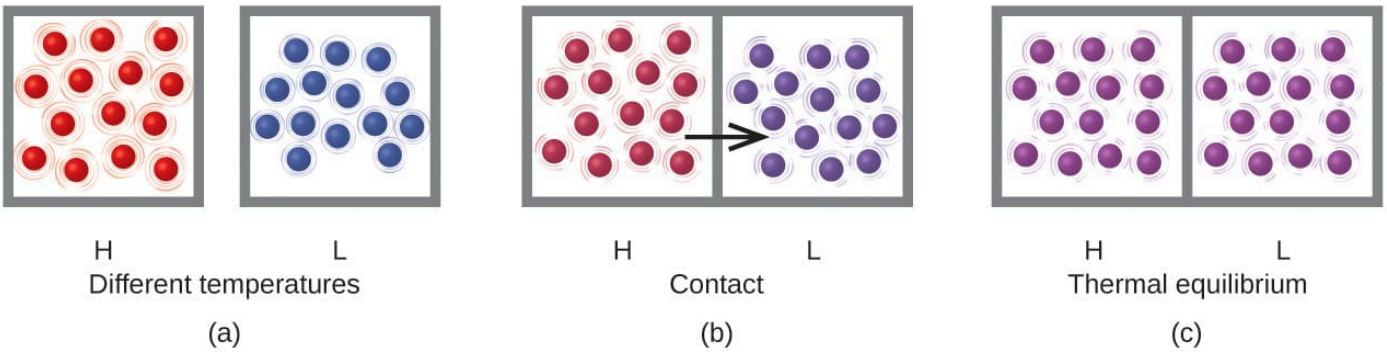


Figure 8.3.2: (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they contact each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach "thermal equilibrium" when both substances are at the same temperature and their molecules have the same average kinetic energy. (CC BY 4.0; via Openstax). Detailed Description of Figure 8.3.2.

Heat (*q*) is the transfer of thermal energy between two bodies at different temperatures.

Link to Learning: Energy Forms and Changes

Click on the **PhET simulation** to explore energy forms and changes. Visit the Energy Systems tab to create combinations of energy sources, transformation methods, and outputs. Click on Energy Symbols to visualize the transfer of energy.

Temperature Scales

As we mentioned above, temperature is a measure of the average amount of energy of motion, or kinetic energy, a system contains. Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure 8.3.3.

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. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

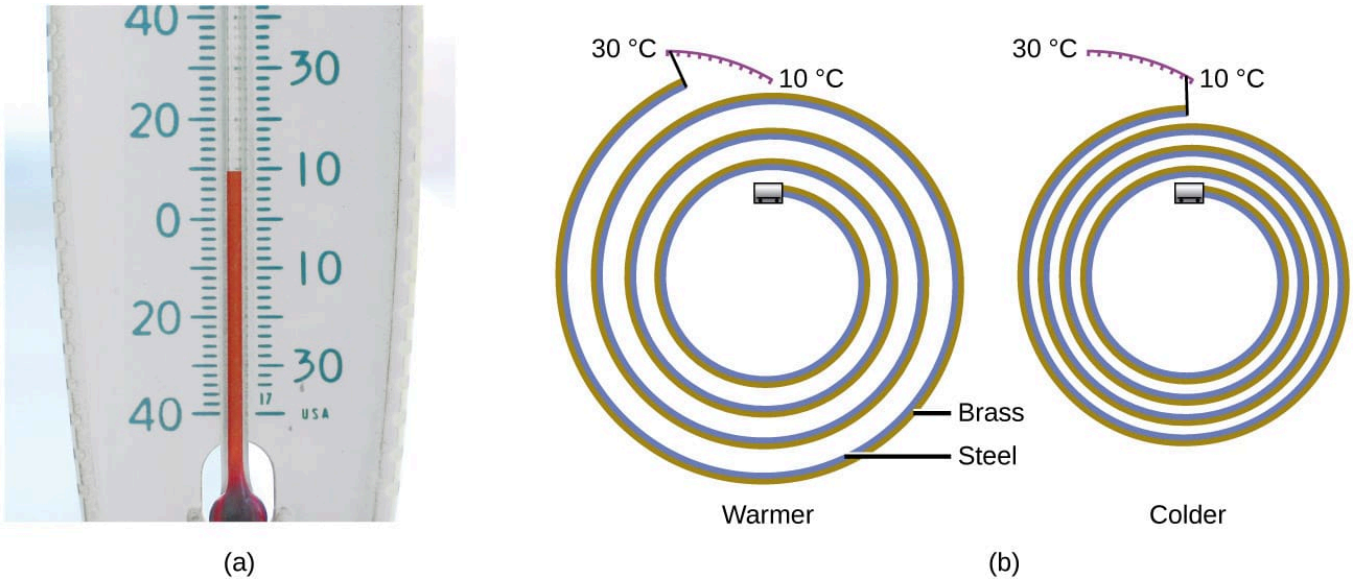


Figure 8.3.3: (a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature. (CC BY 4.0; via Openstax). Detailed Description of Figure 8.3.3.

Temperatures are expressed using scales that use units called degrees. There are three scales used for reporting temperatures:

- **Fahrenheit (expressed as °F)**
- **Celsius (°C)**
- **Kelvin (K)**

Figure 8.3.4 compares the three temperature scales.

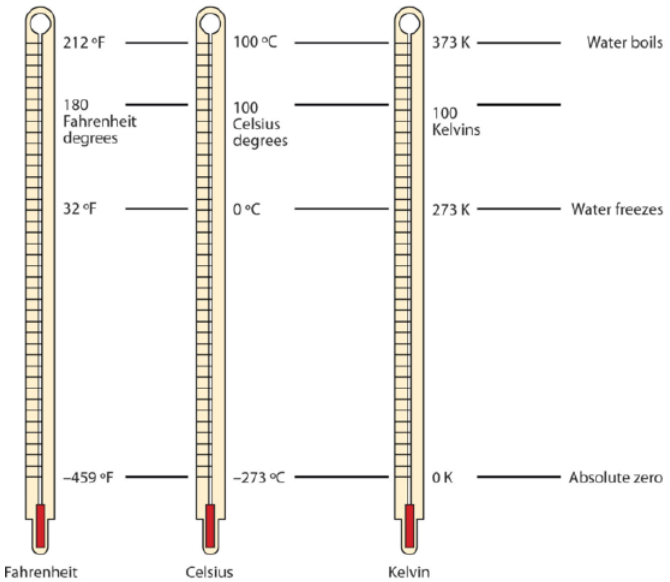


Figure 8.3.4: A comparison of the three temperature scales showing three thermometers. These thermometers have a red-colored liquid which contains alcohol and food coloring. Silver looking thermometers contain mercury, which is a neurotoxin. (CC BY-NC-SA 3.0; Basics of General, Organic, and Biological Chemistry (Ball et al.)). Detailed Description of Figure 8.3.4.

In the United States, the commonly used temperature scale is the Fahrenheit scale (symbolized by °F and spoken as “degrees Fahrenheit”). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32 °F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212 °F.

However, science and scientists use other scales to express temperature.

For example, the Celsius scale (symbolized by °C and spoken as “degrees Celsius”) defines 0°C as the freezing point of water and 100 °C as the boiling point of water. This scale is divided into 100 divisions between these two landmarks and extended higher and lower as well. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}C = (^{\circ}F - 32) \times \frac{5}{9} \tag{8.3.1}$$

$$^{\circ}F = \left(^{\circ}C \times \frac{9}{5}\right) + 32 \tag{8.3.2}$$

The fundamental unit of temperature in SI is the Kelvin (K). The Kelvin temperature scale (note that the name of the scale capitalizes the word Kelvin, but the unit itself is lowercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$K =^{\circ} C + 273.15 \tag{8.3.3}$$

> Note that science uses the Celsius and Kelvin scales almost exclusively

The reason that the Kelvin scale is defined this way is that there exists a minimum possible temperature called absolute zero (zero Kelvin). The Kelvin temperature scale is set so that 0 K is absolute zero, and the temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example 8.3.1.

Virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.

Example 8.3.1: Temperature Conversions

a. What is 98.6°F in degrees Celsius?
b. What is 25.0°C in degrees Fahrenheit?

Solution

a. Using Equation 8.3.1, we have

$$\begin{aligned} ^{\circ}C &= (98.6 - 32) \times \frac{5}{9} \\ &= 66.6 \times \frac{5}{9} \\ &= 37.0^{\circ}C \end{aligned}$$

b. Using Equation 8.3.2, we have

$$\begin{aligned} ^{\circ}F &= (25.0 \times \frac{9}{5}) + 32 \\ &= 45.0 + 32 \\ &= 77.0^{\circ}F \end{aligned}$$

Note that the Kelvin scale does not use the word degrees; a temperature of 295 K is spoken of as “two hundred ninety-five kelvin” and not “two hundred ninety-five degrees Kelvin.”

Example 8.3.2: Room Temperature

If the normal room temperature is 72.0°F, what is room temperature in degrees Celsius and kelvin?

Solution

First, we use Equation 8.3.1 to determine the temperature in degrees Celsius:

$$\begin{aligned} {}^{\circ}C &= (72.0 - 32) \times \frac{5}{9} \\ &= 40.0 \times \frac{5}{9} \\ &= 22.2^{\circ}C \end{aligned}$$

Then we use Equation 8.3.3 to determine the temperature in the Kelvin scale:

$$\begin{aligned} K &= 22.2^{\circ}C + 273.15 \\ &= 295.4K \end{aligned}$$

So, room temperature is about 295 K.

Health Application: Body temperature

Normal body temperature is defined as being 98.6°F (+/- 1.0°F). To determine body temperature, thermometers can be placed inside or on the surface of the body. The two best methods of obtaining body temperature are placing the thermometer either under the tongue or inside the rectum. Typically, children are capable of holding a thermometer in their mouths around the age of four (have fun before that age).

Fever is defined as body temperature being above 100°F (adults). High fever status occurs at and above 104°F. For adults, these adults should seek medical attention immediately if the fever exceeds 104°F. For children, these values are much lower.

Hyperthermia (inability of the body to regulate heat) occurs when normal body temperature is exceeded. Conditions that can cause hyperthermia are fever (infection), heat stroke, thyroid disorders, heart attack, or traumatic injury. Medications for cancer, arthritis, and thyroid patients can cause the body temperature to rise. Symptoms of hyperthermia include sweating, confusion, nausea, and dizziness.

Hypothermia (exposure to cold environments) occurs when normal body temperature dips below 95.0°F. When this occurs, the affected person(s) should seek immediate medical attention. During hypothermia, the body has problems producing heat. Medical conditions like diabetes, infection, and thyroid dysfunction can cause hypothermia. Watch this video of American Marines attempting to survive extreme environments. Symptoms of this condition involve shivering, confusion, and sluggish behavior.

Treating hyperthermia could involve hydrating a patient. If a fever is due to infection, analgesics (like Tylenol, Advil, aspirin, or Aleve) can help alleviate fever. Placing someone in a cool bath can also relieve symptoms.

To combat hypothermia, one must remove wet clothing, redress in warm materials, and participate in physical movement.

Heat Transfer

Heat is a familiar manifestation of energy. When we touch a hot object, energy flows from the hot object into our fingers, and we

perceive that incoming energy as the object being “hot.” Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as “cold.” In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

Suppose we consider the transfer of heat from the opposite perspective—namely, what happens to a system that gains or loses heat? Generally, the system’s temperature changes. (We will address a few exceptions later.) The greater the original temperature difference, the greater the transfer of heat, and the greater the ultimate temperature change. The relationship between the amount of heat transferred and the temperature change can be written as

$$\text{heat} \propto \Delta T \tag{8.3.4}$$

where $\propto$ means “is proportional to” and **ΔT is the change in temperature of the system**. Any change in a variable is always defined as “the final value minus the initial value” of the variable, so **ΔT is ($T_{final} - T_{initial}$)**. In addition, the greater the mass of an object, the more heat is needed to change its temperature. We can include a variable representing mass (m) to the proportionality as follows:

$$\text{heat} \propto m\Delta T \tag{8.3.5}$$

To change this proportionality into an equality, we include a proportionality constant, c. The proportionality constant is called the specific heat and is commonly symbolized by c as we will discuss below:

$$\text{heat} = mc\Delta T \tag{8.3.6}$$

Heat Capacity

We now introduce two concepts useful in describing heat flow and temperature change.

The heat capacity (*C*) of a body of matter is the quantity of heat (q) it absorbs or releases when it experiences a temperature change (ΔT) of 1 degree Celsius (or equivalently, 1 Kelvin):

$$C = \frac{q}{\Delta T} \tag{8.3.7}$$

Heat capacity is determined by both **the type and amount of substance** that absorbs or releases heat. It is therefore an extensive property—its value is proportional to the amount of the substance.

The specific heat (c) of a substance is the amount of energy that must be transferred to or from 1 g of that substance to change its temperature by 1°C. Every substance has a characteristic specific heat, which is reported in units of cal/g•°C or cal/g•K, depending on the units used to express ΔT. Table $8.3.1$ lists the specific heats for various materials.

The direction of heat flow is not shown in heat = mc ΔT. If energy goes into an object, the total energy of the object increases, and the values of heat ΔT are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and ΔT are negative.

Water has a relatively high specific heat (about 4.2 J/g °C for the liquid and 2.09 J/g °C for the solid); most metals have much lower specific heats (usually less than 1 J/g °C). The specific heat of a substance varies somewhat with temperature. However, this variation is usually small enough that we will treat specific heat as constant over the range of temperatures that will be considered in this chapter. Specific heats of some common substances are listed in Table $8.3.1$.

Table 8.3.1: Specific Heats of Common Substances at 25 °C and 1 bar

Substance	Symbol (*state*)	Specific Heat (J/g °C)
helium	He(*g*)	5.193
water	$H_2O(l)$	4.184
ethanol	$C_2H_6O(l)$	2.376
ice	H_2O(s)	2.093 (at –10 °C)
water vapor	$H_2O(g)$	1.864
nitrogen	$N_2(g)$	1.040
air		1.007
oxygen	$O_2(g)$	0.918
aluminum	Al(s)	0.897
carbon dioxide	$CO_2(g)$	0.853
argon	Ar(*g*)	0.522
iron	Fe(s)	0.449
copper	Cu(s)	0.385
lead	Pb(s)	0.130
gold	Au(s)	0.129
silicon	Si(s)	0.712

If we know the mass of a substance and its specific heat, we can determine the amount of heat, *q*, entering or leaving the substance by measuring the temperature change before and after the heat is gained or lost:

$$\begin{aligned} q &= (\text{ specific heat }) \times (\text{ mass of substance }) \times (\text{ temperature change }) \\ &= c \times m \times \Delta T \\ &= c \times m \times (T_{\text{final}} - T_{\text{initial}}) \end{aligned} \tag{8.3.8}$$

In this equation, *c* is the specific heat of the substance, *m* is its mass, and ΔT (which is read "delta T") is the temperature change, $T_{\text{final}} - T_{\text{initial}}$. If a substance gains thermal energy, its temperature increases, its final temperature is higher than its initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a positive value, and the value of *q* is positive. If a substance loses thermal energy, its temperature decreases, the final temperature is lower than the initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a negative value, and the value of *q* is negative.

Example 8.3.3: Calculating q

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Solution

We can use Equation 8.3.6 to determine the amount of heat, but first we need to determine ΔT. Because the final temperature of the iron is 73.3°C and the initial temperature is 25.0°C, ΔT is as follows:

$$\begin{aligned} \Delta T = T_{final} - T_{initial} \\ &= 73.3^oC - 25.0^oC \\ &= 48.3^oC \end{aligned}$$

The mass is given as 150.0 g, and Table 8.3.1 gives the specific heat of iron as 0.108 cal/g•°C. Substitute the known values into Equation 8.3.6 and solve for amount of heat:

$$\text{heat} = (150.0\ \text{g})\left(0.108\ \frac{\text{cal}}{\text{g}\cdot{}^\circ\text{C}}\right)(48.3^\circ\text{C}) = 782\ \text{cal}$$

Note how the gram and °C units cancel algebraically, leaving only the calorie unit, which is a unit of heat. Because the temperature of the iron increases, energy (as heat) must be flowing into the metal.

Example 8.3.4: Calculating the Specific Heat of a Substance

A 10.3 g sample of a reddish-brown metal gave off 71.7 cal of heat as its temperature decreased from 97.5°C to 22.0°C. What is the specific heat of the metal? Can you identify the metal from the data in Table 8.3.1?

Solution

The question gives us the heat, the final and initial temperatures, and the mass of the sample. The value of ΔT is as follows:

$$\Delta T = T_{final} - T_{initial} = 22.0^\circ\text{C} - 97.5^\circ\text{C} = -75.5^\circ\text{C}$$

If the sample gives off 71.7 cal, it loses energy (as heat), so the value of heat is written as a negative number, –71.7 cal. Substitute the known values into heat = mcΔT and solve for c:

$$-71.7\ \text{cal} = (10.3\ \text{g})(c)(-75.5^\circ\text{C})$$

$$c = \frac{-71.7\ \text{cal}}{(10.3\ \text{g})(-75.5^\circ\text{C})}$$

$$c = 0.0923\ \text{cal/g}\bullet{}^\circ\text{C}$$

This value for specific heat is very close to that given for copper in Table 8.3.1.

Example 8.3.5: Measuring Heat

A flask containing $8.0\ 10^2$ g of water is heated, and the temperature of the water increases from 21 °C to 85 °C. How much heat did the water absorb?

Solution

To answer this question, consider these factors:

- the specific heat of the substance being heated (in this case, water)
- the amount of substance being heated (in this case, 8.0×10^2 g)
- the magnitude of the temperature change (in this case, from 21 °C to 85 °C).

The specific heat of water is 4.184 J/g °C, so to heat 1 g of water by 1 °C requires 4.184 J. We note that since 4.184 J is required to heat 1 g of water by 1 °C, we will need *800 times as much* to heat 8.0×10^2 g of water by 1 °C. Finally, we observe that since 4.184 J are required to heat 1 g of water by 1 °C, we will need *64 times as much* to heat it by 64 °C (that is, from 21 °C to 85 °C).

This can be summarized using Equation 8.3.8:

$$\begin{aligned} q = c \times m \times \Delta T &= c \times m \times (T_{\text{final}} - T_{\text{initial}}) \\ &= \left(4.184 J/\frac{g}{\circ}{}^{\circ}C\right) \times \left(8.0 \times 10^2 g\right) \times (85 - 21)^{\circ}C \\ &= (4.184 J/g^{\circ}C) \times \left(8.0 \times 10^2 g\right) \times (64)^{\circ}C \\ &= 210,000 J \left(= 2.1 \times 10^2 kJ\right) \end{aligned} \tag{8.3.9}$$

Because the temperature increased, the water absorbed heat and q is positive.

Note that the relationship between heat, specific heat, mass, and temperature change can be used to determine any of these quantities (not just heat) if the other three are known or can be deduced.

Example 8.3.6: Determining Other Quantities

A piece of unknown metal weighs 348 g. When the metal piece absorbs 6.64 kJ of heat, its temperature increases from 22.4 °C to 43.6 °C. Determine the specific heat of this metal (which might provide a clue to its identity).

Solution

Since mass, heat, and temperature change are known for this metal, we can determine its specific heat using the relationship in Equation 8.3.8:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

Substituting the known values:

$$6640 J = c \times (348 g) \times (43.6 - 22.4)^{\circ}C$$

Solving:

$$c = \frac{6640 J}{(348 g) \times (21.2^{\circ}C)} = 0.900 J/g^{\circ}C$$

Comparing this value with the values in Table 8.3.1, this value matches the specific heat of aluminum, which suggests that the unknown metal may be aluminum.

Key Takeaways

- Heat transfer is related to temperature change.
- The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C.

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8.4: Phase Changes

State of Matter

The amount of energy in molecules or atoms determines **the state of matter** which depends on the balance between the *kinetic energy* of the individual particles and the attractive forces that exist between them. These forces are called **intermolecular forces (IMF)**. The kinetic energy which is a function of the temperature of the substance, keeps the molecules apart and moving around, whereas the intermolecular forces draw the particles together.

A phase is a certain state of matter that includes a specific set of physical properties (Figure 8.4.1). That is, the atoms, the molecules, or the ions that make up the phase do so in a consistent manner throughout the phase. There are three stable phases of matter: gas, liquid, or solid. Matter exits in one of these different phases.

- In the **gas** phase, atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- In the **liquid** phase, atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- In the **solid** phase, atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

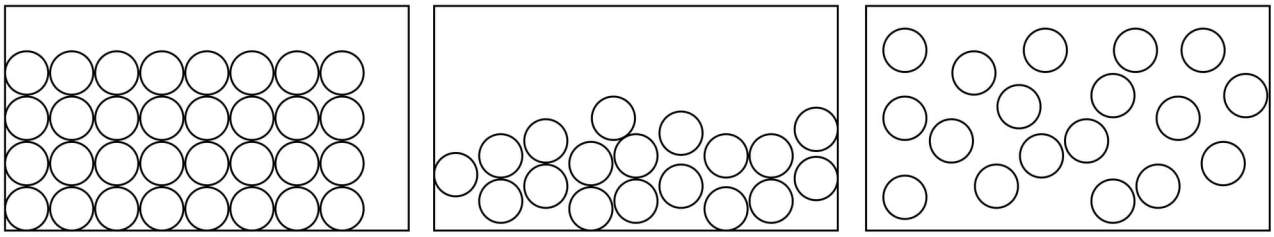


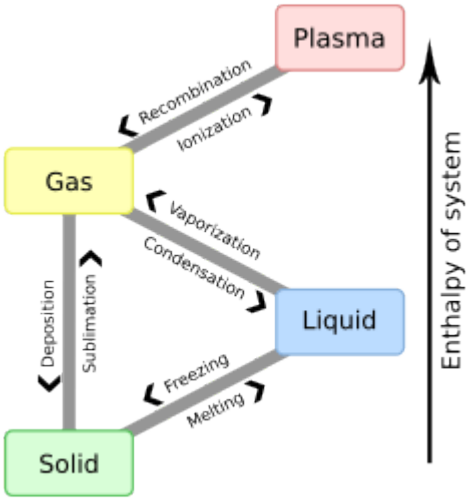
Figure 8.4.1: The solid state (left box) is depicted by closely packed particles arranged in a fixed, orderly grid pattern. The particles are uniformly aligned with no gaps between them, representing a solid's rigid structure, fixed shape, and fixed volume. The liquid state (middle box), shows particles closely packed but not arranged in a fixed pattern, with the ability to flow and adapt to the container's shape. The gas state (right box) illustrates widely spaced particles with large gaps between them. They are in random motion, where matter expands to fill the entire container, having neither a fixed shape nor a fixed volume. (CC BY-SA 4.0; via Wikimedia Commons). Detailed Description of Figure 8.4.1.

Phase Changes

A phase change is a physical process in which a substance goes from one phase to another. Figure 8.4.2 illustrates different phase changes. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance.

The melting point is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid).

The boiling point is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid).



Phase Changes	
Solid → Liquid	Melting or fusion
Gas → Liquid	Condensation
Liquid → Solid	Freezing
Solid → Gas	Sublimation
Liquid → Gas	Vaporization

Figure 8.4.2: This figure illustrates the phase changes between the three primary states of matter: solid, liquid, and gas. Arrows represent the transitions, such as melting (solid to liquid), vaporization (liquid to gas), freezing (liquid to solid), and condensation (gas to liquid). The connections highlight the reversible transformations of matter under different conditions. (CC BY-SA 4.0; penubag Flanker via Wikimedia Commons). Detailed Description of Figure 8.4.2.

- If heat is added to a substance, such as in melting, vaporization, and sublimation, the process is **endothermic**. In this instance, heat is increasing the speed of the molecules causing them move faster (examples: solid to liquid; liquid to gas; solid to gas).
- If heat is removed from a substance, such as in freezing and condensation, then the process is **exothermic**. In this instance, heat is decreasing the speed of the molecules causing them move slower (examples: liquid to solid; gas to liquid). These changes **release heat** to the surroundings.

> The nature of the phase change depends on the direction of the heat transfer. The amount of heat needed to change a sample from solid to liquid would be the same to reverse from liquid to solid. The only difference is the direction of heat transfer.

Heating Curves

Two key points are worth emphasizing.

First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O. If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O. A similar process can occur at 100°C: adding heat increases the amount of gaseous H_2O, while removing heat increases the amount of liquid H_2O (Figure 8.4.3).

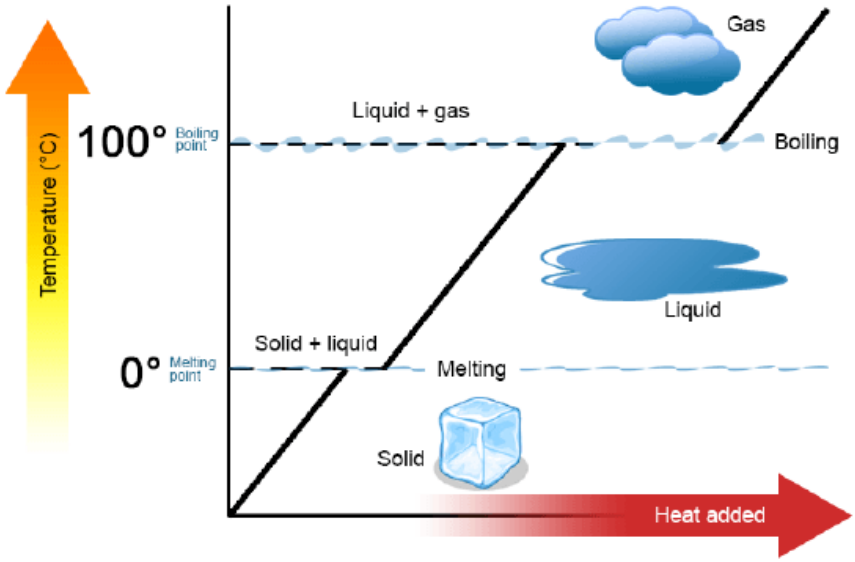


Figure 8.4.3:**Heating curve** for water. (CC BY 3.0; Community College Consortium for Bioscience Credentials via Wikimedia Commons). Detailed Description of Figure 8.4.3.

Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well. As heat is added to solid water (ice), the temperature increases until it reaches 0 °C, the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase.

Second, as shown in Figure 8.4.3, **the temperature of a substance does not change as the substance goes from one phase to another**. In other words, phase changes are isothermal (isothermal means "constant temperature"). Again, consider H_2O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H_2O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

For each phase change of a substance, there is a characteristic quantity of heat needed to perform the phase change per gram (or per mole) of material.

The heat of fusion (ΔH_{fus}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the melting point.

The heat of vaporization (ΔH_{vap}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the boiling point.

If you know the total number of grams or moles of material, you can use the ΔH_{fus} or the ΔH_{vap} to determine the total heat being transferred for melting or solidification using these expressions:

$$\text{heat} = n \times \Delta H_{fus} \tag{8.4.1}$$

where n is the number of moles and ΔH_{fus} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{fus} \tag{8.4.2}$$

where m is the mass in grams and ΔH_{fus} is expressed in energy/gram.

For the boiling or condensation, use these expressions:

$$\text{heat} = n \times \Delta H_{vap} \tag{8.4.3}$$

where n is the number of moles) and ΔH_{vap} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{vap} \tag{8.4.4}$$

where m is the mass in grams and ΔH_{vap} is expressed in energy/gram.

Remember that a phase change depends on the direction of the heat transfer. If heat transfers in, solids become liquids, and liquids become solids at the melting and boiling points, respectively. If heat transfers out, liquids solidify, and gases condense into liquids. At these points, there are no changes in temperature as reflected in the above equations.

Example 8.4.1: Heat transfer during a phase change

How much heat is necessary to melt 55.8 g of ice (solid H_2O) at 0°C? The heat of fusion of H_2O is 79.9 cal/g.

Solution

We can use the relationship between heat and the heat of fusion (Equation 8.4.2) to determine how many cal of heat are needed to melt this ice:

$$\text{heat} = \text{m} \times \Delta \text{H}_{\text{fus}}$$

$$\text{heat} = (55.8 \cancel{\text{g}}) \left(\frac{79.9 \text{ cal}}{\cancel{\text{g}}} \right) = 4,460 \text{ cal}$$

Table 8.4.1 lists the heats of fusion and vaporization for some common substances. Note the units on these quantities; when you use these values in problem solving, make sure that the other variables in your calculation are expressed in units consistent with the units in the specific heats or the heats of fusion and vaporization.

Table 8.4.1: Heats of Fusion and Vaporization for Selected Substances

Substance	ΔH_{fus} (cal/g)	ΔH_{vap} (cal/g)
aluminum (Al)	94.0	2,602
gold (Au)	15.3	409
iron (Fe)	63.2	1,504
water (H_2O)	79.9	540
sodium chloride (NaCl)	123.5	691
ethanol (C_2H_5OH)	45.2	200.3
benzene (C_6H_6)	30.4	94.1

Link to Learning: Sublimation

There is also a phase change where a solid goes directly to a gas:

$$\text{solid} \rightarrow \text{gas} \tag{8.4.5}$$

This phase change is called **sublimation.** Each substance has a characteristic heat of sublimation associated with this process. For example, the heat of sublimation (ΔH_{sub}) of H_2O is 620 cal/g.

We encounter sublimation in several ways. You may already be familiar with dry ice, which is simply solid carbon dioxide (CO_2). At –78.5°C (–109°F), solid carbon dioxide sublimes, changing directly from the solid phase to the gas phase:

$$\text{CO}_2(\text{s}) \xrightarrow{-78.5^\circ\text{C}} \text{CO}_2(\text{g}) \tag{8.4.6}$$

Solid carbon dioxide is called dry ice because it does not pass through the liquid phase. Instead, it does directly to the gas phase. (Carbon dioxide *can* exist as liquid but only under high pressure.)

Dry ice has many practical uses, including the long-term preservation of medical samples.



Figure 8.4.4: Dr Ice Sublimating. The dry ice emits dense white vapor as it transitions directly from a solid to a gas. The vapor spreads out over the bowl and across the surface of the table, demonstrating the sublimation process. The surrounding environment is dimly lit, emphasizing the mist-like effect created by the vapor. (Public Domain; Sarathtly via Wikimedia Commons). Detailed Description of Figure 8.4.4.

Even at temperatures below 0°C, solid H_2O will slowly sublime. For example, a thin layer of snow or frost on the ground may slowly disappear as the solid H_2O sublimes, even though the outside temperature may be below the freezing point of water. Similarly, ice cubes in a freezer may get smaller over time. Although frozen, the solid water slowly sublimes, redepositing on the colder cooling elements of the freezer, which necessitates periodic defrosting (frost-free freezers minimize this redeposition). Lowering the temperature in a freezer will reduce the need to defrost as often.

Under similar circumstances, water will also sublime from frozen foods (e.g., meats or vegetables), giving them an unattractive, mottled appearance called freezer burn. It is not really a “burn,” and the food has not necessarily gone bad, although it looks unappetizing. Freezer burn can be minimized by lowering a freezer’s temperature and by wrapping foods tightly so water does not have any space to sublime into.

Key Takeaway

There is an energy change associated with any phase change.

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8.5: Bond Energies and Chemical Reactions

Introduction

What happens when you take a basketball, place it halfway up a playground slide, and then let it go? The basketball rolls down the slide. What happens if you do it again? Does the basketball roll down the slide? It should. If you were to perform this experiment over and over again, do you think the basketball would ever roll up the slide? Probably not. Why not? Well, for starters, in all of our experience, the basketball has always moved to a lower position when given the opportunity. The gravitational attraction of Earth exerts a force on the basketball, and given the chance, the basketball will move down. We say that the basketball is going to a lower gravitational potential energy. The basketball can move up the slide, but only if someone exerts some effort (that is, work) on the basketball. A general statement can be made based on countless observations over centuries of study:

> All objects tend to move spontaneously to a position of minimum energy unless acted on by some other force or object.

A similar statement can be made about atoms in compounds:

> Atoms bond together to form compounds because in doing so they attain lower energies than they possess as individual atoms.

In making compounds, atoms act like a basketball on a playground slide; they move in the direction of decreasing energy. Thus, when atoms combine to make a compounds, a quantity of energy, equal to the difference between the energies of the bonded atoms and the energies of the separated atoms, is released, usually as heat.

We can reverse this process, just as with the basketball. If we put energy into a molecule, we can cause its bonds to break, separating a molecule into individual atoms. The energy needed to break the bond is the same amount of energy that is liberated when the atoms made the chemical bond in the first place.

Bond Energy

The term **bond energy** is usually used to describe the strength of interactions between atoms that make covalent bonds.

Bonds between certain specific elements usually have a characteristic energy. For atoms in ionic compounds attracted by opposite charges, the term **lattice energy** is used. For now, we will deal with covalent bonds in molecules.

Although each molecule has its own characteristic bond energy, some generalizations are possible. For example, although the exact value of a C–H bond energy depends on the particular molecule, all C–H bonds have a bond energy of roughly the same value because they are all C–H bonds. It takes roughly 100 kcal of energy to break 1 mol of C–H bonds, so we speak of the bond energy of a C–H bond as being about 100 kcal/mol.

A C–C single bond has an approximate bond energy of 80 kcal/mol, while a C=C double bond has a bond energy of about 145 kcal/mol. The C=C bond is stronger than C-C.

Table 8.5.1 lists the approximate bond energies of various covalent bonds.

Table 8.5.1: Approximate Bond Energies

Bond	Bond Energy (kcal/mol)
C–H	100
C–O	86

Bond	Bond Energy (kcal/mol)
C=O	190
C–N	70
C–C	85
C=C	145
C≡C	200
N–H	93
H–H	105
Br-Br	46
Cl–Cl	58
O–H	110
O=O	119
H–Br	87
H–Cl	103

Enthalpy Change or Heat of Reaction, ΔH

When a chemical reaction occurs, the atoms in the reactants rearrange their chemical bonds to make products. The new arrangement of bonds does not have the same total energy as the bonds in the reactants. Therefore, when chemical reactions occur, **there will always be an accompanying energy change.**

Energy changes in chemical reactions are usually measured as changes in enthalpy. In this process, one adds energy to the reaction to break bonds, and extracts energy for the bonds that are formed.

$$\text{enthalpy change, } \Delta H \approx \sum(\text{bonds broken}) - \sum(\text{bonds formed}) \tag{8.5.1}$$

The enthalpy change, ΔH, for a given reaction can be calculated using the bond energy values from Table 8.5.1 by using Equation 8.5.1 . The ≈ sign is used because we are adding together average bond energies (i.e., over many different molecules). Hence, this approach does not give exact values for the enthalpy change, ΔH.

Example 8.5.1: Formation of H_2O

Let's consider the reaction in Equation 8.5.2 of 2 mols of hydrogen gas (H_2) with 1 mol of oxygen gas (O_2) to give 2 mol water:

$$2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g) \tag{8.5.2}$$

or shown graphically in Figure 8.5.1:

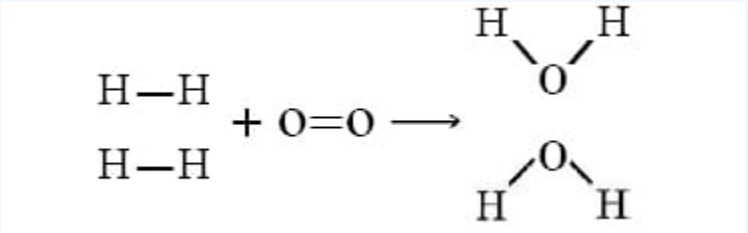


Figure 8.5.1: Reaction of Hydrogen Gas with Oxygen Gas. This reaction produces two water molecules. In this reaction, 2 H–H bonds and 1 O=O bonds are broken, while 4 O–H bonds (two for each H_2O) are formed. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.1.

Average bond energies involved with this reaction are:

- H−H = 105 kcal/mol
- O=O = 119 kcal/mol
- O−H = 110 kcal/mol

In this reaction, 2 H–H bonds and 1 O=O bonds are broken, while 4 O–H bonds (two for each H_2O) are formed. The energy changes can be tabulated and calculated as follows Table 8.5.2

Table 8.5.2: Bond Energy Changes

Bonds Broken (kcal/mol)		Bonds Formed (kcal/mol)	
2 H–H	2 mol x 105 kcal/mol = 210 kcal	4 O–H	4 mols x 110 kcal/mol = 440 kcal
1 O=O	1 mol x 119 kcal/mol = 119 kcal		
	Total = 329 kcal		**Total** = 440 kcal

$$\begin{aligned}\Delta H &\approx \sum(\text{bonds broken}) - \sum(\text{bonds formed}) \\ &\approx 329\ \text{kcal} - 440\ \text{kcal} \\ &\approx -111\ \text{kcal}\end{aligned}$$

The enthalpy change (ΔH) of the reaction is approximately **–111 kcal/mol.**

We can also re-write the reaction equation with the ΔH information (see below).

$$2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g) \qquad \Delta H = -111\ \text{kcal}$$

Note that an **exothermic** reaction has a **negative ΔH** value.This means that bonds in the products (440 kcal) are stronger than the bonds in the reactants (329 kcal) by about 111 kcal/mol. Because the bonds in the products are stronger than those in the reactants, the reaction releases more energy than it consumes. This excess energy is **released** as heat, so the reaction is **exothermic**. Hence, we can re-write the reaction with the heat released (111 kcal) on the **product side** of the equation (Figure 8.5.2):

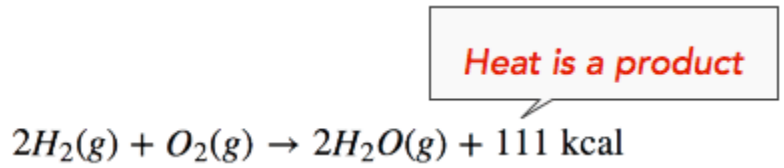


Figure 8.5.2: Heat Expressed as a Product. The heat released as 111 kcal, and can be expressed as a product on the chemical equation. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.2.

An **endothermic** reaction on the other hand will have a **positive ΔH** value, because energy in the form of heat is absorbed from the surroundings, and the absorbed heat will be written on the **reactant side** Figure 8.5.5.

Example 8.5.2: Determining the Enthalpy Change of a Reaction

What is the enthalpy change for the reaction shown in Equation 8.5.3? Is the reaction exothermic or endothermic?

$$H_2\,(g) + Br_2\,(g) \rightarrow 2\,HBr\,(g) \tag{8.5.3}$$

Solution

Step 1- First look at the equation and identify which bonds exist on in the reactants (bonds broken).

- one H-H bond and

- one Br-Br bond

Step 2- Do the same for the products (bonds formed)

- two H-Br bonds

Step 3- Identify the bond energies of these bonds from Table 8.5.1:

- H-H bonds: 105 kcal/mol
- Br-Br bonds: 46 kcal/mol

Step 4- Set up the table (Table 8.5.3) and apply the formula for enthalpy change.

Table 8.5.3: Solutions to Example 7.4.1

Bonds Broken (kcal/mol)		Bonds Formed (kcal/mol)	
1 H–H	1 mol x 105 kcal/mol = 105 kcal	2 H–Br	2 mols x 87 kcal/mol = 174 kcal
1 Br–Br	1 mol x 46 kcal/mol = 46 kcal		
	Total = 151 kcal		**Total** = 174 kcal

$$\Delta H \approx 151 \text{ kcal} - 174 \text{ kcal}$$
$$\approx -23 \text{ kcal}$$

Step 5- Since ΔH is negative (–23 kcal), the reaction is exothermic.

Endothermic and Exothermic Reactions

As mentioned above, endothermic and exothermic reactions can be thought of as having energy as either a reactant or a product.

Endothermic reactions require energy, so **energy is a reactant**. Heat flows from the surroundings to the system (reaction mixture) and the enthalpy of the system increases (ΔH is positive).

Exothermic reactions on the other hand release energy, so **energy is a product.** The enthalpy of the system decreases (ΔH is negative).

- In the course of an **endothermic** process, the system gains heat from the surroundings and so the **temperature of the surroundings decreases (gets cold).**
- A chemical reaction is **exothermic** if heat is released by the system into the surroundings. Because the surroundings is gaining heat from the system, **the temperature of the surroundings increases.** See Figure 8.5.3.

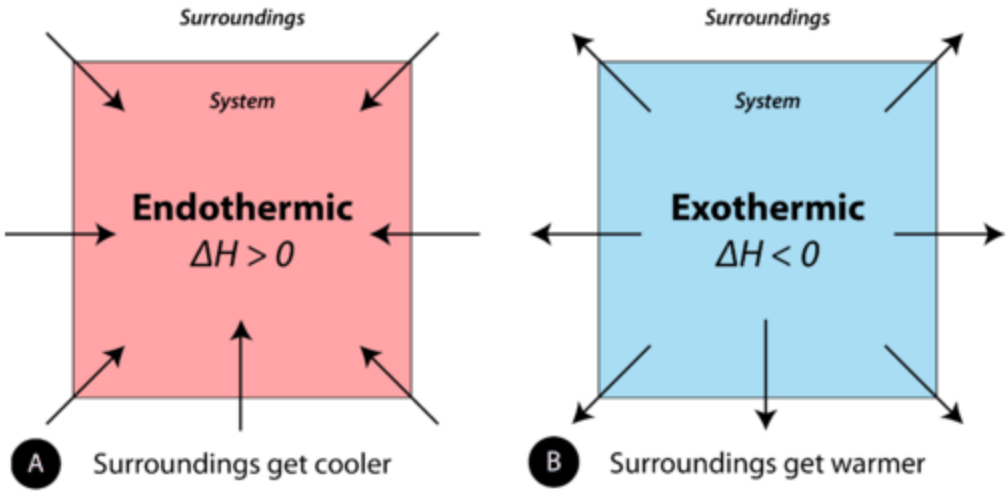


Figure 8.5.3: Endothermic and Exothermic Reactions. (A) Endothermic reaction. (B) Exothermic reaction. Endothermic reaction: surroundings get cooler and delta H is greater than 0, Exothermic reaction: surroundings get warmer and delta H is less than 0. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.3.

We will look at two examples:

- **Endothermic Reaction**: When $1\ \text{mol}$ of calcium carbonate decomposes into $1\ \text{mol}$ of calcium oxide and $1\ \text{mol}$ of carbon dioxide, $177.8\ \text{kJ}$ of heat is absorbed. Because the heat is absorbed by the system, the $177.8\ \text{kJ}$ is written as a **reactant** Figure 8.5.4.

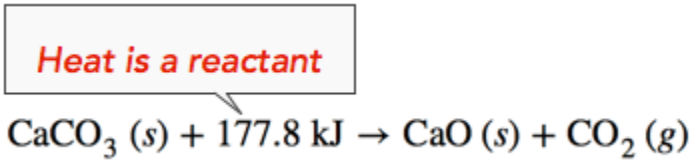


Figure 8.5.4: Figure 8.5.4: Heat as a Reactant. The heat absorbed as 177.8 kJ, and can be expressed as a reactant on the chemical equation. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.4.

> The **ΔH is positive** for an endothermic reaction.

$$\text{CaCO}_3\,(s) \rightarrow \text{CaO}\,(s) + \text{CO}_2\,(g) \quad \Delta H = +177.8\ \text{kJ}$$

- **Exothermic Reaction**: When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of $1\ \text{mol}$ of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of the balanced equation in two ways. First, the amount of heat released can be written in the **product** side of the reaction Figure 8.5.5. Another way is to write the **heat of reaction** (ΔH) information with a **negative** sign, $-890.4\ \text{kJ}$.

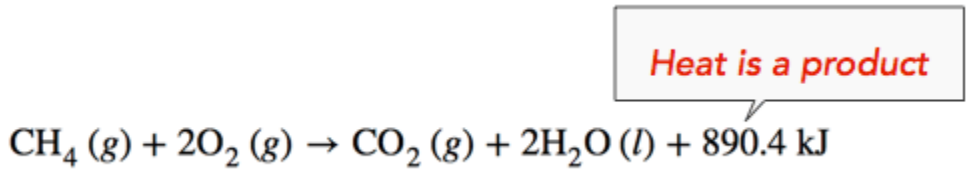


Figure 8.5.5: Heat as a Product. The heat released as 890.4 kJ, and can be expressed as a product on the chemical equation. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.5.

> The **ΔH is negative** for an exdothermic reaction.

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) \quad \Delta H = -890.4\,\mathrm{kJ}$$

Example 8.5.3: Determining whether a Reaction is Exothermic or Endothermic

Is each chemical reaction exothermic or endothermic?

a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 213$ kcal
b. $N_2(g) + O_2(g) + 45$ kcal $\rightarrow 2NO(g)$

Solution

a. Because energy (213 kcal) is a product, energy is given off by the reaction. Therefore, this reaction is exothermic.
b. Because energy (45 kcal) is a reactant, energy is absorbed by the reaction. Therefore, this reaction is endothermic.

Energy Diagrams

Endothermic and exothermic reactions can be visually represented by **energy-level diagrams** like the ones in Figure $8.5.6Aand6B$.

In endothermic reactions, the reactants have higher bond energy (stronger bonds) than the products. **Strong bonds have lower potential energy than weak bonds**. Hence, the **energy of the reactants is lower than that of the products**. This type of reaction is represented by an "uphill" energy-level diagram shown in Figure $8.5.6A$. For an endothermic chemical reaction to proceed, the reactants must absorb energy from their environment to be converted to products.

In an exothermic reaction, the bonds in the product have higher bond energy (stronger bonds) than the reactants. In other words, the **energy of the products is lower than the energy of the reactants,** hence is energetically downhill, shown in Figure $8.5.6B$. Energy is given off as reactants are converted to products. The energy given off is usually in the form of heat (although a few reactions give off energy as light). In the course of an exothermic reaction, heat flows from the system to its surroundings, and thus, gets warm.

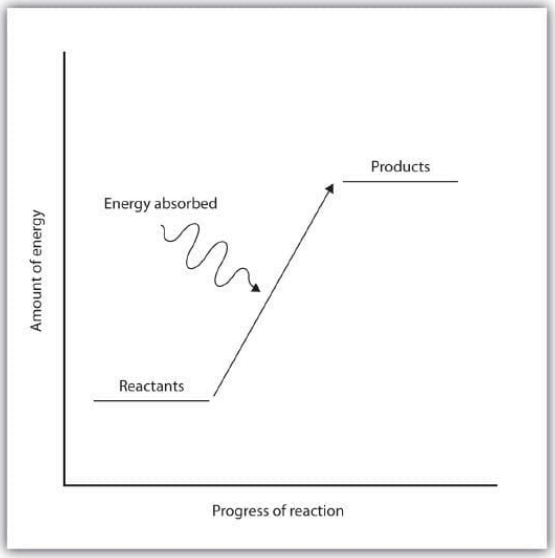


Figure $8.5.6A$: Endothermic Reaction Diagram. Endothermic Reactions (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure $8.5.6A$.

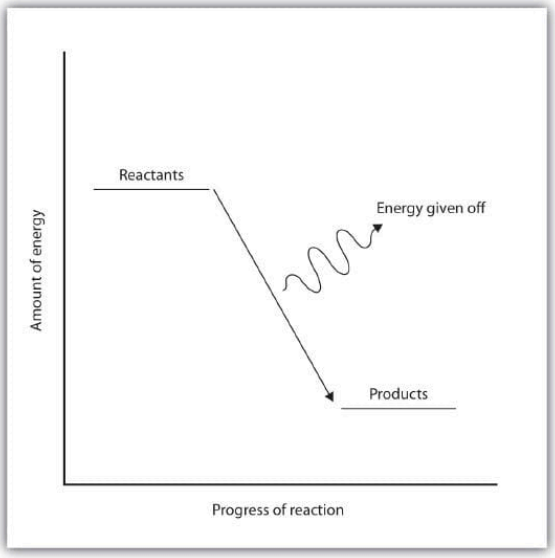


Figure 8.5.6B:Exothermic Reaction Diagram. Exothermic Reactions (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.5.6B.

Table 8.5.4: Endothermic and Exothermic Reactions

Endothermic Reactions	Exothermic Reactions
Heat is absorbed by reactants to form products.	Heat is released.
Heat is absorbed from the surroundings; as a result, the surroundings get cold.	Heat is released by the reaction to surroundings; surroundings feel hot.
ΔH_{rxn} is positive	ΔH_{rxn} is negative
The bonds broken in the reactants are stronger than the bonds formed in the products	The bonds formed in the products are stronger than the bonds broken in the reactants
The reactants are lower in energy than the products	The products are lower in energy than the reactants
Represented by an "uphill" energy diagram	Represented by an "downhill" energy diagram

Key Takeaways

- Atoms are held together by a certain amount of energy called bond energy.
- Energy is required to break bonds. Energy is released when chemical bonds are formed because atoms become more stable.
- Chemical processes are labeled as exothermic or endothermic based on whether they give off or absorb energy, respectively.

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8.6: The Energy of Biochemical Reactions

Introduction

The chemistry of the human body, or any living organism, is very complex. Even so, the chemical reactions found in the human body follow the same principles of energy that other chemical reactions follow.

Energy of Biochemical Reactions

Where does the energy that powers our bodies come from? The details are complex, but we can look at some simple processes at the heart of cellular activity.

An important reaction that provides energy for our bodies is the oxidation of glucose ($C_6H_{12}O_6$):

$$C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \rightarrow 6\,CO_2\,(g) + 6\,H_2O(\ell) + 670\ \text{kcal} \tag{8.6.1}$$

Considering that 1 mol of $C_6H_{12}O_6\,(s)$ has a volume of about 115 mL, we can see that glucose is a compact source of energy.

Glucose and other sugars are examples of **carbohydrates**, which are one of the three main dietary components of a human diet. All carbohydrates supply approximately 4 kcal/g. You can verify that by taking the heat of reaction for glucose oxidation and dividing it by its molar mass. **Proteins**, the building blocks of structural tissues like muscle and skin, also supply about 4 kcal/g. Other important energy sources for the body are **fats**, which are largely hydrocarbon chains. Fats provide even more energy per gram, about 9 kcal/g.

Another important reaction is the conversion of adenosine triphosphate (ATP) to adenosine diphosphate (ADP), which is shown in Figure 8.6.1. [See detailed description in alt text document]. Under physiological conditions, the breaking of an O–P bond and the formation of an O–P and two O–H bonds gives off about 7.5 kcal/mol of ATP. This may not seem like much energy, especially compared to the amount of energy given off when glucose reacts. It is enough energy, however, to fuel other biochemically important chemical reactions in our cells.

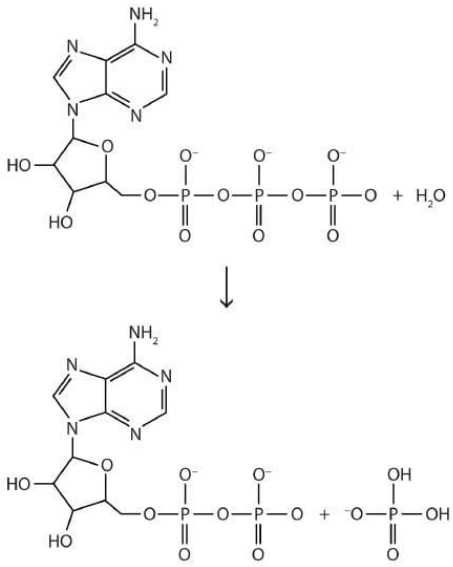


Figure 8.6.1: ATP to ADP. The conversion of ATP (top) to ADP (bottom) provides energy for the cells of the body. (CC BY-NC-SA 3.0; anonymous, via Basics of General Organic and Biological Chemistry). Detailed Description of Figure 8.6.1.

> Even complex biological reactions must obey the basic rules of chemistry.

Career Focus: Dietitian

A dietitian is a nutrition expert who communicates food-related information to the general public. In doing so, dietitians promote the general well-being among the population and help individuals recover from nutritionally related illnesses.

Our diet does not just supply us with energy. We also get vitamins, minerals, and even water from what we eat. Eating too much, too little, or not enough of the right foods can lead to a variety of problems. Dietitians are trained to make specific dietary recommendations to address particular issues relating to health. For example, a dietitian might work with a person to develop an overall diet that would help that person lose weight or control diabetes. Hospitals employ dietitians in planning menus for patients, and many dietitians work with community organizations to improve the eating habits of large groups of people.

Key Takeaway

Energy to power the human body comes from chemical reactions.

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8.7: Summary

Energy is the ability to do work. The transfer of energy from one place to another is heat. Heat and energy are measured in units of joules, calories, or kilocalories (equal to 1,000 calories). The amount of heat gained or lost when the temperature of an object changes can be related to its mass and a constant called the specific heat of the substance.

The transfer of energy can also cause a substance to change from one phase to another. During the transition, called a phase change, heat is either added or lost. Despite the fact that heat is going into or coming out of a substance during a phase change, the temperature of the substance does not change until the phase change is complete; that is, phase changes are isothermal. Analogous to specific heat, a constant called the heat of fusion of a substance describes how much heat must be transferred for a substance to melt or solidify (that is, to change between solid and liquid phases), while the heat of vaporization describes the amount of heat transferred in a boiling or condensation process (that is, to change between liquid and gas phases).

Every chemical change is accompanied by an energy change. This is because the interaction between atoms bonding to each other has a certain bond energy, the energy required to break the bond (called lattice energy for ionic compounds), and the bond energies of the reactants will not be the same as the bond energies of the products. Reactions that give off energy are called exothermic, while reactions that absorb energy are called endothermic. Energy-level diagrams can be used to illustrate the energy changes that accompany chemical reactions.

Even complex biochemical reactions have to follow the rules of simple chemistry, including rules involving energy change. Reactions of carbohydrates and proteins provide our bodies with about 4 kcal of energy per gram, while fats provide about 9 kcal per gram.

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8.8: Key Terms

Key Terms

- **Energy**: The capacity to do work or transfer heat.
- **Potential Energy**: Stored energy that an object possesses due to its position or composition.
- **Kinetic Energy**: The energy an object has due to its motion.
- **Joule**: The SI unit of energy, defined as the energy transferred when a force of one newton moves an object one meter.
- **Calorie**: A unit of energy, commonly used in food, equivalent to 4.184 joules.
- **Specific Heat Capacity:** The amount of heat energy required to raise the temperature of one gram of a substance by one degree Celcius (°C) or one Kelvin (K) without a phase change.
- **Specific Heat**: The amount of heat required to raise the temperature of one gram of a substance by one degree Celsius.
- **Temperature**: A measure of the average kinetic energy of the particles in a substance.
- **Heat**: The transfer of thermal energy between substances due to a temperature difference.
- **State of Matter**: The distinct forms in which matter exists, including solid, liquid, and gas.
- **Intermolecular Forces**: The forces of attraction or repulsion between molecules that influence physical properties.
- **Phase**: A distinct physical state of matter, characterized by uniform physical and chemical properties.
- **Phase Change**: The transition of matter from one state to another, such as melting or evaporation.
- **Gas**: A state of matter with no fixed shape or volume, expanding to fill its container.
- **Liquid**: A state of matter with a definite volume but no fixed shape.
- **Solid**: A state of matter with a definite shape and volume.
- **Melting Point**: The temperature at which a solid turns into a liquid.
- **Boiling Point**: The temperature at which a liquid turns into a gas.
- **Endothermic**: A process that absorbs heat from the surroundings.
- **Exothermic**: A process that releases heat to the surroundings.
- **Bond Energy**: The amount of energy required to break one mole of chemical bonds in a molecule.
- **Heat of Fusion**: The energy required to change one gram of a solid to a liquid at its melting point.
- **Heat of Vaporization**: The energy required to convert one gram of a liquid to a gas at its boiling point.
- **Lattice Energy**: The energy required to separate one mole of an ionic solid into its gaseous ions.
- **Carbohydrates**: Organic molecules composed of carbon, hydrogen, and oxygen, serving as a primary energy source.
- **Proteins**: Large biomolecules consisting of amino acids, essential for cellular function and structure.
- **Fats**: Organic compounds made of glycerol and fatty acids, serving as energy storage molecules.
- **Melting/Fusion**: The process of changing from a solid to a liquid by heat absorption.
- **Condensation**: The process where a gas turns into a liquid by releasing heat.
- **Freezing**: The process where a liquid turns into a solid by heat loss.
- **Sublimation**: The direct transition from solid to gas without passing through the liquid state.
- **Vaporization**: The transition from liquid to gas, including boiling and evaporation.
- **Heating Curve**: A graph that shows temperature changes of a substance as it is heated over time.
- **Energy Diagram**: A graphical representation of the energy changes during a chemical reaction.

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8.9: Alternative text for Chapter 8

Figure 8.2.1: Energy in our Daily Lives

The image is divided into three sections, each illustrating different examples of energy sources or transformations. On the left is a close-up view of a cheeseburger partially wrapped in white paper. The cheeseburger contains a sesame seed bun, layers of lettuce, a slice of tomato, red onion rings, a beef patty, and melted cheese, representing the chemical energy stored in food, which is used by the human body to perform work when metabolized. In the center, a busy multi-lane highway is shown filled with vehicles, including cars, buses, and vans, moving in both directions. The road is lined with trees on both sides, and the scene appears to be on an overcast or rainy day. The vehicles symbolize the conversion of chemical energy from fossil fuels into mechanical energy to power transportation, highlighting human reliance on energy for mobility. On the right, an industrial facility is depicted, featuring a large furnace or coke oven used in metal production processes. The structure includes metal piping, industrial machinery, and a chute filled with glowing hot coke or molten material, with a worker in protective clothing standing near the machinery. This scene represents the energy-intensive process of converting raw materials into usable metals, with thermal energy playing a critical role in industrial manufacturing.

Figure 8.2.2: Energy Conversion

The image consists of two panels side by side, showcasing examples of water in motion and its use in energy production. The left panel features a view of Victoria Falls, a massive waterfall located on the border of Zambia and Zimbabwe. The water flows powerfully over a wide cliff edge, plunging into a gorge below, creating a misty spray that rises into the air. A rainbow arcs across the scene, visible due to the refraction of sunlight through the mist. The surroundings include lush greenery and a clear sky, emphasizing the natural beauty and raw power of the waterfall. This illustrates the potential energy of water at a height, which can be converted into kinetic energy as it falls. The right panel displays the Hoover Dam, an iconic hydroelectric structure located in the United States. The dam is nestled between steep, rocky canyon walls and holds back a large reservoir of water. Below the dam, water flows through turbines housed in a series of rectangular structures at the base of the dam, generating electricity. The reservoir and water channels are clearly visible, with roads and infrastructure around the dam highlighting its scale and human engineering. This image demonstrates the use of water's potential energy for generating renewable energy through hydroelectric power.

Figure 8.2.3: Combustion Engine

The image shows a detailed cross-sectional view of an internal combustion engine, specifically a V-type engine with two cylinders arranged in a V configuration. The components are color-coded to distinguish various parts. The image provides clear contrast for the mechanical components. This visual highlights the intricate design and operation of an internal combustion engine, emphasizing how chemical energy from fuel is converted into mechanical energy.

Figure 8.2.4: Nutritional Energy

The image is a detailed nutrition facts label for a food product. At the top, the heading "Nutrition Facts" is displayed in bold, followed by the information that the package contains three servings per container. The serving size is specified as 75 grams (75g). The next section focuses on the calorie content, which is prominently displayed as 20 calories per serving, written in bold with a larger font for emphasis. Below the calorie information, a horizontal line separates the macronutrient and micronutrient details from the top section. The label provides the following nutrient information per serving: Total Fat is listed as 0 grams (0% of the daily value), with subcategories for Saturated Fat and Trans Fat also at 0 grams. Sodium content is 140 milligrams, accounting for 6% of the daily value. Total Carbohydrate content is 4 grams, which is 1% of the daily value, with Dietary Fiber listed at 2 grams (7% of the daily value) and Total Sugars at 0 grams, including 0 grams of added sugars. Protein content is 2 grams, representing 4% of the daily value. A note at the bottom of the label specifies that the product is not a significant source of cholesterol, vitamin D, calcium, iron, or potassium. A footnote explains that the percentages of daily value are based on a 2,000-calorie diet, which is used as a general reference for daily nutrition advice. The layout is structured and designed to clearly present the nutritional information for consumers.

Figure 8.3.1: Temperature Measure

Two molecular drawings are shown and labeled a and b. Drawing (a) is a box containing fourteen spheres that are surrounded by lines indicating that the particles are moving rapidly. This drawing has a label that reads “Hot liquid.” Drawing (b) depicts another box of equal size that also contains fourteen spheres. They are all surrounded by smaller lines that depict some particle motion, but not as much as in drawing a. This drawing has a label that reads “Cold liquid.”

Figure 8.3.2: Heat Flow and Transfer

The image consists of three panels labeled (a), (b), and (c), illustrating the process of heat transfer between two substances and the establishment of thermal equilibrium. In panel (a), two separate boxes labeled "H" and "L" represent substances at different temperatures. The box labeled "H" on the left contains particles surrounded by wavy lines, symbolizing high-energy, rapidly moving particles indicative of a higher temperature. The box labeled "L" on the right contains particles with fewer wavy lines, representing lower-energy, slower-moving particles at a cooler temperature. The caption beneath this panel reads "Different temperatures." Panel (b) shows the two boxes labeled "H" and "L" in contact, with an arrow pointing from the high-temperature box to the low-temperature box. This arrow represents the transfer of thermal energy from the hotter substance to the cooler substance. The particles in the "H" box remain exhibit slightly reduced energy, while the particles in the "L" box begin to indicate an increase in their energy as they absorb heat. The caption beneath this panel reads "Contact." Panel (c) depicts the two boxes after thermal equilibrium has been reached. Both boxes now contain particles with identical energy levels, as indicated by the uniform appearance of the wavy lines around the particles in both boxes. This illustrates that heat transfer has stopped, and the temperatures of both substances are equal. The caption beneath this panel reads "Thermal equilibrium." This image visually explains the principles of heat transfer, energy exchange, and the eventual equalization of temperature between two substances.

Figure 8.3.3: Measuring Temperature

A picture labeled a is shown as well as a pair of drawings labeled b. Picture a shows the lower portion of an alcohol thermometer. The thermometer has a printed scale to the left of the tube in the center that reads from negative forty degrees at the bottom to forty degrees at the top. It also has a scale printed to the right of the tube that reads from negative thirty degrees at the bottom to thirty five degrees at the top. On both scales, the volume of the alcohol in the tube reads between nine and ten degrees. The two images labeled b both depict a metal strip coiled into a spiral and composed of brass and steel. The left coil, which is loosely coiled, is labeled along its upper edge with the 30 degrees C and 10 degrees C. The end of the coil is near the 30 degrees C label. The right hand coil is much more tightly wound and the end is near the 10 degree C label.

Figure 8.3.4: Temperature Scales

Three thermometers illustrating the Fahrenheit, Celsius, and Kelvin scales, with liquid showing different temperature levels

Figure 8.4.1: States of Matter

A visual representation of the three states of matter: solid, liquid, and gas, displayed in three adjacent boxes. Left Box (Solid): Depicts closely packed particles arranged in a fixed, orderly grid pattern. The particles are uniformly aligned with no gaps between them, representing a solid's rigid structure, fixed shape, and fixed volume. Middle Box (Liquid): Shows particles that are close together but not arranged in a fixed pattern. The particles are scattered with slight gaps and exhibit irregular spacing, indicating that a liquid flows to take the shape of its container while maintaining a consistent volume. Right Box (Gas): Displays particles that are widely spaced and randomly distributed, with large gaps between them. The particles are dispersed throughout the box, illustrating the gaseous state, where matter expands to fill the entire container, having neither a fixed shape nor a fixed volume. This diagram effectively illustrates the differences in particle arrangement and behavior across the three physical states of matter.

Figure 8.4.2: Phase Changes

A flowchart diagram illustrating the phase changes between the three primary states of matter: solid, liquid, and gas. The diagram consists of three rectangular nodes, each labeled with a state of matter. The node for "Solid" is at the bottom left, the node for "Liquid" is in the center right, and the node for "Gas" is in the top left. Arrows connect the nodes to represent the transitions between these states. A horizontal arrow points from "Solid" to "Liquid," indicating melting. A reverse arrow points from "Liquid" to

"Solid," representing freezing. A diagonal arrow points from "Liquid" to "Gas," labeled as vaporization, and its reverse arrow points from "Gas" to "Liquid," labeled condensation. Lastly, a vertical arrow points directly from "Solid" to "Gas," signifying sublimation, while its reverse arrow from "Gas" to "Solid" indicates deposition. Each phase is visually separated by space, and the arrows emphasize the dynamic and reversible nature of these phase changes. A temperature vs. heat energy diagram illustrating phase changes of water. The x-axis represents heat added, marked with an arrow pointing to the right, and the y-axis represents temperature in degrees Celsius, ranging from 0°C at the bottom to 100°C at the top.

Figure 8.4.3: Heating Curve for Water

At the bottom of the diagram, the solid phase (ice) is shown with an image of an ice cube. A horizontal line labeled "Melting point" at 0°C indicates the phase transition from solid to liquid, with a section labeled "Solid + liquid," where both phases coexist. The middle section depicts the liquid phase with an image of a puddle of water. A horizontal line labeled "Boiling point" at 100°C marks the transition from liquid to gas, with a section labeled "Liquid + gas," where both phases coexist. At the top, the gas phase is represented with an image of a cloud. The diagram uses sloped lines to show temperature increases within the pure phases (solid, liquid, and gas) and flat lines during phase transitions (melting and boiling), emphasizing the energy required for phase changes without a temperature increase. The overall diagram conveys the relationship between heat energy input and temperature during state changes.

Figure 8.4.4: Dry Ice Sublimating

A bowl containing dry ice sublimating at room temperature. The dry ice emits dense vapor as it transitions directly from a solid to a gas. The vapor spreads out over the bowl and across the surface of the table, demonstrating the sublimation process. The surrounding environment is dimly lit, emphasizing the mist-like effect created by the vapor.

Figure 8.5.1: Reaction of Oxygen Gas with Hydrogen Gas

A chemical reaction diagram illustrating the formation of water molecules from hydrogen and oxygen molecules. On the left side of the reaction, two diatomic hydrogen molecules (H—H) and one diatomic oxygen molecule (O=O) are shown as reactants. An arrow indicates the reaction process leading to the right side, where two water (H_2O) molecules are formed. Each water molecule is depicted in a V-shape, with an oxygen atom at the center bonded to two hydrogen atoms at an angle. The oxygen atoms are represented with double lone pairs, and the hydrogen atoms are arranged around them, reflecting the bent molecular geometry characteristic of water.

Figure 8.5.2: Heat Expressed as a Product

A chemical equation representing the combustion of hydrogen gas. The equation reads:**$H_2(g) + O_2(g) \rightarrow 2H_2O(g) + 111$ kcal**, indicating that two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor and release 111 kilocalories of energy. A speech bubble with text states, **"Heat is a product"** emphasizing that this is an exothermic reaction where heat is released.

Figure 8.5.3: Endothermic and Exothermic Reactions

Diagram illustrating the differences between endothermic and exothermic reactions. On the left, a square represents an endothermic process where energy is absorbed into the system from the surroundings, indicated by inward-facing arrows pointing toward the labeled 'System.' The label '$\Delta H > 0$' shows the positive change in enthalpy, and text below states, 'Surroundings get cooler.' On the right, another square represents an exothermic process where energy is released from the system into the surroundings, indicated by outward-facing arrows pointing away from the 'System.' The label '$\Delta H < 0$' shows the negative change in enthalpy, and text below states, 'Surroundings get warmer.' The surroundings are labeled in both diagrams to contextualize the flow of energy.

Figure 8.5.4: Heat as a Reactant

A chemical equation representing the decomposition of calcium carbonate. The equation reads: **$CaCO_3(s) + 177.8$ kJ $\rightarrow$ $CaO(s)$**

+ CO_2(g), indicating that solid calcium carbonate absorbs 177.8 kilojoules of energy to break down into solid calcium oxide and gaseous carbon dioxide. A speech bubble states, **"Heat is a reactant,"** emphasizing that this is an endothermic reaction where heat is required for the reaction to proceed.

Figure 8.5.5: Heat as a Product

A chemical equation representing the combustion of methane. The equation reads: **CH_4(g) + $2O_2$(g) → CO_2(g) + $2H_2O$(l) + 890.4 kJ**, indicating that one molecule of methane gas reacts with two molecules of oxygen gas to produce one molecule of carbon dioxide gas, two molecules of liquid water, and releases 890.4 kilojoules of energy. A speech bubble states, **"Heat is a product,"** emphasizing that this is an exothermic reaction where heat is released.

Figure 8.5.6A: Endothermic Reaction Diagram

A line graph illustrating the energy profile of a chemical reaction. The x-axis is labeled "Progress of reaction," and the y-axis is labeled "Amount of energy." At the start of the reaction (labeled as "Reactants"), a horizontal line represents the initial energy level of the reactants. From this point, the graph shows a sharp upward diagonal rise, indicating that energy is absorbed as the reaction progresses. A wavy arrow labeled "Energy absorbed" is shown pointing upwards along the incline.The graph levels off at a higher energy state labeled "Products," representing the final energy level of the products after the reaction is complete. This diagram represents an endothermic reaction, where energy is absorbed, resulting in the products having a higher energy state than the reactants. The image visually depicts the relationship between energy changes and the progress of a reaction.

Figure 8.5.6B: Exothermic Reaction Diagram

A line graph illustrating the energy profile of a chemical reaction. The x-axis is labeled "Progress of reaction," and the y-axis is labeled "Amount of energy." The graph begins with a horizontal line labeled "Reactants," representing the initial energy level of the reactants. From this point, the graph shows a sharp downward diagonal slope, indicating a release of energy as the reaction progresses. A wavy arrow labeled "Energy given off" points downward along the decline, signifying the energy released during the reaction. The graph levels off at a lower energy state labeled "Products," representing the final energy level of the products after the reaction is complete. This diagram represents an exothermic reaction, where energy is released, resulting in the products having a lower energy state than the reactants. The image provides a clear visual of the energy changes occurring in an exothermic reaction.

Figure 8.6.1: ATP to ADP

A chemical reaction showing the chemical conversion of ATP to ADP, releasing energy for cellular processes The diagram shows the chemical reaction of ATP (adenosine triphosphate) hydrolyzing into ADP (adenosine diphosphate) and inorganic phosphate. At the top, the ATP molecule is depicted, which consists of three components: a nitrogenous base called adenine with a double-ring structure, a ribose sugar attached to the adenine, and a chain of three phosphate groups connected by high-energy bonds. The ribose sugar is a five-membered ring with hydroxyl groups attached to the second and third carbon atoms, while each phosphate group contains a phosphorus atom bonded to four oxygen atoms, with two of these oxygens carrying negative charges. A water molecule (H_2O) is shown as a reactant in the reaction. An arrow pointing downward illustrates the hydrolysis process, where ATP is broken down. At the bottom, the products of the reaction are displayed. These include ADP, which retains the adenine, ribose sugar, and two phosphate groups, and an inorganic phosphate (Pi), which contains one phosphorus atom bonded to four oxygen atoms, one of which has a hydroxyl group. The reaction highlights the release of energy as one phosphate group is cleaved from ATP through the addition of water. This diagram visually explains the biochemical process of ATP hydrolysis and the energy dynamics involved.

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9: GASES, LIQUIDS, AND SOLIDS

Introduction

This chapter explores the fundamental concepts of matter and gases. We begin by examining the three states of matter—solids, liquids, and gases—and the physical changes they undergo. Next, we look at the forces between molecules that influence a substance's properties. The chapter then introduces the kinetic molecular theory, which explains the behavior of gases based on particle motion. We also define gas pressure, understand its measurement in units like Pascals, and learn how gases exert force on their containers. Finally, we apply key gas laws to analyze how gases respond to changes in pressure, volume, and temperature.

Learning Objectives

- Describe the three physical states of matter and the transitions between them.
- Identify intermolecular forces and explain how they affect physical properties.
- Explain the kinetic molecular theory and how it accounts for the behavior of gases.
- Define gas pressure in terms of particle collisions and convert between pressure units.
- Use gas laws to solve problems involving pressure, volume, and temperature changes in gases.
- Apply Dalton's Law of Partial Pressures to calculate total and partial pressures in gas mixtures.
- Use Henry's Law to explain the relationship between gas solubility and pressure in liquids.

9.1: Prelude to Solids, Liquids, and Gases

9.2: Intermolecular Forces

9.3: Gases and the Kinetic-Molecular Theory

9.4: Units of Gas Pressure

9.5: Gas Laws

9.6: Summary

9.7: Key Terms

9.8: Alternative text for Chapter 9

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9.1: Prelude to Solids, Liquids, and Gases

Overview

We normally experience **carbon dioxide** (CO_2) as a gas, but if it were cooled down to about –78°C, it would become a solid. The everyday term for solid carbon dioxide is **dry ice**.

Why “dry” ice? Solid carbon dioxide is called dry ice because it converts from a solid to a **gas** directly, without going through the liquid phase, in a process called **sublimation**. Thus, there is no messy **liquid** phase to worry about. Although it is a novelty, dry ice has some potential dangers. Because it is so cold, it can freeze living tissues very quickly, so people handling dry ice should wear special protective gloves. The cold carbon dioxide gas is also heavier than air (because it is cold and more dense), so people in the presence of dry ice should be in a well-ventilated area.



Figure 9.1.1: Dry ice sublimation. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.1.1.

Dry ice has several common uses. Because it is so cold, it is used as a refrigerant to keep other things cold or frozen (for example, meats or ice cream). In the medical field, dry ice is used to preserve medical specimens, blood products, and drugs. It also has dermatological applications (for example, freezing off warts). Organs for transplant are kept cool with dry ice until the recipient of the new organ is ready for surgery. In this respect, carbon dioxide is much like water—more than one phase of the same substance has significant uses in the real world.

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9.2: Intermolecular Forces

Introduction

Consider these two aspects of the molecular-level environments in solid, liquid, and gaseous matter:

- Particles in a solid are tightly packed together and often arranged in a regular pattern; in a liquid, they are close together with no regular arrangement; in a gas, they are far apart with no regular arrangement.
- Particles in a solid vibrate about fixed positions and do not generally move in relation to one another; in a liquid, they move past each other but remain in essentially constant contact; in a gas, they move independently of one another except when they collide.

The differences in the properties of a solid, liquid, or gas reflect the strengths of the attractive forces between the atoms, molecules, or ions that make up each phase. The phase in which a substance exists depends on the relative extents of its **intermolecular forces** (IMFs) and the kinetic energies (KE) of its molecules. IMFs are the various forces of attraction that may exist between the atoms and molecules of a substance due to electrostatic phenomena, as will be detailed in this module. These forces serve to hold particles close together, whereas the particles' KE provides the energy required to overcome the attractive forces and thus increase the distance between particles. Figure 9.2.1 illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

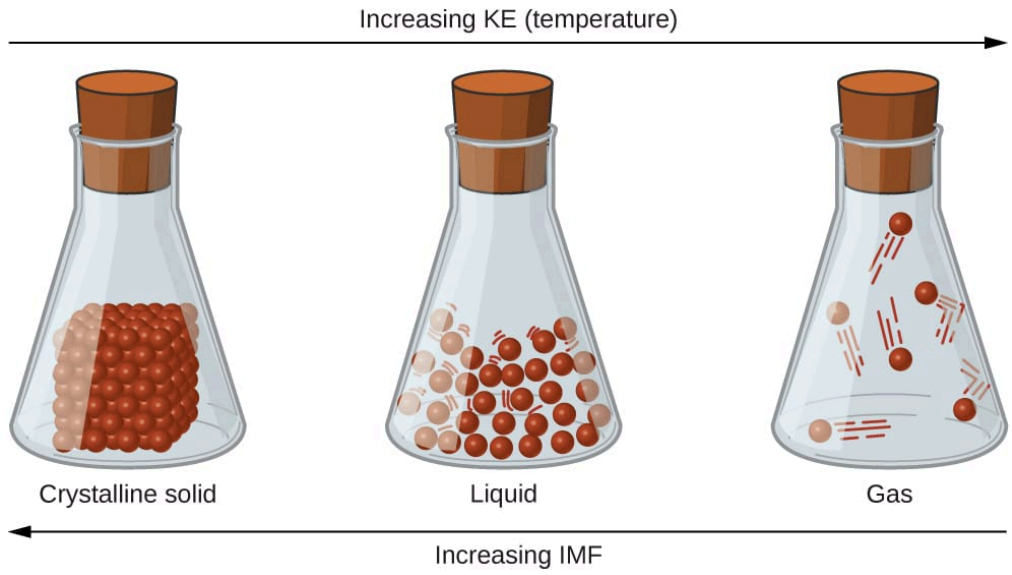


Figure 9.2.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown). (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.1.

As an example of the processes depicted in this figure, consider a sample of water. When gaseous water is cooled sufficiently, the attractions between H_2O molecules will be capable of holding them together when they come into contact with each other; the gas condenses, forming liquid H_2O. For example, liquid water forms on the outside of a cold glass as the water vapor in the air is cooled by the cold glass, as seen in Figure 9.2.2.



Figure 9.2.2: Condensation forms when water vapor in the air is cooled enough to form liquid water. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.2.



Figure 9.2.3: Gaseous butane is compressed within the storage compartment of a disposable lighter, resulting in its condensation to the liquid state. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.3.

Finally, if the temperature of a liquid becomes sufficiently low, or the pressure on the liquid becomes sufficiently high, the molecules of the liquid no longer have enough KE to overcome the IMF between them, and a solid forms. A more thorough discussion of these and other changes of state, or phase transitions, is provided in a later module of this chapter.

Forces between Molecules

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not intramolecular forces. Intramolecular forces are those within the molecule that keep the molecule together, for example, the bonds between the atoms. Intermolecular forces are the attractions between molecules, which determine many of the physical properties of a substance. Figure 9.2.4: illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

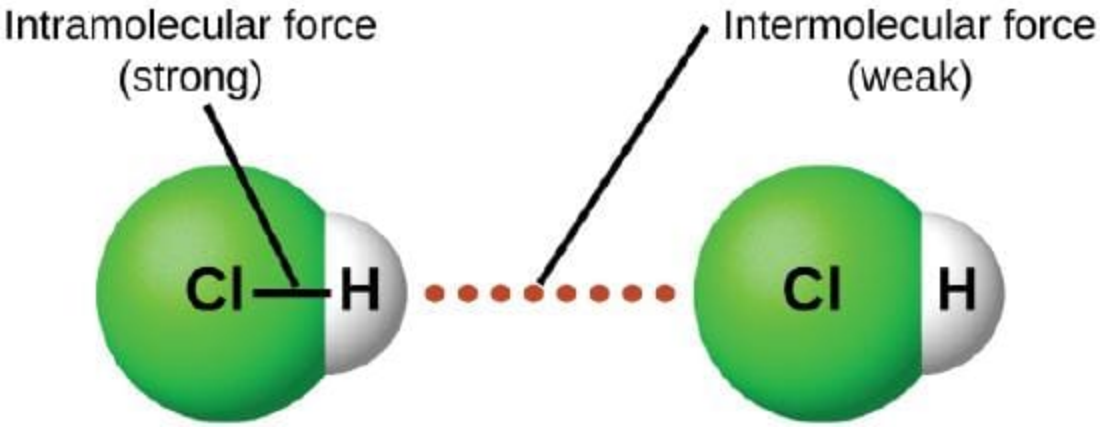


Figure 9.2.4: *Intra*molecular forces keep a molecule intact. *Inter*molecular forces hold multiple molecules together and determine many of a substance's properties. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.4.

All of the attractive forces between neutral atoms and molecules are known as van der Waals forces, although they are usually referred to more informally as intermolecular attraction. We will consider the various types of IMFs in the next three sections of this module.

Dispersion Forces

When talking about dispersion forces, it's worth starting with what a dipole is. A **dipole** is a separation of electrical charge within a molecule or atom, where one part becomes slightly negative (δ⧾) and another part becomes slightly positive (δ⧾). This happens when electrons are unevenly distributed. Dipoles can be permanent, due to differences in electronegativity between atoms, or temporary, caused by the constant motion of electrons.

One of the three van der Waals forces is present in all condensed phases, regardless of the nature of the atoms or molecules composing the substance. This attractive force is called the London dispersion force in honor of German-born American physicist Fritz London who, in 1928, first explained it. This force is often referred to as simply the **dispersion force**. Because the electrons of an atom or molecule are in constant motion (or, alternatively, the electron's location is subject to quantum-mechanical variability), at any moment in time, an atom or molecule can develop a temporary, **instantaneous dipole** if its electrons are distributed asymmetrically. The presence of this dipole can, in turn, distort the electrons of a neighboring atom or molecule, producing an **induced dipole**. These two rapidly fluctuating, temporary dipoles thus result in a relatively weak electrostatic attraction between the species—a so-called dispersion force like that illustrated in Figure 9.2.5.

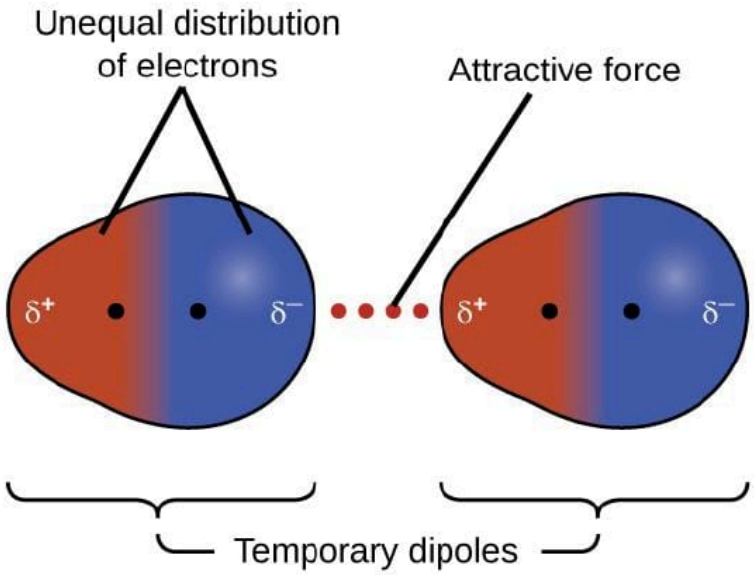


Figure 9.2.5: Dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.5.

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. Larger and heavier atoms and molecules exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 9.2.1.

Table 9.2.1: Melting and Boiling Points of the Halogens

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as polarizability. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

Example 9.2.1: London Forces and Their Effects

Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4, SiH_4, GeH_4, and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be $CH_4 < SiH_4 < GeH_4 < SnH_4$.

A graph of the actual boiling points of these compounds versus the period of the group 14 element (Figure 9.2.6:) shows this prediction to be correct:

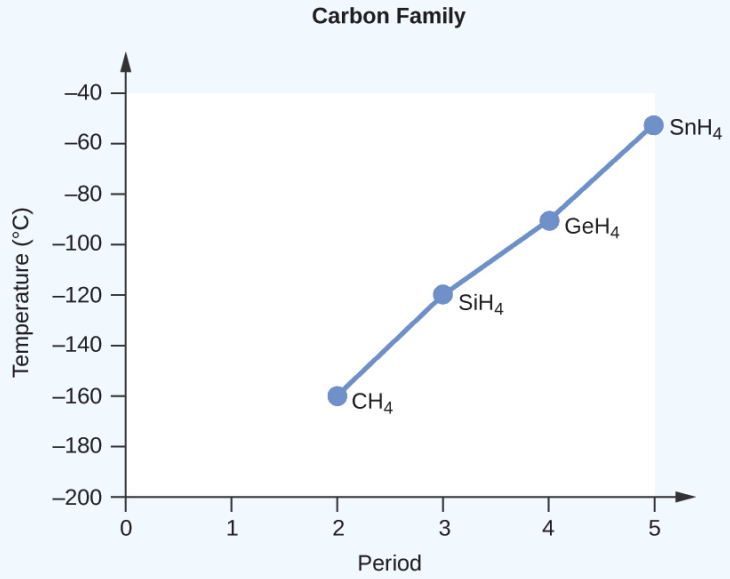


Figure 9.2.6: Boiling points of group 14 tetra hydrides (CH_4, SiH_4, GeH_4, SnH_4) increase down the group, consistent with rising molar mass and polarizability. This trend confirms that dispersion forces—stronger in larger, more polarizable molecules—are the dominant intermolecular force in these nonpolar compounds. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.6.

The shapes of molecules also affect the magnitudes of the dispersion forces between them. For example, boiling points for the isomers *n*-pentane, isopentane, and neopentane (shown in Figure 9.2.7) are 36 °C, 27 °C, and 9.5 °C, respectively. Even though these compounds are composed of molecules with the same chemical formula, C_5H_{12}, the difference in boiling points suggests that dispersion forces in the liquid phase are different, being greatest for *n*-pentane and least for neopentane. The elongated shape of *n*-pentane provides a greater surface area available for contact between molecules, resulting in correspondingly stronger dispersion forces. The more compact shape of isopentane offers a smaller surface area available for intermolecular contact and, therefore, weaker dispersion forces. Neopentane molecules are the most compact of the three, offering the least available surface area for intermolecular contact and, hence, the weakest dispersion forces. This behavior is analogous to the connections that may be formed between strips of VELCRO brand fasteners: the greater the area of the strip's contact, the stronger the connection.

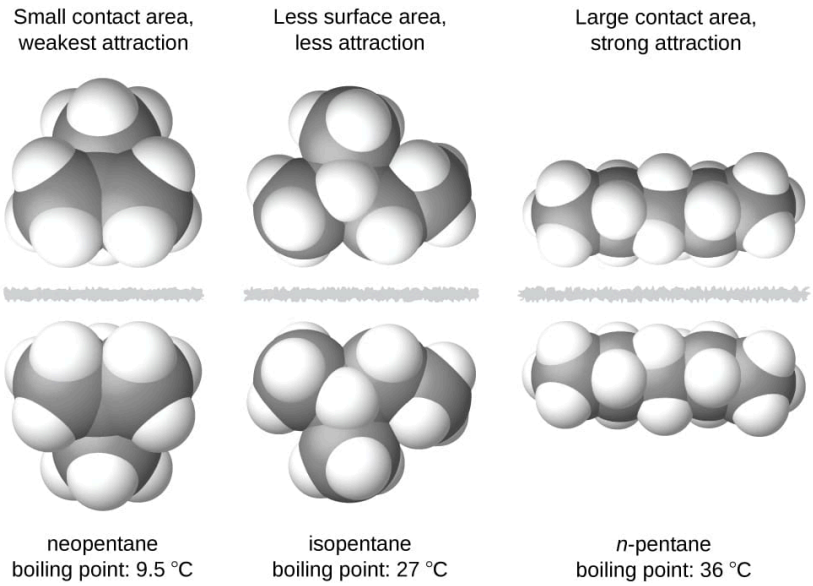


Figure 9.2.7: The strength of the dispersion forces increases with the contact area between molecules, as demonstrated by the boiling points of these pentane isomers. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.7.

Chemistry in Everyday Life: Geckos and Intermolecular Forces

Geckos have an amazing ability to adhere to most surfaces. They can quickly run up smooth walls and across ceilings that have no toe-holds, and they do this without having suction cups or a sticky substance on their toes. And while a gecko can lift its feet easily as it walks along a surface, if you attempt to pick it up, it sticks to the surface. How are geckos (as well as spiders and some other insects) able to do this? Although this phenomenon has been investigated for hundreds of years, scientists only recently uncovered the details of the process that allows geckos' feet to behave this way.

Geckos' toes are covered with hundreds of thousands of tiny hairs known as *setae*, with each seta, in turn, branching into hundreds of tiny, flat, triangular tips called spatulae. The huge numbers of spatulae on its setae provide a gecko, shown in Figure 9.2.8, with a large total surface area for sticking to a surface. In 2000, Kellar Autumn, who leads a multi-institutional gecko research team, found that geckos adhered equally well to both polar silicon dioxide and nonpolar gallium arsenide. This proved that geckos stick to surfaces because of dispersion forces—weak intermolecular attractions arising from temporary, synchronized charge distributions between adjacent molecules. Although dispersion forces are very weak, the total attraction over millions of spatulae is large enough to support many times the gecko's weight.

In 2014, two scientists developed a model to explain how geckos can rapidly transition from "sticky" to "non-sticky." Alex Greaney and Congcong Hu at Oregon State University described how geckos can achieve this by changing the angle between their spatulae and the surface. Geckos' feet, which are normally nonsticky, become sticky when a small shear force is applied. By curling and uncurling their toes, geckos can alternate between sticking and unsticking from a surface, and thus easily move across it. Later research led by Alyssa Stark at University of Akron showed that geckos can maintain their hold on hydrophobic surfaces (similar to the leaves in their habitats) equally well whether the surfaces were wet or dry. Stark's experiment used a ribbon to gently pull the geckos until they slipped, so that the researchers could determine the geckos' ability to hold various surfaces under wet and dry conditions. Further investigations may eventually lead to the development of better adhesives and other applications.

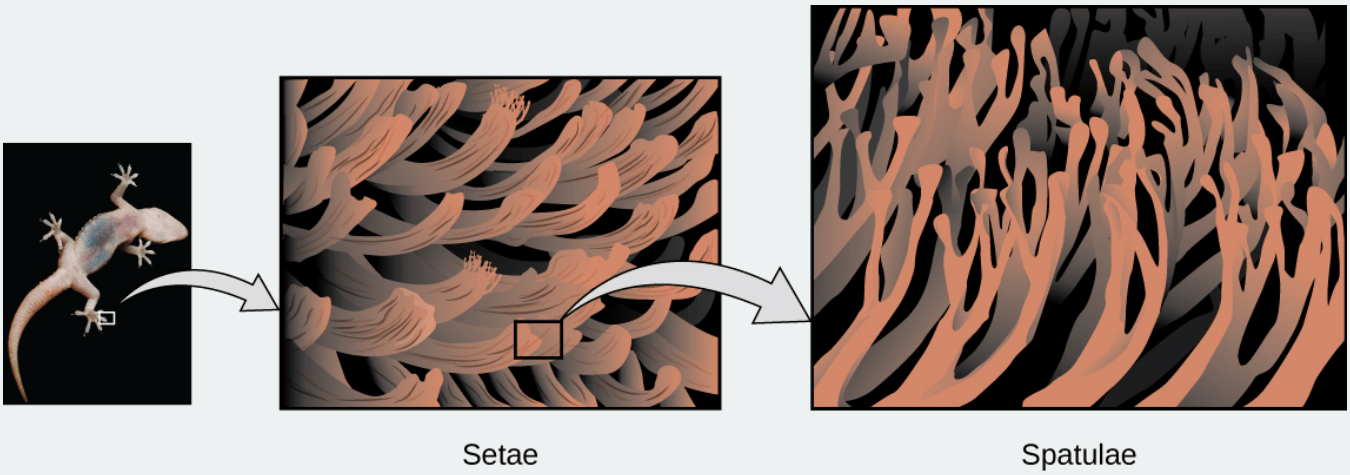


Figure 9.2.8: Geckos' toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko's millions of spatulae. By changing how the spatulae contact the surface, geckos can turn their stickiness "on" and "off." (CC BY 4.0(opens in new window); via OpenStax). Detailed Description of Figure 9.2.8.

Link to Learning

Watch this video to learn more about Kellar Autumn's research that determined that van der Waals forces are responsible for a gecko's ability to cling and climb.

Dipole-Dipole Attractions

Recall from the chapter on chemical bonding and molecular geometry that *polar* molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a dipole. Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative

charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a **dipole-dipole interactions**—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in Figure 9.2.9.

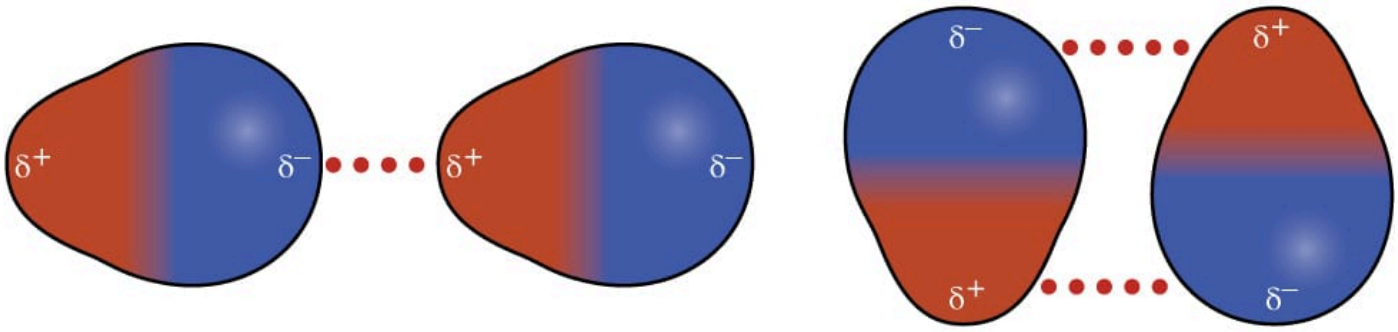


Figure 9.2.9: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.9.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F_2 molecules. Both HCl and F_2 consist of the same number of atoms and have approximately the same molecular mass. At a temperature of 150 K, molecules of both substances would have the same average KE. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to "stick together" to form a liquid, whereas the relatively weaker dispersion forces between nonpolar F_2 molecules are not, and so this substance is gaseous at this temperature. The higher normal boiling point of HCl (188 K) compared to F_2 (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F_2 molecules. We will often use values such as boiling or freezing points, or enthalpies of vaporization or fusion, as indicators of the relative strengths of IMFs of attraction present within different substances.

Example 9.2.2: Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

Solution

CO and N_2 are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N_2 is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N_2 molecules, so CO is expected to have the higher boiling point.

Hydrogen Bonding

Nitrosyl fluoride (ONF, molecular mass 49 amu) is a gas at room temperature. Water (H_2O, molecular mass 18 amu) is a liquid, even though it has a lower molecular mass. We clearly cannot attribute this difference between the two compounds to dispersion forces. Both molecules have about the same shape and ONF is the heavier and larger molecule. It is, therefore, expected to experience more significant dispersion forces. Additionally, we cannot attribute this difference in boiling points to differences in the dipole moments of the molecules. Both molecules are polar and exhibit comparable dipole moments. The large difference between the boiling points is due to a particularly strong dipole-dipole attraction that may occur when a molecule contains a hydrogen atom bonded to a fluorine, oxygen, or nitrogen atom (the three most electronegative elements). The very large difference in electronegativity between the H atom (2.1) and the atom to which it is bonded (4.0 for an F atom, 3.5 for an O atom, or 3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to *highly concentrated partial charges* with these atoms. Molecules with F-H, O-H, or N-H moieties are very strongly attracted to similar moieties in nearby molecules, a particularly strong type of dipole-dipole attraction called **hydrogen bonding**. Examples of hydrogen bonds include HF···HF, H_2O···HOH, and H_3N···HNH_2, in which the hydrogen bonds are denoted by dots. Figure 9.2.10 illustrates hydrogen bonding between water molecules.

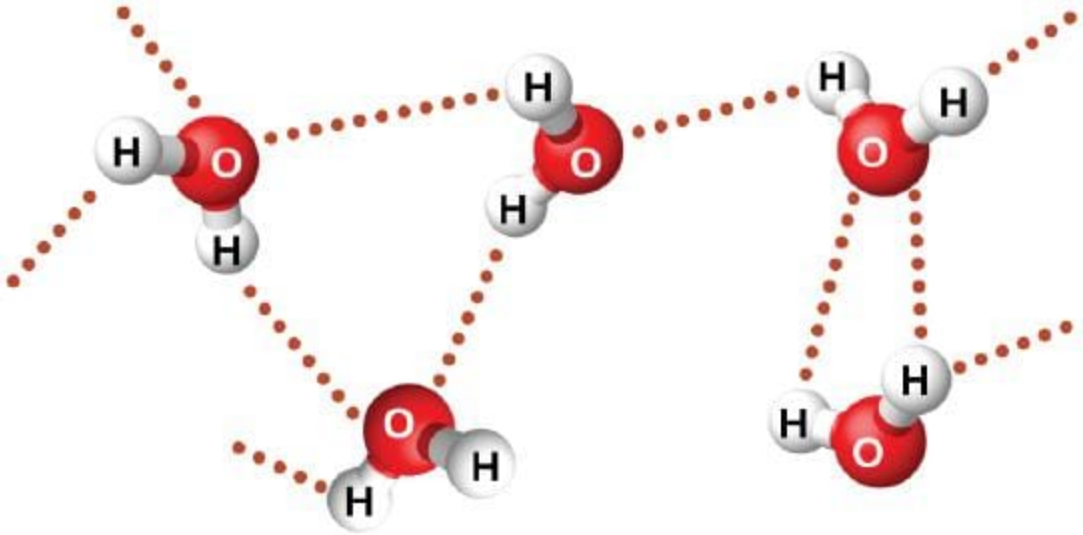


Figure 9.2.10: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.10.

Despite use of the word "bond," keep in mind that hydrogen bonds are *intermoledular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, only about 5 to 10% as strong, but are generally much stronger than other dipole-dipole attractions and dispersion forces.

Hydrogen bonds have a pronounced effect on the properties of condensed phases (liquids and solids). For example, consider the trends in boiling points for the binary hydrides of group 15 (NH_3, PH_3, AsH_3, and SbH_3), group 16 hydrides (H_2O, H_2S, H_2Se, and H_2Te), and group 17 hydrides (HF, HCl, HBr, and HI). The boiling points of the heaviest three hydrides for each group are plotted in Figure 9.2.11. As we progress down any of these groups, the polarities of the molecules decrease slightly, whereas the sizes of the molecules increase substantially. The effect of increasingly stronger dispersion forces dominates that of increasingly weaker dipole-dipole attractions, and the boiling points are observed to increase steadily.

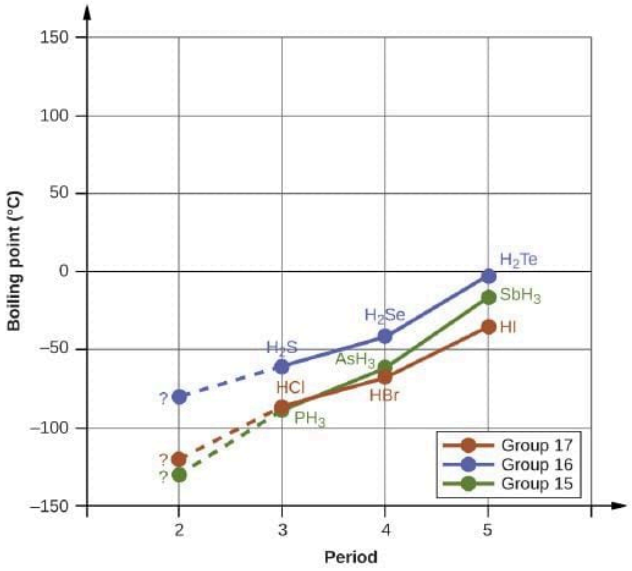


Figure 9.2.11: For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.11.

If we use this trend to predict the boiling points for the lightest hydride for each group, we would expect NH_3 to boil at about −120 °C, H_2O to boil at about −80 °C, and HF to boil at about −110 °C. However, when we measure the boiling points for these compounds, we find that they are dramatically higher than the trends would predict, as shown in Figure 9.2.12. The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

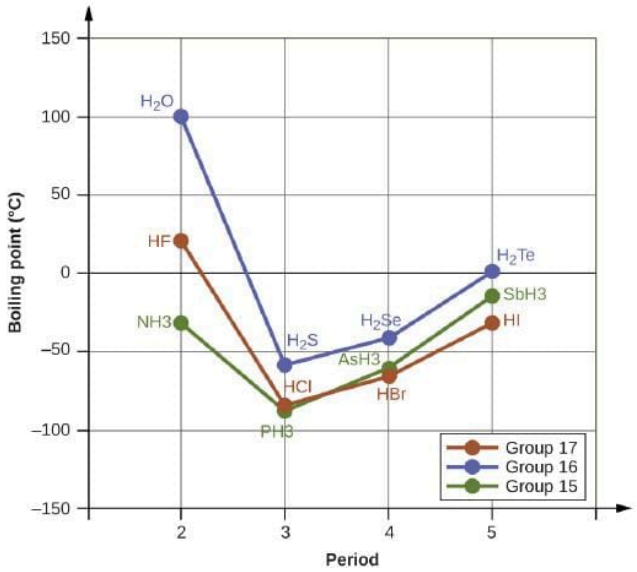


Figure 9.2.12: In comparison to periods 3–5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding. (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.12.

Example 9.2.3: Effect of Hydrogen Bonding on Boiling Points

Consider the compounds dimethylether (CH_3OCH_3), ethanol (CH_3CH_2OH), and propane ($CH_3CH_2CH_3$). Their boiling points, not necessarily in order, are −42.1 °C, −24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

Solution

The VSEPR-predicted shapes of CH_3OCH_3, CH_3CH_2OH, and $CH_3CH_2CH_3$ are similar, as are their molar masses (46 g/mol, 46 g/mol, and 44 g/mol, respectively), so they will exhibit similar dispersion forces. Since $CH_3CH_2CH_3$ is nonpolar, it may exhibit *only* dispersion forces. Because CH_3OCH_3 is polar, it will also experience dipole-dipole attractions. Finally, CH_3CH_2OH has an –OH group, and so it will experience the uniquely strong dipole-dipole attraction known as hydrogen bonding. So the ordering in terms of strength of IMFs, and thus boiling points, is $CH_3CH_2CH_3 < CH_3OCH_3 < CH_3CH_2OH$. The boiling point of propane is −42.1 °C, the boiling point of dimethylether is −24.8 °C, and the boiling point of ethanol is 78.5 °C.

How Sciences Interconnect: Hydrogen Bonding and DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism's characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism's offspring. A DNA molecule consists of two (anti-)parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in Figure 9.2.13.

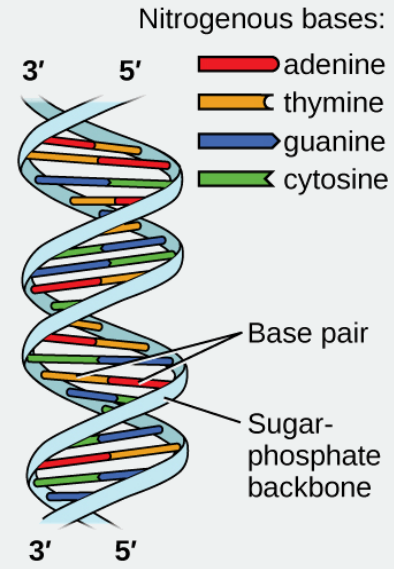


Figure 9.2.13: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (CC BY 4.0;via OpenStax). Detailed Description of Figure 9.2.13.

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure Figure 9.2.14.

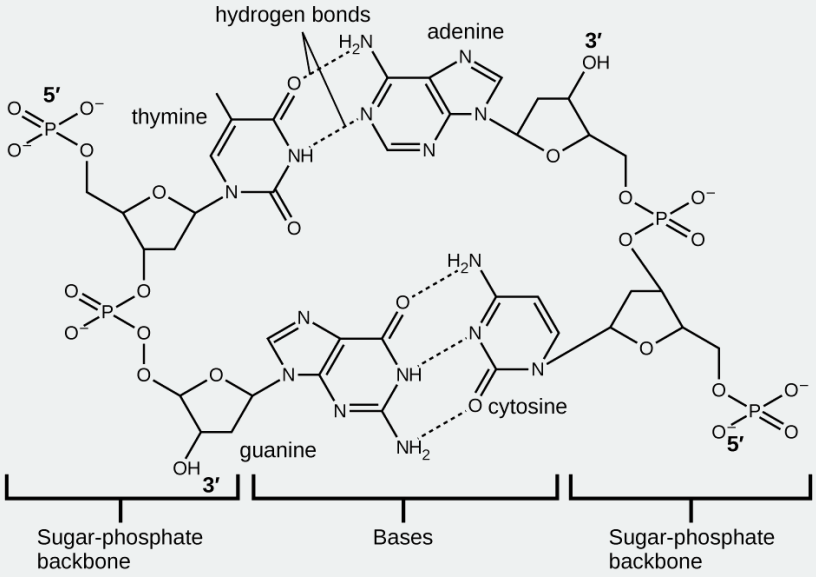


Figure 9.2.14: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called “complementary base pairs.” (CC BY 4.0; via OpenStax). Detailed Description of Figure 9.2.14.

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily “unzip” down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

Key Takeaways

- Intermolecular forces are attractions between molecules that influence physical properties like boiling and melting points.
- Dispersion Forces are the weakest and exist in all molecules, especially nonpolar ones.
- Dipole-dipole interactions occur between polar molecules with permanent dipoles.
- Hydrogen bonding is the strongest, occurring when H is bonded to N, O, or F.

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9.3: Gases and the Kinetic-Molecular Theory

Gases

In this section, you will learn how the kinetic-molecular theory describes the motion and behavior of gas particles. This theory explains how properties like pressure and temperature arise from particle movement and collisions, and provides the foundation for understanding gas laws and ideal gas behavior.

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 Molecular Theory of Gases**. The kinetic theory of gases 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 9.3.1 shows a representation of how we mentally picture the gas phase.

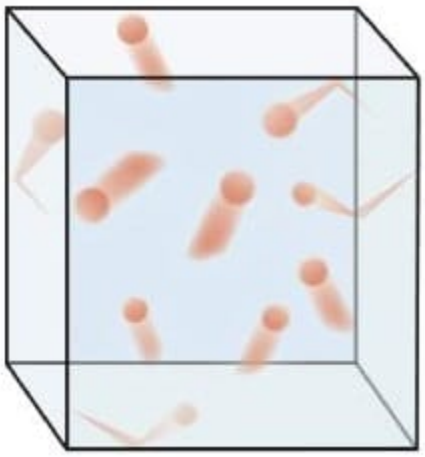


Figure 9.3.1: The Kinetic Theory of Gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion, with a lot of distance between the particles. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.3.1.

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, and 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.

Key Takeaways

- Gas particles move constantly and randomly with negligible volume.
- Collisions between particles and container walls create pressure.
- Temperature reflects the average kinetic energy of gas particles.

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9.4: Units of Gas Pressure

Introduction

At the macroscopic level, a complete physical description of a sample of a gas requires four quantities: temperature (expressed in kelvins), volume (expressed in liters), amount (expressed in moles), and pressure (in atmospheres). As we explain in this section, these variables are not independent. If we know the values of any three of these quantities, we can calculate the fourth and thereby obtain a full physical description of the gas. Temperature, volume, and amount have been discussed in previous chapters. We now discuss pressure and its units of measurement.

Units of Pressure

Any object, whether it is your computer, a person, or a sample of gas, exerts a force on any surface with which it comes in contact. The air in a balloon, for example, exerts a force against the interior surface of the balloon, and a liquid injected into a mold exerts a force against the interior surface of the mold, just as a chair exerts a force against the floor because of its mass and the effects of gravity. If the air in a balloon is heated, the increased kinetic energy of the gas eventually causes the balloon to burst because of the increased pressure (an increase in the collisions between gas molecules) inside the balloon. Pressure (P) is defined as the amount of force (F) per unit area (A):

$$P = \frac{F}{A}$$

Pressure is dependent on both the force exerted and the size of the area to which the force is applied. We know from experience that applying the same force to a smaller area produces a higher pressure. When we use a hose to wash a car, for example, we can increase the pressure of the water by reducing the size of the opening of the hose with a thumb.

The units of pressure are derived from the units used to measure force and area. The SI unit for pressure, derived from the SI units for force (newtons) and area (square meters), is the newton per square meter (N/m^2), which is called the Pascal (Pa), after the French mathematician Blaise Pascal (1623–1662):

$$1\ \text{Pa} = 1\ \text{N/m}^2 \tag{9.4.1}$$

Example 9.4.1

Assuming a paperback book has a mass of 2.00 kg, a length of 27.0 cm, a width of 21.0 cm, and a thickness of 4.5 cm, what pressure does it exert on a surface if it is

a. lying flat?
b. standing on edge in a bookcase?

Given: mass and dimensions of object

Asked for: pressure

Strategy:

a. Calculate the force exerted by the book and then compute the area that is in contact with a surface.
b. Substitute these two values into Equation to find the pressure exerted on the surface in each orientation ($P = \frac{F}{A}$).

Solution:

The force exerted by the book does *not* depend on its orientation. Recall that the force exerted by an object is $F = ma$, where m is its mass and a is its acceleration. In Earth's gravitational field, the acceleration is due to gravity (9.8067 m/s^2 at Earth's surface). In SI units, the force exerted by the book is therefore

$$F = ma = 2.00\ \text{kg} \times 9.8067 \frac{\text{m}}{\text{s}^2} = 19.6 \frac{\text{kg} \cdot \text{m}}{\text{s}^2} = 19.6\ \text{N}$$

a. We calculated the force as 19.6 N. When the book is lying flat, the area is

$$A = 0.270\ \text{m} \times 0.210\ \text{m} = 0.0567\ \text{m}^2.$$

The pressure exerted by the text lying flat is thus

$$P = \frac{F}{A} = \frac{19.6\ \text{N}}{0.0567\ \text{m}^2} = 3.46 \times 10^2 \text{Pa}$$

b. If the book is standing on its end, the force remains the same, but the area decreases:

$$\text{A} = 21.0\ \text{cm} \times 4.5\ \text{cm} = 0.210\ \text{m} \times 0.045\ \text{m} = 9.5 \times 10^{-3}\ \text{m}^2$$

The pressure exerted by the text lying flat is thus

$$P = \frac{19.6\ \text{N}}{9.5 \times 10^{-3}\ \text{m}^2} = 2.06 \times 10^3\ \text{Pa}$$

Atmospheric Pressure

Just as we exert pressure on a surface because of gravity, so does our atmosphere. We live at the bottom of an ocean of gases that becomes progressively less dense with increasing altitude. Approximately 99% of the mass of the atmosphere lies within 30 km of Earth's surface, and half of it is within the first 5.5 km (Figure 9.4.1). Every point on Earth's surface experiences a net pressure called atmospheric pressure. The pressure exerted by the atmosphere is considerable: a 1.0 m^2 column, measured from sea level to the top of the atmosphere, has a mass of about 10,000 kg, which gives a pressure of about 100 kPa.

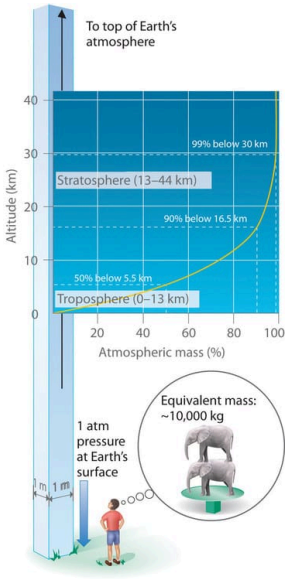


Figure 9.4.1: Each square meter of Earth's surface supports a column of air that is more than 200 km high and weighs about 10,000 kg at Earth's surface, resulting in a pressure at the surface of 1.01 × 105 N/m2. This corresponds to a pressure of 101 kPa = 760 mmHg = 1 atm. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.4.1.

Atmospheric pressure can be measured using a barometer, a device invented in 1643 by one of Galileo's students, Evangelista Torricelli (1608–1647). A barometer may be constructed from a long glass tube that is closed at one end. It is filled with mercury

and placed upside down in a dish of mercury without allowing any air to enter the tube. Some of the mercury will run out of the tube, but a relatively tall column remains inside (Figure 9.4.2). Why doesn't all the mercury run out? Gravity is certainly exerting a downward force on the mercury in the tube, but it is opposed by the pressure of the atmosphere pushing down on the surface of the mercury in the dish, which has the net effect of pushing the mercury up into the tube. Because there is no air above the mercury inside the tube in a properly filled barometer (it contains a vacuum), there is no pressure pushing down on the column. Thus the mercury runs out of the tube until the pressure exerted by the mercury column itself exactly balances the pressure of the atmosphere.

Under normal weather conditions at sea level, the two forces are balanced when the top of the mercury column is approximately 760 mm above the level of the mercury in the dish, as shown in Figure 9.4.2. This value varies with meteorological conditions and altitude. In Denver, Colorado, for example, at an elevation of about 1 mile, or 1609 m (5280 ft), the height of the mercury column is 630 mm rather than 760 mm.

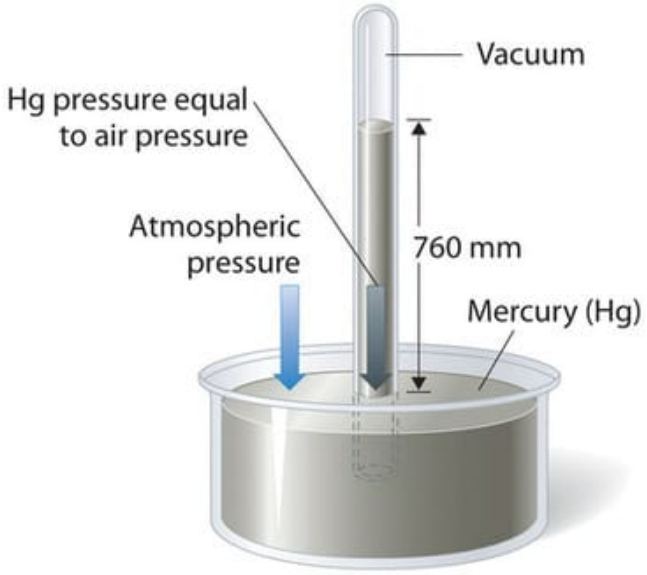


Figure 9.4.2: The pressure exerted by the atmosphere on the surface of the pool of mercury supports a column of mercury in the tube that is about 760 mm tall. Because the boiling point of mercury is quite high (356.73°C), there is very little mercury vapor in the space above the mercury column. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.4.2.

Mercury barometers have been used to measure barometric pressure for so long that they have their own unit for pressure: the millimeter of mercury (mmHg), often called the torr, after Torricelli. Standard barometric pressure is the barometric pressure required to support a column of mercury exactly 760 mm tall; this pressure is also referred to as 1 atmosphere (atm). These units are also related to the pascal:

$$1 \text{ atm} = 760\ mmHg \tag{9.4.2}$$
$$= 760\ torr \tag{9.4.3}$$
$$= 1.01325 \times 10^5\ Pa \tag{9.4.4}$$
$$= 101.325\ kPa \tag{9.4.5}$$

Thus a pressure of 1 atm equals 760 mmHg exactly.

We are so accustomed to living under this pressure that we never notice it. Instead, what we notice are *changes* in the pressure, such as when our ears pop in fast elevators in skyscrapers or in airplanes during rapid changes in altitude. We make use of barometric pressure in many ways. We can use a drinking straw because sucking on it removes air and thereby reduces the pressure inside the straw. The barometric pressure pushing down on the liquid in the glass then forces the liquid up the straw.

Example 9.4.2

One of the authors visited Rocky Mountain National Park several years ago. After departing from an airport at sea level in the eastern United States, he arrived in Denver (altitude 5280 ft), rented a car, and drove to the top of the highway outside Estes Park (elevation 14,000 ft). He noticed that even slight exertion was very difficult at this altitude, where the barometric pressure is only 454 mmHg. Convert this pressure to

a. atmospheres (atm).
b. bar.

Given: pressure in millimeters of mercury

Asked for: pressure in atmospheres and bar

Strategy:

Use the conversion factors in Equation 9.4.5 to convert from millimeters of mercury to atmospheres and kilopascals.

Solution:

From Equation 9.4.5, we have 1 atm = 760 mmHg = 101.325 kPa. The pressure at 14,000 ft in atm is thus

$$P = 454 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \tag{9.4.6}$$
$$= 0.597 \ atm$$

The pressure in bar is given by

$$P = 0.597 \text{ atm} \times \frac{1.01325 \text{ bar}}{1 \text{ atm}} \tag{9.4.7}$$
$$= 0.605 \ bar$$

Barometers measure atmospheric pressure, but **manometers** measure the pressures of samples of gases contained in an apparatus. The key feature of a manometer is a U-shaped tube containing mercury (or occasionally another nonvolatile liquid). A closed-end manometer is shown schematically in part (a) in Figure 9.4.3. When the bulb contains no gas (i.e., when its interior is a near vacuum), the heights of the two columns of mercury are the same because the space above the mercury on the left is a near vacuum (it contains only traces of mercury vapor). If a gas is released into the bulb on the right, it will exert a pressure on the mercury in the right column, and the two columns of mercury will no longer be the same height. The difference between the heights of the two columns is equal to the pressure of the gas.

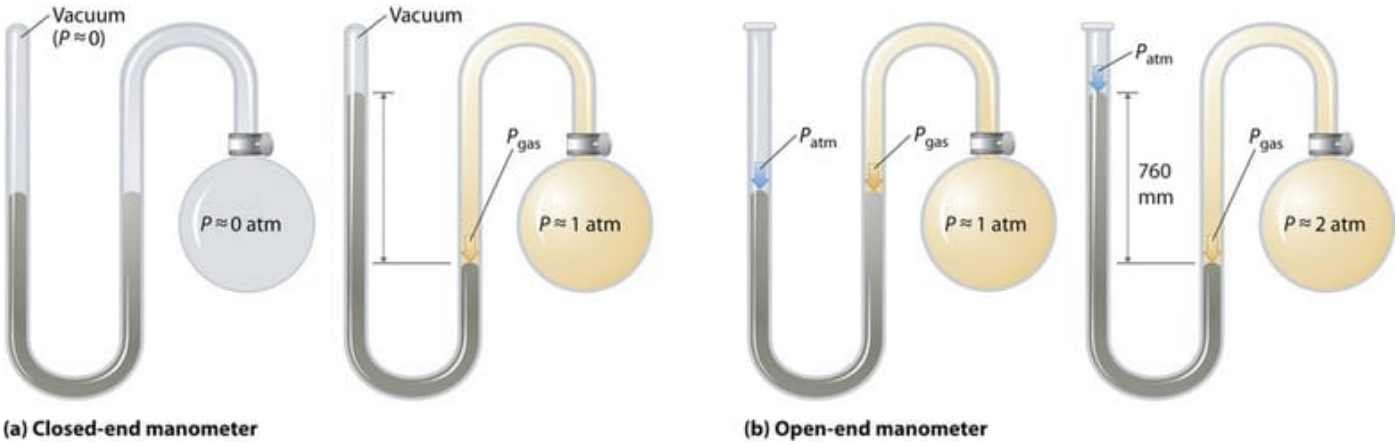


Figure 9.4.3: (a) In a closed-end manometer, the space above the mercury column on the left (the reference arm) is essentially a vacuum (P ≈ 0), and the difference in the heights of the two columns gives the pressure of the gas contained in the bulb directly. (b) In an open-end manometer, the left (reference) arm is open to the atmosphere (P ≈ 1 atm), and the difference in the heights of the two columns gives the difference between barometric pressure and the pressure of the gas in the bulb. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.4.3.

If the tube is open to the atmosphere instead of closed, as in the open-end manometer shown in part (b) in Figure 9.4.3, then the two columns of mercury have the same height only if the gas in the bulb has a pressure equal to the barometric pressure. If the gas in the bulb has a *higher* pressure, the mercury in the open tube will be forced up by the gas pushing down on the mercury in the other arm of the U-shaped tube. The pressure of the gas in the bulb is therefore the sum of the barometric pressure (measured with a barometer) and the difference in the heights of the two columns. If the gas in the bulb has a pressure *less* than that of the atmosphere, then the height of the mercury will be greater in the arm attached to the bulb. In this case, the pressure of the gas in the bulb is the barometric pressure minus the difference in the heights of the two columns.

Key Takeaways

- Pressure is force per unit area, measured in Pascals (Pa).
- A complete gas description needs temperature, volume, amount, and pressure.
- Standard atmospheric pressure is 1 atm = 760 mmHg = 101.3 kpa.
- Barometers and manometers are tools to measure gas pressure.

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9.5: Gas Laws

Overview

In this section, we will explore the fundamental gas laws that describe how gases behave under different conditions. These laws—Boyle's Law, Charles's Law, Gay-Lussac's Law, Avogadro's Law, the Combined Gas Law, and Dalton's Law—each explain the relationships between pressure, volume, temperature, and the number of moles of gas. We will examine them one by one to understand how changes in these variables affect gases.

Boyle's Law - The Relation between Volume and Pressure

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. We say 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 (*k*) for a given amount of gas at a constant temperature:

$$P \times V = k \; (at \; constant \; n \; and \; T)$$

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

$$P_1V_1 = k = P_2V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1V_1 = P_2V_2 \; (at \; constant \; n \; and \; T)$$

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 $9.5.1$ shows two representations of how Boyle's law works.

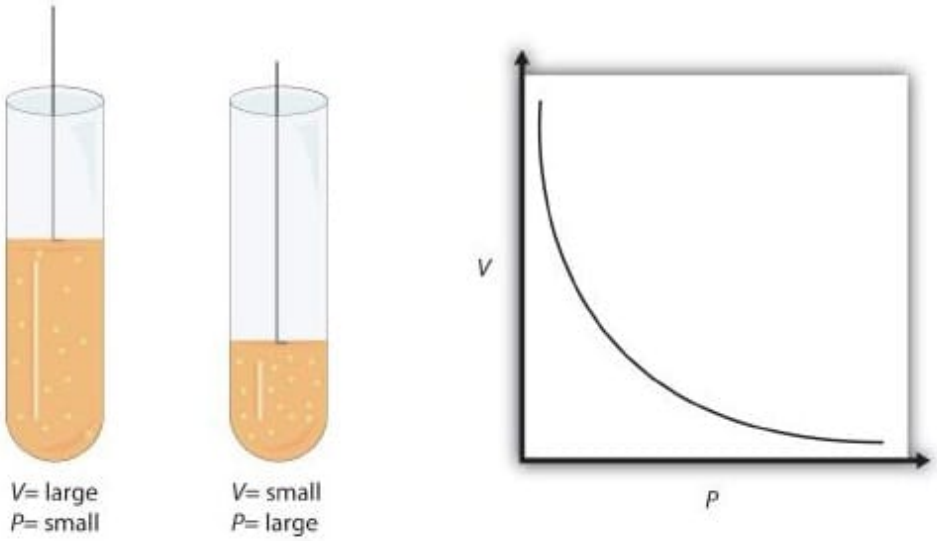


Figure 9.5.1: Boyle's Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot P versus V for a given amount of gas at a certain temperature; such a plot will look like the graph on the right. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.5.1.

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 9.5.1

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

First, determine what quantities we are given. We are given an initial pressure and an initial volume, so let these values be P_1 and V_1:

$$P_1 = 2.44 \text{ atm and } V_1 = 4.01 \text{ L}$$

We are given another quantity, final pressure of 1.93 atm, but not a final volume. This final volume is the variable we will solve for.

$$P_2 = 1.93 \text{ atm and } V_2 = ? \text{ L}$$

Substituting these values into Boyle's law, we get

$$(2.44 \text{ atm})(4.01 \text{ L}) = (1.93 \text{ atm})V_2$$

To solve for the unknown variable, we isolate it by dividing both sides of the equation by 1.93 atm—both the number *and* the unit:

$$\frac{(2.44\,atm)(4.01\,L)}{1.93\,atm} = \frac{(1.93\,atm)\,V_2}{1.93\,atm}$$

Note that, on the left side of the equation, the unit *atm* is in the numerator and the denominator of the fraction. They cancel algebraically, just as a number would. On the right side, the unit *atm* and the number 1.93 are in the numerator and the denominator, so the entire quantity cancels:

$$\frac{(2.44\ \cancel{atm}\ )(4.01\,L)}{1.93\ \cancel{atm}} = \frac{(1.93\ \cancel{atm}\ )\,V_2}{1.93\ \cancel{atm}}$$

What we have left is

$$\frac{(2.44)(4.01\,L)}{1.93} = V_2$$

Now we simply multiply and divide the numbers together and combine the answer with the L unit, which is a unit of volume. Doing so, we get $V_2 = 5.07\ L$

Does this answer make sense? We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 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 answer makes sense based on Boyle's law.

As mentioned, you can use any units for pressure or volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

Example 9.5.2

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

We can still use Boyle's law to answer this, but now the two volume quantities have different units. It does not matter which unit we change, as long as we perform the conversion correctly. Let us change the 0.663 L to milliliters:

$$0.663\,L \times \frac{1000\,ml}{1\,L} = 663\,ml$$

Now that both volume quantities have the same units, we can substitute into Boyle's law:

$$(722\,torr)(88.8\,ml) = P_2(663\,ml)$$

$$\frac{(722\,torr)(88.8)\,ml}{(663\,ml)} = P_2$$

The mL units cancel, and we multiply and divide the numbers to get P_2 = 96.7 torr

The volume is increasing, and the pressure is decreasing, which is as expected for Boyle's law.

Freshly-baked bread 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 stoppage of molecular motion.

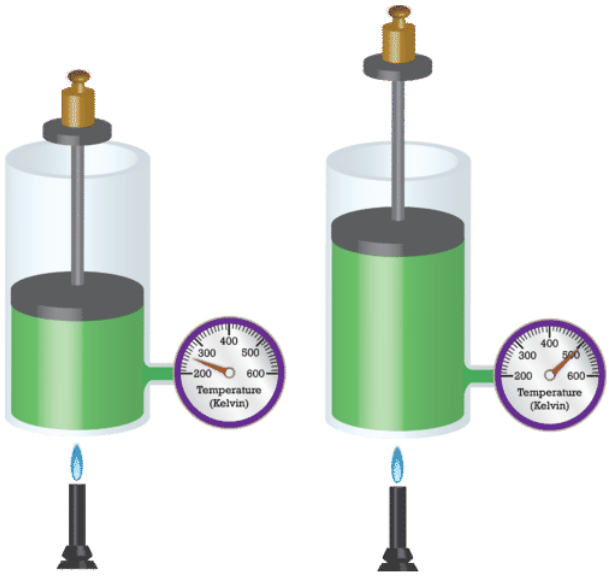


Figure 9.5.2: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume. (CC BY-NC; via CK-12). Detailed Description of Figure 9.5.2.

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 (Table 9.5.1) 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 9.5.1: 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
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 (Figure 9.5.3), the result is a straight line, indicative of a direct relationship, shown in the figure below.

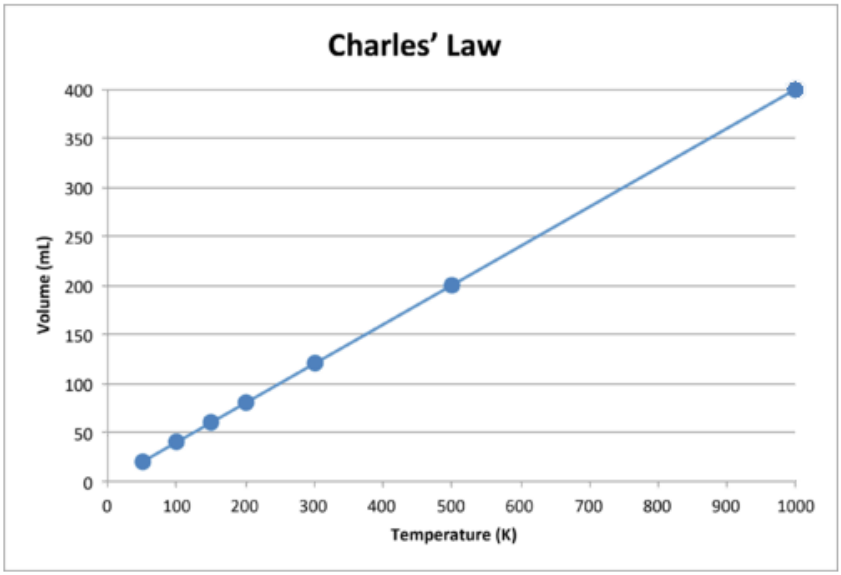


Figure 9.5.3: The volume of a gas increases as the Kelvin temperature increases. (CC BY-NC; via CK-12). Detailed Description of Figure 9.5.3.

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 9.5.3

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.

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

Known

- $V_1 = 2.20\ \mathrm{L}$
- $T_1 = 22^{\circ}\mathrm{C} = 295\ \mathrm{K}$
- $T_2 = 71^{\circ}\mathrm{C} = 344\ \mathrm{K}$

Unknown

Use Charles's Law to solve for the unknown volume (V_2). The temperatures have first been converted to Kelvin.

Step 2: Solve.

First, rearrange the equation algebraically to solve for V_2.

$$V_2 = \frac{V_1 \times T_2}{T_1}$$

Now substitute the known quantities into the equation and solve.

$$V_2 = \frac{2.20\ \text{L} \times 344\ \text{K}}{295\ \text{K}} = 2.57\ \text{L}$$

Step 3: Think about your result.

The volume increases as the temperature increases. The result has three significant figures

Gay-Lussac's Law- The Relationship Between Pressure and Temperature

A third gas law may be derived as a corollary to Boyle's and Charles's laws. Suppose we double the thermodynamic temperature of a sample of gas. According to Charles's law, the volume should double. Now, how much pressure would be required at the higher temperature to return the gas to its original volume? According to Boyle's law, we would have to double the pressure to halve the volume. Thus, if the volume of gas is to remain the same, doubling the temperature will require doubling the pressure. This law was first stated by the Frenchman Joseph Gay-Lussac (1778 to 1850). According to **Gay-Lussac's law**, for a given amount of gas held at constant volume, the pressure is proportional to the absolute temperature. Mathematically,

$$P \propto T \text{ or } P = k \times T \text{ or } \frac{P}{T} = k$$

where $\propto$ means "is proportional to," and *k* is a proportionality constant that depends on the identity, amount, and volume of the gas.

In terms of two sets of data: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$. This equation is useful for pressure-temperature calculations for a confined gas at constant volume. Note that temperatures must be on the kelvin scale for any gas law calculations (0 on the kelvin scale and the lowest possible temperature is called absolute zero). (Also note that there are at least three ways we can describe how the pressure of a gas changes as its temperature changes: We can use a table of values, a graph, or a mathematical equation.)

Example 9.5.4

A container is designed to hold a pressure of 2.5 atm. The volume of the container is 20.0 cm^3, and it is filled with air at room temperature (20°C) and normal atmospheric pressure. Would it be safe to throw the container into a fire where temperatures of 600°C would be reached?

Solution

Using the common-sense method, we realize that the pressure will increase at the higher temperature, and so:

$$P_2 = 1.0\ \text{atm}\ \times \frac{(273.15 + 600)\ \text{K}}{(273.15 + 20)\ \text{K}} = 3.0\ \text{atm}$$

This would exceed the safe strength of the container. Note that the volume of the container was not needed to solve the problem.

This concept works in reverse, as well. For instance, if we subject a gas to lower temperatures than their initial state, the external atmosphere can actually force the container to shrink. The following video demonstrates how a sample of hot gas, when cooled will collapse a container. A syringe barrel is filled with hot steam (vaporized water) and a plunger placed to cap off the end. The syringe is then placed in a beaker of ice water to cool the internal gas. When the temperature of the water vapor decreases, the pressure exerted by the vapor decreases as well. This leads to a difference in pressure between the vapor inside the barrel and the atmosphere. Atmospheric pressure then pushes the plunger into the barrel.

Combined Gas Law

One thing we notice about all gas laws, collectively, is that 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 kelvins.

Notice that each of the previous gas laws introduced, can be derived from the combined gas law:

At constant T, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ *gives* Boyle's Law: $P_1V_1 = P_2V_2$

At constant P, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ *gives* Charles's Law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

At constant V, $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ *gives* Gay-Lussac's Law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

In other words, if you know the equation for the combined gas law, you can calculate relationships between pressure, volume, or temperature of a fixed amount of gas.

Example 9.5.5

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

We can use the combined gas law directly; all the units are consistent with each other, and the temperatures are given in Kelvin. Substituting,

$$\frac{(1.82\,atm)(8.33\,L)}{286\,K} = \frac{P_2(5.72\,L)}{355\,K}$$

We rearrange this to isolate the P_2 variable all by itself. When we do so, certain units cancel:

$$\frac{(1.82\,atm)(8.33\,\cancel{L})(355\,\cancel{K})}{(286\,\cancel{K})(5.72\,\cancel{L})} = P_2$$

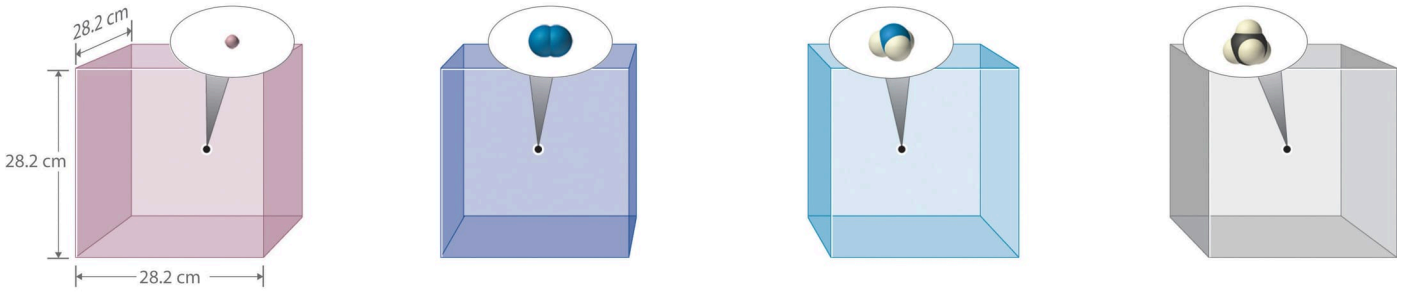
Multiplying and dividing all the numbers, we get

$$P_2 = 3.29\,atm$$

Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas were changing.

Avogadro's Law - The Relation between Volume and Molar Amount

We can demonstrate the relationship between the volume and the amount of a gas by filling a balloon; as we add more gas, the balloon gets larger. The specific quantitative relationship was discovered by the Italian chemist Amadeo Avogadro, who recognized the importance of Gay-Lussac's work on combining volumes of gases. In 1811, Avogadro postulated that, at the same temperature and pressure, equal volumes of gases contain the same number of gaseous particles (Figure 9.5.4). This is the historic "Avogadro's hypothesis."



He
$V = 22.4$ L
$P = 1$ atm
$T = 0$°C
Mass: 4.003 g/mol
$n = 1$ mol

N_2
$V = 22.4$ L
$P = 1$ atm
$T = 0$°C
Mass: 28.013 g/mol
$n = 1$ mol

NH_3
$V = 22.4$ L
$P = 1$ atm
$T = 0$°C
Mass: 17.031 g/mol
$n = 1$ mol

CH_4
$V = 22.4$ L
$P = 1$ atm
$T = 0$°C
Mass: 16.043 g/mol
$n = 1$ mol

Figure 9.5.4: Equal volumes of four different gases at the same temperature and pressure contain the same number of gaseous particles. Because the molar mass of each gas is different, the *mass* of each gas sample is different even though all contain 1 mol of gas. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 9.5.4.

Because the number of particles is related to the number of moles (1 mol = 6.022 × 10^{23} particles), **Avogadro's law** essentially states that equal volumes of different gases, at the same temperature and pressure, contain the same amount (moles, particles) of gas. At constant temperature and pressure, the volume (V) of a sample of gas is directly proportional to the number of moles (n) of gas in the sample. Stated mathematically,

$$V \propto n \text{ or } V = k \times n \text{ or } \frac{V}{n} = k$$

where $\propto$ means "is proportional to," and *k* is a proportionality constant that is the same for all gases.

In terms of two sets of data: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$.

This relationship is valid for most gases at relatively low pressures, but deviations from strict linearity are observed at elevated pressures. Mathematical relationships can also be determined for the other variable pairs, such as *P* versus *n*, and *n* versus *T*.

Example 9.5.6

At constant temperature and pressure, 5.00 L of nitrogen gas contains 0.200 mol. What volume would 0.500 mol of nitrogen gas occupy under the same conditions?

Solution

Use the formula:

V1n1=V2n2\frac{V_1}{n_1} = \frac{V_2}{n_2}n1V1=n2V2

Substitute the known values:

5.00 L0.200 mol=V20.500 mol\frac{5.00\ \text{L}}{0.200\ \text{mol}} = \frac{V_2}{0.500\ \text{mol}}0.200 mol5.00 L=0.500 molV2

Solve for V2V_2V2:

V2=5.00 L0.200 mol×0.500 mol=12.5 LV_2 = \frac{5.00\ \text{L}}{0.200\ \text{mol}} \times 0.500\ \text{mol} = 12.5\ \text{L}V2=0.200 mol5.00 L×0.500 mol=12.5 L

Standard Temperature and Pressure

It should be obvious by now that some physical properties of gases depend strongly on the conditions. What we need is a set of standard conditions so that properties of gases can be properly compared to each other. **Standard Temperature and Pressure (STP)** is defined as exactly 100 kPa of pressure (0.986 atm) and 273 K (0°C). For simplicity, we will use 1 atm as standard pressure. Defining STP allows us to more directly compare the properties of gases that differ from one another.

One property shared among gases is a molar volume. The **molar volume** is the volume of 1 mol of a gas. At STP, the molar volume of a gas can be easily determined by using the ideal gas law:

$$(1\,atm)V = (1\,mol)\left(0.08205\frac{L.atm}{mol.K}\right)(273\,K)$$

All the units cancel except for L, the unit of volume. So *V* = 22.4 L

Note that we have not specified the identity of the gas; we have specified only that the pressure is 1 atm and the temperature is 273 K. This makes for a very useful approximation: any gas at STP has a volume of 22.4 L per mole of gas; that is, the molar volume at STP is 22.4 L/mol (Figure 9.5.1). This molar volume makes a useful conversion factor in stoichiometry problems if the conditions are at STP. If the conditions are not at STP, a molar volume of 22.4 L/mol is not applicable. However, if the conditions are at STP, the combined gas law can be used to calculate what the volume of the gas would be if at STP; then the 22.4 L/mol molar volume can be used.

Example 9.5.6

How many moles of Ar are present in 38.7 L at STP?

Solution

We can use the molar volume, 22.4 L/mol, as a conversion factor, but we need to reverse the fraction so that the L units cancel and mol units are introduced. It is a one-step conversion:

$$38.7\,\cancel{L} \times \frac{1\,mol}{22.4\,\cancel{L}} = 1.73\,mol$$

Example 9.5.7

What volume of $\mathrm{H_2}$ is produced at STP when 55.8 g of Zn metal react with excess HCl?

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Solution

This is a stoichiometry problem with a twist: we need to use the molar volume of a gas at STP to determine the final answer. The first part of the calculation is the same as in a previous example:

$$55.8\,\cancel{g\,Zn} \times \frac{1\,\cancel{mol\,Zn}}{65.41\,\cancel{g\,Zn}} \times \frac{1\,mol\,H_2}{1\,\cancel{mol\,Zn}} = 0.853\,H_2$$

Now we can use the molar volume, 22.4 L/mol, because the gas is at STP:

$$0.853\,\cancel{mol\,H_2} \times \frac{22.4\,L}{1\,\cancel{mol\,H_2}} = 19.1\,L\,H_2$$

Alternatively, we could have applied the molar volume as a third conversion factor in the original stoichiometry calculation.

Ideal Gas Law

So far, the gas laws we have considered have all required that the gas change its conditions; then we can predict a resulting change in one of its properties. Are there any gas laws that relate the physical properties of a gas at any given time? Consider a further extension of the combined gas law to include *n*. By analogy to Avogadro's law, *n* is positioned in the denominator of the fraction, opposite the volume. So,

$$\frac{PV}{nT} = constant$$

Because pressure, volume, temperature, and amount are the only four independent physical properties of a gas, the constant in the above equation is truly a constant. Indeed, because we do not need to specify the identity of a gas to apply the gas laws, this constant is the same for all gases. We define this constant with the symbol *R*, so the previous equation is written as

$$\frac{PV}{nT} = R$$

which is usually rearranged as

$$PV = nRT$$

This equation is called the **ideal gas law**. It relates the four independent properties of a gas at any time. The constant \(*R*\) is called the **ideal gas law constant**. Its value depends on the units used to express pressure and volume.

Table 9.5.2: Values of the Ideal Gas Law Constant lists the numerical values of R.

Numerical Value	Units
0.08205	$\frac{L \cdot atm}{mol \cdot K}$
62.36	$\frac{L \cdot torr}{mol \cdot K} = \frac{L \cdot mmHg}{mol \cdot K}$
8.314	$\frac{J}{mol \cdot K}$

The ideal gas law is used like any other gas law, with attention paid to the unit and expression of the temperature in kelvin. However, *the ideal gas law does not require a change in the conditions of a gas sample*. The ideal gas law implies that if you know any three of the physical properties of a gas, you can calculate the fourth property.

Example 9.5.8

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

The first step is to convert temperature to kelvins:

$$34 + 273 = 307\,\mathrm{K}$$

Now we can substitute the conditions into the ideal gas law:

$$(1.21 atm)(V) = (4.22\,mol)\left(0.08205 \frac{L.atm}{mol.K}\right)(307\,K)$$

The *atm* unit is in the numerator of both sides, so it cancels. On the right side of the equation, the *mol* and *K* units appear in the numerator and the denominator, so they cancel as well. The only unit remaining is *L*, which is the unit of volume that we are looking for. We isolate the volume variable by dividing both sides of the equation by 1.21:

$$V = \frac{(4.22)(0.08205)(307)}{1.21} L$$

Then solving for volume, we get V = 87.9 L

Example 9.5.9

At a given temperature, 0.00332 g of Hg in the gas phase has a pressure of 0.00120 mmHg and a volume of 435 L. What is its temperature?

Solution

We are not given the number of moles of Hg directly, but we are given a mass. We can use the molar mass of Hg to convert to the number of moles.

$$0.00332\,\cancel{g\,Hg} \times \frac{1\,mol\,Hg}{200.59\,\cancel{g\,Hg}} = 0.0000165\,mol = 1.65 \times 10^{-5} mol$$

Pressure is given in units of millimeters of mercury. We can either convert this to atmospheres or use the value of the ideal gas constant that includes the mmHg unit. We will take the second option. Substituting into the ideal gas law,

$$(0.00332\,mm\,Hg)(435\,L) = (1.65 \times 10^{-5} mol)(62.36 \frac{L.mmHg}{mol.K})T$$

The mmHg, L, and mol units cancel, leaving the K unit, the unit of temperature. Isolating T on one side, we get

$$T = \frac{(0.00332)(435)}{(1.65 \times 10^{-5})(62.36)} K$$

Then solving for K, we get T = 1,404 K.

The ideal gas law can also be used in stoichiometry problems.

Example 9.5.10

What volume of $\mathrm{H_2}$ is produced at 299 K and 1.07 atm when 55.8 g of Zn metal react with excess HCl?

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Solution

Here we have a stoichiometry problem where we need to find the number of moles of H2 produced. Then we can use the ideal gas law, with the given temperature and pressure, to determine the volume of gas produced. First, the number of moles of H2 is calculated:

$$55.8\,\cancel{g\,Zn} \times \frac{1\,\cancel{mol\,Zn}}{65.41\,\cancel{g\,Zn}} \times \frac{1\,mol\,H_2}{1\,\cancel{mol\,Zn}} = 0.853\,H_2$$

Now that we know the number of moles of gas, we can use the ideal gas law to determine the volume, given the other conditions:

$$(1.07atm)V = (0.853\,mol)\left(0.08205 \frac{L.atm}{mol.K}\right)(299\,K)$$

All the units cancel except for L, for volume, which means V = 19.6 L

Partial Pressure and Dalton's Law

One of the properties of gases is that they mix with each other. When they do so, they become a solution—a homogeneous mixture. Some of the properties of gas mixtures are easy to determine if we know the composition of the gases in the mix.

In gas mixtures, each component in the gas phase can be treated separately because, according to the kinetic theory of gases, there are little to no attractive or repulsive forces between particles. Each component of a gas mixture shares the same temperature and volume, (remember that gases expand to fill the volume of their container; gases in a mixture continue to do that as well). However, each individual gas will have its *own pressure* within a mixture. This is called the **partial pressure** of a gas, P_i. Partial pressures are expressed in torr, millimeters of mercury, or atmospheres like any other gas pressure; however, we use the term *pressure* when talking about pure gases and the term *partial pressure* when we are talking about the individual gas components in a mixture.

Dalton's law of partial pressures states that the total pressure of a gas mixture, P_{tot}, is equal to the sum of the partial pressures of the components, P_i:

which is expressed algebraically as

$$P_{total} = P_1 + P_2 + P_3... = \sum_i P_i$$

or, equivalently

$$P_{total} = \frac{RT}{V} \sum_i n_i$$

where i counts over all gases in mixture.

Although this law may seem trivial, it reinforces the idea that gases behave independently of each other.

Example 9.5.11

A mixture of H_2 at 2.33 atm and N_2 at 0.77 atm is in a container. What is the total pressure in the container?

Solution

Dalton's law of partial pressures states that the total pressure is equal to the sum of the partial pressures. We simply add the two pressures together:

P_{tot} = 2.33 atm + 0.77 atm = 3.10 atm

Example 9.5.12

A 2.00 L container with 2.50 atm of H_2 is connected to a 5.00 L container with 1.90 atm of O_2 inside. The containers are opened, and the gases mix. What is the final pressure inside the containers?

Solution

Because gases act independently of each other, we can determine the resulting final pressures using Boyle's law and then add the two resulting pressures together to get the final pressure. The total final volume is 2.00 L + 5.00 L = 7.00 L. First, we use Boyle's law to determine the final pressure of H_2:

$$(2.50 \text{ atm})(2.00 \text{ L}) = P_2(7.00 \text{ L})$$

Solving for P_2, we get P_2 = 0.714 atm = partial pressure of H_2.

Now we do that same thing for the O2:

(1.90 atm)(5.00 L) = *P*2(7.00 L)*P*2 = 1.36 atm = partial pressure of O2

The total pressure is the sum of the two resulting partial pressures:

*P*tot = 0.714 atm + 1.36 atm = 2.07 atm

Dalton's law states that in a gas mixture (P_{total}) each gas will exert a pressure independent of the other gases (P_n) and each gas will behave as if it alone occupies the total volume. By extension, the partial pressure of each gas can be calculated by multiplying the total pressure (P_{total}) by the gas percentage (%).

$$P_{Total} = P_1 + P_2 + P_3 + P_4 + ... + P_n$$

or

$$P_n = \frac{^n}{P_{Total}}$$

Table 9.5.3: Partial Pressures for the gases in air on a typical day

Gas	Partial Pressure (mm Hg)	Percentage (%)
Nitrogen, (N_2\)	P_{N_2} = 594	78
Oxygen, O_2	P_{O_2}= 160	21
Carbon Dioxide, CO_2	P_{CO_2} = 0.25	0.033
Water Vapor, H_2O	P_{H_2O} = 5.7	0.75
Other trace gases	P_{Other} = 0.05	0.22
Total air	P_{Total} = 760	1

Example 9.5.13

Oxygen gas was collected over water as shown above. The atmospheric pressure was 754 torr, the temperature was 22°C, and the volume of the gas was 155 mL. The vapor pressure of water at 22°C is 19.8 torr. Use this information to estimate the number of moles of O_2 produced.

Solution

From Dalton's law,

$$P_{O_2} = P_{total} – P_{H_2O} = 754 – 19.8 = 734\ torr = 0.966\ atm$$

Now use the Ideal Gas Law to convert to moles

$$n = \frac{PV}{RT} = \frac{(0.966\ atm)(0.155\ L)}{(0.082\ Latmmol^{-1}K^{-1})(295\ K)} = 0.00619\ mol$$

Link to learning: Chemistry in Everyday Life

Carbonated beverages—sodas, beer, sparkling wines—have one thing in common: they have CO_2 gas dissolved in them in such sufficient quantities that it affects the drinking experience. Most people find the drinking experience pleasant—indeed, in the United States alone, over 1.5 × 10^9 gal of soda are consumed each year, which is almost 50 gal per person! This figure does not include other types of carbonated beverages, so the total consumption is probably significantly higher.

All carbonated beverages are made in one of two ways. First, the flat beverage is subjected to a high pressure of $\mathrm{CO_2}$ gas, which forces the gas into solution. The carbonated beverage is then packaged in a tightly-sealed package (usually a bottle or a can) and sold. When the container is opened, the $\mathrm{CO_2}$ pressure is released, resulting in the well-known hiss of an opening container, and $\mathrm{CO_2}$ bubbles come out of solution. This must be done with care: if the $\mathrm{CO_2}$ comes out too violently, a mess can occur!



Figure 9.5.5: Wide glass of a carbonated beverage with carbon dioxide bubbles rising and escaping from the liquid. (Unsplash License; Tina Vanhove via Unsplash). Detailed Description of Figure 9.5.5.

The second way a beverage can become carbonated is by the ingestion of sugar by yeast, which then generates $\mathrm{CO_2}$ as a digestion product. This process is called fermentation. The overall reaction is

$$\mathrm{C_6H_{12}O_6\,(aq) \longrightarrow 2\,C_2H_5OH\,(aq) + 2\,CO_2\,(aq)}$$

When this process occurs in a closed container, the $\mathrm{CO_2}$ produced dissolves in the liquid, only to be released from solution when the container is opened. Most fine sparkling wines and champagnes are turned into carbonated beverages this way. Less-expensive sparkling wines are made like sodas and beer, with exposure to high pressures of $\mathrm{CO_2}$ gas.

Henry’s Law

Henry's law is one of the gas laws formulated by William Henry in 1803. It states: "At a constant temperature, the amount of a given gas that dissolves in a given type and volume of liquid is directly proportional to the partial pressure of that gas in equilibrium with that liquid." An equivalent way of stating the law is that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.

To explain this law, Henry derived the equation:

$$C = kP_{gas}$$

where

Henry’s Law tells us that the greater the pressure of gas above the surface of a liquid, the higher the concentration of the gas in the liquid. Also, Henry’s law tells us that gases diffuse from areas of high gas concentration to areas of low gas concentration.

Applicability of Henry's Law

- Henry's law only works if the molecules are at equilibrium.

- Henry's law does not work for gases at high pressures (e.g., $N_{2\ (g)}$ at high pressure becomes very soluble and harmful when in the blood supply).
- Henry's law does not work if there is a chemical reaction between the solute and solvent (e.g., $HCl_{(g)}$ reacts with water by a dissociation reaction to generate H_3O^+ and Cl^- ions).

Application of Henry's Law: Scuba diving

Our respiratory systems are designed to maintain the proper oxygen concentration in the blood when the partial pressure of O_2 is 0.21 atm, its normal sea-level value. Below the water surface, the pressure increases by 1 atm for each 10.3 m increase in depth; thus a scuba diver at 10.3 m experiences a total of 2 atm pressure pressing on the body. In order to prevent the lungs from collapsing, the air the diver breathes should also be at about the same pressure.



Figure 9.5.6: Scuba Diver Underwater at Depth. (CC BY-NC-SA 3.0; via LibreTexts). Detailed Description of Figure 9.5.6.

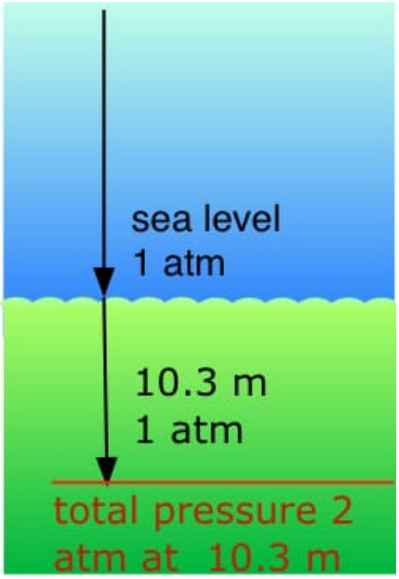


Figure 9.5.7: Pressure Increase with Depth in Water. (CC BY-NC-SA 3.0; via LibreTexts). Detailed Description of Figure 9.5.7.

But at a total pressure of 2 atm, the partial pressure of O_2 in ordinary air would be 0.42 atm; at a depth of 100 ft (about 30 m), the O_2 pressure of 0.8 atm would be far too high for health. For this reason, the air mixture in the pressurized tanks that scuba divers wear must contain a smaller fraction of O_2. This can be achieved most simply by raising the nitrogen content, but high partial pressures of N_2 can also be dangerous, resulting in a condition known as nitrogen narcosis. The preferred diluting agent for sustained deep diving is helium, which has very little tendency to dissolve in the blood even at high pressures.

Chemistry in Everyday Life

Certain diseases—such as emphysema, lung cancer, and severe asthma—primarily affect the lungs. Respiratory therapists help patients with breathing-related problems. They can evaluate, help diagnose, and treat breathing disorders and even help provide emergency assistance in acute illness where breathing is compromised.

Most respiratory therapists must complete at least two years of college and earn an associate's degree, although therapists can assume more responsibility if they have a college degree. Therapists must also pass state or national certification exams. Once certified, respiratory therapists can work in hospitals, doctor's offices, nursing homes, or patient's homes. Therapists work with equipment such as oxygen tanks and respirators, may sometimes dispense medication to aid in breathing, perform tests, and educate patients in breathing exercises and other therapy.

Because respiratory therapists work directly with patients, the ability to work well with others is a must for this career. It is an important job because it deals with one of the most crucial functions of the body.

Key Takeaways

- Boyle's Law: At constant temperature, pressure (P) and volume (V) are inversely related. $P \propto 1/V$ (T constant). $P_1V_1 = P_2V_2$
- Charles's Law: At constant pressure, volume (V) and temperature (T) are directly related. $V \propto T$ (P constant). $T_1V_2 = T_2V_1$
- Gay-Lussac's Law: At constant volume, pressure (P) and temperature (T) are directly related. $P \propto T$ (V constant). $P_1T_2 = P_2T_1$
- Avogadro's Law: At constant temperature and pressure, volume (V) is directly proportional to the number of moles (n). $V \propto n$ (T and P constant). $V_1n_2 = V_2n_1$
- Combined Gas Law: Combines Boyle's, Charles's, and Gay-Lussac's Laws to describe changes in P, V, and T. $P_1V_1 / T_1 = P_2V_2 / T_2$
- Ideal Gas Law: Relates pressure, volume, temperature, and amount of gas in one equation: PV = nRT; Where R is the gas constant.
- Dalton's Law of Partial Pressures: The total pressure of a gas mixture equals the sum of the partial pressures of each individual gas. $P_{total} = P_1 + P_2 + P_3 + ...$
- The solubility of a gas in a liquid increases with its partial pressure, and this relationship is defined by Henry's Law, where the constant k depends on the gas, solvent, and temperature. $C = kP_g$

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9.6: Summary

Matter exists in three main states: solid, liquid, and gas, each with distinct physical properties. Matter changes from one state to another through processes such as melting, freezing, vaporization, condensation, sublimation, and deposition. Intermolecular forces, including dispersion forces, dipole-dipole interactions, and hydrogen bonding, influence physical properties like boiling point and viscosity. Gases consist of particles in constant, random motion as described by the kinetic molecular theory. Gas pressure results from collisions of gas particles with container walls and is measured in units such as Pascals (Pa). Boyle's Law states that pressure and volume of a gas are inversely proportional at constant temperature. Charles's Law states that volume and temperature are directly proportional at constant pressure. Gay-Lussac's Law shows that pressure and temperature are directly proportional at constant volume. Avogadro's Law states that volume is directly proportional to the number of moles of gas at constant temperature and pressure. The Combined Gas Law relates pressure, volume, and temperature when the amount of gas remains constant. Additionally, the Ideal Gas Law, expressed as PV = nRT, combines several gas laws to relate pressure, volume, temperature, and the number of moles of a gas using the gas constant R. Henry's Law states that the solubility of a gas in a liquid is directly proportional to the partial pressure of that gas above the liquid, meaning that as pressure increases, more gas dissolves in the liquid. Finally, Dalton's Law states that the total pressure of a gas mixture equals the sum of the partial pressures of each individual gas.

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9.7: Key Terms

- **Dry Ice:** The solid form of carbon dioxide (CO_2), which sublimates directly into gas without becoming a liquid.
- **Sublimation:** A process where a solid changes directly into a gas without passing through the liquid phase.
- **Carbon Dioxide (CO_2):**A gas found in the atmosphere that becomes dry ice when frozen at around –78°C.
- **Solid:** A state of matter with a fixed shape and volume; in this case, solid CO_2 is dry ice.
- **Gas:** A state of matter without a fixed shape or volume; carbon dioxide gas is released when dry ice sublimates.
- **Liquid**: A state of matter with definite volume but no fixed shape; particles are close but can move past one another.
- **Phase**: is a distinct form in which matter can exist, commonly as a solid, liquid, or gas.
- **Intermolecular Forces**: Attractive forces that occur between molecules and affect physical properties like boiling and melting points.
- **Dipole**: A molecule with a permanent separation of charge due to differences in electronegativity between bonded atoms, resulting in a positive and a negative end.
- **Dispersion Forces**: Weak, temporary forces caused by momentary shifts in electrons; present in all molecules, especially nonpolar ones.
- **Instantaneous Dipole**: A temporary separation of charge in a nonpolar molecule caused by random movement of electrons at a given moment.
- **Dipole**: A molecule with a permanent separation of charge due to differences in electronegativity between bonded atoms, resulting in a positive and a negative end.
- **Dipole-Dipole Interactions**: Attractions between the positive and negative ends of polar molecules.
- **Hydrogen Bonding**: A strong type of dipole-dipole force that occurs when hydrogen is bonded to N, O, or F and interacts with a lone pair on another N, O, or F atom.
- **Manometers**: Instruments used to measure the pressure of gases by comparing it to atmospheric pressure or a vacuum, typically using a liquid column in a U-shaped tube.
- **Kinetic Molecular Theory of Gases**: Explains gas behavior based on particles moving randomly with no attractions and elastic collisions.
- **Ideal Gas**: A gas that follows all gas laws and kinetic theory assumptions perfectly.
- **Boyle's Law**: Pressure is inversely proportional to volume at constant temperature.
- **Charles's Law**: Volume is directly proportional to temperature at constant pressure.
- **Gay-Lussac's Law**: Pressure is directly proportional to temperature at constant volume.
- **Avogadro's Law**: Volume is directly proportional to the number of moles at constant temperature and pressure.
- **Combined Gas Law**: Relates pressure, volume, and temperature when the amount of gas is constant.
- **Standard Temperature and Pressure (STP)**: A set of reference conditions for gases—0°C (273.15 K) temperature and 1 atm pressure—used to compare gas behavior.
- **Molar Volume**: The volume occupied by 1 mole of an ideal gas at STP, equal to 22.4 liters.

- **Dalton's Law of Partial Pressures**: Total pressure of a gas mixture equals the sum of the partial pressures of individual gases.
- **Pressure (P)**: Force per unit area, measured in units like atm, Pa, or mmHg.
- **Volume (V)**: The space occupied by a gas, usually measured in liters (L).
- **Temperature (T)**: A measure of the average kinetic energy of gas particles, in Kelvin (K).
- **Moles (n)**: The amount of gas particles, typically measured in mol.

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9.8: Alternative text for Chapter 9

Figure 9.1.1: Sublimation of Dry Ice (Solid Carbon Dioxide) into Carbon Dioxide Gas

Figure showing stacked blocks of white dry ice (solid carbon dioxide) subliming. The blocks are arranged in a small pile with frosty, opaque surfaces and sharp edges. A visible cloud of white vapor forms around the blocks as the dry ice changes directly from solid to gas. The gas appears to flow downward and pool near the base, illustrating that cold carbon dioxide is denser than air. The background is plain and uncluttered to emphasize the sublimation process.

Figure 9.2.1: Transitions Between Solid, Liquid, and Gaseous States

The image visually demonstrates how the physical state of a substance—solid, liquid, or gas—is determined by the balance between intermolecular forces (IMFs) and kinetic energy (KE). At the center are three stoppered Erlenmeyer flasks arranged horizontally. Each flask contains the same substance in a different phase. The flask on the left holds a crystalline solid, where particles are tightly packed in an orderly, rigid structure, reflecting strong IMFs and low kinetic energy. The middle flask contains a liquid, with particles still close together but more disordered and mobile. This phase represents a balance between moderate IMFs and increased kinetic energy. The flask on the right contains a gas, in which the particles are far apart and moving rapidly in random directions. In this phase, kinetic energy dominates, overcoming the weaker IMFs entirely.

Above the three flasks, a horizontal arrow points to the right and is labeled “increasing kinetic energy with temperature.” This arrow indicates that as temperature rises, so does the kinetic energy of the molecules, making phase transitions from solid to liquid to gas more likely. At the bottom of the image, another horizontal arrow points to the left and is labeled “increasing intermolecular forces.” This arrow shows that stronger IMFs favor the formation of more ordered, condensed phases such as liquids and solids. Together, the arrangement of flasks and directional arrows illustrates the dynamic balance between kinetic energy and intermolecular forces. A substance’s phase is determined by which of these two opposing factors is more influential under given conditions.

Figure 9.2.2: Condensation Forms When Water Vapor in The Air is Cooled

This figure visually demonstrates condensation, the phase change in which water vapor transitions into liquid water due to cooling. In image (a), a cold beverage glass sits in a warm, humid environment. The outer surface of the glass is covered with numerous visible droplets of liquid water. These droplets form when water vapor in the surrounding air comes into contact with the cold glass surface. The drop in temperature reduces the kinetic energy of the vapor molecules, allowing intermolecular forces—particularly hydrogen bonding—to draw the molecules together into the liquid phase. In image (b), condensation is shown in the form of fog. Here, tiny liquid water droplets appear suspended in the air, creating a misty appearance near the ground or over a surface. This occurs when warm, moist air cools rapidly, often due to contact with cooler air or surfaces, causing water vapor to condense without needing a solid surface. In both scenes, the decrease in temperature allows the attractive forces between H_2O molecules to overcome their motion, pulling them into the more condensed, liquid state and illustrating how condensation can occur both on surfaces and in open air.

Figure 9.2.3: Butane Gas Compression and Liquefaction in a Lighter

If the temperature of a liquid becomes sufficiently low or the pressure on it becomes high enough, the molecules lose enough kinetic energy (KE) that they can no longer overcome the intermolecular forces (IMFs) between them, causing the liquid to solidify into a fixed, orderly structure. The image shows a green disposable lighter containing liquid butane inside its fuel compartment. Normally a gas at standard temperature and pressure, butane is compressed inside the lighter, which forces the molecules closer together. This increased pressure strengthens the intermolecular attractions relative to their kinetic energy, causing the butane gas to condense into liquid form. This example indicates how gases like butane can be liquefied by compression if the temperature remains sufficiently low. A more detailed discussion of these and other phase transitions will be provided later in the chapter.

Figure 9.2.4: Intramolecular Forces Keep a Molecule Intact

The image in Figure 9.2.4 shows two hydrogen chloride (HCl) molecules positioned near each other to demonstrate the difference between intramolecular and intermolecular forces. Each molecule contains a small white cloud representing a hydrogen atom (H) and a larger green cloud representing a chlorine atom (Cl). In the left molecule, a solid straight line connects the hydrogen and chlorine atoms, labeled "intramolecular force (strong)," indicating the covalent bond that holds the atoms together within the molecule. Between the two molecules, a brown dashed line connects the hydrogen atom of one molecule to the chlorine atom of the other. This line is labeled "intermolecular force (weak)" and represents the dipole-dipole attraction between the molecules. The image emphasizes that intramolecular covalent bonds are significantly stronger—requiring about 430 kilojoules per mole to break—while the intermolecular forces between molecules are much weaker, requiring only about 17 kilojoules per mole to overcome.

Figure 9.2.5: Dispersion Forces Result From The Formation of Temporary Dipoles

The image illustrates the formation of dispersion forces between two nonpolar diatomic molecules. Both molecules are shown side by side, each consisting of two atoms with electron clouds surrounding them. In this image, only an uneven distribution of electrons is depicted—there is no initial stage showing even distribution. The electron cloud around each molecule is shown as slightly distorted, with more electron density concentrated on one side. This distortion is indicated by a directional line across each molecule, representing a temporary dipole, with one end labeled δ⁻ (partial negative) and the opposite end labeled δ⁺ (partial positive).

The temporary dipole in one molecule induces a similar dipole in the neighboring molecule, which is also labeled with δ⁻ and δ⁺ to reflect the shift in electron density. Between the two molecules, a brown dashed line connects the oppositely charged ends, indicating the intermolecular attraction known as the dispersion force. This visual representation shows how fluctuations in electron distribution can create temporary dipoles that lead to weak, short-lived electrostatic attractions, even in molecules that are overall nonpolar.

Figure 9.2.6: Boiling Points of Group 14 Tetrahydrides Increase Down The Group

A graph of the boiling points of group 14 tetrahydrides—CH_4, SiH_4, GeH_4, and SnH_4—plotted against the period number of the central group 14 element. The vertical axis represents temperature (boiling point), and the horizontal axis represents the period number. Each compound is represented by a blue sphere placed at its corresponding period on the x-axis, and the spheres are connected by a blue line to illustrate the trend.

The graph displays a clear upward trend, with CH_4 at period 2 showing the lowest boiling point and SnH_4 at period 5 showing the highest. This pattern reflects the increasing molar masses of the compounds: CH_4 (16 g/mol), SiH_4 (32 g/mol), GeH_4 (77 g/mol), and SnH_4 (123 g/mol). As the size and molar mass of the central atom increase down the group, so does the polarizability of the molecules.

This increase in polarizability leads to stronger dispersion forces, which in turn raise the boiling points. Since all four compounds are nonpolar and do not experience dipole-dipole interactions or hydrogen bonding, dispersion forces are the dominant intermolecular forces. The graph confirms the predicted order of boiling points—$CH_4 < SiH_4 < GeH_4 < SnH_4$—and demonstrates how dispersion forces become stronger in larger, more polarizable molecules.

Figure 9.2.7: The Strength of The Dispersion Forces Increases With Surface Area

The image presents a comparative visualization of the three isomers of C_5H_{12}—neopentane, isopentane, and n-pentane—arranged from left to right in order of increasing boiling point. Each molecule is shown with its mirror image placed directly above it to emphasize structural symmetry and shape. On the far left is **neopentane**, the most compact and spherical isomer, with its boiling point labeled below as 9.5 °C. Above its mirror image, a label reads "Small contact area → Weakest attraction," indicating that its compact shape limits the extent of intermolecular contact, resulting in the weakest London dispersion forces. In the center

is **isopentane**, with a moderately branched structure and a boiling point of 27 °C. The label above it reads "Less surface area → Less attraction," showing that while it has more surface contact than neopentane, its dispersion forces remain weaker than those of n-pentane. On the far right is **n-pentane**, with a fully extended, linear structure. Its boiling point is listed as 36 °C, and the label above its mirror image states "Large contact area → Stronger attraction," highlighting that its elongated shape allows for maximum surface contact between molecules, resulting in the strongest dispersion forces among the three isomers. The image conveys that although all three compounds share the same molecular formula (C_5H_{12}), differences in molecular shape directly affect their intermolecular interactions and boiling points. This concept is analogous to VELCRO strips, where a larger contact area produces a stronger connection.

Figure 9.2.8: Geckos' Toes Contain Large Numbers of Tiny Hairs (setae)

The image visually explains the remarkable adhesive ability of geckos, progressing from a whole-animal view on the left to microscopic structures on the right. On the left, a full gecko is shown climbing a vertical or upside-down surface, illustrating its ability to stick without the use of suction or glue. In the center, a close-up of the gecko's toe reveals hundreds of thousands of tiny hair-like structures called setae. These setae are exaggerated in size for clarity, emphasizing their role in surface contact. On the **right**, an even closer view shows that each seta branches into hundreds of even tinier, flat, triangular structures called spatulae. These are also enlarged for visualization. Arrows or labels may indicate how the massive number of spatulae increases the gecko's total surface contact, enabling strong adhesion via dispersion forces —weak intermolecular forces that arise from temporary charge distributions.

Figure 9.2.9: Dipole–Dipole Interactions Between HCl Molecules in Different Orientations

Figure illustrating dipole–dipole interactions between two hydrogen chloride (HCl) molecules in two different orientations. On the left, the two HCl molecules are positioned side by side horizontally, each showing a partial positive charge on hydrogen (δ+) and a partial negative charge on chlorine (δ−); a horizontal dotted line connects the δ+ end of one molecule to the δ− end of the other, indicating electrostatic attraction. On the right, the two HCl molecules are positioned vertically next to each other (not stacked), with opposite ends aligned so that the δ+ region of one molecule is adjacent to the δ− region of the other; two vertical dotted lines indicate dipole–dipole attractions between the opposing partial charges.

Figure 9.2.10: Water Molecules Participate in Multiple Hydrogen-Bonding Interactions

Figure 9.2.10 illustrates dipole-dipole interactions using five water (H_2O) molecules as an example. Each molecule is composed of two white spheres labeled “H” for hydrogen and one red sphere labeled “O” for oxygen, reflecting the standard color coding used to represent water. The partial charges are not explicitly shown, but the polarity is understood: the oxygen atom is more electronegative and thus carries a partial negative character, while the hydrogen atoms carry a partial positive character.

In the image, the five water molecules are arranged to demonstrate how dipole-dipole interactions occur between them. The molecules are oriented so that the hydrogen atoms of one water molecule (white H spheres) are positioned near the oxygen atoms (red O spheres) of neighboring molecules. Between these interacting atoms, red dashed lines are drawn to indicate the intermolecular forces—specifically, dipole-dipole interactions, which in the case of water are also classified as hydrogen bonds.

Figure 9.2.11: Boiling Points of Group 15, 16, and 17 Hydrides Across Periods 3–5

The image is a line graph showing the boiling points of binary hydrides from groups 15, 16, and 17 plotted against the period number on the x-axis, spanning Period 2 to Period 5. The y-axis represents boiling points in degrees Celsius. Each group's hydrides are represented by distinct colored lines for easy identification: green for group 15 (NH_3, PH_3, AsH_3, SbH_3), blue for group 16 (H_2O, H_2S, H_2Se, H_2Te), and red for group 17 (HF, HCl, HBr, HI).

Moving from Period 2 to Period 5, the boiling points for the heavier three hydrides in each group rise steadily. This trend reflects increasing molecular size and stronger dispersion forces as the dominant factor, despite a slight decrease in molecular polarity.

For example, the green line representing group 15 shows PH_3, AsH_3, and SbH_3 boiling points increasing, while the blue and red lines for groups 16 and 17 show similar steady upward trends for their heavier hydrides.

However, the Period 2 hydrides—NH_3 (green), H_2O (blue), and HF (red)—are marked with question marks (?) at their data points to indicate that their boiling points deviate significantly from the expected trend. These points lie well above the smooth curves formed by the heavier hydrides, signaling an anomaly. This elevated boiling point is caused by strong hydrogen bonding, which is much stronger than the dispersion and dipole-dipole forces that dominate in the heavier hydrides.

The question marks highlight the unexpected nature of these Period 2 hydrides' boiling points, emphasizing the key role of hydrogen bonding due to the highly electronegative atoms (N, O, and F) bonded to hydrogen. This graph effectively contrasts the smooth increase in boiling points driven by dispersion forces in heavier hydrides with the pronounced hydrogen-bonding effect seen in the lighter, Period 2 compounds.

Figure 9.2.12: In Comparison to Periods 3–5, The Binary Hydrides of Period 2

The image is a line graph that plots the boiling points of binary hydrides from groups 15, 16, and 17 against their corresponding periods on the x-axis, from Period 2 to Period 5. The y-axis represents boiling point in degrees Celsius. Each group of hydrides is represented by a distinct colored curve: green for group 15 (NH_3, PH_3, AsH_3, SbH_3), blue for group 16 (H_2O, H_2S, H_2Se, H_2Te), and red for group 17 (HF, HCl, HBr, HI).

For all three groups, the curves display a general upward trend from Period 3 to Period 5, reflecting the increasing boiling points of heavier hydrides. This smooth rise is attributed to stronger dispersion forces that come with increasing molecular size and polarizability, even as molecular polarity slightly decreases. For example, in the green curve (group 15), the boiling points rise steadily from PH_3 to AsH_3 to SbH_3. Similarly, the blue and red curves for groups 16 and 17 show smooth increases from H_2S to H_2Te and from HCl to HI, respectively.

However, at Period 2, each curve makes a sharp upward bend, forming a noticeable spike for NH_3 (green), H_2O (blue), and HF (red). These Period 2 hydrides have significantly higher boiling points than expected based on the trend from heavier analogs. Rather than continuing the smooth progression seen in the rest of each curve, the lines bend sharply to accommodate these anomalously high values, creating a distinct "kink" or spike in each group's trend line.

These sharp upward bends visually emphasize the exceptional impact of hydrogen bonding in NH_3, H_2O, and HF. Because nitrogen, oxygen, and fluorine are highly electronegative and small in size, they form strong hydrogen bonds with neighboring molecules—intermolecular forces much stronger than the dipole-dipole or dispersion forces that dominate in the heavier hydrides. The graph powerfully illustrates how hydrogen bonding causes a dramatic increase in boiling point, breaking the otherwise smooth trend set by molecular weight and dispersion forces alone.

Figure 9.2.13: Two Separate DNA Molecules Form a Double-Stranded Helix

The image shows a detailed representation of the DNA double helix, emphasizing its structural components and molecular organization. The two strands of DNA spiral around each other, forming the well-known right-handed helical shape. Each strand features a sugar-phosphate backbone, visually represented by a light blue ribbon that curves smoothly along the outer edges of the helix. This ribbon traces the backbone of each strand, composed of alternating sugar (deoxyribose) and phosphate units.

Extending inward from the backbone are the nitrogenous bases, which are color-coded for clarity: adenine is red, thymine is orange, guanine is blue, and cytosine is green. These bases pair specifically—adenine with thymine, and guanine with cytosine—through hydrogen bonds, which are shown as horizontal lines connecting the bases across the two strands. This base-pairing forms the "rungs" of the DNA "ladder."

The image also includes directional labels: one end of each strand is marked 5′ (five-prime) and the other 3′ (three-prime), indicating the antiparallel orientation of the strands. One strand runs from 5′ to 3′, while the opposite strand runs from 3′ to 5′. This directionality is critical for DNA replication and transcription. Together, the light blue ribbon backbone, colored bases, and labeled ends give a clear and informative visualization of DNA's double-stranded, complementary, and antiparallel structure.

Figure 9.2.14: The Geometries of The Base Molecules Result in Maximum Hydrogen Bonding

The image shows a vertical molecular diagram of complementary base pairing in DNA, with clear representation of the structural orientation of the two antiparallel strands. At the top, the base pair between thymine (T) and adenine (A) is illustrated. Thymine, a single-ring pyrimidine, is on the left, and adenine, a double-ring purine, is on the right. These bases are connected by two dashed lines, representing the two hydrogen bonds that stabilize the A–T pairing.

At the bottom, the base pair between guanine (G) and cytosine (C) is shown, with guanine on the left and cytosine on the right. These are held together by three dashed hydrogen bonds, indicating a stronger interaction than the A–T pair. All hydrogen bonds are shown in the center of the image, highlighting the specific and directional pairing between purines and pyrimidines.

Flanking the base pairs on both sides are vertical straight-line structures representing the sugar-phosphate backbone of the DNA strands. These backbones are not depicted as ribbons but as simple linear bonds, emphasizing the chemical framework rather than the spatial twist of the helix. The left strand is labeled 3′ to 5′, with the direction running bottom to top, while the right strand is labeled 5′ to 3′, also oriented bottom to top. This reflects the antiparallel arrangement of DNA strands, crucial for DNA replication and base pairing.

Figure 9.3.1: The Kinetic Theory of Gases

The image shows a transparent cubic box representing a closed container filled with gas particles. Inside the box are many small spheres, each symbolizing an individual gas particle. These spheres are spaced widely apart, highlighting that gas particles occupy a very small fraction of the container's volume and are separated by large distances.

Each particle is shown with arrows pointing in different directions, indicating their constant, random motion within the container. The arrows vary in length, symbolizing the different speeds of the particles as they move independently of one another. Several particles are depicted colliding with the container walls or with each other, demonstrating the concept of elastic collisions where kinetic energy is conserved and there is no net energy loss.

The image intentionally shows no clustering or attraction between particles, reflecting the principle that there are no intermolecular forces (neither attraction nor repulsion) acting between gas particles. The container itself is static, with no changes in size or shape, indicating a rigid boundary confining the gas.

Overall, the image visualizes the fundamental postulates of the Kinetic Molecular Theory: gases consist of tiny particles in continuous, random motion; these particles are separated by large distances relative to their size; they collide elastically with each other and the container walls; and they do not exert attractive or repulsive forces on each other. This model explains the physical behavior of gases under many conditions.

Figure 9.4.1: Atmospheric Pressure and How Atmospheric Mass is Distributed With Altitude

This image visually represents the concept of atmospheric pressure and how atmospheric mass is distributed with altitude. On the left side of the image is a tall vertical column of air extending from sea level up to 40 kilometers, representing a 1-square-meter column of the atmosphere. This column symbolizes the entire weight of the air pressing down on a unit area at Earth's surface. At the base of the column is a label indicating a total mass of 10,000 kilograms, which corresponds to the weight of the air above one square meter and is responsible for exerting a pressure of 1 atmosphere (1 atm), equivalent to 101 kilopascals, 760 mmHg, or 1.01×10^5 newtons per square meter. A bold downward arrow labeled "1 atm" points from the top of the column to the surface, emphasizing that atmospheric pressure results from the weight of all the air above.

To the right of the column, a person is shown standing at sea level, illustrating that we live at the bottom of this "ocean of air." Near the person are silhouettes of two elephants, used as a visual comparison to help convey the magnitude of atmospheric pressure—the weight of the air column is equivalent to having two elephants standing on each square meter of Earth's surface. Above the person is a graph displaying the relationship between cumulative atmospheric mass and altitude. The vertical axis of the graph shows altitude in kilometers, ranging from 0 to 40 km, while the horizontal axis shows atmospheric mass as a percentage, increasing from 0% to 100%.

A yellow curve runs diagonally across the graph, showing an exponential relationship where atmospheric mass increases rapidly as altitude decreases. The curve begins at 0% atmospheric mass at 40 km and rises smoothly to 100% at sea level. Key data points are labeled along the curve: 50% of the atmosphere's mass lies below 5.5 km, 90% lies below 16.5 km, and 99% lies below 30 km. These values highlight how the vast majority of atmospheric mass is concentrated in the lowest part of the atmosphere. The graph also labels the troposphere, which extends from 0 to 13 km, and the stratosphere, which ranges from 13 to 44 km, indicating how the atmosphere is layered. While the stratosphere continues beyond the top of the graph, the figure focuses on the first 40 km, where nearly all atmospheric mass is located.

Together, the visual components of the figure—the air column, the person and elephants, the pressure arrow, and the graph—illustrate that atmospheric pressure is the direct result of the mass of air above, and that most of this mass is located within the first few kilometers of the atmosphere. The steadily rising yellow curve visually reinforces that the accumulation of atmospheric mass follows an exponential trend with decreasing altitude, reaching nearly 100% at around 40 km above Earth's surface.

Figure 9.4.2: A Mercury Barometer

A diagram of a mercury barometer shows a long glass tube closed at the top, filled with mercury, and inverted into a dish of mercury. Some mercury stays in the tube, forming a tall column, while a vacuum exists above it. Arrows indicate atmospheric pressure pushing down on the mercury in the dish and gravity pulling the mercury column downward, illustrating how the column height balances atmospheric pressure.

Figure 9.4.3: Closed-End and Open-End Manometers: Measuring Gas Pressure

Figure showing two types of mercury manometers used to measure gas pressure. In part (a), a closed-end manometer is shown with a U-shaped tube partially filled with mercury; one arm is connected to a gas bulb and the other arm is sealed, creating a vacuum above the mercury in the reference arm ($P \approx 0$). The difference in mercury heights between the two arms directly represents the pressure of the gas in the bulb. In part (b), an open-end manometer is shown with one arm connected to a gas bulb and the other arm open to the atmosphere ($P \approx 1$ atm). The mercury levels differ depending on the relative pressures: if the gas pressure is greater than atmospheric pressure, mercury is pushed down in the gas-connected arm and rises in the open arm; if the gas pressure is lower than atmospheric pressure, mercury is higher in the gas-connected arm. The height difference between the columns indicates the pressure difference between the gas and atmospheric pressure.

Figure 9.5.1: Demonstration of Boyle's Law

Figure illustrates Boyle's law with three parts. On the left, a piston-and-cylinder contains very small spots of gas at large volume and low pressure. In the middle, the piston is compressed, reducing the volume by half and increasing the pressure of the gas. On the far right, a graph plots volume versus pressure, with volume on the y-axis and pressure on the x-axis, showing the inverse relationship between pressure and volume for the gas at constant temperature.

Figure 9.5.2: Demonstration of Charles's Law

Figure illustrates Charles's law. Two piston-and-cylinder setups are shown, each with the same mass weight on top of the piston to maintain constant pressure, a temperature gauge attached to the cylinder, and a burner beneath to heat the gas. On the left, the gas has a lower temperature and smaller volume, and on the right, after heating, the gas expands to a larger volume, demonstrating that the volume of a gas increases with temperature at constant pressure.

Figure 9.5.3: Gas Volume versus Absolute Temperature Relationship

Image shows a graph of gas volume versus absolute temperature using data from Table 13.5.1. Volume (in mL) is on the y-axis, extending from 0 to 400 mL, and temperature (in K) is on the x-axis, extending from 0 to 1000 K. The straight-line graph demonstrates a direct relationship between volume and temperature, illustrating Charles's law, with the constant K remaining at 0.4.

Figure 9.5.4: Demonstrates Avogadro's Hypothesis

Figure demonstrates Avogadro's hypothesis. Four cubic containers are shown from left to right holding helium (He), nitrogen (N_2), ammonia (NH_3), and methane (CH_4). Each container contains one spot representing a gas particle. The gases are at the same volume, pressure, and temperature (0 K). Each gas contains the same number of particles (1 mole). At the bottom of each container, the volume (V), pressure (P), temperature (T), number of moles, and mass of the gas are indicated, showing that while all other properties are equal, the mass differs due to differing molar masses.

Figure 9.5.5: Carbonated Beverage: Dissolved CO_2 Bubbles

The figure shows a glass filled with a carbonated beverage. Small CO_2 bubbles can be seen rising through the liquid, illustrating the gas dissolved in the beverage.

Figure 9.5.6: Scuba Diver Underwater at Depth

A scuba diver is shown underwater positioned below the sea surface. The diver represents being at depth in the ocean, emphasizing increased surrounding water pressure compared to the surface.

Figure 9.5.7: Combined Atmospheric and Hydrostatic Pressure with Depth in Water

Reference diagram associated with the scuba diver figure (the above one), displayed below it due to horizontal layout constraints. The diagram uses color coding to distinguish regions above and below sea level. Above the surface is labeled atmospheric pressure (1 atm). Below the surface, a vertical column extends downward to a depth of 10.3 meters, labeled as contributing an additional 1 atm of hydrostatic pressure from the water column. A boundary line marks sea level between the two regions. A horizontal line at the bottom combines the two pressure contributions, showing a total pressure of 2 atm at 10.3 meters depth (1 atm atmospheric pressure plus 1 atm water pressure).

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10: SOLUTIONS

Introduction

Welcome to a journey into one of chemistry's most essential and practical realms: solutions. The focus is on the critical role solutions play not only in scientific study but in everyday life and human health. Beginning with the fundamental concepts of solutes and solvents, especially highlighting water's unique role, it unpacks the mechanisms of dissolution and the significance of colligative properties and aqueous reactions. Readers will also explore methods for preparing, analyzing, and quantifying solutions through stoichiometry. By delving into these topics, this book emphasizes why a solid understanding of solutions is indispensable for grasping the broader landscape of chemical processes.

Learning Objectives

- Understand what causes solutions to form.
- Express the amount of solute in a solution in various concentration units.
- Use molarity to determine quantities in chemical reactions.
- Determine the resulting concentration of a diluted solution.
- Describe the dissolution process at the molecular level.
- Describe how the properties of solutions differ from those of pure solvents.

10.1: Prelude to Solutions

10.2: Solutions

10.3: The Dissolution Process

10.4: Solubility

10.5: Quantifying Solutions- Concentration Units

10.6: Properties of Solutions

10.7: Summary

10.8: Key Terms

10.9: Alternative text for Chapter 10

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10.1: Prelude to Solutions

Solutions All Around Us: Why They Matter?

Solutions are all around us. Air, for example, is a solution. If you live near a lake, a river, or an ocean, that body of water is not pure H_2O but most probably a solution. Much of what we drink—for example, soda, coffee, tea, and milk—is at least in part a solution. Solutions are a large part of everyday life.

If you watch any of the medical dramas on television, you may have heard a doctor (actually an actor) call for an intravenous solution of "Ringer's lactate" (or "lactated Ringer's"). So what is Ringer's lactate?

Intravenous (IV) solutions are administered for two main reasons:

1. To introduce necessary substances into the bloodstream, such as ions for proper body function, sugar and other food substances for energy, or drugs to treat a medical condition, and
2. To increase the volume of the bloodstream. Many people with acute or long-term medical conditions have received some type of IV solution.

One basic IV solution, called a normal saline solution, is simply a dilute solution of sodium chloride (NaCl) dissolved in water. Normal saline is 9.0 g of sodium chloride (NaCl) dissolved in each liter of solution.

Ringer's lactate is a normal saline solution that also contains small amounts of potassium and calcium ions. In addition, it contains about 2.5 g of lactate ions ($C_3H_5O_3^-$) per liter of solution. The liver metabolizes lactate ions into bicarbonate (HCO_3^-) ions, which help maintain the acid-base balance of blood. Many medical problems, such as heart attacks and shock, affect the acid-base balance of blood, and the presence of lactate in the IV solution eases problems caused by this imbalance.

Physicians can select from a range of premade IV solutions, in accordance with a patient's particular needs. Ringer's lactate is commonly used when a patient's blood volume needs to be increased rapidly. Another frequently used IV solution, called D5W, is a 5% solution of dextrose (a form of sugar) in water (Figure 10.1.1).



Figure 10.1.1: Intravenous (IV) solutions like Ringer's lactate help restore fluids, electrolytes, and maintain acid-base balance in patients. (Public Domain; Numbersinstitute via Wikimedia Commons). Detailed Description of Figure 10.1.1.

A significant portion of the chemistry occurring around us takes place in solution. Much of the chemistry that happens in our own bodies takes place in solution, and many solutions—such as the Ringer's lactate IV solution—are important for our health. In our understanding of chemistry, we need to have a basic understanding of solutions. In this chapter, you will learn about the special characteristics of solutions, how solutions are characterized, and some of their properties.

Key Takeaways

- Solutions are common in everyday life, from air to beverages.
- Understanding solutions is essential in chemistry and medicine.

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10.2: Solutions

Mixtures and Solutions

Matter can be classified based on its composition into pure substances and mixtures.

Pure substances include elements, covalent compounds, and ionic compounds. These substances have a fixed composition and consistent properties. However, most of the matter we encounter in daily life is not pure—it is a mixture of two or more pure substances.

Heterogeneous mixtures have a non-uniform composition, meaning the individual components can often be seen or separated physically.

Homogeneous mixtures, also known as solutions, have a uniform composition throughout, with the different substances evenly distributed at the molecular level (Figure 10.2.1).

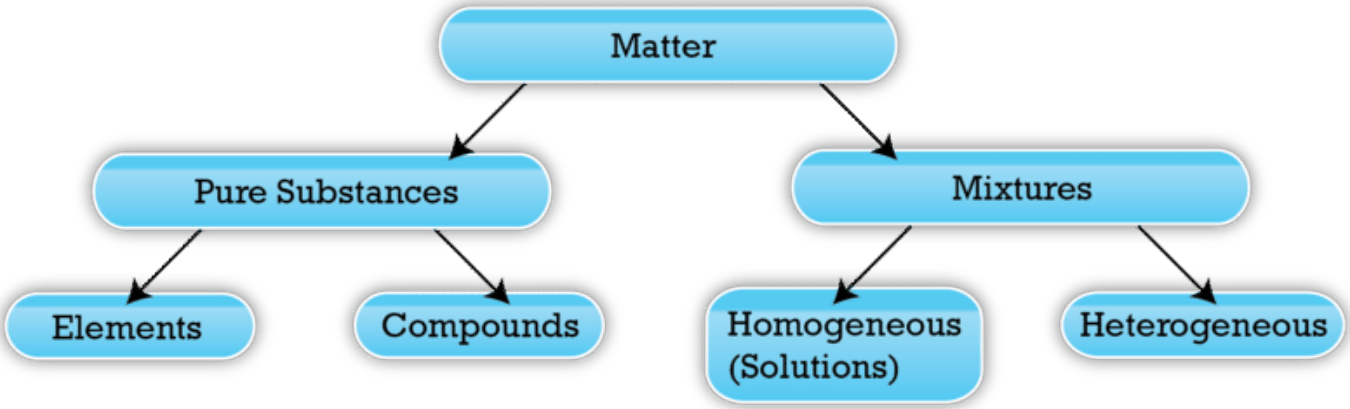


Figure 10.2.1: This diagram shows how matter is categorized into pure substances—elements and compounds—and mixtures, which can be either homogeneous (solutions) or heterogeneous. (CC BY-NC 3.0; Christopher AuYeung via CK-12). Detailed Description of Figure 10.2.1.

Colloids and Suspensions

Suspensions and colloids are two common types of mixtures whose properties are in many ways intermediate between those of true solutions and heterogeneous mixtures. A suspension is a heterogeneous mixture of particles with diameters of about 1 µm (1000 nm) that are distributed throughout a second phase. Common suspensions include paint, blood, and hot chocolate, which are solid particles in a liquid, and aerosol sprays, which are liquid particles in a gas. If the suspension is allowed to stand, the two phases will separate, which is why paints must be thoroughly stirred or shaken before use. A colloid is also a heterogeneous mixture, but the particles of a colloid are typically smaller than those of a suspension, generally in the range of 2 to about 500 nm in diameter. Colloids include fog and clouds (liquid particles in a gas), milk (solid particles in a liquid), and butter (solid particles in a solid). Other colloids are used industrially as catalysts. Unlike in a suspension, the particles in a colloid do not separate into two phases on standing. The only combination of substances that cannot produce a suspension or a colloid is a mixture of two gases because their particles are so small that they always form true solutions. The properties of suspensions, colloids, and solutions are summarized in Table 10.2.1:

Table 10.2.1: Properties of Liquid Solutions, Colloids, and Suspensions

Type of Mixture	Approximate Size of Particles (nm)	Characteristic Properties	Examples
solution	< 2	not filterable; does not separate on standing; does not scatter visible light	air, white wine, gasoline, salt water
colloid	2–500	scatters visible light; translucent or opaque; not filterable; does not separate on standing	smoke, fog, ink, milk, butter, cheese
suspension	500–1000	cloudy or opaque; filterable; separates on standing	muddy water, hot cocoa, blood, paint

Understanding Solutions: Solutes, Solvents, and Solutions

A **solution** is another name for a homogeneous mixture. A *mixture* as a material composed of two or more substances. In a solution, the combination is so intimate that the different substances cannot be differentiated by sight, even with a microscope. Compare, for example, a mixture of salt and pepper and another mixture consisting of salt and water. In the first mixture, we can readily see individual grains of salt and the flecks of pepper. A mixture of salt and pepper is not a solution. However, in the second mixture, no matter how carefully we look, we cannot see two different substances. Salt dissolved in water is a solution.

The major component of a solution, called the **solvent**, is typically the same phase as the solution itself. Each minor component of a solution (and there may be more than one) is called the **solute**. In most of the solutions we will describe in this textbook, there will be no ambiguity about whether a component is the solvent or the solute. For example, in a solution of salt in water, the solute is salt, and the solvent is water.

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts (Figure 10.2.2).

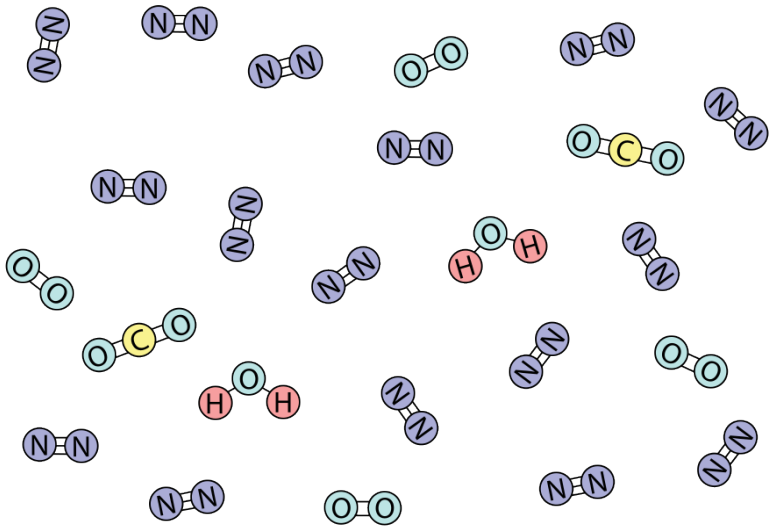


Figure 10.2.2: Air is a solution primarily composed of nitrogen (N_2), oxygen (O_2), carbon dioxide (CO_2), water vapor (H_2O), and trace gases. (Public Domain; RoRo via Wikimedia Commons). Detailed Description of Figure 10.2.2.

An **alloy** is a type of solid solution where a primary metal, such as iron, contains one or more additional elements—either metals or nonmetals—dispersed throughout its structure at the atomic level. These added elements can improve properties like strength, corrosion resistance, or flexibility. Steel is a well-known example of an alloy, consisting mainly of iron combined with carbon and small amounts of other elements, which together enhance its mechanical and chemical properties. Table 10.2.2 lists some common types of solutions, with examples of each.

Table 10.2.2: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	saltwater
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

What causes a solution to form? The simple answer is that the solvent and the solute must have similar intermolecular interactions. When this is the case, the individual particles of solvent and solute can easily mix so intimately that each particle of solute is surrounded by particles of solvent, forming a solution. However, if two substances have very different intermolecular interactions, large amounts of energy are required to force their individual particles to mix intimately, so a solution does not form. Thus, two alkanes like *n*-heptane, C_7H_{16}, and *n*-hexane, C_6H_{14}, are completely miscible in all proportions. The C_7H_{16} and C_6H_{14} molecules are so similar that there are only negligible differences in intermolecular forces.

For a similar reason, methanol, CH_3OH, is completely miscible with water. In this case, both molecules are polar and can form hydrogen bonds among themselves, and so there are strong intermolecular attractions within each liquid. However, CH_3OH dipoles can align with H_2O dipoles, and CH_3OH molecules can hydrogen bond to H_2O molecules, and so the attractions among unlike molecules in the solution are similar to those among like molecules in each pure liquid.

This process leads to a simple rule of thumb: *like dissolves like*. Solvents that are very polar will dissolve solutes that are very polar or even ionic. Solvents that are nonpolar will dissolve nonpolar solutes. Thus water, being polar, is a good solvent for ionic compounds and polar solutes like ethanol (C_2H_5OH).

Key Takeaways

- Matter is classified as pure substances (elements and compounds) or mixtures, which can be homogeneous (solutions) or heterogeneous.
- Solutions, colloids, and suspensions differ in particle size and behavior; solutions are uniform, while colloids and suspensions scatter light or separate over time.
- Solutes and solvents form solutions when their intermolecular forces are similar—summarized by the rule "like dissolves like."
- Solutions exist in all phases (gas, liquid, solid), and examples include air (gas solution), seawater (liquid solution), and alloys (solid solution).
- Solubility depends on molecular interactions; polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes.

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10.3: The Dissolution Process

How Substances Dissolve: The Molecular View

What occurs at the molecular level to cause a solute to dissolve in a solvent? The answer depends in part on the solute, but there are some similarities common to all solutes.

Recall the rule that *like dissolves like*. This means that substances must have similar intermolecular forces to form solutions. When a soluble solute is introduced into a solvent, the particles of solute can interact with the particles of solvent. In the case of a solid or liquid solute, the interactions between the solute particles and the solvent particles are so strong that the individual solute particles separate from each other and, surrounded by solvent molecules, enter the solution. (Gaseous solutes already have their constituent particles separated, but the concept of being surrounded by solvent particles still applies.) This process is called **solvation** and is illustrated in Figure 10.3.1. When the solvent is water, the word **hydration**, rather than solvation, is used.

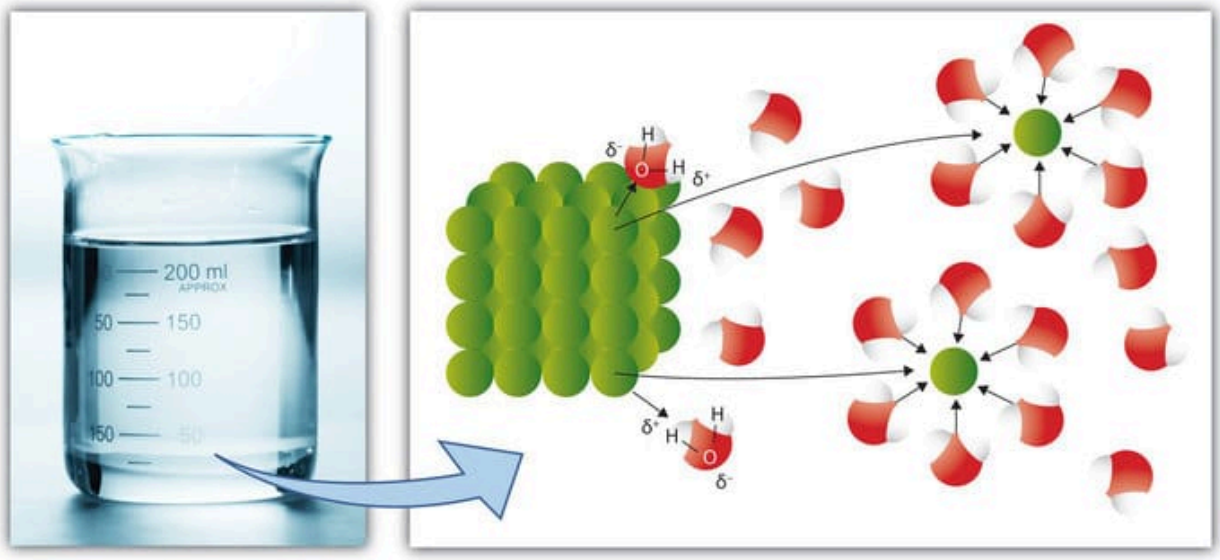


Figure 10.3.1: When a solute dissolves, the individual particles of solute become surrounded by solvent particles. Eventually, the particle detaches from the remaining solute, surrounded by solvent molecules in solution. (CC BY 3.0; Andy Schmitz via Wikimedia Commons). Detailed Description of Figure 10.3.1.

Ionic Compounds and Covalent Compounds as Solutes

In the case of molecular solutes like glucose, the solute particles are individual molecules. However, if the solute is ionic, the individual ions separate from each other and become surrounded by solvent particles. That is, the cations and anions of an ionic solute separate when the solute dissolves. This process is referred to as **dissociation** (Figure 10.3.1).

The dissociation of soluble ionic compounds gives solutions of these compounds an interesting property: they conduct electricity. Because of this property, soluble ionic compounds are referred to as **electrolytes**. Many ionic compounds dissociate completely and are therefore called **strong electrolytes**. Sodium chloride is an example of a strong electrolyte. Some compounds dissolve but dissociate only partially, and solutions of such solutes may conduct electricity only weakly. These solutes are called **weak electrolytes**. Acetic acid (CH_3COOH), the compound in vinegar, is a weak electrolyte. Solutes that dissolve into individual neutral molecules without dissociation do not impart additional electrical conductivity to their solutions and are called nonelectrolytes. Table sugar ($C_{12}H_{22}O_{11}$) is an example of a **nonelectrolyte**.

The term electrolyte is used in medicine to mean any of the important ions that are dissolved in aqueous solution in the body. Important physiological electrolytes include Na^+, K^+, Ca^{2+}, Mg^{2+}, and Cl^-.

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 allows for the 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 10.3.2).

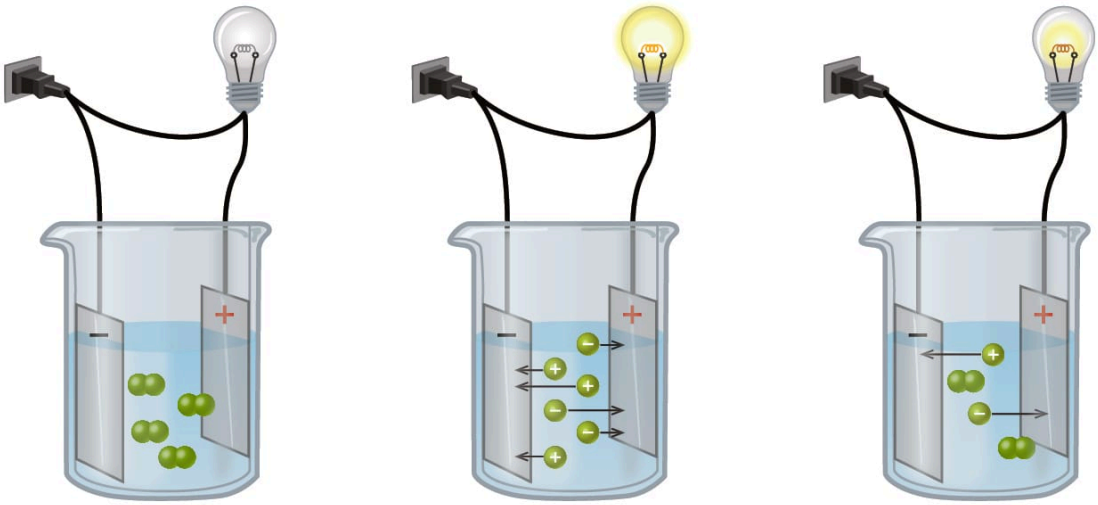


Figure 10.3.2: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.. (CC BY 4.0; contributing authors via OpenStax). Detailed Description of Figure 10.3.2.

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.

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^-):

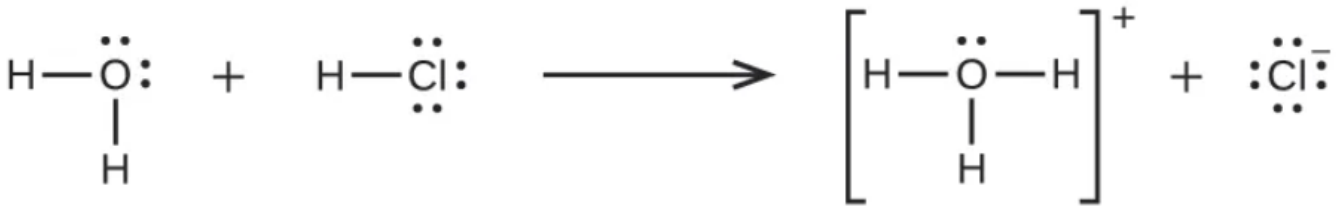


Figure 10.3.3: Reaction of HCl with Water. When hydrogen chloride (HCl) dissolves in water, it donates a proton (H⃝) to a water molecule, forming a hydronium ion (H_3O⃝) and a chloride ion (Cl⃝). This ionization makes HCl a strong electrolyte, as it produces free-moving ions that conduct electricity in solution. (CC BY 4.0; OpenStax via source content).

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.

Key Takeaways

- Solvation occurs when solvent molecules surround and separate solute particles. "Like dissolves like" — similar intermolecular forces promote solubility.
- When water is the solvent, the process is called hydration.
- Ionic compounds dissociate into ions in water and conduct electricity; they are called electrolytes. Strong electrolytes dissociate completely, while weak electrolytes only partially.
- Conductivity tests (e.g., using a light bulb) help classify substances as strong, weak, or nonelectrolytes. Covalent compounds that do not form ions are nonelectrolytes and do not conduct electricity.
- Acids like HCl act as strong electrolytes by donating protons and forming ions in solution.

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10.4: Solubility

Overview

We use the word **soluble** to describe a solute that dissolves in a particular solvent, and the word **insoluble** for a solute that does not dissolve in a solvent. Thus, we say that sodium chloride is soluble in water but insoluble in hexane (C_6H_{14}). If the solute and the solvent are both liquids and soluble in any proportion, we use the word **miscible**, and the word **immiscible** if they are not.

There is usually a limit to how much solute will dissolve in a given amount of solvent. This limit is called the **solubility** of the solute. Some solutes have a very small solubility, while other solutes are soluble in all proportions. Table 10.4.1 lists the solubilities of various solutes in water. Solubilities vary with temperature, so Table 10.4.1 includes the temperature at which the solubility was determined.

Table 10.4.1: Solubilities of Various Solutes in Water at 25°C (Except as Noted).

Substance	Solubility (g in 100 mL of H_2O)
AgCl(s)	0.019
$C_6H_6(\ell)$ (benzene)	0.178
$CH_4(g)$	0.0023
$CO_2(g)$	0.150
$CaCO_3(s)$	0.058
$CaF_2(s)$	0.0016
$Ca(NO_3)_2(s)$	143.9
$C_6H_{12}O_6$ (glucose)	120.3 (at 30°C)
KBr(s)	67.8
$MgCO_3(s)$	2.20
NaCl(s)	36.0
$NaHCO_3(s)$	8.41
$C_{12}H_{22}O_{11}$ (sucrose)	204.0 (at 20°C)

If a solution contains so much solute that its solubility limit is reached, the solution is said to be **saturated**, and its concentration is known from information contained in Table 10.4.1. If a solution contains less solute than the solubility limit, it is **unsaturated**. Under special circumstances, more solute can be dissolved even after the normal solubility limit is reached; such solutions are called **supersaturated** and are not stable. If the solute is solid, excess solute can easily recrystallize. If the solute is a gas, it can bubble out of solution uncontrollably, like what happens when you shake a soda can and then immediately open it (Figure 10.4.1).



Figure 10.4.1: White crystals forming at the bottom of a beaker during crystallization from a supersaturated solution. Crystals begin to form when a supersaturated solution becomes unstable, leading to the solute precipitating out. This image shows the crystallization of ammonium nitrate from a solution that can no longer hold the excess solute in dissolved form. (CC BY-SA 4.0; Alit Variatos via Wikimedia Commons).

The Effects of Temperature on Solubility

Experimentally, it is found that the solubility of most compounds depends strongly on temperature. For most ionic and molecular solids, solubility generally increases as temperature increases. The dependence of solubility on temperature for a number of solids in water is shown by the solubility curves in Figure 10.4.2. Reviewing these data indicates a general trend of increasing solubility with temperature, although there are exceptions, as illustrated by the ionic compound cerium sulfate.

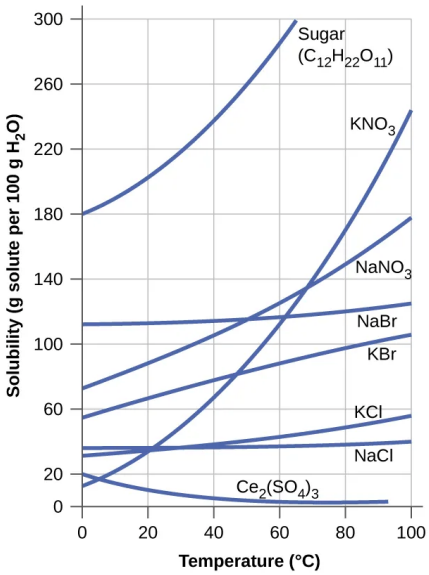


Figure 10.4.2: This graph illustrates how the solubility of different solid solutes changes with temperature. Most solids, such as sugar and potassium nitrate (KNO_3), become more soluble as temperature increases. However, cerium(III) sulfate [$Ce_2(SO_4)_3$] is an exception, with decreasing solubility as the temperature rises. (CC BY 4.0; contributing authors via OpenStax). Detailed Description of Figure 10.4.2.

The temperature dependence of solubility can be exploited to prepare supersaturated solutions of certain compounds. A solution may be saturated with the compound at an elevated temperature (where the solute is more soluble) and subsequently cooled to a lower temperature without precipitating the solute. The resultant solution contains solute at a concentration greater than its equilibrium solubility at the lower temperature (i.e., it is supersaturated) and is relatively stable. Precipitation of the excess solute

can be initiated by adding a seed crystal or by mechanically agitating the solution. Some hand warmers, such as the one pictured in Figure 10.4.3, take advantage of this behavior.



Figure 10.4.3: This hand warmer demonstrates the exothermic crystallization of a supersaturated sodium acetate solution. When the internal metal disc is clicked, it triggers rapid crystal formation, releasing heat and turning the solution from clear to opaque. This process illustrates how supersaturated solutions can release energy upon returning to a stable solid state. (CC BY 4.0; contributing authors via OpenStax). Detailed Description of Figure 10.4.3.

The solubility of gases in liquids decreases with increasing temperature, as shown in Figure 10.4.4. When a gas dissolves in a liquid, weak attractions form between the gas molecules and the solvent, releasing some heat. Heating the solution adds energy to the system, which weakens these attractions. As temperature rises, the kinetic energy of the gas particles increases, causing more of them to escape into the gas phase. As a result, fewer gas molecules stay dissolved in the liquid at higher temperatures.

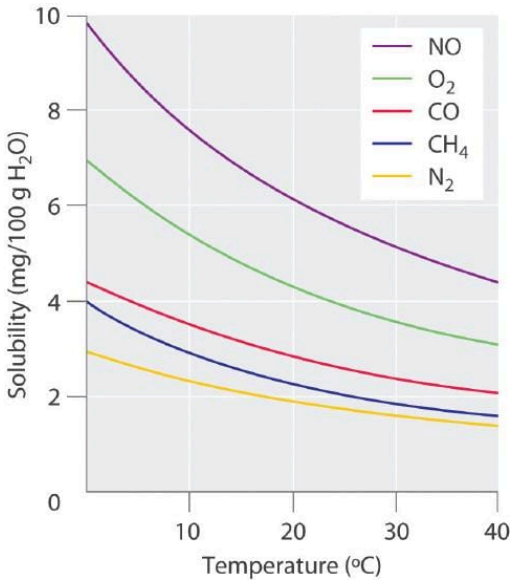


Figure 10.4.4: This graph shows that the solubility of gases in water decreases with increasing temperature. As temperature rises, gas molecules gain kinetic energy and are more likely to escape into the gas phase, reducing their concentration in solution. (CC BY-NC-SA 3.0; anonymous via General Chemistry: Principles, Patterns, and Applications (v. 1.0)). Detailed Description of Figure 10.4.4.

As the temperature of rivers, lakes, or streams increases, the solubility of oxygen in the water decreases. This reduction in dissolved oxygen can have serious consequences for aquatic ecosystems. Warmer water holds less oxygen, and when oxygen levels drop too low, marine and freshwater organisms may not receive the amount they need to survive. In extreme cases, this can lead to widespread die-offs of fish and other aquatic life, severely disrupting the balance of the ecosystem (Figure 10.4.5).



Figure 10.4.5: When oxygen levels in water drop too low, often due to rising temperatures or pollution, aquatic life cannot survive. This image shows a fish kill, a dramatic consequence of oxygen depletion in a body of water. (CC BY 4.0; OpenStax via U.S. Fish and Wildlife Service). Detailed Description of Figure 10.4.5.

The Effects of Pressure on Solubility

External pressure has very little effect on the solubility of liquids and solids. In contrast, the solubility of gases increases as the partial pressure of the gas above a solution increases. This point is illustrated in Figure 10.4.6, which shows the effect of increased pressure on the dynamic equilibrium that is established between the dissolved gas molecules in solution and the molecules in the gas phase above the solution. Because the concentration of molecules in the gas phase increases with increasing pressure, the concentration of dissolved gas molecules in the solution at equilibrium is also higher at higher pressures.

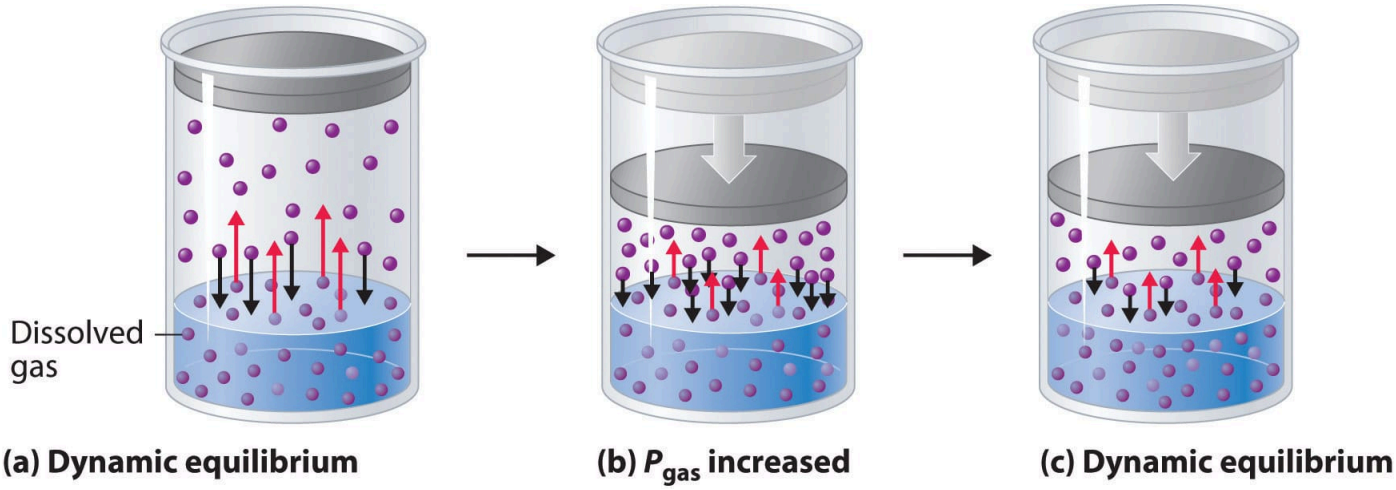


Figure 10.4.6: Diagram showing how increasing gas pressure shifts the equilibrium so that more gas dissolves in the liquid before a new dynamic equilibrium is reached. (CC BY-NC-SA 3.0; anonymous via General Chemistry: Principles, Patterns, and Applications (v. 1.0)). Detailed Description of Figure 10.4.6.

The relationship between pressure and the solubility of a gas is described quantitatively by Henry's law, which is named for its discoverer, the English physician and chemist, William Henry (1775–1836):

$$\mathrm{C} = \mathrm{k} \times \mathrm{P}_{\mathrm{gas}} \tag{10.4.1}$$

where

- C is the concentration of dissolved gas at equilibrium,
- P is the partial pressure of the gas, and
- k is the Henry's law constant, which must be determined experimentally for each combination of gas, solvent, and temperature.

When the partial pressure of a gas above a solution changes while the temperature remains constant, the resulting change in gas solubility can be determined using a simple relationship. Since the ratio C/P remains constant at a fixed temperature, Henry's law can be expressed in a form that clearly shows how a change in one variable affects the other, as shown in the equation:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} \tag{10.4.2}$$

Thus, the higher the pressure, the higher the solubility of a gas in a solvent. Henry's law has important applications. For example, bubbles of CO_2 form as soon as a carbonated beverage is opened because the drink was bottled under CO_2 at a pressure greater than 1 atm. When the bottle is opened, the pressure of CO_2 above the solution drops rapidly, and some of the dissolved gas escapes from the solution as bubbles. Henry's law also explains why scuba divers have to be careful to ascend to the surface slowly after a dive if they are breathing compressed air. At the higher pressures under water, more N_2 from the air dissolves in the diver's internal fluids. If the diver ascends too quickly, the rapid pressure change causes small bubbles of N_2 to form throughout the body, a condition known as "the bends." These bubbles can block the flow of blood through the small blood vessels, causing great pain and even proving fatal in some cases.

Key Takeaways

- Solubility is the maximum amount of solute that can dissolve in a solvent at a given temperature.
- Solutions can be saturated, unsaturated, or supersaturated depending on how much solute is present. Supersaturated solutions are unstable and can crystallize easily.
- Solid solubility usually increases with temperature, though this depends on the specific substance. Gas solubility decreases with increasing temperature and increases with pressure.
- Henry's law explains how gas solubility relates to pressure at constant temperature.

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- 9.2: Concentration by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

10.5: Quantifying Solutions- Concentration Units

Concentration

To define a solution accurately, it is necessary to specify its **concentration**, which refers to the amount of solute dissolved in a given quantity of solvent. Terms such as *dilute* and *concentrated* are commonly used to describe solutions with relatively low or high amounts of solute, respectively; however, these are relative descriptors and their interpretation can vary depending on the context.

Percent Composition

There are several ways of expressing the concentration of a solution by using a percentage. The **mass/mass percent concentration, (m/m)%,** is defined as the mass of a solute divided by the mass of a solution times 100 (Equation 10.5.1):

$$(\text{m/m})\% = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100\% \tag{10.5.1}$$

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.

Example 10.5.1

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:

$$(\text{m/m})\% = \frac{36.5\ \text{g}}{355\ \text{g}} \times 100\% = 10.3\%$$

For gases and liquids, volumes are relatively easy to measure, so the concentration of a liquid or a gas solution can be expressed as a **volume/volume percent concentration, (v/v)%,** the volume of a solute divided by the volume of a solution times 100 (Equation 10.5.2):

$$(\text{v/v})\% = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100\% \tag{10.5.2}$$

Again, the units of the solute and the solution must be the same. A hybrid concentration unit, **mass/volume percent concentration, (m/v)%,** is commonly used for intravenous (IV) fluids (Figure 10.5.1). It is defined as the mass in grams of a solute, divided by volume in milliliters of solution times 100 (Equation 10.5.3):

$$(\text{m/v})\% = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\% \tag{10.5.3}$$



Figure 10.5.1: Mass/Volume Percent0.9% Sodium Chloride— A Common Medical Solution. The 0.9% NaCl concentration on this IV bag is mass/volume percent. An intravenous (IV) bag containing a 0.9% sodium chloride solution, commonly known as normal saline. This isotonic solution is used to restore fluids, maintain electrolyte balance, and deliver medications directly into the bloodstream. (CC BY-SA 3.0; Matanya via Wikimedia Commons). Detailed Description of Figure 10.5.1.

Using Percent Composition in Calculations

The percent concentration can be used to produce a conversion factor between the amount of solute and the amount of solution. As such, concentrations can be useful in a variety of stoichiometry problems as discussed in Chapter 6. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

As an example, if the given concentration is 5% (v/v) solution of alcohol, this means that there are 5 mL of alcohol dissolved in every 100 mL solution.

5 mL alcohol in 100 mL solution

The two possible conversion factors are written as follows:

$$\frac{5\text{ mL alcohol}}{100\text{ mL solution}} \text{ or } \frac{100\text{ mL solution}}{5\text{ mL alcohol}}$$

Use the first conversion factor to convert from a given amount of solution to the amount of solute. The second conversion factor is used to convert from a given amount of solute to the amount of solution. Given any two quantities in any percent composition, the third quantity can be calculated, as the following example illustrates.

Example 10.5.2

A sample of 45.0% v/v solution of ethanol (C_2H_5OH) in water has a volume of 115 mL. What volume of ethanol solute does the sample contain?

Solution

A percentage concentration is simply the number of parts of solute per 100 parts of solution. Thus, the percent concentration of 45.0% v/v implies the following:

$$45.0\%\text{ (v/v)} \rightarrow \frac{45\text{ mL } C_2H_5OH}{100\text{ mL solution}}$$

That is, there are 45 mL of C_2H_5OH for every 100 mL of solution. We can use this fraction as a conversion factor to determine the amount of C_2H_5OH in 115 mL of solution:

$$115 \text{ mL solution} \times \frac{45 \text{ mL } C_2H_5OH}{100 \text{ mL solution}} = 51.8 \text{ mL } C_2H_5OH$$

Example 10.5.3

What volume of a 12.75% m/v solution of glucose ($C_6H_{12}O_6$) in water is needed to obtain 50.0 g of $C_6H_{12}O_6$?

Solution

$$50.0 \text{ g } C_6H_12O_6 \times \frac{100 \text{ mL solution}}{12.75 \text{ g } C_6H_12O_6} = 392 \text{ mL solution}$$

Example 10.5.4

A normal saline IV solution contains 9.0 g of NaCl in every liter of solution. What is the mass/volume percent of normal saline?

Solution

We can use the definition of mass/volume percent, but first we have to express the volume in milliliter units:

1 L = 1,000 mL

Because this is an exact relationship, it does not affect the significant figures of our result.

$$(\text{m/v})\% = \frac{9.0 \text{ g NaCl}}{1,000 \text{ mL solution}} \times 100\% = 0.90\% \text{ (m/v)}$$

Example 10.5.5

The chlorine bleach that you might find in your laundry room is typically composed of 27.0 g of sodium hypochlorite (NaOCl), dissolved to make 500.0 mL of solution. What is the mass/volume percent of the bleach?

Solution

$$(\text{m/v})\% = \frac{27.0 \text{ g NaOCl}}{500.0 \text{ mL solution}} \times 100\% = 5.40\% \text{ (m/v)}$$

In addition to percentage units, the units for expressing the concentration of extremely dilute solutions are **parts per million (ppm)** and **parts per billion (ppb)**. Both of these units are mass based and are defined as follows:

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000 \tag{10.5.4}$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000,000 \tag{10.5.5}$$

Concentrations of *trace elements* in the body—elements that are present in extremely low concentrations but are nonetheless necessary for life—are commonly expressed in parts per million or parts per billion. Concentrations of poisons and pollutants are also described in these units. For example, cobalt is present in the body at a concentration of 21 ppb, while the State of Oregon's Department of Agriculture limits the concentration of arsenic in fertilizers to 9 ppm. In aqueous solutions, 1 ppm is essentially equal to 1 mg/L, and 1 ppb is equivalent to 1 µg/L.

Example 10.5.6

If the concentration of cobalt in a human body is 21 ppb, what mass in grams of Co is present in a body having a mass of 70.0 kg?

Solution

A concentration of 21 ppb means "21 g of solute per 1,000,000,000 g of solution." Written as a conversion factor, this concentration of Co is as follows:

$$21 \text{ ppb Co} \rightarrow \frac{21 \text{ g Co}}{1,000,000,000 \text{ g solution}}$$

We can use this as a conversion factor, but first we must convert 70.0 kg to gram units:

$$70.0 \text{ kg} \times \frac{1,000 \text{ g}}{1 \text{ kg}} = 7.00 \times 10^4 \text{ g}$$

Now we determine the amount of Co:

$$7.00 \times 10^4 \text{ g solution} \times \frac{21 \text{ g Co}}{1,000,000,000 \text{ g solution}} = 0.0015 \text{ g Co}$$

This is only 1.5 mg.

Molarity

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. 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.

Molarity is defined as the number of moles of a solute dissolved per liter of solution:

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}}$$

Molarity is abbreviated M (often referred to as "molar"), and the units are often abbreviated as mol/L. 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 therefore

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl}$$

which is read as "three point oh molar sodium chloride." Sometimes (aq) is added when the solvent is water, as in "3.0 M NaCl(aq)."

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.

If the amount of solute is given in mass units, it must first be converted to moles before using the definition of molarity to calculate concentration. For example, to determine the molar concentration of a solution containing 22.4 g of HCl dissolved in 1.56 L of solution, we use a flowchart showing mass of solute leading to moles of solute, which then leads to molarity of solution with the addition of volume in liters (Figure 10.5.2).

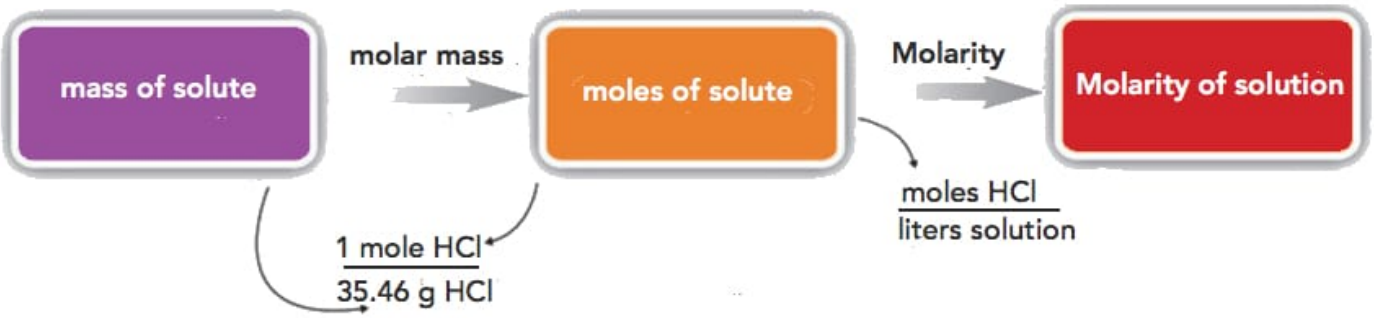


Figure 10.5.2: This flowchart outlines the process of calculating the molarity of a solution by converting the mass of solute to moles using molar mass, then dividing by the volume of solution in liters. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.2.

Step 1: Convert the mass of solute to moles using the molar mass of HCl (36.46 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ HCl}{36.46\ \cancel{g HCl}} = 0.614\ mol\ HCl$$

Step 2: Use the definition of molarity to determine the concentration:

$$M\ =\ \frac{0.614\ mol\ HCl}{1.56 L\ solution} = 0.394\ M HCl$$

Example 10.5.7

What is the molarity of an aqueous solution of 25.0 g of NaOH in 750 mL?

Solution

Before we substitute these quantities into the definition of molarity, we must convert them to the proper units. The mass of NaOH must be converted to moles of NaOH. The molar mass of NaOH is 40.00 g/mol:

$$25.0\ \text{g NaOH} \times \frac{1\ \text{mol NaOH}}{40.00\ \text{g NaOH}} = 0.625\ \text{mol NaOH}$$

Next, we convert the volume units from milliliters to liters:

$$750\ \text{mL} \times \frac{1\ \text{L}}{1,000\ \text{mL}} = 0.750\ \text{L}$$

Now that the quantities are expressed in the proper units, we can substitute them into the definition of molarity:

$$\text{M} = \frac{0.625\ \text{mol NaOH}}{0.750\ \text{L}} = 0.833\ \text{M NaOH}$$

Using Molarity in Calculations

The definition of molarity can also be used to calculate a needed volume of solution, given its concentration and the number of moles desired, or the number of moles of solute (and subsequently, the mass of the solute), given its concentration and volume. As in the percent concentration, molarity can also be expressed as a conversion factor.

Molarity is defined as moles solute per liter solution. There is an understood 1 in the denominator of the conversion factor. For example, a 3.0 M solution of sucrose means that there are three moles of sucrose dissolved in every liter of solution. Mathematically, this is stated as follows:

3.0 moles sucrose in 1 L solution

Dividing both sides of this expression by either side, we generate two possible conversion factors:

$$\frac{3.0\text{ mol sucrose}}{1\text{ L solution}} \text{ or } \frac{1\text{ L solution}}{3.0\text{ mol sucrose}}$$

The first conversion factor can be used to convert from *volume (L) of solution to moles solute*, and the *second converts from moles of solute to volume (L) of solution* (Figure 10.5.3).

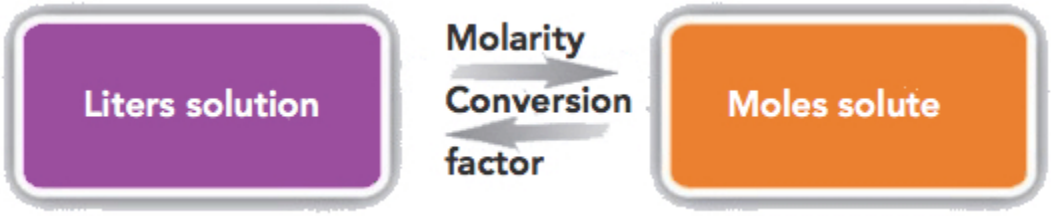


Figure 10.5.3: Molarity is used to convert between liters of solution and moles of solute, depending on the direction of the calculation. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.3.

For example, suppose we are asked how many moles of sucrose are present in 0.108 L of a 3.0 M sucrose solution. The given volume (0.108 L) is multiplied by the first conversion factor to cancel the L units, and find that 0.32 moles of sucrose are present.

$$0.108\ \cancel{L\ solution} \times \frac{3.0\ mol\ sucrose}{1\cancel{L\ solution}} = 0.32\ mol\ sucrose$$

How many liters of 3.0 M sucrose solution are needed to obtain 4.88 mol of sucrose? In such a conversion, we multiply the given (4.88 moles sucrose) with the second conversion factor. This cancels the moles units and converts it to liters of solution.

$$4.88\ \cancel{mol\ sucrose} \times \frac{1\ L\ solution}{3.0\cancel{mol\ sucrose}} = 1.63\ L\ solution$$

Example 10.5.8

What volume of a 0.0753 M solution of dimethylamine [$(CH_3)_2NH$] is needed to obtain 0.450 mol of the compound?

Solution

To solve for the volume, multiply the "given" (0.450 mol of dimethylamine) with the molarity conversion factor (0.0753 M). Use the proper conversion factor to cancel the unit "mol" and get the unit volume (L) of solution:

$$0.450\text{ mol dimethylamine} \times \frac{1\text{ L solution}}{0.0753\text{ mol dimethylamine}} = 5.98\text{ L solution}$$

Example 10.5.9

Ethylene glycol ($C_2H_6O_2$) is mixed with water to make auto engine coolants. How many grams of $C_2H_6O_2$ are in 5.00 L of a 6.00 M aqueous solution?

Solution

1. To solve for the volume, multiply the "given" (0.450 mol of dimethylamine) with the molarity conversion factor (0.0753 M). Use the proper conversion factor to cancel the unit "mol" and get the unit volume (L) of solution:

$$0.450\text{ mol dimethylamine} \times \frac{1\text{ L solution}}{0.0753\text{ mol dimethylamine}} = 5.98\text{ L solution}$$

2. The strategy in solving this problem is to convert the given volume (5.00 L) using the 6.00 M (conversion factor) to solve for moles of ethylene glycol, which can then be converted to grams.

Step 1: Convert the given volume (5.00 L) to moles ethylene glycol.

$$5.00\ \text{L solution} \times \frac{6.00\ \text{mol}\ C_2H_6O_2}{1\ \text{L solution}} = 30.0\ \text{mol}\ C_2H_6O_2$$

Step 2: Convert 30.0 mols $C_2H_6O_2$ to grams $C_2H_6O_2$. Molar mass of $C_2H_6O_2$= 62.08 g/mol

$$30.0\ \text{mol}\ C_2H_6O_2 \times \frac{62.08\ \text{g}\ C_2H_6O_2}{1\ \text{mol}\ C_2H_6O_2} = 1,860\ \text{g}\ C_2H_6O_2$$

The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows (Figure 10.5.4):

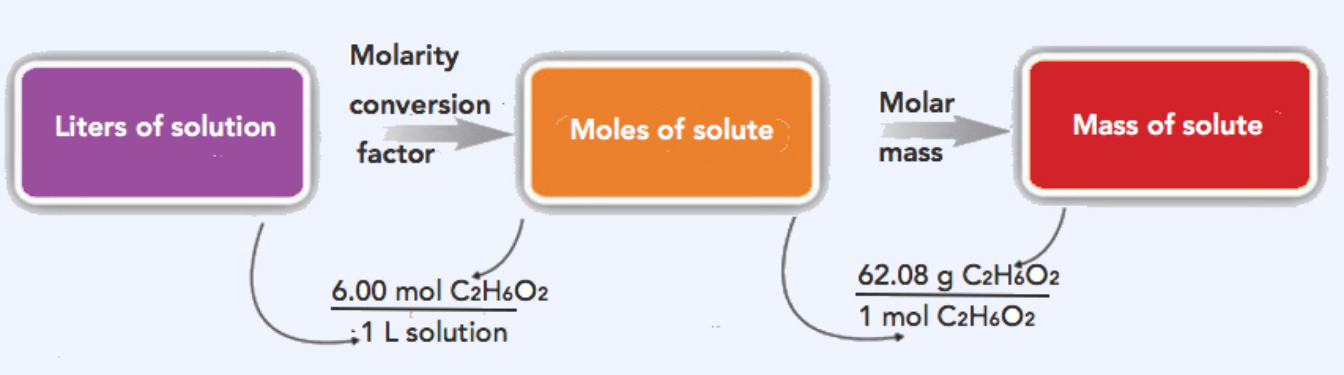


Figure 10.5.4: Molarity and molar mass are used together to convert a solution's volume into moles of solute, and then into mass of solute. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.4.

$$5.00\ \text{L solution} \times \frac{6.00\ \text{mol}\ C_2H_6O_2}{1\ \text{L solution}} \times \frac{62.08\ \text{g}\ C_2H_6O_2}{1\ \text{mol}\ C_2H_6O_2} = 1,860\ \text{g}\ C_2H_6O_2$$

The final answer is rounded off to 3 significant figures. Thus, there are 1,860 g of $C_2H_6O_2$ in the specified amount of engine coolant.

Note: Dimethylamine has a "fishy" odor. In fact, organic compounds called amines cause the odor of decaying fish.

Example 10.5.10

What volume of a 0.0753 M solution of dimethylamine [$(CH_3)_2NH$] is needed to obtain 0.450 mol of the compound?

Solution

This is an alternative method in case you don't want to use the conversion factor for molarity. In both parts, we will use the definition of molarity to solve for the desired quantity.

$$0.0753\ \text{M} = \frac{0.450\ \text{mol}\ (CH_3)_2NH}{\text{volume of solution}}$$

To solve for the volume of solution, we multiply both sides by volume of solution and divide both sides by the molarity value to isolate the volume of solution on one side of the equation:

$$\text{volume of solution} = \frac{0.450\ \text{mol}\ (CH_3)_2NH}{0.0753\ \text{M}} = 5.98\ \text{L}$$

Note that because the definition of molarity is mol/L, the division of mol by M yields L, a unit of volume.

Example 10.5.11

Ethylene glycol ($C_2H_6O_2$) is mixed with water to make auto engine coolants. How many grams of $C_2H_6O_2$ are in 5.00 L of a 6.00 M aqueous solution?

Solution

The molar mass of $C_2H_6O_2$ is 62.08 g/mol., so

$$6.00\ \text{M} = \frac{\text{moles of solute}}{5.00\ \text{L}}$$

To solve for the number of moles of solute, we multiply both sides by the volume:

$$\text{moles of solute} = (6.00\ \text{M})(5.00\ \text{L}) = 30.0\ \text{mol}$$

Note that because the definition of molarity is mol/L, the product M × L gives mol, a unit of amount. Now, using the molar mass of $C_3H_8O_3$, we convert mol to g:

$$30.0\ \text{mol} \times \frac{62.08\ \text{g}}{\text{mol}} = 1,860\ \text{g}$$

Thus, there are 1,860 g of $C_2H_6O_2$ in the specified amount of engine coolant.

Using Molarity in Stoichiometry Problems

Of all the ways of expressing concentration, molarity is the one most commonly used in stoichiometry problems because it is directly related to the mole unit. Consider the following chemical equation:

$$\text{HCl}\,(\text{aq}) + \text{NaOH}\,(\text{aq}) \longrightarrow \text{H}_2\text{O}(\ell) + \text{NaCl}\,(\text{aq}) \tag{10.5.6}$$

Suppose we want to know how many liters of aqueous HCl solution will react with a given mass of NaOH. A typical approach to answering this question is as follows (Figure 10.5.5):

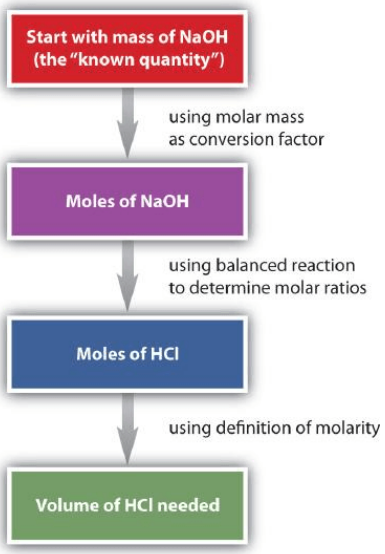


Figure 10.5.5: To calculate the volume of HCl needed to react with a given mass of NaOH, convert the mass to moles, apply the mole ratio from the balanced equation, and use molarity to find the required volume. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.5.

Example 10.5.12

How many milliliters of a 2.75 M HCl solution are needed to react with 185 g of NaOH? The balanced chemical equation for this reaction is as follows:

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow H_2O(\ell) + NaCl\,(aq)}$$

Solution

We will follow the flowchart to answer this question. First, we convert the mass of NaOH to moles of NaOH using its molar mass, 40.00 g/mol:

$$185\ \mathrm{g\ NaOH} \times \frac{1\ \mathrm{mol\ NaOH}}{40.00\ \mathrm{g\ NaOH}} = 4.63\ \mathrm{mol\ NaOH}$$

Using the balanced chemical equation, we see that there is a one-to-one ratio of moles of HCl to moles of NaOH. We use this to determine the number of moles of HCl needed to react with the given amount of NaOH:

$$4.63\ \mathrm{mol\ NaOH} \times \frac{1\ \mathrm{mol\ HCl}}{1\ \mathrm{mol\ NaOH}} = 4.63\ \mathrm{mol\ HCl}$$

Finally, we use the definition of molarity to determine the volume of 2.75 M HCl needed:

$$2.75\ \mathrm{M\ HCl} = \frac{4.63\ \mathrm{mol\ HCl}}{\mathrm{volume\ of\ HCl\ solution}}$$

$$\mathrm{volume\ of\ HCl} = \frac{4.63\ \mathrm{mol\ HCl}}{2.75\ \mathrm{M\ HCl}} = 1.68\ \mathrm{L} \times \frac{1,000\ \mathrm{mL}}{1\ \mathrm{L}} = 1,680\ \mathrm{mL}$$

We need 1,680 mL of 2.75 M HCl to react with the NaOH.

The same multi-step problem can also be worked out in a single line, rather than as separate steps, as follows (Figure 10.5.6):

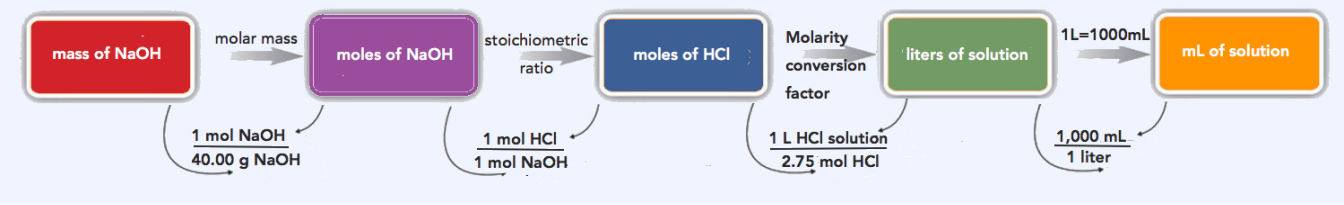


Figure 10.5.6: To determine the volume of HCl solution needed, convert the given mass of NaOH to moles, use the mole ratio from the balanced equation to find moles of HCl, then apply molarity to calculate volume in liters and convert to milliliters. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.6.

$$185\ \mathrm{g\ NaOH} \times \frac{1\ \mathrm{mol\ NaOH}}{40.00\ \mathrm{g\ NaOH}} \times \frac{1\ \mathrm{mol\ HCl}}{1\ \mathrm{mol\ NaOH}} \times \frac{1\ \mathrm{L\ HCl\ solution}}{2.75\ \mathrm{mol\ HCl}} \times \frac{1000\ \mathrm{mL\ HCl\ solution}}{1\ \mathrm{L\ HCl\ solution}} = 1,680\ \mathrm{mL\ HCl\ solution}$$

Our final answer (rounded off to three significant figures) is 1,680 mL HCl solution.

The general steps for performing stoichiometry problems such as this are shown in Figure 10.5.7. You may want to consult this figure when working with solutions in chemical reactions. The double arrows in Figure 10.5.7 indicate that you can start at either end of the chart and, after a series of simple conversions, determine the quantity at the other end.

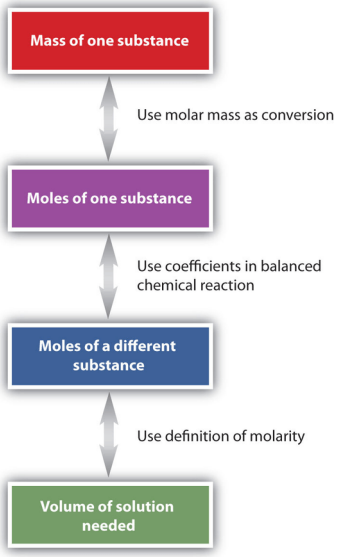


Figure 10.5.7: To find the volume of a solution needed for a reaction, start with the mass of a known substance, convert to moles, use mole ratios from the balanced chemical equation, and apply molarity to calculate volume. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.7.

Many of the fluids found in our bodies are solutions. The solutes range from simple ionic compounds to complex proteins with typical concentrations (Table 10.5.1).

Table 10.5.1: Approximate Concentrations of Various Solutes in Some Solutions in the Body

Solution	Solute	Concentration (M)
blood plasma	Na^+	0.138
	K^+	0.005
	Ca^{2+}	0.004
	Mg^{2+}	0.003
	Cl^-	0.110
	HCO_3^-	0.030
stomach acid	HCl	0.10
urine	NaCl	0.15
	PO_4^{3-}	0.05
	NH_2CONH_2 (urea)	0.30

Note: Concentrations are approximate and can vary widely (Table 10.5.1).

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the number of moles of the solute does not change, while the volume of the solution increases. We can set up an equality between the moles of the solute before the dilution (1) and the moles of the solute after the dilution (2).

$$\text{mol}_1 = \text{mol}_2$$

Since the moles of solute in a solution is equal to the molarity multiplied by the liters, we can set those equal.

$$M_1 \times L_1 = M_2 \times L_2$$

Finally, because the two sides of the equation are set equal to one another, the volume can be in any units we choose, as long as that unit is the same on both sides. Our equation for calculating the molarity of a diluted solution becomes:

$$M_1 \times V_1 = M_2 \times V_2$$

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.

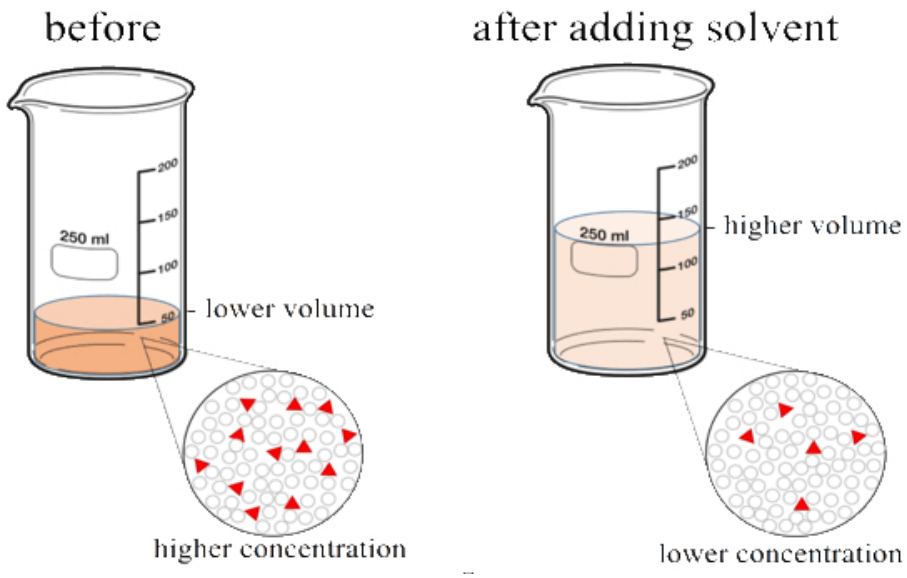


Figure 10.5.8: Adding more solvent to a solution increases its volume and decreases the concentration of solute particles without changing the total amount of solute. (CC BY-NC-SA 4.0; anonymous via Introductory Biochemistry). Detailed Description of Figure 10.5.8.

Dilution Calculations: Molarity and Percent Solutions

Any units of concentration and volume can be used, as long as both concentrations and both volumes have the same unit. For example, if we are using M (molarity), then we can express the equation as follows:

$$\text{Molarity}_{\text{initial}} \times \text{Volume}_{\text{initial}} = \text{Molarity}_{\text{final}} \times \text{Volume}_{\text{final}} \tag{10.5.7}$$

$$\text{M}_1 \times \text{V}_1 = \text{M}_2 \times \text{M}_2 \tag{10.5.8}$$

If we are using percent, the dilution equation is as follows:

$$\text{Concentration}_{\text{initial}} \times \text{Volume}_{\text{initial}} = \text{Concentration}_{\text{final}} \times \text{Volume}_{\text{final}} \tag{10.5.9}$$

$$\text{C}_1 \times \text{V}_1 = \text{C}_2 \times \text{M}_2 \tag{10.5.10}$$

Example 10.5.13

A 125 mL sample of 0.900 M NaCl is diluted to 1,125 mL. What is the final concentration of the diluted solution?

Solution

Because the volume units are the same, and we are looking for the molarity of the final solution, we can use (concentration × volume)$_{initial}$ = (concentration × volume)$_{final}$:

(0.900 M × 125 mL) = (concentration × 1,125 mL)

We solve by isolating the unknown concentration by itself on one side of the equation. Dividing by 1,125 mL gives

$$\text{concentration} = \frac{0.900\ \text{M} \times 125\ \text{mL}}{1,125\ \text{mL}} = 0.100\ \text{M}$$

as the final concentration.

Key Takeaways

- Percent concentration can be expressed as % m/m, % v/v, or % m/v, depending on the solution's components.
- ppm and ppb are used for very dilute solutions, especially in environmental and biological contexts.
- Molarity (mol/L) is a standard unit for concentration, linking moles of solute to volume of solution.
- Molarity is useful in stoichiometric calculations to convert between volume and moles.
- Dilution decreases concentration by increasing volume, following the equation:

$$\mathrm{M_1}\ \times \mathrm{V_1}\ = \mathrm{M_2}\ \times \mathrm{M_2}$$

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10.6: Properties of Solutions

Colligative Properties

Solutes affect some properties of solutions that depend only on the concentration of the dissolved particles. These properties are called **colligative properties**. Four important colligative properties that we will examine here are vapor pressure depression, boiling point elevation, freezing point depression, and osmotic pressure.

Molecular compounds separate into individual molecules when they are dissolved, so for every 1 mol of molecules dissolved, we get 1 mol of particles. In contrast, ionic compounds separate into their constituent ions when they dissolve, so 1 mol of an ionic compound will produce more than 1 mol of dissolved particles. For example, every mole of NaCl that dissolves yields 1 mol of Na^+ ions and 1 mol of Cl^- ions, for a total of 2 mol of particles in solution. Thus, the effect on a solution's properties by dissolving NaCl may be twice as large as the effect of dissolving the same amount of moles of glucose ($C_6H_{12}O_6$).

Vapor Pressure Depression

All liquids evaporate. In fact, given enough volume, a liquid will turn completely into a vapor. If enough volume is not present, a liquid will evaporate only to the point where the rate of evaporation equals the rate of vapor condensing back into a liquid. The pressure of the vapor at this point is called the **vapor pressure** of the liquid.

The presence of a dissolved solid lowers the characteristic vapor pressure of a liquid so that it evaporates more slowly. (The exceptions to this statement are if the solute itself is a liquid or a gas, in which case the solute will also contribute something to the evaporation process. We will not discuss such solutions here.) This property is called **vapor pressure depression** and is depicted in Figure 10.6.1.

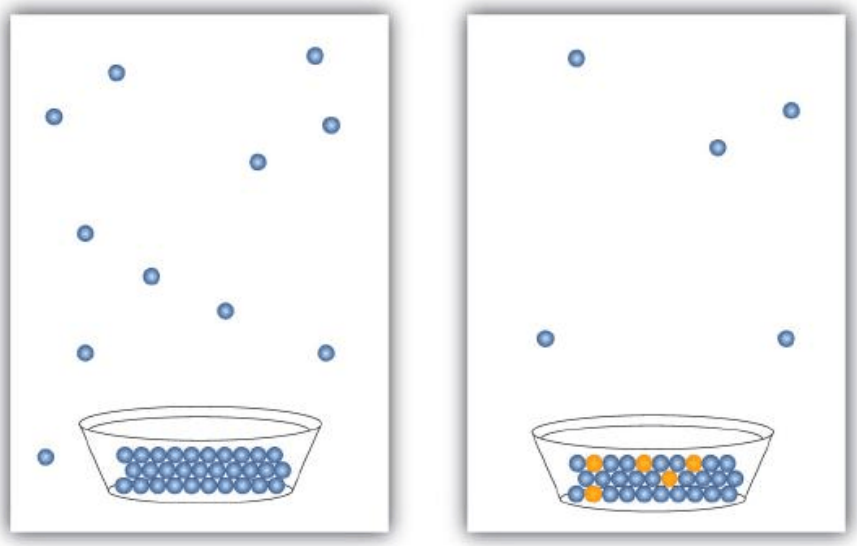


Figure 10.6.1: Adding a nonvolatile solute to a solvent lowers the vapor pressure by reducing the number of solvent molecules at the surface that can escape into the gas phase. (CC BY-NC-SA 3.0, anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 10.6.1.

Boiling Point and Freezing Point Effects

A related property of solutions is that their boiling points are higher than the boiling point of the pure solvent. Because the presence of solute particles decreases the vapor pressure of the liquid solvent, a higher temperature is needed to reach the

boiling point. This phenomenon is called **boiling point elevation**. For every mole of particles dissolved in a liter of water, the boiling point of water increases by about 0.5°C. The addition of one mole of sucrose (molecular compound) in one liter of water will raise the boiling point from 100°C to 100.5°C but the addition of one mole of NaCl in one liter of water will raise the boiling point by 2 x 0.5°C = 1°C. Furthermore, the addition of one mole of CaCl2CaCl2 in one liter of water will raise the boiling point by 3 x 0.5°C = 1.5°C.

The presence of solute particles has the opposite effect on the freezing point of a solution. When a solution freezes, only the solvent particles come together to form a solid phase, and the presence of solute particles interferes with that process. Therefore, for the liquid solvent to freeze, more energy must be removed from the solution, which lowers the temperature. Thus, solutions have lower freezing points than pure solvents do. This phenomenon is called **freezing point depression**. For every mole of particles in a liter of water, the freezing point decreases by about 1.9°C (Figure 10.6.2).

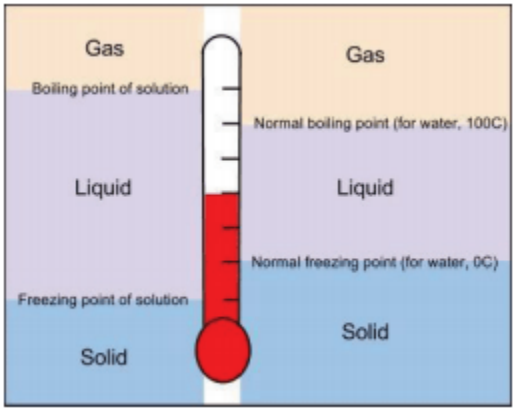


Figure 10.6.2: The presence of a solute raises the boiling point and lowers the freezing point of a solution compared to the pure solvent. (CC BY-NC-SA 4.0; Delmar Larsen via CHEM 114: Introductory Chemistry). Detailed Description of Figure 10.6.2.

Both boiling point elevation and freezing point depression have practical uses. For example, solutions of water and ethylene glycol ($C_2H_6O_2$) are used as coolants in automobile engines because the boiling point of such a solution is greater than 100°C, the normal boiling point of water. In winter, salts like NaCl and $CaCl_2$ are sprinkled on the ground to melt ice or keep ice from forming on roads and sidewalks (Figure 10.6.3). This is because the solution made by dissolving sodium chloride or calcium chloride in water has a lower freezing point than pure water, so the formation of ice is inhibited.



Figure 10.6.3: The salt sprinkled on this sidewalk makes the water on the sidewalk have a lower freezing point than pure water, so it does not freeze as easily. This makes walking on the sidewalk less hazardous in winter. (CC BY-NC-SA 3.0, anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 10.6.3.

Example 10.6.1

Which solution's freezing point deviates more from that of pure water—a 1 M solution of NaCl or a 1 M solution of $\mathrm{CaCl_2}$?

Solution

Colligative properties depend on the number of dissolved particles, so the solution with the greater number of particles in solution will show the greatest deviation. When NaCl dissolves, it separates into two ions, Na^+ and Cl^-. But when $\mathrm{CaCl_2}$ dissolves, it separates into three ions—one Ca^{2+} ion and two Cl^- ions. Thus, mole for mole, $\mathrm{CaCl_2}$ will have 50% more impact on freezing point depression than NaCl.

Example 10.6.2

Estimate the boiling point of 0.2 M $\mathrm{CaCl_2}$ solution.

Solution

The boiling point increases 0.5^0C for every mole of solute per liter of water. For this estimation, let's assume that 1 liter of solution is roughly the same volume as 1 liter of water. A 0.2 M $\mathrm{CaCl_2}$ solution contains 0.2 moles of $\mathrm{CaCl_2}$ solution formula units per liter of solution. Each $\mathrm{CaCl_2}$ unit separates into three ions.

$$0.2 \text{ mol } \mathrm{CaCl_2} \times \frac{3 \text{ mol ions}}{1 \text{ mol } \mathrm{CaCl_2}} \times \frac{0.5 \text{ deg C}}{1 \text{ mol ion}} = 0.3 \text{ deg C}$$

The normal boiling point of water is 100^0C, so the boiling point of the solution is raised to 100.3^0C.

Osmotic Pressure

Certain natural and synthetic materials exhibit **selective permeation**, allowing only specific molecules or ions—based on size, shape, polarity, or charge—to pass through. Biological cell membranes are excellent natural examples of selective permeation, while dialysis tubing used in medical treatments serves as a simplified synthetic counterpart. Such materials are generally known as **semipermeable membranes**.

Osmosis is the process by which solvent molecules move through a semipermeable membrane, while solute particles are unable to pass. When solutions of different concentrations are separated by such a membrane, solvent molecules move from the side with lower solute concentration (more dilute) to the side with higher solute concentration (more concentrated). This movement continues until equilibrium is reached. The driving force behind this movement is known as **osmotic pressure**, and it can be counteracted by applying external pressure to the more concentrated solution. The pressure required to stop the net movement of solvent is equal to the solution's osmotic pressure.

In a typical setup, such as the one shown in Figure 10.6.4, a semipermeable membrane separates a pure solvent from a solution. Although solvent molecules move in both directions across the membrane, the net flow is from the pure solvent into the solution, due to the higher solvent concentration on the pure side. This directional movement of solvent molecules, driven by concentration differences, defines the process of osmosis.

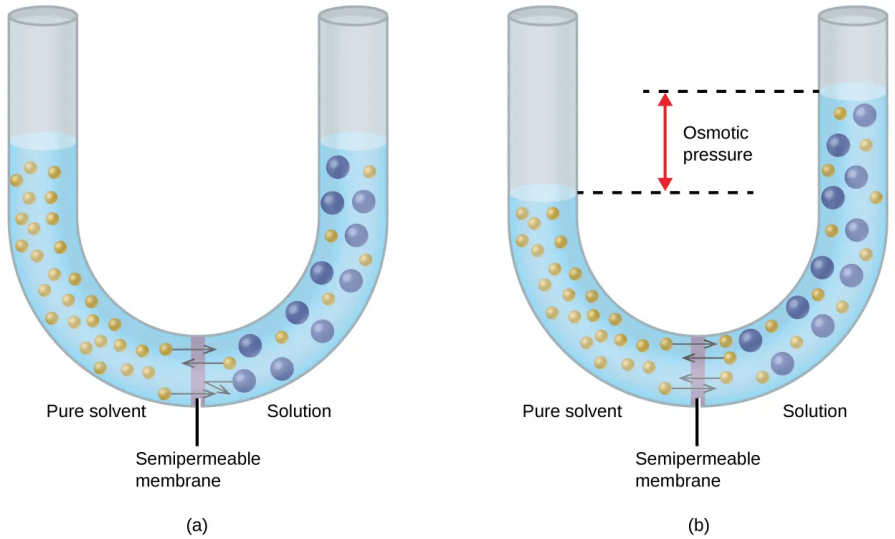


Figure 10.6.4: When a solution and pure solvent are separated by a semipermeable membrane, solvent flows into the solution, creating osmotic pressure. (CC BY 4.0; OpenStax via source content). Detailed Description of Figure 10.6.4.

A solution composed of 100 mL of water and 15 g of table sugar consists of a solute (sugar) dissolved in a solvent (water). A solution is defined as a homogeneous mixture in which one or more solutes are uniformly distributed within a solvent.

A second solution prepared with 100 mL of water and 45 g of table sugar also consists of sugar as the solute and water as the solvent, but has a higher solute concentration than the first. When two solutions have unequal solute concentrations, the one with the higher solute concentration is termed **hypertonic**, while the one with the lower solute concentration is termed **hypotonic**. If two solutions contain equal solute concentrations, they are described as **isotonic**. In this case, the first solution is hypotonic relative to the second, and the second solution is hypertonic relative to the first.

When these two solutions are placed on opposite sides of a semipermeable membrane that allows the passage of water molecules but not sugar molecules, a concentration gradient of water is established. Water moves from the hypotonic solution (higher water concentration) to the hypertonic solution (lower water concentration). This net movement of water continues until equilibrium is reached and both solutions become isotonic.

Red blood cells behave the same way (see figure below). When red blood cells are in a hypertonic (higher concentration) solution, water flows out of the cell faster than it comes in. This results in *crenation* (shriveling) of the blood cell. On the other extreme, a red blood cell that is hypotonic (lower concentration outside the cell) will result in more water flowing into the cell than out. This results in swelling of the cell and potential *hemolysis* (bursting) of the cell. In an isotonic solution, the flow of water in and out of the cell is happening at the same rate.

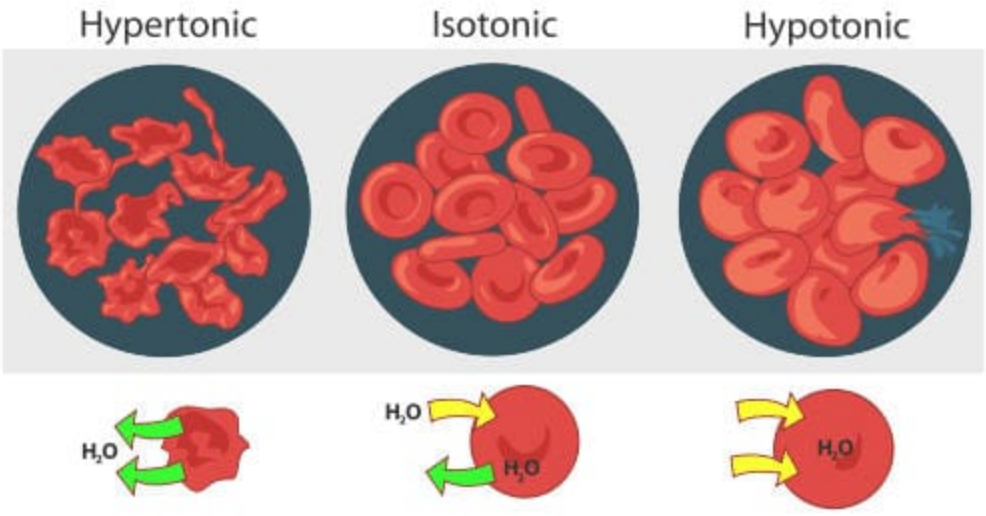


Figure 10.6.5: Red blood cells respond differently in hypertonic, isotonic, and hypotonic solutions due to water movement across their membranes. (CC BY-NC-SA 4.0; Theodore Chan via Introductory Biochemistry). Detailed Description of Figure 10.6.5.

Osmosis is the diffusion of water molecules across a semipermeable membrane from an area of lower concentration solution (i.e., higher concentration of water) to an area of higher concentration solution (i.e., lower concentration of water). Water moves into and out of cells by osmosis.

- If a cell is in a hypertonic solution, the solution has a lower water concentration than the cell cytosol, and water moves out of the cell until both solutions are isotonic.
- Cells placed in a hypotonic solution will take in water across their membranes until both the external solution and the cytosol are isotonic.

A red blood cell will swell and undergo hemolysis (burst) when placed in a hypotonic solution. When placed in a hypertonic solution, a red blood cell will lose water and undergo *crenation* (shrivel). Animal cells tend to do best in an isotonic environment, where the flow of water in and out of the cell is occurring at equal rates.

Osmolarity (osmol) is a way of reporting the total number of particles in a solution to determine osmotic pressure. It is defined as the molarity of a solute times the number of particles a formula unit of the solute makes when it dissolves (represented by i):

$$osmol = M \times i \tag{10.6.1}$$

If more than one solute is present in a solution, the individual osmolarities are additive to get the total osmolarity of the solution. Solutions that have the same osmolarity have the same osmotic pressure. If solutions of differing osmolarities are present on opposite sides of a semipermeable membrane, solvent will transfer from the lower-osmolarity solution to the higher-osmolarity solution. Counterpressure exerted on the high-osmolarity solution will reduce or halt the solvent transfer. An even higher pressure can be exerted to force solvent from the high-osmolarity solution to the low-osmolarity solution, a process called *reverse osmosis*. Reverse osmosis is used to make potable water from saltwater where sources of fresh water are scarce.

Example 10.6.3

A 0.50 M NaCl aqueous solution and a 0.30 M $Ca(NO_3)_2$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

Solution

The solvent will flow into the solution of higher osmolarity. The NaCl solute separates into two ions—Na^+ and Cl^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol (NaCl)} = 0.50\ \text{M} \times 2 = 1.0\ \text{osmol}$$

The $Ca(NO_3)_2$ solute separates into three ions—one Ca^{2+} and two NO_3^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol } [Ca(NO_3)_2] = 0.30\ \text{M} \times 3 = 0.90\ \text{osmol}$$

The osmolarity of the $Ca(NO_3)_2$ solution is lower than that of the NaCl solution, so water will transfer through the membrane from the $Ca(NO_3)_2$ solution to the NaCl solution.

Link to Learning: Chemistry in Everyday Life, Dialysis

The main function of the kidneys is to filter the blood to remove wastes and extra water, which are then expelled from the body as urine. Some diseases rob the kidneys of their ability to perform this function, causing a buildup of waste materials in the bloodstream. If a kidney transplant is not available or desirable, a procedure called dialysis can be used to remove waste materials and excess water from the blood. In one form of dialysis, called hemodialysis, a patient's blood is passed though a length of tubing that travels through an artificial kidney machine (also called a dialysis machine). A section of tubing composed of a semipermeable membrane is immersed in a solution of sterile water, glucose, amino acids, and certain electrolytes. The osmotic pressure of the blood forces waste molecules and excess water through the membrane into the sterile solution. Red and white blood cells are too large to pass through the membrane, so they remain in the blood. After being cleansed in this way, the blood is returned to the body.

Dialysis is a continuous process, as the osmosis of waste materials and excess water takes time. Typically, 5–10 lb of waste-containing fluid is removed in each dialysis session, which can last 2–8 hours and must be performed several times a week. Although some patients have been on dialysis for 30 or more years, dialysis is always a temporary solution because waste materials are constantly building up in the bloodstream. A more permanent solution is a kidney transplant (Figure 10.6.6).



Figure 10.6.6: Dialysis uses a semipermeable membrane to filter waste and excess fluids from the blood of patients with kidney failure. Blood flows through tubing into the dialysis machine, where solutes diffuse across the membrane while essential components remain in circulation. (CC BY-NC-SA 3.0, anonymous via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 10.6.6.

Cell walls are semipermeable membranes, so the osmotic pressures of the body's fluids have important biological consequences. If solutions of different osmolarity exist on either side of the cells, solvent (water) may pass into or out of the cells, sometimes with disastrous results. Consider what happens if red blood cells are placed in a hypotonic solution, meaning a solution of lower osmolarity than the liquid inside the cells. The cells swell up as water enters them, disrupting cellular activity and eventually causing the cells to burst. This process is called hemolysis. If red blood cells are placed in a hypertonic solution, meaning one having a higher osmolarity than exists inside the cells, water leaves the cells to dilute the external solution, and the red blood cells shrivel and die. This process is called crenation. Only if red blood cells are placed in isotonic solutions that have the same

osmolarity as exists inside the cells are they unaffected by negative effects of osmotic pressure. Glucose solutions of about 0.31 M, or sodium chloride solutions of about 0.16 M, are isotonic with blood plasma. The concentration of an isotonic sodium chloride (NaCl) solution is only half that of an isotonic glucose ($C_6H_{12}O_6$) solution because NaCl produces two ions when a formula unit dissolves, while molecular glucose, $C_6H_{12}O_6$ produces only one particle when a formula unit dissolves. The osmolarities are therefore the same even though the concentrations of the two solutions are different.

Osmotic pressure explains why you should not drink seawater if you are abandoned in a life raft in the middle of the ocean. Its osmolarity is about three times higher than most bodily fluids. You would actually become thirstier as water from your cells was drawn out to dilute the salty ocean water you ingested. Our bodies do a better job coping with hypotonic solutions than with hypertonic ones. The excess water is collected by our kidneys and excreted. Osmotic pressure effects are used in the food industry to make pickles from cucumbers and other vegetables and in brining meat to make corned beef. It is also a factor in the mechanism of getting water from the roots to the tops of trees!

Link to Learning: Career Focus, Perfusionist

A perfusionist is a medical technician trained to assist during any medical procedure in which a patient's circulatory or breathing functions require support. The use of perfusionists has grown rapidly since the advent of openheart surgery in 1953. Most perfusionists work in operating rooms, where their main responsibility is to operate heart-lung machines. During many heart surgeries, the heart itself must be stopped. In these situations, a heart-lung machine keeps the patient alive by aerating the blood with oxygen and removing carbon dioxide. The perfusionist monitors both the machine and the status of the blood, notifying the surgeon and the anesthetist of any concerns and taking corrective action if the status of the blood becomes abnormal. Despite the narrow parameters of their specialty, perfusionists must be highly trained. Certified perfusion education programs require a student to learn anatomy, physiology, pathology, chemistry, pharmacology, math, and physics. A college degree is usually required. Some perfusionists work with other external artificial organs, such as hemodialysis machines and artificial livers.

Key Takeaways

- Colligative properties depend on the number of dissolved particles, not their type.
- Adding solutes lowers vapor pressure and raises boiling point.
- Solutes also lower the freezing point of solutions.
- Osmosis is the movement of solvent through a semipermeable membrane toward higher solute concentration.
- Cell shape changes in hypertonic, hypotonic, and isotonic solutions due to water movement.

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10.7: Summary

A solution is a homogeneous mixture. The major component is the solvent, while the minor component is the solute. Solutions can have any phase; for example, an alloy is a solid solution. Solutes are soluble or insoluble, meaning they dissolve or do not dissolve in a particular solvent. The terms miscible and immiscible, instead of soluble and insoluble, are used for liquid solutes and solvents. The statement like dissolves like is a useful guide to predicting whether a solute will dissolve in a given solvent.

The amount of solute in a solution is represented by the concentration of the solution. The maximum amount of solute that will dissolve in a given amount of solvent is called the solubility of the solute. Such solutions are saturated. Solutions that have less than the maximum amount are unsaturated. Most solutions are unsaturated, and there are various ways of stating their concentrations. Mass/mass percent, volume/volume percent, and mass/volume percent indicate the percentage of the overall solution that is solute. Parts per million (ppm) and parts per billion (ppb) are used to describe very small concentrations of a solute. Molarity, defined as the number of moles of solute per liter of solution, is a common concentration unit in the chemistry laboratory. Equivalents express concentrations in terms of moles of charge on ions. When a solution is diluted, we use the fact that the amount of solute remains constant to be able to determine the volume or concentration of the final diluted solution.

Dissolving occurs by solvation, the process in which particles of a solvent surround the individual particles of a solute, separating them to make a solution. For water solutions, the word hydration is used. If the solute is molecular, it dissolves into individual molecules. If the solute is ionic, the individual ions separate from each other, forming a solution that conducts electricity. Such solutions are called electrolytes. If the dissociation of ions is complete, the solution is a strong electrolyte. If the dissociation is only partial, the solution is a weak electrolyte. Solutions of molecules do not conduct electricity and are called nonelectrolytes.

Solutions have properties that differ from those of the pure solvent. Some of these are colligative properties, which are due to the number of solute particles dissolved, not the chemical identity of the solute. Colligative properties include vapor pressure depression, boiling point elevation, freezing point depression, and osmotic pressure. Osmotic pressure is particularly important in biological systems. It is caused by osmosis, the passage of solvents through certain membranes like cell walls. The osmolarity of a solution is the product of a solution's molarity and the number of particles a solute separates into when it dissolves. Osmosis can be reversed by the application of pressure; this reverse osmosis is used to make fresh water from saltwater in some parts of the world. Because of osmosis, red blood cells placed in hypotonic or hypertonic solutions lose function through either hemolysis or crenation. If they are placed in isotonic solutions, however, the cells are unaffected because osmotic pressure is equal on either side of the cell membrane.

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10.8: Key Terms

- **Solution:** Another name for a homogeneous mixture.
- **Solvent:** The major component of a solution.
- **Solute:** The minor component of a solution.
- **Alloy:** A solid solution of a metal with other substances dissolved in it.
- **Solvation:** The process by which solute particles are surrounded by solvent particles.
- **Hydration:** Solvation by water molecules.
- **Dissociation:** The process of cations and anions of an ionic solute separating when the solute dissolves.
- **Electrolytes:** An ionic compound that dissolves in water.
- **Strong electrolytes:** An ionic compound that ionizes completely when it dissolves.
- **Weak electrolytes:** An ionic compound that does not ionize completely when it dissolves.
- **Nonelectrolytes:** A compound that does not ionize at all when it dissolves.
- **Soluble:** A solute that dissolves in a particular solvent.
- **Insoluble:** A solute that does not dissolve in a particular solvent.
- **Miscible:** Liquids that dissolve in each other.
- **Immiscible:** Liquids that do not dissolve in each other.
- **Solubility:** The limit of how much solute can be dissolved in a given amount of solvent.
- **Saturated:** A solution whose solute is at its solubility limit.
- **Unsaturated:** A solution whose solute is less than its solubility limit.
- **Supersaturated:** A solution that contains more dissolved solute than it can normally hold at a given temperature, making it unstable.
- **Concentration:** How much solute is dissolved in a certain amount of solvent.
- **mass/mass percent concentration, (m/m)%:** A concentration unit that relates the mass of the solute to the mass of the solution.
- **volume/volume percent concentration, (v/v)%:** A concentration unit that relates the volume of the solute to the volume of the solution.
- **mass/volume percent concentration, (m/v)%:** A concentration unit that relates the mass of the solute to the volume of the solution.
- **parts per million (ppm):** The mass of a solute compared to the mass of a solution times 1,000,000.
- **parts per billion (ppb):** The mass of a solute compared to the mass of a solution times 1,000,000,000.
- **Molarity:** Number of moles of solute per liter of solution.
- **Colligative properties:** A characteristic of solutions that depends only on the number of dissolved particles.
- **Vapor pressure:** The pressure of a vapor that is in equilibrium with its liquid phase.
- **Vapor pressure depression:** The lowering of the vapor pressure of a solution versus the pure solvent.
- **Boiling point elevation:** The raising of the boiling point of a solution versus the pure solvent.
- **Freezing point depression:** The lowering of the freezing point of a solution versus the pure solvent.
- **Selective permeation:** Selective permeation is the ability of a membrane to allow certain molecules or ions—typically small or uncharged—to pass through while blocking others based on size, charge, or polarity.
- **Semipermeable membrane:** A semipermeable membrane is a barrier that allows certain particles, usually solvent molecules, to pass through while blocking others, such as larger solute particles.
- **Osmosis:** The process by which solvent molecules can pass through certain membranes but solute particles cannot.
- **Osmotic pressure:** The tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal.
- **Hypertonic:** A hypertonic solution has a higher solute concentration compared to another solution, causing water to move out of cells.
- **Hypotonic:** A hypotonic solution has a lower solute concentration compared to another solution, causing water to enter cells.
- **Isotonic:** An isotonic solution has the same solute concentration as another solution, resulting in no net movement of water across a membrane.
- **Osmolarity:** A way of reporting the total number of particles in a solution to determine the osmotic pressure.

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10.9: Alternative text for Chapter 10

Figure 10.1.1: IV Solutions-Delivering Essential Compounds

A close-up of an intravenous (IV) drip bag in a hospital setting, with a clear fluid-filled bag hanging above a hospital bed. A tube runs from the IV bag down to a patient lying in the bed. The patient, an older adult with gray hair, is wearing a teal hospital gown and is resting under a striped blanket. The focus of the image is on the IV equipment in the foreground, while the patient is slightly blurred in the background, emphasizing the medical treatment being administered.

Figure 10.2.1: Classification of Matter.

This diagram organizes types of matter into categories. At the top is matter. It branches into two main groups: pure substances and mixtures. Pure substances are further divided into elements and compounds. Mixtures are further divided into homogeneous solutions and heterogeneous mixtures. The image presents a hierarchical classification of matter.

Figure 10.2.2: Air-A Mixture of Gases in Solution

A molecular diagram illustrating air as a gaseous solution composed of multiple types of molecules. The image features small, color-coded circles with element labels to represent different gases. Nitrogen (N_2) molecules are shown as pairs of purple circles, oxygen (O_2) as pairs of light blue circles, carbon dioxide (CO_2) as one yellow and two blue circles, and water vapor (H_2O) as two red circles bonded to one blue circle. The diagram emphasizes the uniform distribution of these molecules, representing the homogeneous nature of air as a solution of gases.

Figure 10.3.1: Solvation

The image shows the process of solvation at the molecular level. On the left, there is a photograph of a 200 mL beaker filled with clear liquid. An arrow points to the right, where a diagram illustrates solute particles (represented as green spheres in a lattice structure) being surrounded by water molecules (red and white bent structures representing H_2O). Water molecules are shown interacting with the surface ions of the solute. The oxygen atoms in water (red spheres) face the green solute ions, indicating a negative dipole attraction to the positively charged ions. This visual demonstrates how solvent molecules surround and separate solute particles during dissolution.

Figure 10.3.2: Comparing Conductivity of Electrolyte and Nonelectrolyte Solutions

The image shows three side-by-side beakers with electrical conductivity setups to compare the behavior of different solutions. Each beaker contains a pair of electrodes connected to a power source and a light bulb, which indicates the level of electrical conductivity based on how brightly it glows.

- In the left beaker labeled "ethanol – No conductivity," green molecules are floating in the liquid, but no ions are present. The light bulb remains off, indicating no electrical current.
- In the middle beaker labeled "KCl – High conductivity," green spheres with plus and minus signs (representing K⊕ and Cl⊖ ions) move between the electrodes, and the light bulb glows brightly, showing strong conductivity.
- In the right beaker labeled "acetic acid solution – Low conductivity," fewer ions are shown, with some molecules remaining neutral. The light bulb glows dimly, indicating weak conductivity.

The illustration demonstrates how ionic dissociation affects a solution's ability to conduct electricity.

Figure 10.4.2: Effect of Temperature on Solubility of Solids in Water

A graph showing the solubility of various solid solutes in water (g solute per 100 g H_2O) as a function of temperature (0°C to 100°C). Most solutes—including sugar ($C_{12}H_{22}O_{11}$), KNO_3, $NaNO_3$, NaBr, KBr, KCl, and NaCl—show increasing solubility with temperature. $Ce_2(SO_4)_3$ is the exception, showing decreasing solubility as temperature increases.

Figure 10.4.3: Heat Release from Crystallization of a Supersaturated Solution

Three sequential images showing the activation of a commercial hand warmer containing a supersaturated solution of sodium acetate. In the first image, the clear liquid solution is at rest inside a sealed plastic pouch. In the second image, after a metal disc inside is clicked, white crystalline solids begin to form within the liquid, initiating crystallization. In the third image, the entire pouch has turned opaque white as the sodium acetate has fully crystallized, releasing heat in the process.

Figure 10.4.4: Effect of Temperature on the Solubility of Gases in Water

A line graph showing how the solubility of five different gases—NO (purple), O_2 (green), CO (red), CH_4 (blue), and N_2 (yellow)—changes with temperature in water. The x-axis represents temperature from 0°C to 40°C, and the y-axis represents solubility in mg of gas per 100 g of H_2O. All gases show a decreasing trend in solubility as temperature increases, with NO having the highest solubility and N_2 the lowest across the entire temperature range.

Figure 10.4.5: Impact of Low Oxygen Levels on Aquatic Life

A photograph showing a number of dead fish floating near the icy edge of a body of water. The water is partially covered with ice, suggesting cold weather conditions. The image visually represents a mass die-off of fish, which can result from low levels of dissolved oxygen in the water, often caused by temperature changes or pollution.

Figure 10.4.6: A Model Depicting Why the Solubility of a Gas Increases as the Partial Pressure Increases at Constant Temperature

A three-part diagram illustrating the effect of pressure on the solubility of a gas in a liquid:

- (a) Dynamic Equilibrium: A container with a liquid at the bottom and gas particles above it. Some gas molecules (purple) are entering the liquid (black arrows pointing down), while others are escaping (red arrows pointing up), indicating a state of dynamic equilibrium between dissolved and gaseous particles.
- (b) P_{gas} Increased: The same system with a piston applying pressure to the gas. The gas molecules are now more densely packed in the gas phase, and more black arrows (into the liquid) than red arrows (escaping) are shown, indicating increased dissolution due to higher pressure.
- (c) New Dynamic Equilibrium: After time passes, the system reaches a new dynamic equilibrium with a higher concentration of gas molecules dissolved in the liquid. The number of gas molecules entering and leaving the liquid is balanced again, but at a higher solubility level than in (a).

Figure 10.5.1: Mass/Volume Percent-0.9% Sodium Chloride— A Common Medical Solution

A clear plastic intravenous (IV) fluid bag labeled “Sodium Chloride 0.9% for IV Injection” from TEVA Medical. The bag contains a transparent liquid solution used in medical treatments. The label includes dosage information, expiration date (July 2005), and instructions in English and Hebrew. At the bottom of the bag, there are two ports: one blue and one white/yellow, used for connecting the bag to medical tubing. The bag is partially collapsed, indicating it has been used or is ready for use.

Figure 10.5.2: Steps to Calculate Molarity from Mass of Solute

A three-step flowchart showing how to calculate the molarity of a solution. The first box, labeled "mass of solute", leads to the second box, "moles of solute", via an arrow labeled "molar mass", with a conversion shown below: 1 mole HCl / 35.46 g HCl. The second box leads to the third box, "Molarity of solution", via an arrow labeled "Molarity", with the formula moles HCl / liters solution. The chart illustrates the need to convert mass to moles before applying the molarity formula.

Figure 10.5.3: Using Molarity as a Conversion Factor

A two-box flowchart illustrating the use of molarity as a conversion factor. The left box (purple) is labeled "Liters solution". The right box (orange) is labeled "Moles solute". A double-headed arrow between the two boxes is labeled "Molarity Conversion factor", indicating that molarity can be used to convert from volume to amount of solute or vice versa. The diagram represents how the relationship between moles of solute and volume of solution can be used in both directions for solution calculations.

Figure 10.5.4: Calculating Mass of Solute from Solution Volume and Molarity

A three-step flowchart showing how to calculate the mass of solute from a known volume of solution using molarity and molar mass. The first box (purple) is labeled "Liters of solution". An arrow labeled "Molarity conversion factor" points to the second box (orange) labeled "Moles of solute", with a conversion shown below: 6.00 mol $C_2H_6O_2$ / 1 L solution. A second arrow labeled "Molar mass" leads to the third box (red) labeled "Mass of solute", with the conversion 62.08 g $C_2H_6O_2$ / 1 mol $C_2H_6O_2$. This diagram demonstrates the step-by step process of using molarity and molar mass to find the mass of ethylene glycol in a given volume of solution.

Figure 10.5.5: Stoichiometry Flowchart: Mass of NaOH to Volume of HCl

The image shows a vertical flow diagram representing the steps to calculate the volume of hydrochloric acid (HCl) needed to react with a known mass of sodium hydroxide (NaOH). It begins with a red box stating “Start with mass of NaOH (the 'known quantity'),” followed by a purple box labeled “Moles of NaOH,” which is calculated using molar mass as a conversion factor. Next is a blue box labeled “Moles of HCl,” determined by using the balanced chemical equation to find the appropriate molar ratio. The final step is shown in a green box labeled “Volume of HCl needed,” found using the definition of molarity. Arrows connect each step, showing the flow from mass to volume through mole-based stoichiometry.

Figure 10.5.6: Converting Mass of NaOH to Volume of HCl in Milliliters

The image shows a multi-step calculation process for determining the volume of HCl solution needed to react with a given mass of NaOH. It begins with a red box labeled "mass of NaOH," which is converted to moles using the molar mass of NaOH (1 mol NaOH / 40.00 g NaOH). The next step is a purple box labeled "moles of NaOH." Using the stoichiometric ratio from the balanced chemical equation (1 mol HCl / 1 mol NaOH), the value is converted into "moles of HCl," shown in a blue box. The next green box, "liters of solution," is calculated using molarity as a conversion factor (1 L HCl solution / 2.75 mol HCl). Finally, the volume in liters is converted to milliliters in the orange box labeled "mL of solution" using the factor 1,000 mL / 1 liter. Arrows and conversion factors between boxes guide the sequence from mass to volume in mL.

Figure 10.5.7: Relating Mass of One Substance to the Volume of Another Using Molarity

The image shows a vertical, five-step diagram outlining how to calculate the volume of a solution required for a chemical reaction, starting from the mass of a different substance. At the top is a red box labeled “Mass of one substance,” which connects to a purple box labeled “Moles of one substance” using molar mass as a conversion factor. The purple box leads to a blue box labeled “Moles of a different substance,” with the conversion based on the coefficients from a balanced chemical equation. Finally, the blue box connects to a green box labeled “Volume of solution needed,” using the definition of molarity to complete the calculation. Double-headed arrows between each step indicate the bidirectional nature of these relationships.

Figure 10.5.8: Effect of Dilution on Solution Concentration

The image shows two beakers side by side, labeled "before" and "after adding solvent," demonstrating the effect of dilution. The left beaker represents the original solution, which has a lower volume and higher concentration, illustrated by a dense grouping of red triangle-shaped solute particles in a zoomed-in circle. The right beaker shows the same solution after more solvent has been added, resulting in a higher total volume. The zoomed-in view of this diluted solution displays solute particles more spread out, indicating a lower concentration. The total amount of solute remains unchanged, but the concentration decreases due to the increased volume.

Figure 10.6.1: Effect of Solute on Vapor Pressure

The image shows two side-by-side diagrams illustrating vapor pressure depression. The left panel depicts a pure solvent, with a beaker containing only blue spheres (representing solvent particles) and many of the same particles escaping into the space above as vapor. The right panel shows a solution where orange spheres (representing nonvolatile solute particles) are mixed among the blue solvent particles in the beaker. Fewer solvent particles are seen in the vapor phase above the solution. This demonstrates that the presence of solute particles reduces the number of solvent molecules at the surface available to evaporate, resulting in a lower vapor pressure.

Figure 10.6.2: Boiling Point Elevation and Freezing Point Depression in Solutions

A diagram shows a vertical thermometer in the center, comparing phase changes for a pure solvent and a solution. On both sides, three horizontal regions represent the solid, liquid, and gas phases. On the right, labeled for the pure solvent (water), the normal freezing point is shown at 0°C and the normal boiling point at 100°C. On the left, for the solution, the freezing point is shown lower than 0°C and the boiling point higher than 100°C. The red fluid level in the thermometer spans from below the normal freezing point to above the normal boiling point, illustrating that solutes cause freezing point depression and boiling point elevation in solutions.

Figure 10.6.3: Real-Life Application of Freezing Point Depression

The photo shows a close-up of a red and gray brick sidewalk with white salt pellets scattered across the surface. The salt particles are irregularly distributed along the grooves and tops of the bricks. This demonstrates a practical use of freezing point depression, where salt is applied to outdoor surfaces in cold weather to lower the freezing point of water, thereby melting existing ice or preventing its formation. The contrast between the reddish-brown bricks and white salt highlights the visibility and widespread distribution of the salt crystals used for de-icing.

Figure 10.6.4: Osmosis and the Development of Osmotic Pressure

The diagram shows a U-shaped tube divided into two halves by a semipermeable membrane at the bottom center. The left side of each image contains pure solvent (yellow particles), and the right side contains a solution (a mix of yellow solvent particles and larger purple solute particles). In part (a), the solvent begins to move through the membrane from the pure solvent side to the solution side due to osmosis. In part (b), the movement of solvent continues until a height difference forms between the two sides. The right side, containing the solution, rises higher than the left. A red arrow labeled “Osmotic pressure” shows the pressure difference caused by the net flow of solvent into the solution side. Solute particles cannot pass through the membrane, so only the solvent moves, resulting in the buildup of osmotic pressure.

Figure 10.6.5: Effect of Osmotic Conditions on Red Blood Cells

The illustration shows three panels comparing the behavior of red blood cells in hypertonic, isotonic, and hypotonic solutions. In the hypertonic solution on the left, red blood cells appear shriveled due to water exiting the cells, as indicated by green arrows pointing out. In the center panel, red blood cells in an isotonic solution maintain their normal shape, with water moving in and out equally, represented by yellow arrows in both directions. On the right, red blood cells in a hypotonic solution appear swollen, as water flows into the cells, shown by yellow arrows pointing inward. The diagram highlights how osmotic pressure affects cell

volume depending on the solute concentration of the surrounding solution.

Figure 10.6.6: Dialysis in Medical Treatment

A patient undergoing dialysis treatment. The image shows a close-up of a person lying on a hospital bed with a dialysis machine beside them. The patient's arm is connected to the machine via tubes, with visible taped insertion points. The dialysis machine features a digital control panel, tubing for blood flow, and filters used to clean the blood. The setup illustrates a real-world medical application of selective permeation and semipermeable membranes, where waste products and excess fluids are removed from the blood.

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11: ACIDS AND BASES

Introduction

Many of us are familiar with the group of chemicals called acids. But do you know what it takes for a compound to be an acid? Actually, there are several different definitions of acid that chemistry uses, and each definition is appropriate under different circumstances. Less familiar—but just as important to chemistry and ultimately to us—is the group of chemicals known as bases. Both acids and bases are important enough that we devote an entire chapter to them—their properties and their reactions.

Learning Objectives

- To describe characteristics of acids and bases.
- To define and recognize the differences between Arrhenius acids and bases and Brønsted-Lowry acid or bases.
- To write equations of neutralization reactions by illustrating the proton transfer process that defines a Brønsted-Lowry acid-base reaction.
- Describe the difference between strong and weak acids and bases by exploring how a chemical reaction reaches chemical equilibrium.
- Define the pH scale and use it to describe acids and bases.
- To write chemical equations for water acting as an acid and as a base.
- To understand acid-base reactions.
- To define buffers and describe how buffers react with an acid or a base.

11.1: Prelude to Acids and Bases

11.2: Definitions of Acids and Bases

11.3: Water - Both an Acid and a Base

11.4: Strength of Acids and Bases

11.5: Dissociation of Water

11.6: Acid Base Reactions

11.7: Buffers

11.8: Summary

11.9: Key Terms

11.10: Alternative Text for Chapter 11

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11.1: Prelude to Acids and Bases

Introduction

Although a minimal amount of carbohydrate digestion occurs in the mouth, chemical digestion really gets underway in the stomach. The stomach participates in almost all digestive activities. It vigorously churns food and it secretes gastric juices that break down food and absorbs certain drugs, including aspirin and some alcohol. The stomach begins the digestion of protein and continues the digestion of carbohydrates and fats. It stores food as an acidic liquid called chyme, and releases it gradually into the small intestine through the pyloric sphincter.

One of the most concentrated acids in the body is stomach acid, which can be approximated as a 0.05 M hydrochloric acid solution. Special cells in the stomach wall secrete this acid, along with special enzymes, as part of the digestion process (Figure 11.1.1).

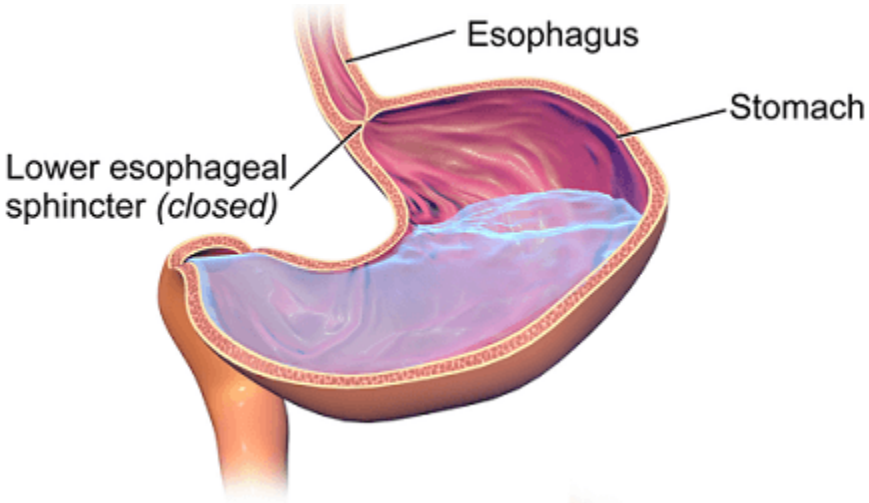


Figure 11.1.1: Stomach Acid and the Lower Esophageal Sphincter. The stomach contains strong hydrochloric acid (HCl) to aid digestion, while the lower esophageal sphincter prevents this acid from refluxing into the esophagus. (CC BY-SA 4.0; anonymous via The Language of Medical Terminology). Detailed Description of Figure 11.1.1.

In a laboratory, a 0.05 M solution of hydrochloric acid would dissolve some metals. How does the stomach survive the presence of such a reactive acid?

Actually, the stomach has several mechanisms for withstanding this chemical onslaught. First, the lining of the stomach is coated with a thin layer of mucus that contains some bicarbonate ions (HCO_3^-). These react with the hydrochloric acid to produce water, carbon dioxide, and harmless chloride ions. If any acid penetrates through the mucus, it can attack the surface layer of stomach cells, called the gastric epithelium. Cells in the gastric epithelium are being constantly shed, so damaged cells are quickly removed and replaced with healthy cells.

However, if the gastric epithelium is destroyed faster than it can be replaced, the acid may reach the wall of the stomach, resulting in ulcers. If an ulcer grows large enough, it can expose blood vessels in the stomach wall, causing bleeding. In extreme situations, the loss of blood through a severe ulcer can threaten a person's health.

Ulcers can also result from the presence of a certain bacterium—Helicobacter pylori—in the stomach. The mechanism for this ulcer formation is not the same as that for ulcers caused by stomach acid and is not completely understood. However, there are two main treatments for ulcers: (1) antacids to react chemically with excess hydrochloric acid in the stomach and (2) antibiotics to destroy the H. pylori bacteria in the stomach.

In this chapter, we will explore the fundamental concepts and chemical behavior of acids and bases.

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11.2: Definitions of Acids and Bases

General Properties of Acids and Bases

Acids and **bases** are two fundamental classes of compounds that play essential roles in both chemistry and biology, from digestion in the human body to reactions in laboratory experiments. One way to define a class of compounds is by describing the various characteristics its members have in common. In the case of the compounds known as acids, the common characteristics include a sour taste, the ability to change the color of the vegetable dye **litmus** to red, and the ability to dissolve certain metals and simultaneously produce hydrogen gas. For the compounds called bases, the common characteristics are a slippery texture, a bitter taste, and the ability to change the color of litmus to blue. Acids and bases also react with each other to form compounds generally known as **salts**. In this section we are going to explore and learn about the properties of acids and bases as well as various definitions.

We commonly encounter acids and bases in our foods –some foods are acidic, and others are basic (alkaline) as illustrated in Figure 11.2.1. To better understand how acids and bases behave in everyday contexts and chemical reactions, it is helpful to first examine their general physical and chemical properties.

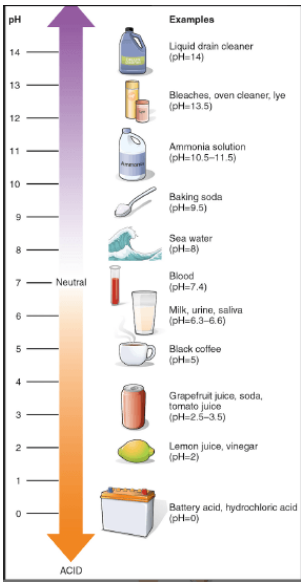


Figure 11.2.1: The pH scale ranges from 0 (most acidic) to 14 (most basic), showing everyday examples such as battery acid, lemon juice, coffee, blood, seawater, baking soda, ammonia, and drain cleaner. (CC BY 3.0, OpenStax College via Wikimedia Commons). Detailed Description of Figure 11.2.1.

The general properties of acids and bases are the following:

1. Acids taste sour, e.g., citrus fruits taste sour because of citrus acid and ascorbic acid, i.e., vitamin C, in them. Basic (alkaline) substances, on the other hand, taste bitter.
2. Basic (alkaline) substances feel soapy, while acidic substances may sting.
3. Acids turn blue litmus paper to red but do not change the color of red litmus paper. Bases turn red litmus paper blue but do not change the color of blue litmus paper, as illustrated in Figure 11.2.2.
4. Acids and bases neutralize each other.
5. Hydrochloric acid is found in the stomach that helps digestion. Excess hydrochloric acid may cause acid burns—antacids like milk of magnesia are bases that help by neutralizing excess acid in the stomach.

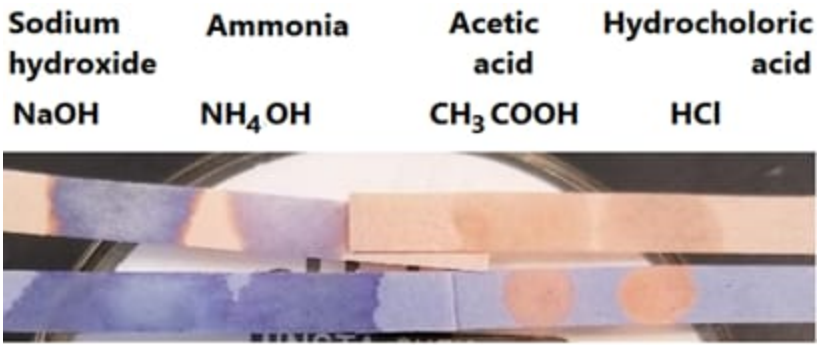


Figure 11.2.2: Demonstration of acids turn red litmus paper blue and bases turn blue litmus paper red - solution listed at the top of the image was spotted on the red litmus paper (top) or blue litmus paper (bottom). (Public Domain, Muhammad Arif Malik via LibreText Chemistry). Detailed Description of Figure 11.2.2.

Arrhenius's Definition of Acids and Bases

Instead of relying on sensory characteristics, however, chemists define acids and bases in chemical terms. The Swedish chemist Svante Arrhenius developed the first chemical definitions of acids and bases in the late 1800s. Arrhenius defined an acid as a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution. Similarly, Arrhenius defined a base as a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.

In essence, Arrhenius's definition states that:

- An acid is a substance that generates hydrogen ions (H^+) when dissolved in water, and
- A base is a substance that generates hydroxide ions (OH^-) when dissolved in water.

For example, hydrochloric acid (HCl) is an acid because it forms H^+ when it dissolves in water.

$$HCl(g) \overset{Water}{\longrightarrow} H^+(aq) + Cl^-(aq)$$

Similarly, sodium hydroxide (NaOH), is a base because it forms OH^- when it dissolves in water.

$$NaOH(s) \overset{Water}{\longrightarrow} Na^+(aq) + OH^-(aq)$$

Note that hydrogen ion H^+ does not exist in reality. It bonds with water molecules and exists as an ion named the **hydronium ion** $H_3O^+(aq)$.

$$H^+(aq) + H_2O \rightarrow H_3O^+(aq)$$

However, $H^+(aq)$ is often written interchangeably with H_3O^+ (aq).

Naming Arrhenius Acids and Bases

Naming Arrhenius acids

Aqueous solutions of acids have their own naming rules:

1. The names of **binary acids** (compounds with hydrogen and one other element in their formula) are based on the root of the name of the other element preceded by the prefix *hydro-* and followed by the suffix *-ic acid*. Thus, an aqueous solution of HCl [designated "HCl (aq)"] is called hydrochloric acid, H_2S (aq) is called hydrosulfuric acid, and so forth.
2. Acids composed of more than two elements (typically hydrogen and oxygen and some other element) have names based on the name of the other element, followed by the suffix *-ic acid* or *-ous acid*, depending on the number of oxygen atoms in the

acid's formula.

- If the anion is an **oxyanion** with the last syllable *–ate*, change the last syllable with *–ic*. Do not use the prefix hydro-, but add the last word "acid." If there is a prefix *per-* in the name of the oxyanion, retain the prefix in the acid name.
    - For example, NO_3^- is a nitrate, and HNO_3 is nitric acid.
    - Another example, ClO_4^- is a perchlorate, and $HClO_3$ is perchloric acid.
- If the anion is an oxyanion with the last syllable *–ite*, change the last syllable with *–ous*. Do not use the prefix hydro-, but add the last word "acid." If there is a prefix *hypo-* in the name of the oxyanion, retain the prefix in the acid name.
    - For example, NO_2^- is nitrite, and HNO_2 is nitrous acid.
    - Another example, ClO^- is hypochlorite, and HClO is hypochlorous acid.

Unfortunately, there is no strict rule for the number of oxygen atoms that are associated with the *-ic acid* suffix; *the names of these acids are best memorized*. Table 11.2.1 lists the names and formulas of some of the common acids and their anions.

Note that acids have hydrogen written first, as if it were the cation, while most bases have the negative hydroxide ion, if it appears in the formula, written last.

Table 11.2.1: Names of some common acids and their anions.

Acid formula	Acid name	Anion	Anion name
HCl	Hydrochloric acid	Cl^-	Chloride
HBr	Hydrobromic acid	Br^-	Bromide
HI	Hydroiodic acid	I^-	Iodide
HCN	Hydrocyanic acid	CN^-	Cyanide
HNO_3	Nitric acid	NO_3^-	Nitrate
HNO_2	Nitrous acid	NO_2^-	Nitrite
H_2SO_4	Sulfuric acid	SO_4^{2-}	Sulfate
H_2SO_3	Sulfurous acid	SO_3^{2-}	Sulfite
H_2CO_3	Carbonic acid	CO_3^-	Carbonate
CH_3COOH	Acetic acid	CH_3COO^-	Acetate
H_3PO_4	Phosphoric acid	PO_4^{3-}	Phosphate
H_3PO_3	Phosphorous acid	PO_3^{3-}	Phosphite
$HClO_4$	Perchloric acid	ClO_4^-	Perchlorate
$HClO_3$	Chloric acid	ClO_3^-	Chlorate
$HClO_2$	Chlorous acid	ClO_2^-	Chlorite

Acid formula	Acid name	Anion	Anion name
HClO	Hypoclorous acid	ClO^-	Hypochlorite

Naming Arrhenius Bases

Many bases and their aqueous solutions are named using the normal rules of ionic compounds that were presented previously; that is, they are named as hydroxide compounds. The Arrhenius bases are ionic compounds of metal and hydroxide ion, and their name starts with the name of the metal element followed by the name of the anion, i.e., hydroxide. For example, the base NaOH is named sodium hydroxide.

Table 11.2.2 lists the names and formulas of some of the common Arrhenius bases.

Table 11.2.2: Names of some of the common Arrhenius bases

Formula	Name
LiOH	Lithium hydroxide
NaOH	Sodium hydroxide
KOH	Potassium hydroxide
$Ca(OH)_2$	Calcium hydroxide
$Sr(OH)_2$	Strontium hydroxide
$Ba(OH)_2$	Barium hydroxide

Example 11.2.1

Name each substance.

a. HF(aq)
b. $Sr(OH)_2$(aq)

Solution

a. This acid has only two elements in its formula, so its name includes the *hydro-* prefix. The stem of the other element's name, fluorine, is *fluor*, and we must also include the *-ic acid* ending. Its name is hydrofluoric acid.
b. This base is named as an ionic compound between the strontium ion and the hydroxide ion: strontium hydroxide.

Brønsted–Lowry's Definition of Acids and Bases

Some bases do not have hydroxide ions in their formula, yet they act as bases and **neutralize** acids. For example, ammonia (NH_3) and calcium carbonate ($CaCO_3$) do not contain hydroxide ions, but they neutralize acids. In fact, the Arrhenius definitions of an acid and a base focus on hydrogen ions and hydroxide ions. Further, Arrhenius's definition limits the acid-base reactions in the water medium. The acid-base reactions can take place in other mediums also, e.g., HCl –an acid, and NH_3 –a base can react with and neutralize each other in the gas phase also.

How can acids and bases be defined more fundamentally?

In 1923, the Danish scientist Johannes Brønsted and the English scientist Thomas Lowry independently proposed new definitions for acids and bases. Rather than considering both hydrogen and hydroxide ions, they focused on the hydrogen ion only. A Brønsted-Lowry acid is a compound that supplies a hydrogen ion in a reaction. A Brønsted-Lowry base, conversely, is a compound that accepts a hydrogen ion in a reaction. Thus, the Brønsted-Lowry definitions of an acid and a base focus on the movement of hydrogen ions in a reaction, rather than on the production of hydrogen ions and hydroxide ions in an aqueous solution.

Brønsted–Lowry's definition states that:

1. An acid is a **proton** donor, and
2. A base is a proton acceptor.

HCl is a Brønsted–Lowry acid because it donates a proton to the water solvent. Water is a base in the reaction represented below, because it accepts a proton from the acid.

$$\mathrm{HCl}(g) + \mathrm{H_2O}(l) \rightarrow \mathrm{Cl^-}(aq) + \mathrm{H_3O^+}(aq)$$

In a reaction between HCl and NH_3, NH_3 is the base, for the same reason that it accepts a proton from the Brønsted–Lowry acid, HCl.

$$\mathrm{HCl(g) + NH_3(\ g) \rightarrow NH_4Cl(s)}$$

Brønsted–Lowry's acids have **ionizable** protons that they donate to bases. Therefore, Brønsted–Lowry's acid is generally written as HA, where H^+ is the donatable proton, and A^- is the anion of the acid. Examples of acids are HCl, H_2SO_4, HNO_3, and CH_3COOH. Note that acetic acid has only one acidic proton that is attached to the O atom in the carboxylic acid group (–COOH). The rest of the protons attached to carbon atoms are not acidic. All organic acids have a carboxylic acid group (–COOH).

Brønsted–Lowry's acid may have net (+) charge, no charge, or net (–) charge on it. For example, H_3O^+, HCl, and HSO_4^- are all acids because they can donate a proton to a base.

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:

$$\mathrm{NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)}$$

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 as seen in Figure 11.2.3:

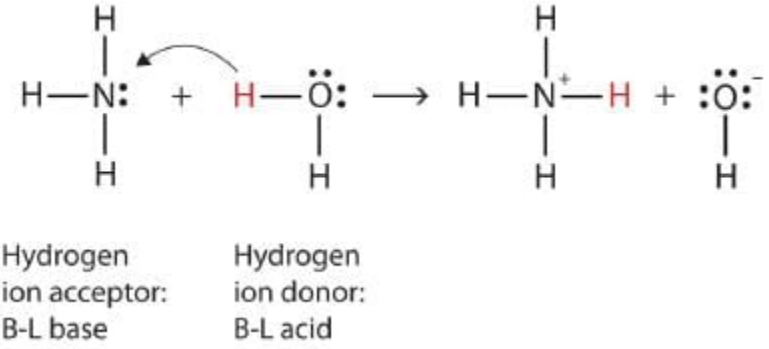


Figure 11.2.3: Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense. (CC BY-NC-SA 3.0; Andy Schmitz via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 11.2.3.

Let's look at an Arrhenius acid like hydrochloric acid closely to show that it is still an acid in the Brønsted-Lowry sense.

This 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, as we mentioned above, 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.

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+. 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 as shown below.

$$\mathrm{HCl(g) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)}$$

We can illustrate this process using Lewis electron dot diagrams as seen in Figure 11.2.4:

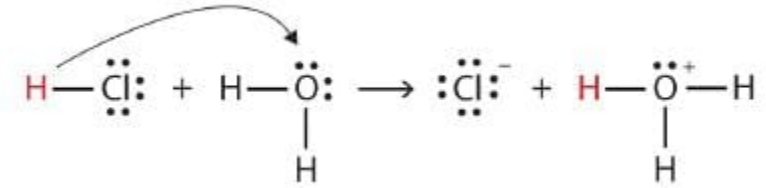


Figure 11.2.4: A hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. (CC BY-NC-SA 3.0; Andy Schmitz via Introduction to Chemistry: General, Organic, and Biological (v. 1.0)). Detailed Description of Figure 11.2.4.

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.

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 11.2.2

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:

$$\mathrm{C_6H_5NH_2(aq) + H_2O(\ell) \rightarrow C_6H_5NH_3^+(aq) + OH^-(aq)}$$

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.

The Brønsted-Lowry definitions of an acid and a base can be applied to chemical reactions that occur in **solvents** other than water as well. The following Example 11.2.3 illustrates this.

Example 11.2.3

Sodium amide ($NaNH_2$) dissolves in methanol (CH_3OH) and separates into sodium ions and amide ions (NH_2^-). The amide ions react with methanol to make ammonia and the methoxide ion (CH_3O^-). Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

Solution

The equation for the reaction is between NH_2^- and CH_3OH to make NH_3 and CH_3O^- is as follows:

$$NH_2^-(solv) + CH_3OH(\ell) \rightarrow NH_3(solv) + CH_3O^-(solv)$$

The label (solv) indicates that the species are dissolved in some solvent, in contrast to *(aq)*, which specifies an aqueous (H_2O) solution. In this reaction, we see that the NH_2^- ion accepts a proton from a CH_3OH molecule to make an NH_3 molecule. Thus, as the proton acceptor, NH_2^- is the Brønsted-Lowry base. As the proton donor, CH_3OH is the Brønsted-Lowry acid.

Brønsted-Lowry Acid-Base Reactions in Pharmaceuticals

There are many interesting applications of Brønsted-Lowry acid-base reactions in the pharmaceutical industry. For example, drugs often need to be water soluble for maximum effectiveness. However, many complex organic compounds are not soluble or are only slightly soluble in water. Fortunately, those drugs that contain proton-accepting nitrogen atoms (and there are a lot of them) can be reacted with dilute hydrochloric acid [HCl (aq)]. The nitrogen atoms—acting as Brønsted-Lowry bases—accept the hydrogen ions from the acid to make an ion, which is usually much more soluble in water. The modified drug molecules can then be isolated as chloride salts:

\[\mathrm{RN(sl\: aq) + H^+(aq) \rightarrow RNH^+(aq) \xrightarrow{Cl^-(aq)} RNHCl(s)} \label{Eq3} \]

where RN represents some organic compound containing nitrogen. The label *(sl aq)* means “slightly aqueous,” indicating that the compound RN is only slightly soluble. Drugs that are modified in this way are called *hydrochloride salts*. Examples include the powerful painkiller codeine, which is commonly administered as codeine hydrochloride. Acids other than hydrochloric acid are also used. Hydrobromic acid, for example, gives *hydrobromide salts*. Dextromethorphan, an ingredient in many cough medicines, is dispensed as dextromethorphan hydrobromide. The accompanying figure shows another medication (lidocaine) as a hydrochloride salt (Figure 11.2.5).

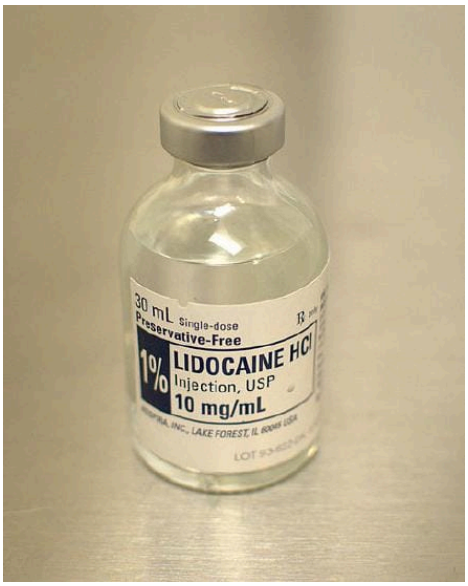


Figure 11.2.5: Lidocaine is a hydrochloride salt, one of the most commonly used local anaesthetics in dentistry among other things. (CC BY-SA 2.0; JL Johnson via Wikimedia Commons). Detailed Description of Figure 11.2.5.

Mono-, di- and tri-protic acids

Acids that have only one acidic proton are **mono-protic**, e.g., HCl, HNO_3, CH_3COOH, are mono-protic acids where the acid proton is shown in bold font. Some acids have two acidic protons –they are **di-protic**, e.g., H_2SO_4 and H_2CO_3 are di-protic acids. The acids with three acidic protons are tri-protic, e.g., H_3PO_3 is a **tri-protic** acid. For example, phosphoric acid (H_3PO_4) can dissociate and donate three protons, as shown in chemical reactions below:

$$\begin{aligned} &H_3PO_4 + H_2O \rightleftharpoons H_3O^+ + H_2PO_4^- \\ &H_2PO_4^- + H_2O \rightleftharpoons H_3O^+ + HPO_4^{2-} \\ &HPO_4^{2-} + H_2O \rightleftharpoons H_3O^+ + PO_4^{3-} \end{aligned} \tag{11.2.1}$$

The base accepts proton by making a bond with it. The bond is a pair of bonded electrons. Since the proton is a hydrogen atom without an electron, both electrons in the bond come from the base. The base must have a lone pair of electrons on it. The base is usually represented as **:**B to emphasize a lone pair of electrons on it that is shown as a pair of dots.

Conjugate acid-base pairs

We have defined, according to the Bronsted-Lowry theory of acids and bases, that an acid is a proton donor and a base is a proton acceptor. Once an acid has given up a proton, the part that remains is called the acid's **conjugate base**. This species is a base because it can accept a proton (to re-form the acid). The conjugate base of HF (first example below) is fluoride ion, F^-.

$$\text{Acid} = H^+ + \text{Conjugate base of Acid}^-$$

$$HF \rightleftharpoons H^+ + F^-$$

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$NH_4^+ \rightleftharpoons H^+ + NH_3$$

Similarly, the part of the base that remains after a base accepts a proton is called the base's **conjugate acid**. This species is an acid because it can give up a proton (and thus re-form the base). The conjugate acid of fluoride ion, F^- (first example below) is HF.

$$H^+ + \text{Base} = \text{Conjugate acid of Base}^+$$

$$H^+ + F^- \rightleftharpoons HF$$

$$H^+ + OH^- \rightleftharpoons H_2O$$

$$H^+ + H_2O \rightleftharpoons H_3O^+$$

$$H^+ + NH_3 \rightleftharpoons NH_4^+$$

To summarize, the conjugate base of HF is fluoride ion, F^-, and the conjugate acid of fluoride ion, F^-, is HF. The HF/F- pair is referred to as a **conjugate acid-base pair**. The difference in the formulas of a conjugate acid-base pair (example: HF and F^-) is H^+. The table below, Table 11.2.3, lists conjugate acid-base pairs for your reference so that you can figure out the strategy of identifying them. For any given acid or base, you should be able to give its conjugate base or conjugate acid. The formula of an acid's conjugate base is generated by removing a proton (H^+) from the acid formula. The formula of the base's conjugate acid is formed by adding a proton (H^+) to the formula of the base.

Table 11.2.3. Conjugate acid-base pairs.

Conjugate Acid	Conjugate Base
H_3O^+	H_2O
H_2O	OH^-

Conjugate Acid	Conjugate Base
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
NH_4^+	NH_3
NH_3	NH_2^-
CH_3COOH	CH_3COO^-
$CH_3NH_3^+$	CH_3NH_2

Example 11.2.4 : Conjugate Pairs

Write the formula of the conjugate base of (a) HCl and (b) HCO_3^-.

Write the formula of the conjugate acid of (c) CH_3NH_2 and (d) OH^-.

Solution:

A conjugate base is formed by removing a proton (H^+). A conjugate acid is formed by accepting a proton (H^+).

a. After HCl donates a proton, a Cl^- ion is produced, and so Cl^- is the conjugate base.
b. After hydrogen carbonate ion, HCO_3^-, donates a proton, its conjugate base, CO_3^{2-} is produced.
c. After accepting a proton (H^+), CH_3NH_2 is converted to $CH_3NH_3^+$, its conjugate acid.
d. After accepting a proton (H^+), OH^- is converted to H_2O, its conjugate acid.

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,(Figure 11.2.6).

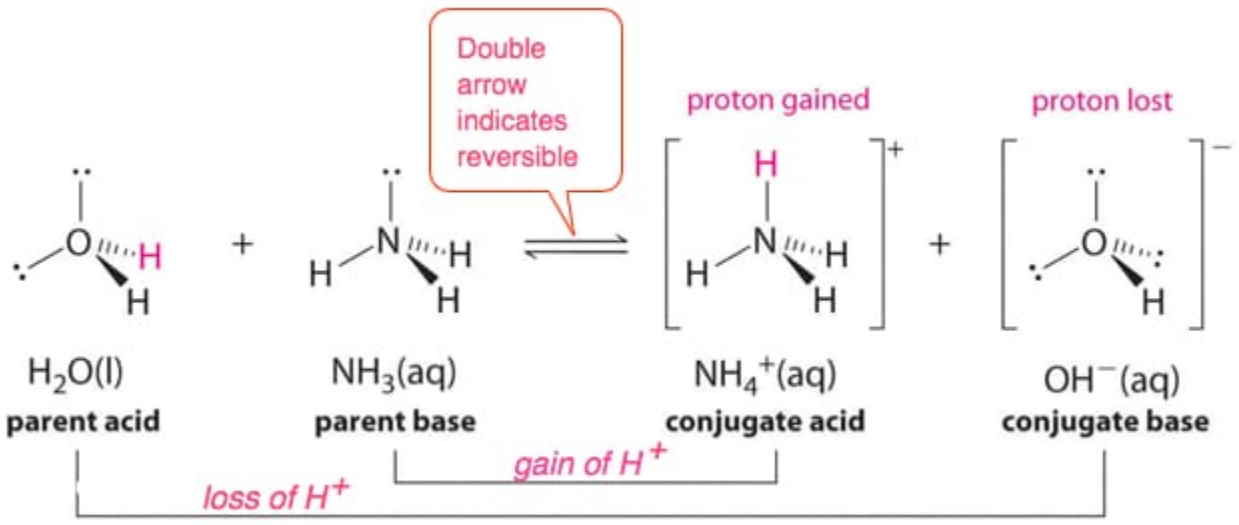


Figure 11.2.6: In the forward reaction, the parent acid is H_2O and and the parent base is NH_3 (shown in the illustration below). The acid H_2O loses a proton (H^+) to form its conjugate base OH^-. The base NH_3 gains a proton, to produce its conjugate acid NH_4^+. In the reverse reaction, the acid NH_4^+ loses a proton (H^+) to form its conjugate base NH_3. The base OH^- gains a proton, to produce its conjugate acid H_2O. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.2.6.

The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-. This means that the conjugate acid of the base

NH_3 is NH_4^+ while the conjugate base of the acid NH_4^+ is NH_3. Similarly, the conjugate base of the acid H_2O is OH^-, and the conjugate acid of the base OH^- is H_2O (Figure 11.2.7).

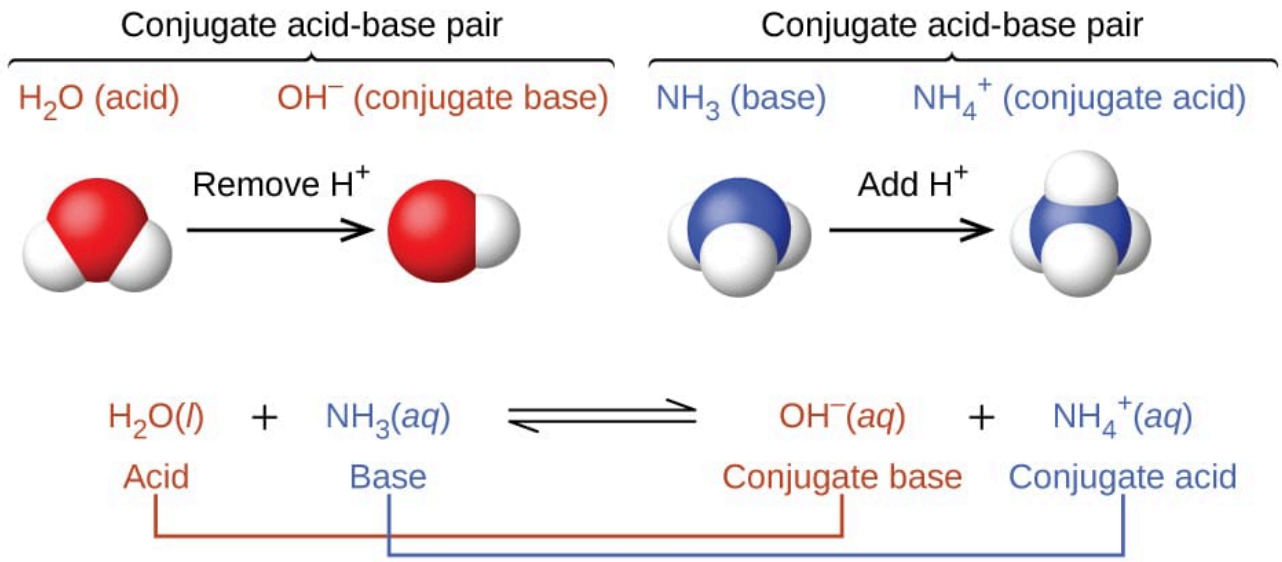


Figure 11.2.7: Conjugate Acid-base Pairs of the Reaction of Ammonia with Water. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.2.7.

Let's look at another example, Figure 11.2.8. When hydrogen fluoride (HF) dissolves in water and ionizes, protons are transferred from hydrogen fluoride (parent acid) molecules to water (parent base) molecules, yielding hydronium ions (conjugate acid of water) and fluoride ions (conjugate base of HF):

$$\mathrm{HF} + H_2O \rightleftharpoons H_3O^+ + \mathrm{F}^-$$

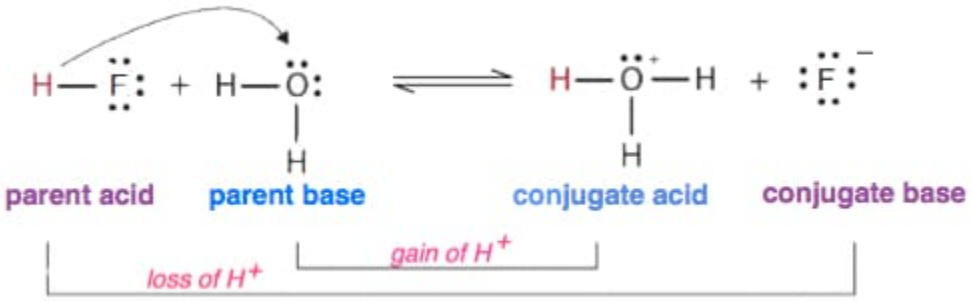


Figure 11.2.8: Dissociation of HF in Water. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.2.8.

Example 11.2.5

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 (Figure 11.2.9), 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).

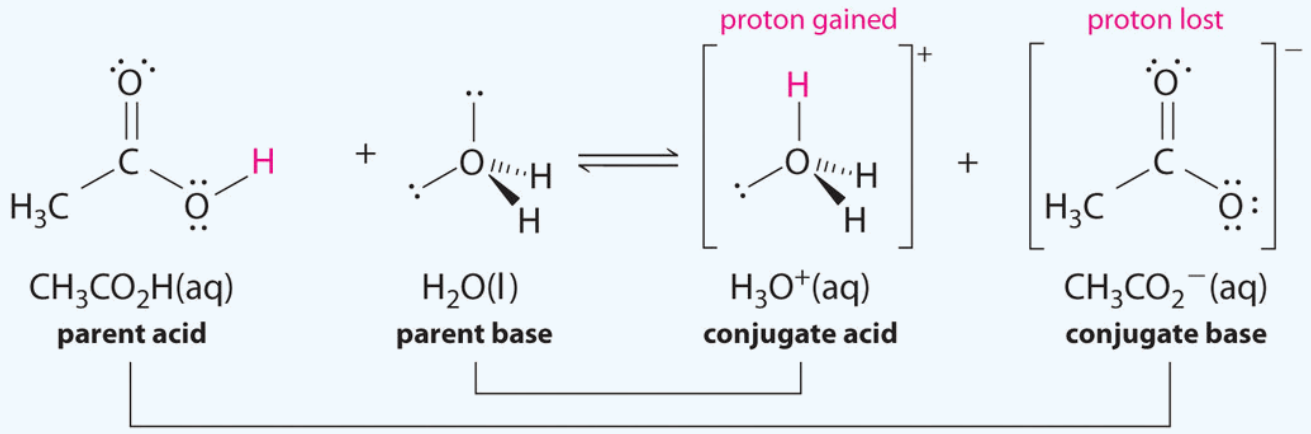


Figure 11.2.9: Reaction of acetic acid with water (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Image.

Example 11.2.6

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H_2O and OH^-, where H_2O 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 $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

The use of conjugate acid-base pairs allows us to make a very simple statement about relative strengths of acids and bases.

> The stronger an acid, the weaker its conjugate base, and, conversely, the stronger a base, the weaker its conjugate acid.

We will take a deeper look at acid-base strength in the next section.

Key Takeaways

- An Arrhenius acid increases the H^+ ion concentration in water, while an Arrhenius base increases the OH^- ion concentration in water.
- A Brønsted-Lowry acid is a proton donor, and a Brønsted-Lowry base is a proton acceptor.
- Brønsted-Lowry acid-base reactions are essentially proton transfer reactions.
- The stronger an acid, the weaker its conjugate base, and, conversely, the stronger a base, the weaker its conjugate acid.

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11.3: Water - Both an Acid and a Base

Amphoteric Substances

Water (H_2O) is an interesting compound in many respects. Here, we will consider its ability to behave as an acid or a base.

Water acts as a base in some reactions, e.g., with HF, and as an acid in some reactions, e.g., with NH_3.

> Substances like water that can act as an acid and also as a base are called amphoteric substances.

Other examples of **amphoteric** substances include HSO_4^-, HCO_3^-, and NH_3, as illustrated in Figure 11.3.1.

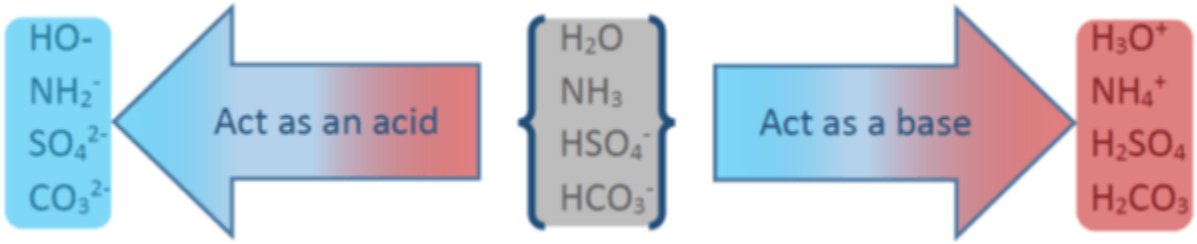


Figure 11.3.1: They can act as an acid in some reactions and also as a base in other reactions. (Public Domain, Muhammad Arif Malik via LibreText Chemistry). Detailed Description of Figure 11.3.1.

Water: Both an Acid and a Base

In some circumstances, a water molecule will accept a proton and thus act as a Brønsted-Lowry base. We saw an example in the dissolving of HCl in H_2O:

$$\mathrm{HCl(g) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)}$$

In other circumstances, a water molecule can donate a proton and thus act as a Brønsted-Lowry acid. For example, in the presence of the amide ion, a water molecule donates a proton, making ammonia as a product:

$$\mathrm{H_2O(l) + NH_2^-(aq) \rightarrow OH^-(aq) + NH_3(aq)}$$

In this case, NH_2^- is a Brønsted-Lowry base (the proton acceptor).

So, depending on the circumstances, H_2O can act as either a Brønsted-Lowry acid or a Brønsted-Lowry base. Water is not the only substance that can react as an acid in some cases or a base in others, but it is certainly the most common example—and the most important one. A substance that can either donate or accept a proton, depending on the circumstances, is called an **amphiprotic** compound**.**

Autoionization of Water

A water molecule can act as an acid or a base even in a sample of pure water. About 6 in every 100 million (6 in 10^8) water molecules undergo the following reaction:

$$\mathrm{H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)}$$

This process is called the **autoionization of water** (Figure 11.3.2) and occurs in every sample of water, whether it is pure or part of a solution. Autoionization occurs to some extent in any amphiprotic liquid. (For comparison, liquid ammonia undergoes

autoionization as well, but only about 1 molecule in a million billion (1 in 10^{15}) reacts with another ammonia molecule.)

It is rare to truly have pure water. Water exposed to air will usually be slightly acidic because dissolved carbon dioxide gas, or carbonic acid, decreases the pH slightly below 7. Alternatively, dissolved minerals, like calcium carbonate (limestone), can make water slightly basic.

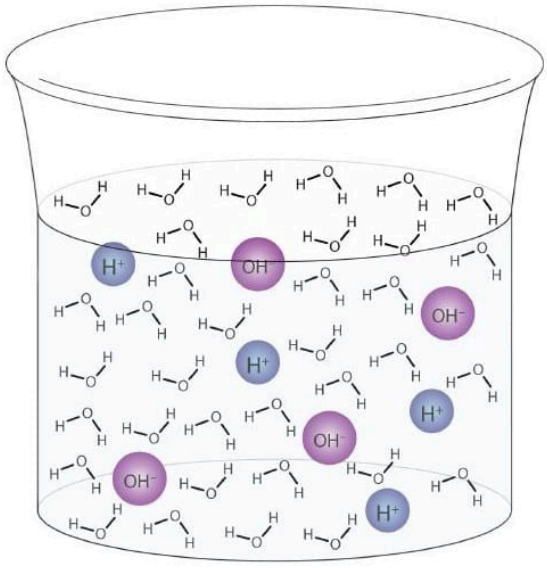


Figure 11.3.2: A small fraction of water molecules—approximately 6 in 100 million—ionize spontaneously into hydronium ions and hydroxide ions. This picture necessarily overrepresents the amount of autoionization that really occurs in pure water. (CC BY-NC-SA 3.0, Anonymous via LibreTexts). Detailed Description of Figure 11.3.2.

Example 11.3.1

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base for the following rections.

1. $$\mathrm{H_2O(l) + NO_2^-(aq) \rightarrow HNO_2(aq) + OH^-(aq)}$$
2. $$\mathrm{HC_2H_3O_2(aq) + H_2O(l) \rightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)}$$

Solution

1. In this reaction, the water molecule donates a proton to the NO_2^- ion, making OH^-(aq). As the proton donor, H_2O acts as a Brønsted-Lowry acid.
2. In this reaction, the water molecule accepts a proton from $HC_2H_3O_2$, becoming H_3O^+(aq). As the proton acceptor, H_2O is a Brønsted-Lowry base.

Key Takeaway

- Amphoteric substances can act both like acids and bases
- Water molecules can act as both an acid and a base, depending on the conditions.

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11.4: Strength of Acids and Bases

Introduction

Not all acids and bases behave the same way in solution. Some release many ions, while others release only a few. This difference is what we refer to as **acid strength**. A **strong acid** or **strong base** breaks apart almost completely in water, producing a high concentration of ions. In contrast, **weak acids** and **weak bases** only partially break apart, leaving a mixture of ions and unreacted molecules.

Understanding the strength of acids and bases is important because it affects how substances react in the body, how medications are absorbed, and how chemical reactions proceed in biological systems. The strength of acid HA is the extent to which the acid dissociates into H^+ and A^- ions, and it is illustrated in Figure 11.4.1.

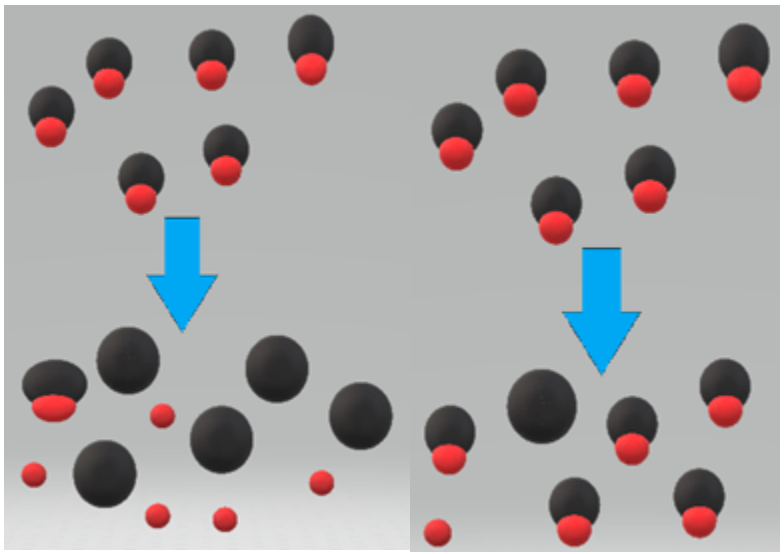


Figure 11.4.1: (a) A strong acid fully ionizes in water, as shown by the complete separation of molecules into hydrogen ions (red) and conjugate base ions (black). (CC0 1.0; Cwszot via Wikimedia Commons). (b) A weak acid only partially ionizes in water, as shown by the presence of mostly intact molecules with few free hydrogen ions (CC0 1.0; Cwszot via Wikimedia Commons). Detailed Description of Figure 11.4.1.

Strong acids

> Strong acids, like HCl, almost 100% dissociate into ions when they dissolve in water.

$$\mathrm{HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)}$$

> One arrow is used to indicate that the reaction is nearly 100% complete.

Strong acids include $HClO_4$, H_2SO_4, HI, HBr, HCl, and HNO_3

Weak acids

Weak acids dissolve in water but partially dissociate into ions.

For example, acetic acid (CH_3COOH) is a weak acid, 1 M acetic acid dissolves in water, but only 0.4% of the dissolved molecules dissociate into ions, the remaining 99.6% remain undissociated. The chemical equation of the acid dissociation is shown below.

$$CH_3COOH(aq) + H_2O(I) \rightleftarrows H_3O^+(aq) + CH_3COO^-(aq)$$

Two arrows pointing in opposite directions are used for the dissociation of weak acids to indicate that the reaction is an equilibrium, i.e., two ways.

Often the arrows are not equal in size -the longer arrow points to acid-base pair that is weaker and present in a larger concentration at equilibrium than their conjugate pair.

Figure 11.4.2 illustrates the extent of the dissociation of strong acid like HCl and a weak acid like HAc into their ions.

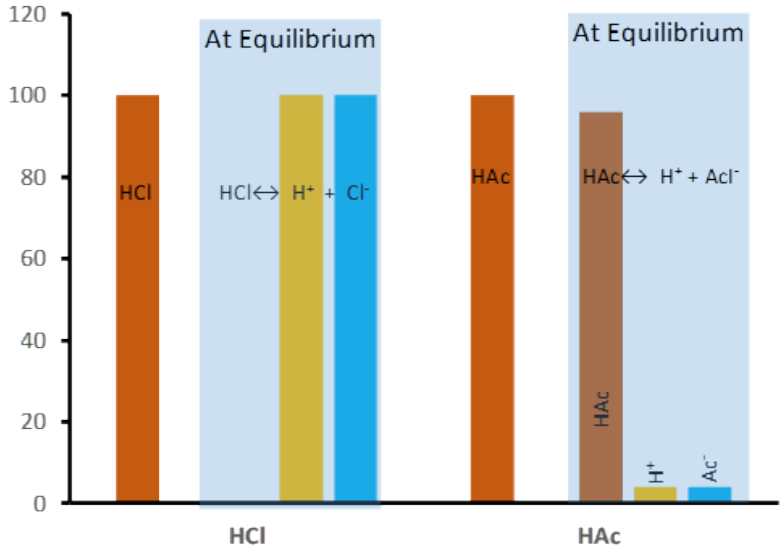


Figure 11.4.2: After dissolution in water. HCl –the strong acid is 100% dissociated into H^+ and Cl^- ions leaving no dissolved HCl molecules in water, but acetic acid (HAc) –the weak acid has a high concentration of HAc molecules and low concentration of H^+ and Ac^- ions. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.4.2.

Strong bases

Strong bases almost %100 dissociate into ions when dissolved in water. For example, NaOH is a strong base, and it dissociates almost 100% into ions in water.

$$NaOH(s) \xrightarrow{Water} Na^+(aq) + OH^-(aq)$$

One arrow is used for the dissolution of strong bases to indicate that the reaction is almost complete.

Strong bases include hydroxides of alkali metals, i.e., LiOH, NaOH, KOH, RbOH, CsOH, and hydroxides of heavy alkaline earth

metals, i.e., $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$. The last three, i.e., the hydroxides of heavy alkaline earth metals, have low solubility in water, but the dissolved fraction exists as ions.

Weak bases

> Weak bases partially dissociate into ions when dissolved in water.

For example, ammonia is a weak base –only 0.42% of the dissolved ammonia molecules dissociate into ammonium ions and hydroxide ions in water from a 1 M solution of ammonia.

$$NH_3(aq) + H_2O(I) \rightleftarrows {NH_4}^+(aq) + {OH}^-(aq)$$

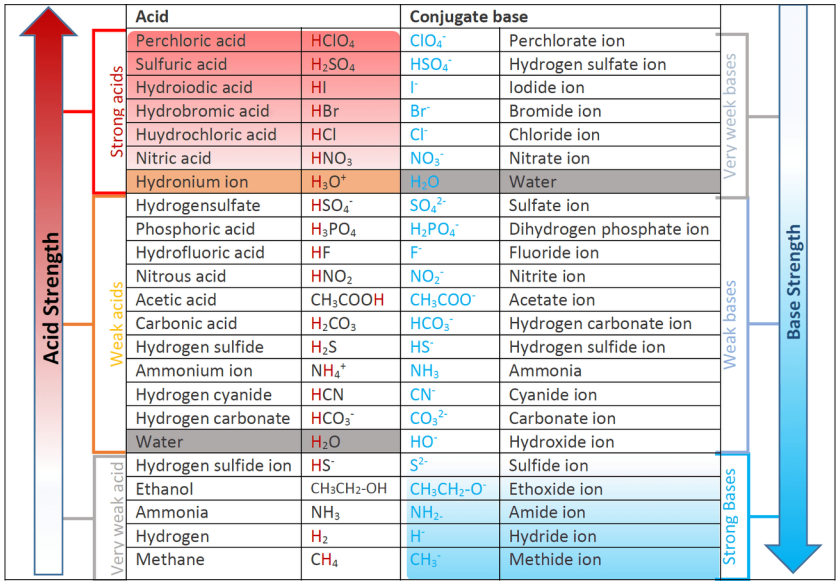
Weak bases in household use include ammonia (NH_3) in window cleaners, NaClO in bleach, Na_2CO_3 and Na_3PO_4 in laundry detergent, $NaHCO_3$ in tooth past, Na_2CO_3 in baking powder, $CaCO_3$ for use in lawns, $Mg(OH)_2$ and $Al(OH)_3$ in antacids and laxatives.

> The solubility and the strength of acids and bases are two different things. A strong base may be less soluble, and a weak base may be more soluble or vice versa, but a dissolved strong base exists as ions only, and a dissolved weak base exists both as molecules and ions.

The relative strength of acid-conjugate base pair

> A general rule is that the stronger the acid, the weaker the conjugate base, and vice versa.

The conjugate bases of strong acids have negligible base strength, and the conjugate acids of strong basses have negligible acid strength. Figure 11.4.3 illustrates the relative strengths of some acids and their conjugated bases.



Acid		Conjugate base	
Perchloric acid	$HClO_4$	ClO_4^-	Perchlorate ion
Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion
Hydroiodic acid	HI	I^-	Iodide ion
Hydrobromic acid	HBr	Br^-	Bromide ion
Huydrochloric acid	HCl	Cl^-	Chloride ion
Nitric acid	HNO_3	NO_3^-	Nitrate ion
Hydronium ion	H_3O^+	H_2O	Water
Hydrogensulfate	HSO_4^-	SO_4^{2-}	Sulfate ion
Phosphoric acid	H_3PO_4	$H_2PO_4^-$	Dihydrogen phosphate ion
Hydrofluoric acid	HF	F^-	Fluoride ion
Nitrous acid	HNO_2	NO_2^-	Nitrite ion
Acetic acid	CH_3COOH	CH_3COO^-	Acetate ion
Carbonic acid	H_2CO_3	HCO_3^-	Hydrogen carbonate ion
Hydrogen sulfide	H_2S	HS^-	Hydrogen sulfide ion
Ammonium ion	NH_4^+	NH_3	Ammonia
Hydrogen cyanide	HCN	CN^-	Cyanide ion
Hydrogen carbonate	HCO_3^-	CO_3^{2-}	Carbonate ion
Water	H_2O	HO^-	Hydroxide ion
Hydrogen sulfide ion	HS^-	S^{2-}	Sulfide ion
Ethanol	CH_3CH_2-OH	CH_3CH_2-O^-	Ethoxide ion
Ammonia	NH_3	NH_2^-	Amide ion
Hydrogen	H_2	H^-	Hydride ion
Methane	CH_4	CH_3^-	Methide ion

Figure 11.4.3: Strength of Acid-conjugate Base Pairs Relative to Water as a Reference. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.4.3.

Chemistry in Everyday Life-Household Acids and Bases

Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for hydrochloric acid [$\mathrm{HCl\,(aq)}$]. Vinegar has already been mentioned as a dilute solution of acetic acid [$\mathrm{HC_2H_3O_2\,(aq)}$]. In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid ($\mathrm{HC_6H_7O_6}$).

One of the more familiar household bases is ammonia ($\mathrm{NH_3}$), which is found in numerous cleaning products. As we mentioned previously, ammonia is a base because it increases the hydroxide ion concentration by reacting with water:

$$\mathrm{NH_3\,(aq) + H_2O(\ell) \rightarrow NH_4^+\,(aq) + OH^-\,(aq)}$$

Many soaps are also slightly basic because they contain compounds that act as Brønsted-Lowry bases, accepting protons from water and forming excess hydroxide ions. This is one reason that soap solutions are slippery.

Perhaps the most dangerous household chemical is the lye-based drain cleaner . Lye is a common name for sodium hydroxide, although it is also used as a synonym for potassium hydroxide. Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and form a clog in a pipe. Unfortunately, lye can also attack tissues and other substances in our bodies. Thus, when we use lye-based drain cleaners, we must be very careful not to touch any of the solid drain cleaner or spill the water it was poured into as seen in Figure 11.4.4. Safer, nonlye drain cleaners use peroxide compounds to react on the materials in the clog and clear the drain.



(a) (b)

Figure 11.4.4: (a) Bottles of alkaline drain cleaners containing sodium hydroxide can dissolve greases and hair. (Public Domain; Skatebiker via Wikimedia Commons). (b) Chemical burns caused by sodium hydroxide solution photographed 44 hours after exposure (CC BY-SA 3.0, Blazius via Wikimedia Commons). Detailed Description of Figure 11.4.4.

Acid-Base Equilibrium

The behavior of weak acids and bases illustrates a key concept in chemistry.

Does the chemical reaction describing the ionization of a weak acid or base just stop when the acid or base is done ionizing?

Actually, no. Rather, the reverse process—the reformation of the molecular form of the acid or base—occurs, ultimately at the same rate as the ionization process. For example, the ionization of the weak acid $\mathrm{HC_2H_3O_2\,(aq)}$ is as follows:

$$\mathrm{HC_2H_3O_2\,(aq) + H_2O(\ell) \rightarrow H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)}$$

The reverse process also begins to occur:

$$\mathrm{H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)\ \rightarrow HC_2H_3O_2\,(aq) + H_2O(\ell)}$$

Eventually, there is a balance between the two opposing processes, and no additional change occurs. The chemical reaction is better represented at this point with a double arrow:

$$\mathrm{HC_2H_3O_2\,(aq) + H_2O(\ell) \rightleftharpoons H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)}$$

The $\rightleftharpoons$ implies that both the forward and reverse reactions are occurring, and their effects cancel each other out. A process at this point is considered to be at **chemical equilibrium** (or equilibrium). It is important to note that the processes do not stop. They balance out each other so that there is no further net change; that is, chemical equilibrium is a **dynamic equilibrium**.

Example 11.4.1: Partial Ionization

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. HNO_2(aq)
b. C_5H_5N(aq)

Solution

a. $$\mathrm{HNO_2\,(aq) + H_2O(\ell)\ \rightarrow H_3O^+\,(aq) + NO_2{}^-\,(aq)}$$

b. $$\mathrm{C_5H_5N\,(aq) + H_2O(\ell)\ \rightarrow OH^-\,(aq) + C_5H_5NH^+\,(aq)}$$

Acid Dissociation Constant, K_a

The dissociation for a general weak acid, HA, can be written as follows:

$$\mathrm{HA}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{A^-}\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid dissociation constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_\mathrm{a} = \frac{[\mathrm{H^+}]\,[\mathrm{A^-}]}{[\mathrm{HA}]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution.

> Therefore, the numerical value of K_a is a reflection of the strength of the acid. Weak acids with relatively higher K_a values are stronger than acids with relatively lower K_a values.

Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the K_a value approaches infinity. For this reason, K_a values are generally reported for weak acids only.

The table below (Table $11.4.1$) is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table 11.4.1: Acid Ionization Constants at 25°C

Name of Acid	Ionization Equation	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$ $HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$	very large 1.3×10^{-2}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.2×10^{-7} 4.8×10^{-11}

Hydrofluoric acid $HF_{(aq)}$ reacts directly with glass (very few chemicals react with glass). Hydrofluoric acid is used in glass etching.

Strong and Weak Bases and Base Ionization Constant, K_b

As with acids, bases can either be strong or weak, depending on their extent of dissociation.

- A strong base is a base, which dissociates completely in an aqueous solution. The most common strong bases are soluble metal hydroxide compounds such as potassium hydroxide.
- A weak base is a base that ionizes only slightly in an aqueous solution.

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

The equilibrium greatly favors the reactants and the extent of ionization of the ammonia molecule is very small.

An equilibrium expression can be written for the reactions of weak bases with water. Because the concentration of water is extremely large and virtually constant, the water is not included in the expression. A **base dissociation constant** (K_b) is the equilibrium constant for the ionization of a base. For ammonia the expression is:

$$K_b = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}$$

The numerical value of K_b is a reflection of the strength of the base. Weak bases with relatively higher K_b values are stronger than bases with relatively lower K_b values. Table 11.4.2 is a listing of base ionization constants for several weak bases.

Table 11.4.2: Base Ionization Constants at 25°C

Name of Base	Ionization Equation	K_b
Methylamine	$CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$	5.6×10^{-4}
Ammonia	$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	1.8×10^{-5}
Pyridine	$C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH^-$	1.7×10^{-9}
Acetate ion	$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$	5.6×10^{-10}
Fluoride ion	$F^- + H_2O \rightleftharpoons HF + OH^-$	1.4×10^{-11}

Name of Base	Ionization Equation	K_b
Urea	$H_2NCONH_2 + H_2O \rightleftharpoons H_2NCONH_3^+ + OH^-$	1.5×10^{-14}

Key Takeaways

- Acids and bases can be strong or weak depending on the extent of ionization in solution.
- Most chemical reactions reach equilibrium at which point there is no net change.

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11.5: Dissociation of Water

Overview

As we have already seen, H_2O is an amphoteric substance which means it can donate a proton act as an acid or accept a proton and act as a base. Within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases.

> About one water molecule in half a billion dissociates into an OH^- ion by losing a proton to another water molecule. The molecule that receives a proton becomes H_3O^+.

The chemical equation is as follows:

$$\underbrace{H_2O\,(l)}_{\text{acid}} + \underbrace{H_2O\,(l)}_{\text{base}} \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq)$$

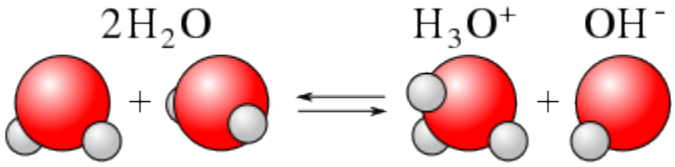


Figure 11.5.1: Water autoionizes to give hydroxide and hydronium ions. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.5.1.

The dissociation of water is an equilibrium reaction in which one water molecule donates its proton to another water molecule (Figure 11.5.1). The water molecule that receives proton is acting as a base, and it converts to conjugate acid H_3O^+. The other water molecule that donates a proton is acting as an acid, and it converts to conjugate base OH^-. The arrows in the reaction show that the base uses one of its lone pairs of electrons to make a bond with proton, and the previous bond pair of electrons turns into a third lone pair of electrons on the oxygen atom of the base. The reaction is reversible, i.e., the conjugate acid (H_3O^+) and the conjugate base (OH^-) react to re-form the two water molecules as illustrated in Figure 11.5.2.

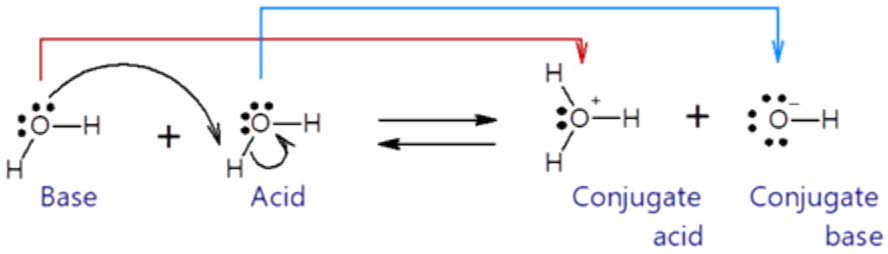


Figure 11.5.2: This is an equilibrium reaction in which one water molecule donates its proton to another water molecule. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.5.2.

The Dissociation Constant for Water, Kw

The ion-product of water, also known as the dissociation constant for water (*K*w) is the mathematical product of the concentration of hydrogen ions and hydroxide ions.

$$K_w = [H^+][OH^-]$$

Note that H_2O is not included in the ion-product expression because it is a pure liquid. The value of Kw is very small, in accordance with a reaction that favors the reactants.

At 25°C, the experimentally determined value of K_w in pure water is 1.0×10^{-14}.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

At 25°C, the experimentally determined value of K_w in pure water is 1.0×10^{-14}.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

In a sample of pure water, the concentrations of hydrogen and hydroxide ions are equal to one another. Pure water or any other aqueous solution in which this ratio holds is said to be neutral. To find the molarity of each ion, the square root of Kw is taken.

$$[H^+] = [OH^-] = 1.0 \times 10^{-7}$$

The product of these two concentrations is 1.0×10^{-14}

$$[H^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

Neutral, Acidic, or Basic Aqueous Solutions

A solution that has an equal concentration of H_3O^+ and OH^-, each equal to 10^{-7} M, is a neutral solution.

An acidic solution has an acid dissolved in water. When an acid dissolves in water it dissociates adding more H_3O^+. The $[OH^-]$ must decrease to keep the K_w constant.

A solution that has $[H_3O^+]$ more than 10^{-7}, and $[OH^-]$ less than 10^{-7} is an acidic solution.

A basic solution has a base dissolved in water. When a base dissolves in water it dissociates adding more OH^-. The $[H_3O^+]$ must decrease to keep the K_w constant.

A solution that has $[H_3O^+]$ less than 10^{-7}, and $[OH^-]$ more than 10^{-7} is a basic solution.

Calculations of $[H_3O^+]$ and $[OH^-]$ based on K_w

The water dissociation constant remains the same whether the aqueous solution is neutral, acidic, or basic, i.e.:

$$K_w = [H_3O^+][OH^-] = (10^{-7})(10^{-7}) = 10^{-14} \text{ at } 25^\circ C$$

Therefore, if the molar concentration of hydronium ions $[H_3O^+]$ is known, the molar concentration of hydroxide ions $[OH^-]$ can be calculated using the following formula:

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{[OH^-]}$$

Similarly, if the molar concentration of hydroxide ions $[OH^-]$ is known, the molar concentration of hydronium ions $[OH^-]$ can be calculated using the following formula:

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{10^{-14}}{[H_3O^+]}$$

When a strong acid like HCl dissolves in water, it dissociates ~100% into ions. Therefore, the $[H_3O^+]$ is equal to the molar

concentration of the acid. The amount H_3O^+ added by dissociation of water molecules is very small compared to that coming from the dissociation of a strong acid and can be neglected. Similarly, when a strong base like NaOH dissolves in water, it dissociates ~100% into ions. Therefore, the $[OH^-]$ is equal to the molar concentration of the base.

> When a weak acid or a weak base dissolves in water, it partially dissociates into ions. Therefore, the $[H_3O^+]$ or the $[OH^-]$ in the cases of weak acids and weak bases has to be determined experimentally for the calculations.

Example 11.5.1

Calculate the concentration of OH^- ions in a 0.10 M HNO_3 solution?

Solution

The HNO_3 is a strong acid. Therefore, $[HNO_3] = 0.10$ M $= [H_3O^+]$. Desired $[OH^-] = ?$

Formula: $[\mathrm{OH^-}] = \frac{10^{-14}}{[\mathrm{H_3O^+}]}$

Plug in values an calculate: $[\mathrm{OH^-}] = \frac{10^{-14}}{0.10} = 10^{-13}\ \mathrm{M}$

Example 11.5.2

A vinegar solution has $[H_3O^+] = 2.0 \times 10^{-3}$. a) What is the hydroxide ion concentration in the vinegar solution? b) is the solution acidic, basic, or neutral?

Solution

a) Given $[H_3O^+] = 2.0 \times 10^{-3}$. Desired $[OH^-] = ?$

Formula: $[\mathrm{OH^-}] = \frac{10^{-14}}{[\mathrm{H_3O^+}]}$

Plug in values and calculate: $[\mathrm{OH^-}] = \frac{10^{-14}}{2.0\times10^{-3}} = 5.0 \times 10^{-12}\mathrm{M}$

b) The solution is acidic because $[H_3O^+] > [OH^-]$.

Example 11.5.3

Calculate the value of $[H^3O^+]$ and $[OH^-]$ in a 0.010 M NaOH solution?

Solution

The NaOH is a strong base. Therefore $[NaOH] = 0.010$ M $= [OH^-]$. Desired $[H_3O^+] = ?$

Formula: $[\mathrm{H_3O^+}] = \mathrm{K_w}/\,[\mathrm{OH^-}] = 10^{-14}/\,[\mathrm{OH^-}]$

Calculations: $[\mathrm{H_3O^+}] = 10^{-14}/0.010 = 10^{-12}\mathrm{M}$

Example 11.5.4

a. Calculate the $[H_3O^+]$ in an ammonia solution that has $[OH^-] = 4.0 \times 10^{-4}$ M? b) Is the solution acidic, basic, or neutral?

Solution

a. Given $[OH^-] = 4.0 \times 10^{-4}$. Desired $[H_3O^+]$ = ?

Formula: $[H_3O^+] = K_w / [OH^-] = 10^{-14} / [OH^-]$

Calculations: $[H_3O^+] = \frac{10^{-14}}{4.0 \times 10^{-4}} = 2.5 \times 10^{-11}\ M$

b. The solution is basic because $[H_3O^+] < [OH^-]$.

The pH Scale

One qualitative measure of the strength of an acid or a base solution is the **pH scale**, which is based on the concentration of the hydronium (or hydrogen) ion in aqueous solution.

$$pH = -\log[H^+]$$

or

$$pH = -\log[H_3O^+]$$

Figure 11.5.3 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.

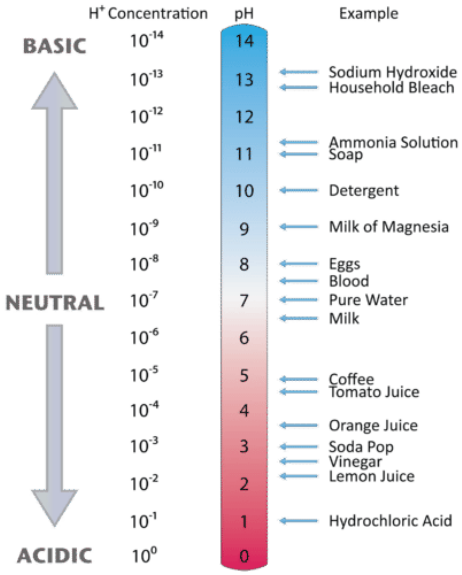


Figure 11.5.3: The relationship between $[H^+]$ and pH Values for Several Common Materials. (CC BY-NC-SA 3.0; Anonymous via LibreTexts). Detailed Description of Figure 11.5.3.

A neutral (neither acidic nor basic) solution has a pH of 7. A pH below 7 means that a solution is acidic, with lower values of pH corresponding to increasingly acidic solutions. A pH greater than 7 indicates a basic solution, with higher values of pH corresponding to increasingly basic solutions. Thus, given the pH of several solutions, you can state which ones are acidic, which ones are basic, and which are more acidic or basic than others. These are summarized in Table \({11{.}5{.}\eatSpaces{1}\endSpaces}.

Table 11.5.1: Acidic, Basic and Neutral pH Values

Classification	Relative Ion Concentrations	pH at 25 °C
acidic	$[H^+] > [OH^-]$	pH < 7
neutral	$[H^+] = [OH^-]$	pH = 7
basic	$[H^+] < [OH^-]$	pH > 7

Example 11.5.5

Find the pH, given the $[H^+]$ of the following:

a. 1×10^{-3} M
b. 2.5×10^{-11} M
c. 4.7×10^{-9} M

Solution

$pH = - \log [H_3O^+]$

Substitute the known quantity into the equation and solve. Use a scientific calculator for b and c.

a. $pH = - \log [1 \times 10^{-3}] = 3.0$ (1 decimal place since 1 has 1 significant figure)
b. $pH = - \log [2.5 \times 10^{-11}] = 10.60$ (2 decimal places since 2.5 has 2 significant figures)
c. $pH = - \log [4.7 \times 10^{-9}] = 8.33$ (2 decimal places since 4.7 has 2 significant figures)

Note on significant figures:

Because the number(s) before the decimal point in the pH value relate to the power on 10, the number of digits *after* the decimal point (underlined) is what determines the number of significant figures in the final answer.

Significance of pH Scale

As mentioned above, the pH scale varies from ~0 to ~14. The pH of 7 is neutral, pH more than 7 is basic, and pH less than 7 is acidic, as illustrated in Table 11.5.1 Different foods have different pH values, as shown in Figure 11.5.4. Similarly, several of the acids and bases in household use have a specific pH range, as shown in Figure 11.5.4.

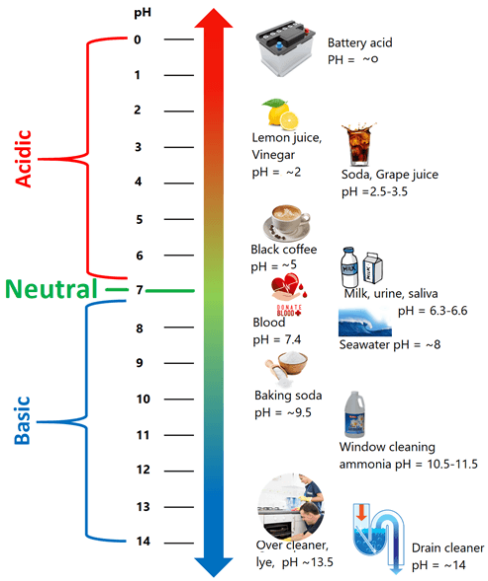


Figure 11.5.4: pH 7 is neutral, pH less than 7 is acidic, and pH more than 7 is basic. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.5.4.

The control of pH is essential in the proper functioning of biological systems. Plants thrive if the soil does not have too acidic or too basic a pH. Lacks and revers have a specific range of pH in which the aquatic life can survive. US natural water has a pH in the range of 6.5 to 8.5 range. Seawater has a specific narrow range of pH 7.5 to pH 8.4 in which the life in the sea can function correctly.

The pH of body fluids

Human body fluids vary in pH, as shown in Figure 11.5.5. Saliva in the mouth is slightly acidic, but the stomach has the lowest pH in the body. The strongly acidic pH in the stomach helps in digesting some foods, and it also helps in killing bacteria that may enter the stomach through the foods. When the food enters the large intestine, the pH changes to basic which helps in digesting the foods that could not be digested in the acidic environment of the stomach. The pH changes to more basic in the small intestine.

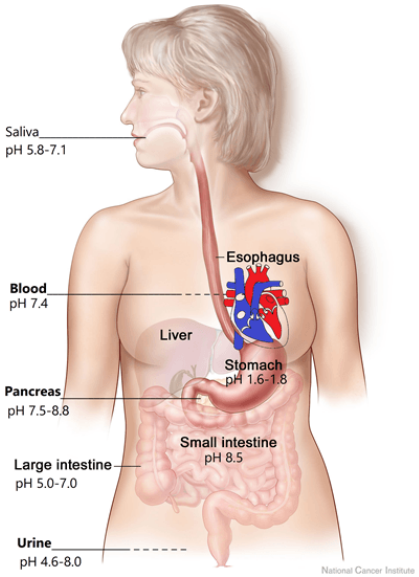


Figure 11.5.5: The pH in Various Human Body Fluids. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.5.5.

The blood has a pH of 7.4, and it can vary in a small range of 7.35 to 7.45. If the blood pH goes outside of the 7.35 to 7.45 range it can result in medical problems. Enzymes in the body need a specific pH range because hydrogen bonding plays a vital role in

the structures needed for proper functioning. pH changes affect the hydrogen bonding and can make the enzymes less active or may inactivate them. The pH of blood is maintained by a complex action of buffers that are described in the later sections. The pH of urine can vary over a broad range from 4.6 to 8, depending on the recent diet and exercises.

Key Takeaways

- Water molecules can act as both an acid and a base, depending on the conditions.
- The pH scale is used to succinctly communicate the acidity or basicity of a solution.

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11.6: Acid Base Reactions

Neutralization

Acid-base reactions are chemical reactions that involve the transfer of a proton (H^+). Examples include neutralization reactions, and the reactions of acids with metals, and carbonates.

The reaction of HCl with NH_3 is an acid/base reaction. HCl is an acid because it donates its proton to NH_3, and the NH_3 is a base because it accepts a proton, as shown in Figure 11.6.1



Figure 11.6.1: This is an example of an acid-base reaction. Hydrochloric acid evaporates from liquid HCl in a beaker reacting with ammonia fumes coming from the test tube to produce ammonium chloride (white smoke). (Public Domain; Walkerma via Wikimedia Commons). Detailed Description of Figure 11.6.1.

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 (Arrhenius base), water is also a product. The general reaction is as follows:

$$\text{acid + base} \rightarrow \text{water + salt}$$

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:

$$\mathrm{HCl(aq) + Na(OH)(aq) \rightarrow NaCl)(aq + H_2O(l)}$$

With coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$\mathrm{2NaOH(aq) + H_2(SO_4)(aq) \rightarrow Na_2(SO_4)(aq) + H_2O(l)}$$

A neutralization reaction takes place between citric acid or acetic acid (proton donors) in lemons or vinegar and putrescine (proton acceptor) in the flesh of fish in (Figure 11.6.2). An example of a neutralization reaction between Acetic Acid and Putrescine is shown below.

$$\underset{\text{Acetic Acid}}{\mathrm{CH_3COOH}} + \underset{\text{Putrescine}}{\mathrm{NH_2CH_2CH_2CH_2CH_2NH_2}} \rightarrow \underset{\text{Acetate Ion}}{\mathrm{CH_3COO^-}} + \underset{\text{Putrescinium Ion}}{\mathrm{NH_3^+CH_2CH_2CH_2CH_2NH_2}}$$



Figure \(\PageIndex{2}\): Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish . It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This reduces the odor of the fish, and also adds a “sour” taste that we seem to enjoy. (CC BY 4.0; Anonymous via OpenStax). Detailed Description of Figure 11.6.2.

Writing a Balanced Chemical Equation for a Neutralization Reactions

The following example shows the steps needed to write the balanced equation of the reaction of an acid with an Arrhenius base.

Example 11.6.1

Write a balanced chemical equation of a reaction between HCl and $Ca(OH)_2$?

Solution

Step 1) Write the formula of acid and base in the reactants and salt and water in the products. All the strong electrolytes dissolve in water, so use (aq) to represent their state.

$$\mathrm{HCl(aq) + Ca(OH)_2(aq) \rightarrow \ Salt(aq) \ + H_2O(l)}$$

Step 2) Balance the H^+ in the acid with the OH^- in the base.

$$\mathrm{2HCl(aq) + Ca(OH)_2(aq) \rightarrow Salt(aq) + H_2O(l)}$$

Step 3) Balance the H_2O with the H^+ in the acid, or with the OH^- in the base.

$$\mathrm{2HCl(aq) + Ca(OH)_2(aq) \rightarrow \ Salt \ (aq) + 2H_2O(I)}$$

Step 4) Write the salt by combing the cations from the base and the anions from the acid. Make sure the charges are balanced in the salt to make it a neutral substance.

$$\mathrm{2HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2H_2O(I)} \qquad (11.6.1)$$

Example 11.6.2

Write a balanced chemical equation of a reaction between H_2SO_4 and $Ba(OH)_2$?

Solution

Step 1) Write the formula of acid and base in the reactants and salt and water in the products. All the strong electrolytes dissolve in water, so use (aq) to represent their state.

$$\mathrm{H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow Salt(aq) + H_2O(l)}$$

Step 2) Balance the H^+ in the acid with the OH^- in the base. They are already balanced in the above equation.

Step 3) Balance the H_2O with the H^+ in the acid, or with the OH^- in the base.

$$\mathrm{H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow Salt(aq) + 2H_2O(l)}$$

Step 4) Write the salt by combing the cations from the base and the anions from the acid. Make sure the charges are balanced in the salt to make it a neutral substance.

$$\mathrm{H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(aq) + 2H_2O(l)}$$

Stoichiometry Calculations of Neutralization Reactions

Once a neutralization reaction is properly balanced, we can use it to perform stoichiometry calculations, such as the ones we practiced earlier**.**

Example 11.6.3

Nitric acid [HNO_3(aq)] can be neutralized by calcium hydroxide [$Ca(OH)_2$(aq)].

a. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.
b. For one reaction, 16.8 g of HNO_3 is present initially. How many grams of $Ca(OH)_2$ are needed to neutralize that much HNO_3?
c. In a second reaction, 805 mL of 0.672 M $Ca(OH)_2$ is present initially. What volume of 0.432 M HNO_3 solution is necessary to neutralize the $Ca(OH)_2$ solution?

Solution

a. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of HNO_3 to provide H^+ ions. The balanced chemical equation is as follows: $Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(\ell)$

   The salt formed is calcium nitrate.

b. This calculation is much like the calculations we did in Chapter 7. First we convert the mass of HNO_3 to moles using its molar mass of 1.01 + 14.01 + 3(16.00) = 63.02 g/mol; then we use the balanced chemical equation to determine the related number of moles of $Ca(OH)_2$ needed to neutralize it; and then we convert that number of moles of $Ca(OH)_2$ to the mass of $Ca(OH)_2$ using its molar mass of 40.08 + 2(1.01) + 2(16.00) = 74.10 g/mol (Figure 11.6.3).

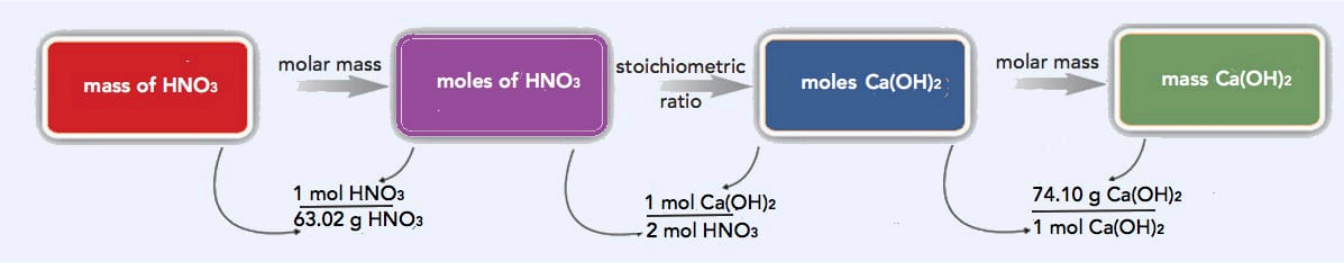


Figure 11.6.3: Concept map of the Calculations. (CC BY-NC-SA 3.0; Anonymous via Libretexts). Detailed Description of Figure 11.6.3.

$$16.8\ \text{g HNO}_3 \times \frac{1\ \text{mol HNO}_3}{63.02\ \text{g HNO}_3} \times \frac{1\ \text{mol Ca(OH)}_2}{2\ \text{mol HNO}_3} \times \frac{74.10\ \text{g Ca(OH)}_2}{1\ \text{mol Ca(OH)}_2} = 9.88\ \text{g Ca(OH)}_2\ \text{needed}$$

c. Having concentration information allows us to employ the skills we developed earlier. We have two alternative solutions: the multi-step process and the combined one-line process (Figure 11.6.4).

First, we use the concentration and volume data to determine the number of moles of $Ca(OH)_2$ present. Recognizing that 805 mL = 0.805 L,

$$0.672\ \text{M Ca(OH)}_2 = \frac{\text{mol Ca(OH)}_2}{0.805\ \text{L soln}}$$

0.672 M $CaOH)_2$ × (0.805 L soln) = mol $Ca(OH)_2$ = 0.541 mol $Ca(OH)_2$

We combine this information with the proper ratio from the balanced chemical equation to determine the number of moles of HNO_3 needed:

$$0.541\ \text{mol Ca(OH)}_2 \times \frac{2\ \text{mol HNO}_3}{1\ \text{mol Ca(OH)}_2} = 1.08\ \text{mol HNO}_3$$

Now, using the definition of molarity one more time, we determine the volume of acid solution needed:

$$0.432\ \text{M HNO}_3 = \frac{1.08\ \text{mol HNO}_3}{\text{volume of HNO}_3}$$

$$\text{volume of HNO}_3 = \frac{1.08\ \text{mol HNO}_3}{0.432\ \text{M HNO}_3} = 2.50\ \text{L} = 2.50 \times 10^3\ \text{mL HNO}_3$$

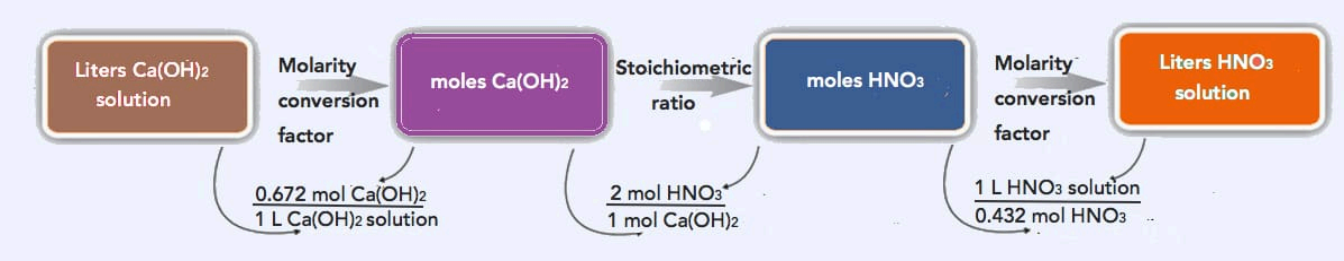


Figure 11.6.4: Concept map of the Calculations. (CC BY-NC-SA 3.0; Anonymous via Libretexts). Detailed Description of Figure 11.6.4.

$$0.805\ \text{L Ca(OH)}_2\ \text{soln} \times \frac{0.672\ \text{mol Ca(OH)}_2}{1\ \text{L Ca(OH)}_2\ \text{soln}} \times \frac{2\ \text{mol HNO}_3}{1\ \text{mol Ca(OH)}_2} \times \frac{1\ \text{L HNO}_3\ \text{soln}}{0.432\ \text{mol HNO}_3} = 2.50\ \text{L HNO}_3\ \text{soln needed}$$

Acid-base Titrations

The **acid-base titration** is an analytical process of determining the concentration of acid, called **analyte**, by *neutralizing* it with a base of known concentration, called the **standard**, or vice versa.

Usually, the base is in a burette and added drop by drop to a known volume of the acid in an Erlenmeyer flask, as illustrated in Figure 11.6.5. A few drops of an acid-base indicator are mixed with the acid. The indicator changes color at a specific narrow pH range.

- The point when the stoichiometric amount of the base has been added to the acid is called the **equivalence point.**
- The point when the indicator changes color is the **end point**.
- Usually, the equivalence point is almost the same as the end point.

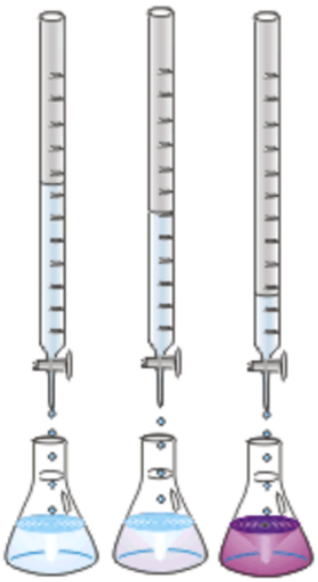


Figure 11.6.5: The analyte color due to the indicator before, at the endpoint, and after being shown by the images on the left, middle, and right, respectively. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.6.5.

The calculation steps after the titration are usually the following.

Steps in acid-base titration calculations

1. The amount, in moles, of the standard is calculated by multiplying the amount in liters with the molarity of the standard.
2. Then, the amount, in moles, of the analyte, is calculated by using the mol-to-mol conversion factor from the balanced chemical equation of the reaction.
3. Finally, the molarity of the analyte is calculated by dividing moles by the volume in liters of the analyte.

Example 11.6.4

What is the molarity of the HCl solution if 50.0 mL of the HCl solution requires 42.0 mL of 0.123 M NaOH solution in the titration?

Solution

Step 1) Given:

$$[\text{NaOH}] = 0.123 \text{ M} = \frac{0.123 \text{ mol } NaOH}{1 \text{ NaOH}}$$

$$\text{Volume of standard} = 42.0 \text{ mL.} = 42.0 \text{ mL NaOH} \times \frac{1 \text{ L NaOH}}{1000 \text{ mL NaOH}} = 0.0420 \ L \text{ NaOH}$$

$$\text{Vol. analyte} = 50.0 \text{ mL} = 50.0 \text{ mL HCl} \times \frac{1 \text{ L HCl}}{1000 \text{ mL HCl}} = 0.0500 \text{ L HCl}$$

Step 2) Calculate the moles of the standard from the volume and the molarity product.

$$\text{Moles of NaOH consumed} = 0.0420 \text{ L NaOH} \times \frac{0.123 \text{ mol NaOH}}{1 \text{ L NaOH}} = 0.00517 \text{ mol NaOH}$$

Step 3) write the balanced chemical equation of the reaction.

$$\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$$

Step 4) write the conversion factor for mole of standard to mole of analyte calculation from the equation.

$$\frac{1\text{mol HCl}}{1\text{ mol NaOH}}$$

Step 5) Calculate the moles of the analyte by multiplying the moles of the standard and the conversion factor.

$$0.00517\text{ mol NaOH} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 0.00517\text{ mol HCl}$$

Step 6) Calculate the molarity of the analyte by dividing the moles with the volume in liters of the analyte.

$$0.00517\text{ mol NaOH} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 0.00517\text{ mol HCl}$$

After learning the steps thoroughly, the calculations can be done in a few steps, as shown in the following.

Calculations:

$$42.0\text{ mL NaOH} \times \frac{1\text{ L NaOH}}{1000\text{ mL NaOH}} \times \frac{0.123\text{ mol NaOH}}{1\text{ L NaOH}} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 0.00517\text{ mol HCl}$$

Molarity of HCl:

$$M = \frac{n(\text{mol})}{V(L)} = \frac{0.00517\text{ mol HCl}}{50.0\text{ mL HCl}} \times \frac{1000\text{ mL HCl}}{1\text{ L HCl}} = 0.103\text{ M HCl}$$

Example 11.6.5

What is the molarity of the H_2SO_4 solution if 50.0 mL of the H_2SO_4 solution requires 32.3 mL of 0.201 M NaOH solution in the titration?

Solution

Step 1) Given:

$$[\text{NaOH}] = 0.201\,\text{M} = \frac{0.201\,\text{mol NaOH}}{1\,\text{L NaOH}}$$

$$\text{Vol. standard} = 32.3\,\text{mL NaOH} \times \frac{1\,\text{L NaOH}}{1000\,\text{mL NaOH}} = 0.0323\,\text{L NaOH} \tag{11.6.2}$$

$$\text{Vol. analyte} = 50.0\,\text{mL } H_2SO_4 \times \frac{1\,\text{L } H_2SO_4}{1000\,\text{mL } H_2SO_4} = 0.0500\,\text{L } H_2SO_4$$

Step 2) Calculate the moles of the standard from the volume and the molarity product.

$$\text{Moles of NaOH consumed} = 0.0323\text{ N NaOH} \times \frac{0.201\text{ mol NaOH}}{1\text{ L NaOH}} = 0.00649\text{ mol } NaOH$$

Step 3) write the balanced chemical equation of the reaction.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Step 4) write the conversion factor for the moles of standard to mole of analyte calculation from the equation.

$$\frac{1\text{ mol } H_2SO_4}{2\text{ mol NaOH}}$$

Step 5) Calculate the moles of the analyte by multiplying the moles of the standard and the conversion factor.

$$0.00649\text{ mol NaOH} \times \frac{1\text{ mol } H_2SO_4}{2\text{ mol NaOH}} = 0.00325\text{ mol } H_2SO_4$$

Step 6) Calculate the molarity of the analyte by dividing the moles with the volume in liters of the analyte.

$$\text{Concentration of } \ H_2SO_4 = \frac{0.00325 \text{ mol } H_2SO_4}{0.0500 \text{ L } H_2SO_4} = 0.0649 \text{ M } H_2SO_4$$

The same calculation in a summary form is in the following.

$$\text{Concentration of } H_2SO_4 = \frac{0.00325 \text{ mol } H_2SO_4}{0.0500 \text{ L } H_2SO_4} = 0.0649 \text{ M } H_2SO_4$$

$$\text{Moles of } H_2SO_4 = \quad 32.3 \text{ mL NaOH} \times \frac{1 \text{ L NaOH}}{1000 \text{ mL NaOH}} \times \frac{0.201 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}} = 0.00326 \text{ mol } H_2SO_4$$

$$\text{Molarity of } H_2SO_4: \quad M = \frac{n(\text{ mol})}{V(L)} = \frac{0.00326 \text{ mol } H_2SO_4}{50.0 \text{ mL } H_2SO_4} \times \frac{1000 \text{ mL } H_2SO_4}{1 \text{ L } H_2SO_4} = 0.0649 \text{ M } H_2SO_4$$

Reactions of Acids with Metals

Metals tend to give out an electron and become cations. The majority of the metals, called reactive metals, give out electrons to protons in the acids and release H_2 gas. For example, Figure 11.6.6 shows magnesium reacting with hydrochloric acid by the following reaction:

$$Mg(s) + 2HCl \rightarrow MgCl_2(aq) + H_2(\text{ g}) \uparrow$$



Figure 11.6.6: Mg Reacting with HCl and H_2 Gas Bubbling Out. (CC BY 4.0; Aaron Huggard via YouTube). Detailed Description of Figure 11.6.6.

Alkali metals like Na give out one electron, alkaline earth metals like Mg give out two electrons, and aluminum gives out three electrons. A proton can accept one electron. Therefore, the number of acidic protons should be equal to the number of electrons given out to balance the electron transfer in these reactions.

Noble metals or jewelry metals like gold, silver, platinum, copper are exceptions –they do not react with acids.

Example 11.6.4

Write a balanced equation for the reaction of Aluminum with HBr?

Solution

Step 1) write the equation with the correct formulas of reactants and products.

$$Al(s) + HBr(aq) \rightarrow AlBr_3(aq) + H_2(\ g)$$

Note that Al gives out three electrons to make Al^{3+}; that is why three Br^- are attached to it in the product to balance the charge.

Step 2) balance the electrons lost by the metal with the number of acidic protons.

$$Al(s) + 3HBr(aq) \rightarrow AlBr_3(aq) + H_2(\ g)$$

Step 3) Balance the rest of the elements by the hit and trial method. There are 3 hydrogen atoms on the left, balance them by adding 3/2 coefficient to the hydrogen on the right.

$$Al(s) + 3HBr(aq) \rightarrow AlBr_3(aq) + \frac{3}{2}H_2(\ g)$$

Step 4. Although the equation is balanced, it is recommended to remove the fraction by multiplying the coefficients in the whole equation with the highest common factor to obtain the final balanced equation.

$$2Al(s) + 6HBr(aq) \rightarrow 2AlBr_3(aq) + 3H_2(\ g)$$

Reactions of Acids with Carbonates and Hydrogen Carbonates

Carbonic acid (H_2CO_3) is a weak acid found in carbonated water. The H_2CO_3 is a product of carbon dioxide (CO_2) and water (H_2O) by the following equilibrium reaction.

$$CO_2\,(aq) + H_2O\,(l) \rightleftharpoons H_2CO_3\,(aq)$$

Hydrogen carbonate (HCO_3^-) and carbonate (CO_3^{2-}) are one and two hydrogen less, respectively than carbonic acid. The salts containing HCO_3^- and CO_3^{2-} accept one and two protons, respectively, from acids to make H_2CO_3. The H_2CO_3 decomposes to carbon dioxide and water by the reverse reaction shown above. Figure 11.6.7 shows the reaction of sodium hydrogen carbonate with hydrochloric acid.

$$NaHCO_3(\ s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(I) + CO_2(\ g)\uparrow$$



Figure 11.6.7: The Reaction of Sodium Carbonate with HCl. Before and after HCl addition. Carbon dioxide bubbles out after HCl addition. (CC BY 3.0; via North Carolina School of Science and Mathematics, YouTube). Detailed Description of Figure 11.6.7.

Similarly, calcium carbonate found in limestone reacts with acids. For example, sulfuric acid is one of the components in acid rain

that reacts with calcium carbonate and damages sculptures made of stone.

$$\mathrm{CaCO_3(\ s) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + H_2O(l) + CO_2(\ g) \uparrow}$$

Chemistry in Everyday Life: Antacids

The stomach sometimes produces excess HCl that may cause heartburn. Antacids are the substances used to neutralize the excess HCl in the stomach.



Figure 11.6.8: Bottle of Antacid Tablets. (CC BY-SA 2.5; Midnightcomm via Wikimedia Commons). Detailed Description of Figure 11.6.8.

Figure 11.6.8 shows antacid tablets. The antacids include Arrhenius bases with very low solubility in water, like $Al(OH)_3$, and $Mg(OH)_2$, or weak bases like $CaCO_2$, and $NaHCO_3$. that react with and neutralize the acid:

$$\mathrm{Al(OH)_3(\ s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)}$$
$$\mathrm{Mg(OH)_2(\ s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)}$$
$$\mathrm{CaCO_3(\ s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(I) + CO_2(\ g)}$$
$$\mathrm{NaHCO_3(\ s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(I) + CO_2(g)}$$

The antacids also have side effects, e.g., aluminum and calcium-containing antacids may cause constipation. Magnesium-containing antacids have a laxative effect. Some formulations use a mixture of aluminum and magnesium-containing antacids to cancel out the side effect of each other. The calcium-containing antacid is not recommended for persons who tend to develop kidney stones, because the kidney stone is usually a salt of calcium. Sodium-containing antacids may increase the sodium level in the body fluid and may raise the blood pH.

Acid rain

Acidic gases like NO, NO_2, N_2O_4, SO_2 are released into the environment primarily during the combustion of fossil fuels. These gases dissolve in the rainwater and make the rain acidic. For example, sulfur dioxide dissolves in water and makes sulfurous acid:

$$\mathrm{SO_2(\ g) + H_2O(I) \rightleftarrows H_2SO_3(aq)}$$

The acid rain damages the environment by making the soil, river water, and lack-water acidic. The acidic soil and water, in turn, affect the plants and aquatic life. The acid water also reacts with calcium carbonate and corrodes metals that are responsible for damage to the sculpture and other structures, as illustrated in Figure 11.6.9, Figure 11.6.10, Figure 11.6.11.

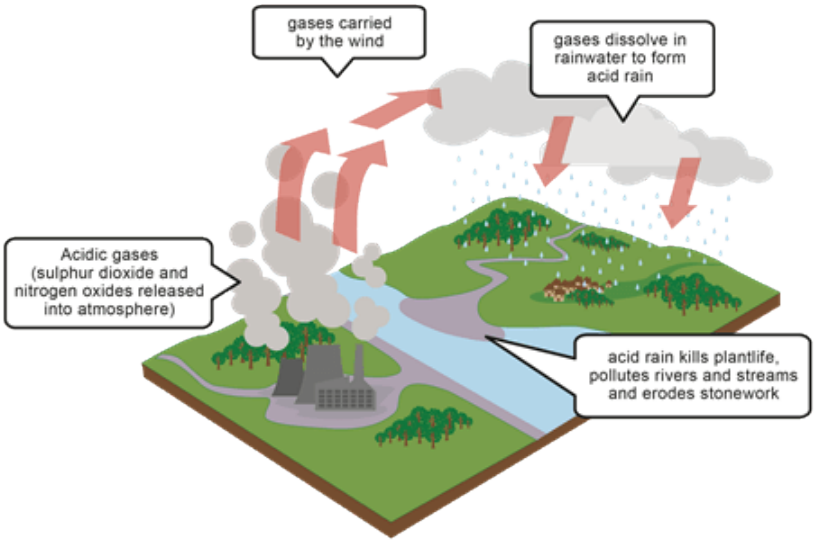


Figure 11.6.9: Illustration of Acid Rain. (CC BY-NC 4.0; Anonymous via HiClipart). Detailed Description of Figure 11.6.9.



(a)



(b)

Figure 11.6.10: (a) Effect of Acid Rain on Forest. (Public Domain; Lovecz via Wikimedia Commons). (b) Effect of Acid Rain on Monuments. (CC BY-SA 3.0; Nino Barbieri via Wikimedia Commons). Detailed Description of Figure 11.6.10.

Applications of Acid-Base Chemistry

There are a number of examples of acid-base chemistry in everyday life as seen in many examples above. It is worth noting a couple more applications .

One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter “rises.” Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq) \tag{11.6.3}$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favors the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavor of the vegetables with the acid making them taste sour.

Key Takeaways

- Acid–base reactions involve proton transfer, with acids donating H^+ and bases accepting H^+.
- Neutralization between acids and bases produces water and an ionic salt, and equations must be balanced carefully.
- Stoichiometry and titrations use balanced equations to calculate unknown amounts or concentrations of acids and bases.
- Acids react with metals to release H_2 gas and with carbonates or bicarbonates to release CO_2.
- Everyday applications include antacids, food chemistry (like baking and pickling), and environmental issues such as acid rain.

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- 10.2: Brønsted-Lowry Definition of Acids and Bases by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
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11.7: Buffers

Definition of Buffer

Buffers are special solutions formulated to keep *pH levels steady* when small amounts of acid or base are introduced. They achieve this by combining a weak acid with its conjugate base (or a weak base with its conjugate acid), which work together to absorb added hydrogen or hydroxide ions. This balancing act makes buffers indispensable in biology and allied health—examples include the bicarbonate system in blood, which helps maintain a narrow pH range essential for proper enzyme function and overall physiological stability.

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 a marked chemical activity. Moreover, the addition of 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 continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

An example of a buffer solution is one that is composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$). Sodium acetate is a salt that dissociates into sodium ions and acetate ions in solution. For as long as acetic acid and acetate ions are present in significant amounts a solution, this can resist dramatic pH changes. Another example of a buffer is a solution containing ammonia (NH_3, a weak base) and ammonium chloride (NH_4Cl). Ammonium acetate is also a salt that dissociates into ammonium ions and chloride ions in solution. The presence of ammonium ions with ammonia molecules satisfies the requisite condition for a buffer solution.

> A buffer is an aqueous solution consisting of a weak acid and its conjugate base or vice versa that lets the pH change to be minimal when a small amount of a strong acid or a strong base adds to it.

The addition of 0.020 mol HCl into 1 L of water changes pH from 7 to 1.7, i.e., about an 80% change in pH. The addition of 0.020 mol NaOH to the same water changes pH from 7 to 12.3, i.e., again, about an 80% change in pH.

In contrast to pure water, 1 L of buffer solution containing 0.50 mol acetic acid (CH_3COOH) and 0.50 mol CH_3COO^-, the conjugate of the acetic acid, changes pH from 4.74 to 4.70 by the addition of the same 0.020 mol HCl, and from 4.74 to 4.77 by the addition of 0.020 mol NaOH, i.e., about 1% change in pH, as illustrated in Figure 11.7.1.

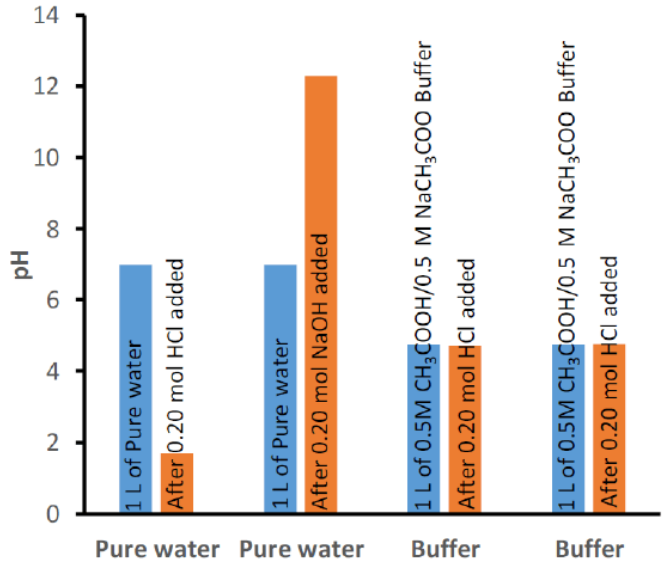


Figure 11.7.1: Acid and base addition on pH change of pure water at pH 7.00 and on acetic acid/sodium acetate buffer at pH 4.74. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.7.1.

Buffers of different initial pH values can be prepared, by varying the ratio of the weak acid to its conjugate base or by using a different set of a weak acid and its conjugate base.

One example is a buffer of initial pH 4.74 comprising 0.5 M acetic acid and 0.5 M sodium acetate shown in Fig. 6.8.1. Another example is a buffer comprising 0.1M dihydrogen phosphate and 0.1M hydrogen phosphate that has an initial pH of 7.21, as shown in Figure 11.7.2.

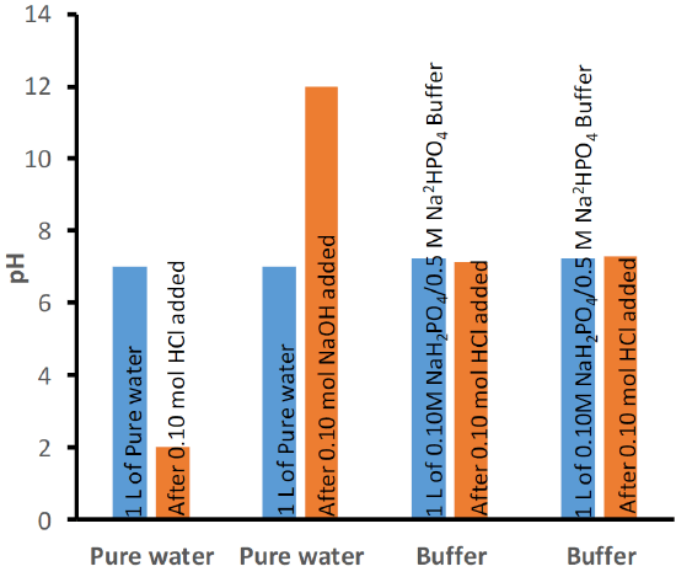


Figure 11.7.2: Effect of acid and base addition on pH change of pure water at pH 7.00 and on Na_2HPO_4/NaH_2PO_4 buffer of pH 7.21. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.7.2.

> A strong acid and its conjugate base or a strong base and its conjugate acid do not make a buffer solution.

How Buffers Work

The essential component of a buffer system is a conjugate acid-base pair whose concentration is fairly high in relation to

the concentrations of added H^+ or OH^- it is expected to buffer against. Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. The acetic acid/sodium acetate buffer has the following equilibrium.

$$CH_3COOH \rightleftarrows H^+ + CH_3COO^-$$

- If a strong base—a source of OH^-(aq) ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$CH_3COOH\,(aq) + OH^-\,(aq) \longrightarrow H_2O(\ell) + CH_3COO^-\,(aq)$$

> 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.

- 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:

$$H^+\,(aq) + CH_3COO^-\,(aq) \longrightarrow CH_3COOH\,(aq)$$

> Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid.

Figure 11.7.3 illustrates both actions of a buffer.

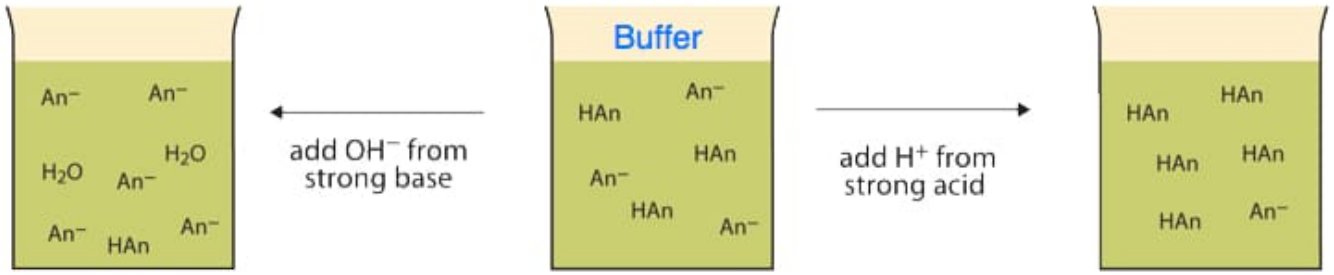


Figure 11.7.3: Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH. (CK-12; Anonymous via LibreTexts). Detailed Description of Figure 11.7.3.

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:

$$NH_3\,(aq) + H^+\,(aq) \longrightarrow NH_4{}^+\,(aq)$$

while the ammonium ion [$NH_4{}^+$(aq)] can react with any hydroxide ions introduced by strong bases:

$$NH_4{}^+\,(aq) + OH^-\,(aq) \longrightarrow NH_3\,(aq) + H_2O\,(l)$$

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(\ell)} \tag{11.7.1}$$

Example 11.7.1

Which solute combinations can make a buffer solution? Assume 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 a salt supplying —formate ion (CHO_2^-), the conjugate base of $HCHO_2$. 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, a weak base. The compound CH_3NH_3Cl is a salt supplying $CH_3NH_3^+$, the conjugate acid of CH_3NH_2. 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.

Buffer Capacity

Buffer capacity refers to how much amount of a strong acid or a strong base the buffer can handle before a drastic change in pH happens.

Higher the amount of weak acid/conjugate base higher the buffer capacity. The buffer that has an equal amount of weak acid and its conjugate base has a higher buffer capacity than the same buffer that has an unequal ratio of the acid and its conjugate base.

Buffers in Living Systems

Buffer solutions are essential components of all living organisms.

- Our blood is buffered to maintain a pH of 7.4 that must remain unchanged as metabolically-generated CO_2 (carbonic acid) is added and then removed by our lungs.
- Lactic acid is produced in our muscles when we exercise. As the lactic acid enters the bloodstream, it is neutralized by the HCO_3^- ion, producing H_2CO_3. An enzyme then accelerates the breakdown of the excess carbonic acid to carbon dioxide and water, which can be eliminated by breathing.
- In addition to the regulating effects of the carbonate buffering system on the pH of blood, the body uses breathing to regulate blood pH. If the pH of the blood decreases too far, an increase in breathing removes CO_2 from the blood through the lungs driving the equilibrium reaction such that [H_3O^+] is lowered. If the blood is too alkaline, a lower breath rate increases CO_2 concentration in the blood, driving the equilibrium reaction the other way, increasing [H^+] and restoring an appropriate pH.
- Buffers in the oceans, in natural waters such as lakes and streams, and within soils help maintain their environmental stability against acid rain and increases in atmospheric CO_2.
- Many industrial processes, such as brewing, require buffer control, as do research studies in biochemistry and physiology that involve enzymes, are active only within certain pH ranges.

The pH in living systems Table 11.7.1 is maintained by buffer systems.

Table 11.7.1 pH in Living Systems

Compartment	pH
Gastric Acid	1
Lysosomes	4.5

Compartment	pH
Granules of Chromaffin Cells$	5.5
Human Skin	5.5
Urine	6
Neutral H_2O at 37 °C	6.81
Cytosol	7.2
Cerebrospinal Fluid	7.3
Blood	7.43-7.45
Mitochondrial Matrix	7.5
Pancreas Secretions	8.1

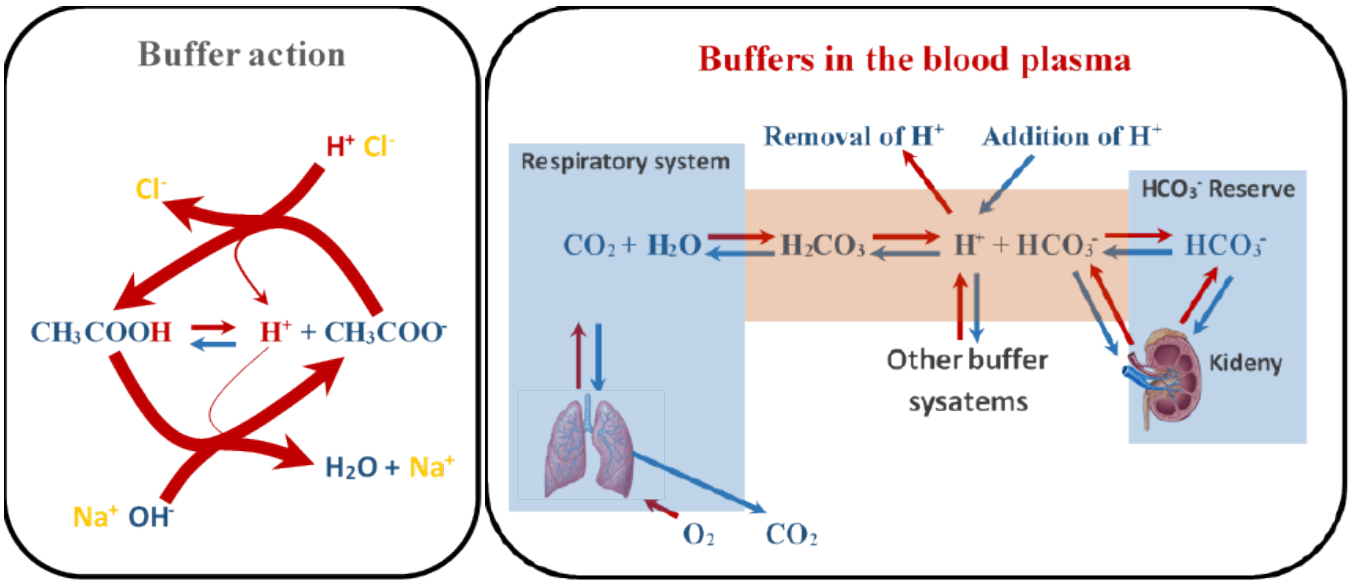
Buffers in the blood

Human blood has a buffering system to minimize extreme changes in pH. The blood maintains its pH of ~7.4 primarily by the carbonic acid/hydrogen carbonate buffer system. The blood pH in the range of 7.45 to 7.35 is considered healthy, but outside of this range causes medical problems. If the blood pH decreases to 6.8 or increases to 8.0, death may occur. It is critical to maintain the blood pH in a narrow range for the cells to function correctly. Specifically, the enzymes and other proteins have secondary, tertiary, and quaternary structures needed for their proper functions. Hydrogen bonding plays a crucial role in defining the protein's structure. pH changes alter the hydrogen bonding making the proteins less effective or ineffective in their functions.

The primary blood buffer system, the carbonic acid/hydrogen carbonate system is illustrated in Figure $11.7.4a$. The primary mechanism for pH regulation by the H_2CO_3/HCO_3^- buffer in the blood is through the lungs.

When the blood pH is acidic compared to the average, the breathing rate increases, exhaling more CO_2 that decreases the concentration of H^+, following the blue arrows in Figure $11.7.4$, thus increasing the pH. The decrease in the breathing rate has the opposite effect.

Kidneys also regulate the blood pH by adding or removing HCO_3^-, but the kidney's response is delayed compared to the response of the lungs.



(a) (b)

Figure 11.7.4: (a) The main blood buffer is H_2CO_3/HCO_3^-. Added acid is consumed by reacting with the HCO_3^- and added base is consumed by consumed by reacting with H_2CO_3, maintaining the pH ~7.4. (b) The lungs regulate the pH by regulating the rate of CO_2 exhale and kidney regulates the pH by regulating the HCO_3^- in the buffer system. (Public Domain; Muhammad Arif Malik via LibreTexts). Detailed Description of Figure 11.7.4.

Chemistry in Everyday Life: The acid that eases the pain

Let's take a look at an acid that is probably the most common medicine: acetylsalicylic acid, also known as aspirin. Aspirin is well known as a pain reliever and antipyretic (fever reducer).

The structure of aspirin is shown in the accompanying figure, Figure 11.7.5. The acid part is circled; it is the H atom in that part that can be donated as aspirin acts as a Brønsted-Lowry acid.

Acetylsalicylic acid is a weak acid. However, it is still an acid, and given that some people consume relatively large amounts of aspirin daily, its acidic nature can cause problems in the stomach lining, despite the stomach's defenses against its own stomach acid.

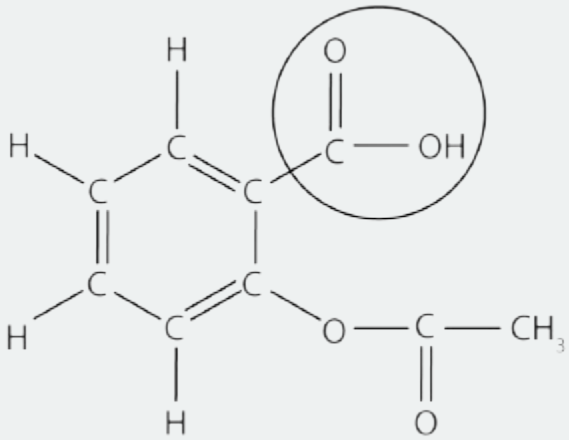


Figure 11.7.5: The circled atoms are the acid part of the molecule. (CC BY-NC-SA 4.0; Anonymous via BCcampus). Detailed Description of Figure 11.7.5.

Because the acid properties of aspirin may be problematic, many aspirin brands offer a "buffered aspirin" form of the medicine, 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. In these cases, the buffering agent regulates the acidity of the aspirin to minimize its acidic side effects.

As useful and common as aspirin is, it was formally marketed as a drug starting in 1899. The US Food and Drug Administration (FDA), the governmental agency charged with overseeing and approving drugs in the United States, wasn't formed until 1906. Some have argued that if the FDA had been formed before aspirin was introduced, aspirin may never have gotten approval due to its potential for side effects-gastrointestinal bleeding, ringing in the ears, Reye's syndrome (a liver problem), and some allergic reactions. However, recently aspirin has been touted for its effects in lessening heart attacks and strokes, so it is likely that aspirin is here to stay.

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 Takeaways

- A buffer is a solution that resists sudden changes in pH.
- Effective buffers consist of a weak acid and its conjugate base, or a weak base and its conjugate acid.
- Biological systems, such as blood, rely on buffer systems like the carbonic acid–bicarbonate pair to function properly.
- Buffer capacity depends on both the concentration and ratio of the buffering components present in the solution.

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- 10.5: Buffers by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

11.8: Summary

The earliest chemical definition of an acid, the Arrhenius definition, says that an acid is a compound that increases the amount of hydrogen ion (H^+) in aqueous solution. An Arrhenius base is a compound that increases the amount of hydroxide ion (OH^-) in aqueous solution. While most bases are named as ionic hydroxide compounds, aqueous acids have a naming system unique to acids. Acids and bases react together in a characteristic chemical reaction called neutralization, in which the products are water and a salt. The principles of stoichiometry, along with the balanced chemical equation for a reaction between an acid and a base, can be used to determine how much of one compound will react with a given amount of the other.

A Brønsted-Lowry acid is any substance that donates a proton to another substance. A Brønsted-Lowry base is any substance that accepts a proton from another substance. The reaction of ammonia with water to make ammonium ions and hydroxide ions can be used to illustrate Brønsted-Lowry acid and base behavior.

Some compounds can either donate or accept protons, depending on the circumstances. Such compounds are called amphiprotic. Water is one example of an amphiprotic compound. One result of water being amphiprotic is that a water molecule can donate a proton to another water molecule to make a hydronium ion and a hydroxide ion. This process is called the autoionization of water and occurs in any sample of water.

Not all acids and bases are equal in chemical strength. A strong acid is an acid whose molecules are all dissociated into ions in aqueous solution. Hydrochloric acid is an example of a strong acid. Similarly, a strong base is a base whose molecules are dissociated into ions in aqueous solution. Sodium hydroxide is an example of a strong base. Any acid or base whose molecules are not all dissociated into ions in aqueous solution is a weak acid or a weak base. Solutions of weak acids and weak bases reach a chemical equilibrium between the un-ionized form of the compound and the dissociated ions. It is a dynamic equilibrium because acid and base molecules are constantly dissociating into ions and reassociating into neutral molecules.

The pH scale is a scale used to express the concentration of hydrogen ions in solution. A neutral solution, neither acidic nor basic, has a pH of 7. Acidic solutions have a pH lower than 7, while basic solutions have a pH higher than 7.

Buffers are solutions that resist dramatic changes in pH when an acid or a base is added to them. They contain a weak acid and a salt of that weak acid, or a weak base and a salt of that weak base. When a buffer is present, any strong acid reacts with the anion of the salt, forming a weak acid and minimizing the presence of hydrogen ions in solution. Any strong base reacts with the weak acid, minimizing the amount of additional hydroxide ions in solution. However, buffers only have limited capacity; there is a limit to the amount of strong acid or strong base any given amount of buffer will react with.

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11.9: Key Terms

- Acids: Substances that increase the concentration of hydrogen ions (H^+) or hydronium ions (H_3O^+) in aqueous solution.
- Bases: Substances that increase the concentration of hydroxide ions (OH^-) or accept hydrogen ions (H^+) in aqueous solution.
- Litmus: a water-soluble mixture of different dyes extracted from lichens. It is often absorbed onto filter paper to produce one of the oldest forms of pH indicator, used to test materials for acidity
- Salts: Ionic compounds formed from the neutralization reaction between an acid and a base.
- Binary Acids: Acids composed of hydrogen and one other nonmetal element, such as HCl or HF.
- Oxyanion: A negatively charged polyatomic ion that contains oxygen and another element, often a nonmetal.
- Neutralize: To react an acid and a base to produce water and a salt, typically bringing the solution closer to a neutral pH.
- Proton: A positively charged subatomic particle found in the nucleus of an atom, and often used to refer to a hydrogen ion (H^+) in acid-base chemistry.
- Ionizable: Describes a substance that can dissociate in solution to produce ions.
- Hydronium ion: The H_3O^+ ion formed when a hydrogen ion (H^+) associates with a water molecule in solution.
- Solvents: The component of a solution that is present in the greatest amount and dissolves the solute.
- Mono-protic: An acid that can donate only one proton (H^+) per molecule.
- Di-protic: An acid that can donate two protons (H^+) per molecule.
- Tri-protic: An acid that can donate three protons (H^+) per molecule.
- Conjugate Acid: The species formed when a base gains a hydrogen ion (H^+).
- Conjugate Base: The species formed when an acid loses a hydrogen ion (H^+).
- Conjugate acid-base pair: Two species that differ by one proton, such as H_2O and OH^- or NH_4^+ and NH_3.
- Amphoteric: A substance that can act as either an acid or a base depending on the chemical environment.
- Amphiprotic: A substance that can both donate and accept a proton (H^+), a type of amphoteric behavior.
- Autoionization of Water: The process by which water molecules spontaneously produce equal concentrations of H_3O^+ and OH^- ions.
- Acid Strength: A measure of how completely an acid dissociates into ions in solution.
- Strong Acid: An acid that completely dissociates into hydrogen ions and its conjugate base in aqueous solution.
- Strong Base: A base that completely dissociates into hydroxide ions and its cation in aqueous solution.
- Weak Acid: An acid that only partially dissociates in solution, existing in equilibrium with its conjugate base.
- Weak Base: A base that only partially accepts protons or produces hydroxide ions in solution.
- Chemical Equilibrium: A state in which the forward and reverse reactions occur at equal rates, and the concentrations of reactants and products remain constant.
- Dynamic Equilibrium: A condition where reactions continue to occur in both directions, but with no net change in concentrations.
- Acid Dissociation Constant: A value that quantifies the extent to which an acid dissociates in solution.
- Base Dissociation Constant: A value that quantifies the extent to which a base accepts protons or generates OH^- in solution.
- pH Scale: A logarithmic scale used to express the acidity or basicity of a solution, ranging from 0 (very acidic) to 14 (very basic), with 7 being neutral.
- Acid-Base Titration: A laboratory technique used to determine the concentration of an acid or base by reacting it with a base or acid of known concentration.
- Analyte: The unknown solution in a titration whose concentration is being determined.
- Standard: A solution of known concentration used in titration to react with the analyte.
- Equivalence Point: The point in a titration at which the number of moles of acid equals the number of moles of base.
- End Point: The observable point in a titration, often indicated by a color change, that signals the equivalence point has been reached or passed.
- Buffers: Solutions that resist significant changes in pH when small amounts of acid or base are added, typically made from a weak acid and its conjugate base.
- Buffer Capacity: The amount of acid or base a buffer can neutralize before a significant change in pH occurs.

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11.10: Alternative Text for Chapter 11

Figure 11.1.1: Closed Lower Esophageal Sphincter Diagram

This medical illustration shows the upper digestive tract, focusing on the esophagus, stomach, and lower esophageal sphincter. The image is a cross-sectional view of the esophagus entering the stomach. The stomach contains blue-tinted fluid representing gastric contents. The lower esophageal sphincter is labeled at the junction of the esophagus and stomach and is shown in the closed position, preventing stomach contents from flowing back into the esophagus. Labels identify the esophagus, stomach, and lower esophageal sphincter.

Figure 11.2.1: Examples of Acids and Alkalines

The image presents a vertical pH scale ranging from 0 to 14, visually depicting the level of acidity or basicity of various substances along with the corresponding hydrogen ion (H^+) concentrations shown in scientific notation. The scale is color-coded, with red at the bottom representing strong acids at pH 0 and blue at the top representing strong bases at pH 14, with a gradual gradient in between.

To the left of the pH scale is a column indicating H^+ concentrations. As the pH value decreases from 14 to 0, the hydrogen ion concentration increases from 10^{-14} to 10^{0}, showing that lower pH values represent higher acidity. The neutral point is marked at pH 7, where the concentration of hydrogen ions is 10^{-7}.

On the right side of the scale, a list of common substances is shown at the corresponding pH levels. At pH 14, sodium hydroxide is listed; at pH 13, household bleach; at pH 11, both ammonia solution and soap; at pH 10, detergent; at pH 9, milk of magnesia; at pH 8, eggs and blood; at pH 7, pure water and milk; at pH 5, coffee and tomato juice; at pH 4, orange juice; at pH 3, soda pop and vinegar; at pH 2, lemon juice; and at pH 0, hydrochloric acid. No examples are provided for pH values 1, 6, and 12.

Each substance is linked to its corresponding location on the pH scale with arrows. On the left side of the image, vertical arrows label the sections of the scale as “Basic” from pH 7 to 14 (pointing upward), “Neutral” at pH 7, and “Acidic” from pH 7 to 0 (pointing downward), emphasizing the transition from neutral to more acidic or basic conditions.

Figure 11.2.2: Litmus Test for Acids and Bases

This image displays the results of a litmus test using four different chemical solutions: sodium hydroxide (NaOH), ammonia (NH_4OH), acetic acid (CH_3COOH), and hydrochloric acid (HCl). Each chemical name is listed above its corresponding chemical formula. Below, two horizontal strips of litmus paper are shown: the upper strip is red litmus, and the lower strip is blue litmus.

Sodium hydroxide causes the red litmus to turn blue, indicating a strong base, while the blue litmus remains unchanged. Ammonia results in a lighter blue color on the red litmus, indicating a weak base, with the blue litmus remaining the same. Acetic acid causes the blue litmus to turn red, indicating a weak acid, with the red litmus showing no change. Hydrochloric acid turns the blue litmus a strong red, indicating a strong acid, while the red litmus remains unchanged. This image illustrates how acids and bases interact with litmus paper to indicate pH.

Figure 11.2.3: Reaction of Ammonia in Water

This image illustrates a Brønsted-Lowry acid-base reaction between ammonia (NH_3) and water (H_2O), demonstrating the transfer of a proton (H^+). On the left side of the reaction, a molecule of ammonia is shown with a lone pair of electrons on the nitrogen atom. It reacts with a water molecule, where the central oxygen atom is bonded to two hydrogen atoms and also has two lone pairs of electrons. A curved arrow indicates the movement of a proton (H^+) from the water molecule to the nitrogen atom of ammonia.

After the reaction, on the right side of the image, the products are an ammonium ion (NH_4^+) and a hydroxide ion (OH^-). The ammonium ion has a positive charge on the nitrogen and four hydrogen atoms bonded to it. The hydroxide ion has a negative charge on the oxygen and one hydrogen atom.

Below the reaction, labels clarify that ammonia acts as the hydrogen ion acceptor and is therefore the Brønsted-Lowry base, while water is the hydrogen ion donor and serves as the Brønsted-Lowry acid. The hydrogen atom involved in the transfer is highlighted in red throughout the image to emphasize its movement during the reaction.

Figure 11.2.4: HCI Dissolving in Water

This image shows a Brønsted-Lowry acid-base reaction between hydrochloric acid (HCl) and water (H_2O), illustrating the transfer of a proton (H^+). On the left side, a hydrogen atom is bonded to a chlorine atom with three lone pairs of electrons. The HCl molecule is shown reacting with a water molecule, which consists of an oxygen atom bonded to two hydrogen atoms and two lone pairs. A curved arrow indicates the movement of the hydrogen ion from HCl to the oxygen atom of the water molecule.

On the right side of the reaction, the products are a chloride ion (Cl^-), shown with four lone pairs of electrons, and a hydronium ion (H_3O^+), which is a water molecule with an extra hydrogen ion bonded to the oxygen, resulting in a positive charge. The transferred proton is colored in red on both sides of the equation to highlight the process. This image demonstrates the acid donating a proton and the base accepting it, forming the conjugate base and conjugate acid.

Figure 11.2.5: Lidocaine

The image shows a small, clear glass vial containing a liquid medication labeled "Lidocaine HCl Injection, USP." The vial has a silver-colored metal cap with a rubber seal and is standing upright on a reflective metallic surface. The label on the vial indicates that it contains 30 milliliters of solution and is marked as "Single-dose" and "Preservative-Free." The concentration of the solution is stated as 1%, which corresponds to 10 milligrams of lidocaine hydrochloride per milliliter. The label also includes additional manufacturer and regulatory information in small print, identifying the product as a local anesthetic commonly used in medical procedures to numb tissue in a specific area.

Figure 11.2.6: Reaction of Ammonia with Water

The image shows a chemical equilibrium between water and ammonia that illustrates the Brønsted–Lowry acid-base reaction concept. On the left side of the reaction, a water molecule (H_2O) is shown donating a proton (H^+) to an ammonia molecule (NH_3). The water molecule is labeled as the parent acid and the ammonia molecule as the parent base. A double-headed arrow between the reactants and products signifies that the reaction is reversible. On the right side, the products are the ammonium ion (NH_4^+), labeled as the conjugate acid, and the hydroxide ion (OH^-), labeled as the conjugate base. Pink text highlights the proton transfer: the water molecule loses a proton (H^+) and becomes OH^-, while the ammonia molecule gains that proton to form NH_4^+. A pink speech bubble emphasizes that the double arrow indicates the reaction is reversible. The image uses curved arrows to show the flow of the proton between the reactants and products, visually representing the acid-base equilibrium and the conjugate acid-base pair relationships.

Figure 11.2.7: Conjugate Acid-Base Pairs of the reaction of Ammonia with Water.

The image shows a visual representation of a Brønsted–Lowry acid-base reaction between water and ammonia, highlighting the formation of conjugate acid-base pairs. On the left side, a water molecule (H_2O) is identified as the acid, and it is shown losing a proton (H^+) to become a hydroxide ion (OH^-), which is labeled as its conjugate base. A molecular model of the water molecule appears as one large red sphere (oxygen) bonded to two smaller white spheres (hydrogen). After the proton is removed, the resulting OH^- ion is shown as a red sphere with a single white sphere, visually simplifying the concept of deprotonation.

On the right side, ammonia (NH_3) is labeled as the base. It gains the proton removed from water to become an ammonium ion (NH_4^+), its conjugate acid. The ammonia molecule is modeled with a large blue sphere (nitrogen) bonded to three smaller white spheres (hydrogens), while the ammonium ion includes an additional hydrogen, forming a total of four bonds around the nitrogen atom. An arrow between each molecular model illustrates the transfer of a proton from water to ammonia.

Below the molecular diagrams, the chemical equation is written: $H_2O(l)$ plus $NH_3(aq)$ in equilibrium with $OH^-(aq)$ and $NH_4^+(aq)$. The terms "acid," "base," "conjugate base," and "conjugate acid" are color-coded and labeled under each species to reinforce the concept of conjugate pairs and the reversible nature of the acid-base reaction. Red is used for acidic components and blue for basic ones, emphasizing the transfer of protons and the dynamic equilibrium between the acid-base pairs.

Figure 11.2.8: Dissociation of HF in Water

The image illustrates the Brønsted–Lowry acid-base reaction between hydrofluoric acid (HF) and water (H_2O), highlighting the exchange of a proton (H^+) and the formation of conjugate acid-base pairs. On the left side, HF is shown as the parent acid with a hydrogen atom bonded to a fluorine atom, which has three lone pairs of electrons. A curved arrow points from the shared hydrogen to the oxygen atom in water, indicating that the hydrogen ion is being transferred from HF to H_2O. Water, depicted with two lone pairs on oxygen and two hydrogen atoms, acts as the parent base.

On the right side of the equilibrium arrow, the products of the reaction are shown: a hydronium ion (H_3O^+) and a fluoride ion (F^-). The hydronium ion is the conjugate acid, formed by the gain of a hydrogen ion by water, and is shown with three hydrogen atoms bonded to oxygen, with a positive charge and one lone pair remaining on the oxygen atom. The fluoride ion is the conjugate base, represented by the fluorine atom now holding an additional electron and bearing a negative charge, with four lone electron pairs.

Text labels beneath the structures indicate the roles of each species: HF is labeled as the parent acid, H_2O as the parent base, H_3O^+ as the conjugate acid, and F^- as the conjugate base. Additional annotations show that HF loses an H^+ and H_2O gains an H^+, visually reinforcing the concept of proton transfer and conjugate pairing. The equation at the top summarizes the reversible reaction as: HF plus H_2O yields H_3O^+ plus F^-. Color coding is used to distinguish between the acid-base roles, with pink and purple highlighting the acid components and blue for the base components.

Figure 11.2.9: Reaction of Acetic Acid with Water

The image shows a Brønsted–Lowry acid-base reaction between acetic acid (CH_3CO_2H) and water (H_2O), illustrating the proton transfer and the formation of conjugate acid-base pairs. On the left side, acetic acid is labeled as the parent acid and is shown with a methyl group (CH_3) attached to a carboxylic acid group (CO_2H). A hydrogen ion (H^+), highlighted in pink, is positioned on the hydroxyl group of the carboxylic acid and is the proton that will be transferred.

Next to acetic acid is a water molecule, labeled as the parent base, represented with two hydrogen atoms bonded to an oxygen atom with two lone pairs. A double-headed arrow between the reactants and products indicates that the reaction is reversible.

On the right side of the equation are the products: the hydronium ion (H_3O^+), labeled as the conjugate acid, and the acetate ion ($CH_3CO_2^-$), labeled as the conjugate base. The hydronium ion has three hydrogen atoms bonded to an oxygen atom, with one of the hydrogens shown in pink to indicate it was gained during the reaction. The acetate ion is shown as the deprotonated form of acetic acid, with a negative charge on one of the oxygen atoms, and the missing hydrogen shown in pink to indicate it was lost.

Text underneath the molecules explains the roles of each species: CH_3CO_2H as the parent acid, H_2O as the parent base, H_3O^+ as the conjugate acid, and $CH_3CO_2^-$ as the conjugate base. The image demonstrates how the acid donates a proton to the base, resulting in the corresponding conjugate acid and base on the product side. Arrows beneath the structures also emphasize the direction of proton loss and gain. The background is light blue, and the hydrogen ion being transferred is consistently shown in pink for clarity.

Figure 11.3.1: Examples of Amphoteric Substances

The image illustrates how certain substances can act either as acids or bases depending on their chemical context. In the center of the image, a vertical gray bracket lists four chemical species—water (H_2O), ammonia (NH_3), hydrogen sulfate ion (HSO_4^-), and hydrogen carbonate ion (HCO_3^-). These species are shown as amphoteric, meaning they can act as either acids or bases.

To the left of the central bracket is a large blue-to-red gradient arrow pointing left and labeled "Act as an acid." It connects to a vertical blue box containing the chemical species hydroxide ion (HO^-), amide ion (NH_2^-), sulfate ion (SO_4^{2-}), and carbonate ion (CO_3^{2-}). These are all conjugate bases that can accept protons and thus, in the presence of stronger bases, can act as acids by donating protons.

To the right of the central bracket is a similar arrow, this time red-to-blue and pointing right, labeled "Act as a base." It connects to a vertical red box listing the hydronium ion (H_3O^+), ammonium ion (NH_4^+), sulfuric acid (H_2SO_4), and carbonic acid (H_2CO_3). These are conjugate acids that can donate protons, and thus, in the presence of stronger acids, they can act as bases by accepting protons.

The gradient coloring of both arrows symbolizes the relative acidity or basicity of the species, highlighting the concept of

conjugate acid-base pairs and the amphoteric nature of the species in the middle. The image visually reinforces the idea that chemical species can behave differently depending on their environment.

Figure 11.3.2: Autoionization of Water

The image shows a beaker containing a liquid that represents water at the molecular level, illustrating the concept of autoionization. Most of the molecules in the beaker are water (H_2O), depicted as two white hydrogen atoms bonded to a red oxygen atom, arranged in a bent shape. These water molecules are distributed throughout the beaker and are connected by dashed lines representing hydrogen bonds between the molecules.

Among the water molecules are several ions. There are a few blue spheres labeled H⊕, representing hydronium ions ($H_3O^{\oplus}$), and a few purple spheres labeled OH⊕, representing hydroxide ions. These ions are spread throughout the solution, interspersed among the water molecules, showing the equilibrium that exists in pure water between its dissociation into hydrogen and hydroxide ions.

The visual emphasizes the low but equal concentrations of H⊕ and OH⊕ ions present in pure water due to the autoionization process. This depiction supports the idea that even in neutral water, there is a dynamic exchange of protons occurring between molecules.

Figure 11.4.1: Acid Strength

The image contains two side-by-side panels labeled (a) and (b), each showing a molecular-level diagram of acid dissociation in solution. Each panel has a gray background and a large downward-pointing blue arrow in the center, indicating dissociation into ions.

Panel (a) illustrates a strong acid in water. The top half shows multiple intact acid molecules, each composed of a large black atom bonded to a smaller red atom. Below the arrow, the bottom half shows complete dissociation: the black spheres (likely representing anions) and red spheres (representing hydrogen ions, H⊕) are all separated and dispersed throughout the solution. This visual emphasizes that strong acids fully dissociate into ions in water.

Panel (b) illustrates a weak acid in water. The top half shows similar intact acid molecules as in panel (a). Below the arrow, the bottom half shows mostly intact acid molecules still bonded together, with only a few red hydrogen ions separated from the parent molecules. This indicates partial dissociation, showing that weak acids only release a small proportion of their hydrogen ions into solution.

The overall comparison demonstrates the difference in ionization between strong acids, which dissociate completely, and weak acids, which dissociate only partially.

Figure 11.4.2: Dissociation of a Strong Acid vs a Weak Acid

The image is a comparative bar graph illustrating the extent of dissociation for hydrochloric acid (HCl) and acetic acid (HAc) in aqueous solution, with the goal of highlighting the difference between strong and weak acids at equilibrium.

On the x-axis, two labeled sections represent the two acids: HCl on the left and HAc on the right. Each section has two sets of vertical bars: the first shows the acid before dissociation, and the second shows the concentrations of species at equilibrium after dissociation.

For HCl, the first dark orange bar labeled "HCl" reaches a height of 100, indicating the initial concentration. In the next group under a shaded area labeled "At Equilibrium," there are two equally tall bars: a yellow bar labeled "H⊕" and a blue bar labeled "Cl⊕", both also reaching 100. This indicates that hydrochloric acid completely dissociates into hydrogen and chloride ions in solution, consistent with its classification as a strong acid.

For HAc, the first brown bar labeled "HAc" also reaches 100, indicating its initial concentration. In the next group, also under a shaded area labeled "At Equilibrium," a shorter brown bar labeled "HAc" remains near the top at about 95, while two much shorter bars are shown beside it: a yellow bar labeled "H⊕" and a blue bar labeled "Ac⊕", each only reaching about 5. This indicates that only a small portion of acetic acid dissociates into hydrogen and acetate ions, demonstrating that it is a weak acid.

The overall comparison shows that strong acids like HCl fully dissociate in water, while weak acids like HAc only partially dissociate, resulting in lower concentrations of free ions at equilibrium.

Figure 11.4.3: Strength of Acid-Conjugate Base Pairs Relative to Water as a Reference

The image is a chart showing a list of acids and their corresponding conjugate bases, organized by their relative acid and base strengths. The chart is divided vertically into two main columns: the left column lists the acids along with their chemical formulas, and the right column lists the conjugate bases formed when those acids lose a hydrogen ion. Both columns are color-coded to indicate the relative strength of each substance.

On the left side of the chart, a vertical arrow labeled "Acid Strength" points upward, showing an increase in acid strength from bottom to top. At the top of the acid list are strong acids like perchloric acid ($HClO_4$), sulfuric acid (H_2SO_4), and hydrochloric acid (HCl). As you move downward, the acids become weaker, transitioning through orange for weak acids like acetic acid (CH_3COOH) and carbonic acid (H_2CO_3), to gray for very weak acids like water (H_2O), ethanol (CH_3CH_2OH), and methane (CH_4).

On the right side of the chart, a vertical arrow labeled "Base Strength" points downward, indicating increasing base strength from top to bottom. Conjugate bases of strong acids, such as perchlorate ion (ClO_4^-), sulfate ion (HSO_4^-), and chloride ion (Cl^-), are near the top and are colored in light gray, indicating very weak bases. As you move downward, the conjugate bases become stronger, transitioning into light blue for weak bases like acetate ion (CH_3COO^-) and carbonate ion (CO_3^{2-}), and eventually into darker blue for strong bases such as ethoxide ion ($CH_3CH_2O^-$), amide ion (NH_2^-), hydride ion (H^-), and methide ion (CH_3^-).

In the middle of the chart, a horizontal section highlights species like hydronium ion (H_3O^+) and water (H_2O) to emphasize their central role in acid-base chemistry. The background shading transitions from red to gray to blue as the acid strength decreases and base strength increases, illustrating the inverse relationship between acid strength and the strength of its conjugate base. Overall, the chart visually reinforces the concept that strong acids have very weak conjugate bases, and weak acids have relatively stronger conjugate bases.

Figure 11.4.4: Safety of Acids and Bases

The image consists of two side-by-side photographs labeled (a) and (b), illustrating the dangers of handling strong bases without proper protection. On the left, image (a) shows two white plastic bottles with white caps placed on a wooden surface. Each bottle has a label clearly marked "caustic soda," which is a common name for sodium hydroxide, a strong base. Warning symbols on the labels indicate that the contents are corrosive and hazardous. On the right, image (b) displays a human hand with severe chemical burns and visible redness, irritation, and raw skin, particularly on the back of the hand and fingers. A bandage is wrapped around the index finger, suggesting an injury. The hand is positioned against a white wall, emphasizing the extent of skin damage. The juxtaposition of the two images highlights the corrosive and dangerous nature of caustic soda, underscoring the importance of proper safety precautions when handling strong chemical bases.

Figure 11.5.1: Autoionization of Water

The image visually represents the autoionization of water through a molecular diagram and a chemical equation. On the left side of the equation are two water molecules, each consisting of one large red sphere representing an oxygen atom and two smaller light gray spheres representing hydrogen atoms. These two water molecules are shown combining in a reversible reaction, as indicated by a double-headed arrow pointing both forward and backward. On the right side of the equation, one product is a hydronium ion (H_3O^+), which is depicted as a red oxygen atom bonded to three gray hydrogen atoms, carrying a positive charge. The second product is a hydroxide ion (OH^-), shown as a red oxygen atom with a single hydrogen attached, and it carries a negative charge. The visual demonstrates the dynamic equilibrium in which water molecules spontaneously dissociate into hydronium and hydroxide ions, a fundamental concept in acid-base chemistry.

Figure 11.5.2: The Dissociation of Water

The image is a structural diagram illustrating a Brønsted-Lowry acid-base reaction using Lewis dot structures and curved arrows. On the left side of the equation are two water molecules. The molecule labeled "Base" shows an oxygen atom with two lone pairs and two hydrogen atoms, one of which is donating a lone pair to a hydrogen ion (H^+) on the second water molecule labeled

"Acid." Curved arrows indicate the flow of electrons: one arrow from the lone pair on the base oxygen to the hydrogen on the acid, and another from the hydrogen-oxygen bond on the acid to its own oxygen, forming a new lone pair.

The products on the right side are labeled "Conjugate acid" and "Conjugate base." The conjugate acid is a hydronium ion (H_3O^+), where the base has accepted a proton, now showing three bonded hydrogen atoms and one lone pair on the oxygen. The conjugate base is a hydroxide ion (OH^-), which results from the acid losing a proton and now has one hydrogen and three lone pairs on the oxygen. Red and blue arrows above the structures visually indicate the relationships between acids and conjugate bases, and bases and conjugate acids, respectively. The double-headed arrow in the middle signifies that this reaction is reversible.

Figure 11.5.3: The Relationship between (H^+) and pH Values for Several Common Materials

The image shows a vertical pH scale that ranges from 0 to 14, accompanied by corresponding hydrogen ion (H^+) concentrations and everyday examples of substances at different pH levels. The left side of the image lists H^+ concentrations in scientific notation, from 10^0 at pH 0 (very acidic) to 10^{-14} at pH 14 (very basic), illustrating that acidity increases as pH decreases. The middle column is the pH scale itself, with pH 7 labeled as neutral. On the right side, various substances are placed at their approximate pH levels: hydrochloric acid at pH 0, lemon juice and vinegar near pH 2, soda pop and orange juice around pH 3 to 4, coffee and tomato juice at pH 5, milk and pure water near pH 6 to 7, eggs and blood slightly above neutral, milk of magnesia around pH 9, ammonia solution and soap near pH 11, and sodium hydroxide and bleach close to pH 13–14. Large arrows on the left side point upward to indicate increasing basicity and downward to indicate increasing acidity, with the central area at pH 7 labeled as neutral. This visual illustrates how pH values relate to H^+ concentration and provides context by connecting familiar substances to their approximate positions on the scale.

Figure 11.5.4: pH Scale and pH of Some Common Substances

The image shows a vertical pH scale ranging from 0 to 14, with an arrow pointing upward for increasing acidity and downward for increasing basicity. Labels divide the scale into three regions: acidic from pH 0 to below 7, neutral at pH 7, and basic from above 7 to 14.

Along the scale, common substances are listed with their approximate pH values. At the top, battery acid is placed near pH 0. Lemon juice and vinegar appear around pH 2, while soda and grape juice are placed between pH 2.5 and 3.5. Black coffee is around pH 5. Milk, urine, and saliva fall within pH 6.3 to 6.6. Blood is shown at pH 7.4, just slightly above neutral. Seawater is labeled at pH 8, and baking soda is shown at pH 9.5. Window cleaning ammonia has a pH range of 10.5 to 11.5. Oven cleaner and lye are listed around pH 13.5, and drain cleaner is noted at pH 14.

The visual uses icons or pictures next to each substance to reinforce their identity: a lemon for lemon juice, a soda glass for soft drinks, a coffee cup for black coffee, a blood drop for blood, a milk carton and water bottle for milk and urine, a wave for seawater, a spoonful of powder for baking soda, a spray bottle and a person cleaning for ammonia and oven cleaner, and a drain symbol for drain cleaner. The diagram helps relate everyday items to their place on the pH scale and illustrates the relative acidity or basicity of common substances.

Figure 11.5.5: The pH in Various Human Body Fluids.

The image is a medical-style illustration of a human female body shown from the front with internal organs partially visible to illustrate the pH levels in various parts of the body. The pH values of different fluids and organs are labeled, indicating their acidity or alkalinity. The illustration includes the following:

Saliva is shown in the mouth and is labeled with a pH range of 5.8 to 7.1, indicating it can be slightly acidic to neutral. The esophagus is labeled but no specific pH is provided for it. Blood is labeled near the chest area with a pH of 7.4, showing it is slightly basic and tightly regulated. The stomach, highlighted in pink with its internal area visible, is labeled with a pH range of 1.6 to 1.8, reflecting its highly acidic environment necessary for digestion. The liver is shown but not labeled with a pH. The pancreas, located behind the stomach and near the upper left abdomen, is labeled with a pH range of 7.5 to 8.8, indicating it secretes basic fluids to neutralize stomach acid. The small intestine is shown below the stomach and labeled with a pH of 8.5, indicating an alkaline environment suitable for enzymatic digestion. The large intestine is labeled with a pH range of 5.0 to 7.0, reflecting a slightly acidic to neutral environment. Finally, urine is labeled near the lower abdomen with a pH range of 4.6 to 8.0, showing that

urine pH can vary significantly depending on diet, hydration, and health status.

The image is labeled with "National Cancer Institute" in the lower right corner, indicating the source of the diagram. The illustration provides a clear overview of the varying pH levels throughout different regions and fluids of the human body, highlighting the importance of pH in physiological processes.

Figure 11.6.1: Proton Transfer from HCI to NH_3

A gloved hand holds a test tube labeled “NH_3” (ammonia) and tilts it toward a small beaker labeled “HCl” (hydrochloric acid), which sits on a dark lab surface. As the two containers approach each other, a visible white smoke or vapor forms in the space between them. This reaction represents the combination of gaseous ammonia (a base) and hydrochloric acid (an acid), producing ammonium chloride (NH_4Cl), which appears as a white smoke due to the formation of solid particles in the air. The scene demonstrates a classic acid-base reaction and highlights the formation of a salt through gas-phase neutralization.

Figure 11.6.2: Example of a Neutralization Reaction

Two raw fish are laid side by side on a sheet of aluminum foil, each stuffed with several slices of lemon. Coarse salt is sprinkled over the fish. Beneath the image is a chemical equation that explains an acid-base reaction relevant to the scene: acetic acid (CH_3COOH) reacts with putrescine ($NH_2CH_2CH_2CH_2CH_2NH_2$), a foul-smelling compound produced during fish spoilage. The reaction produces acetate ion ($CH_3COO^{\ominus}$) and the putrescinium ion ($NH_3^{\oplus}CH_2CH_2CH_2CH_2NH_2$), a protonated, less volatile form of putrescine. This reaction demonstrates how the citric acid in lemon neutralizes the fishy odor by converting the basic putrescine into a non-smelly salt form, offering a real-life application of acid-base chemistry in cooking.

Figure 11.6.3: Concept Map of the Calculations

The image is a step-by-step visual flowchart illustrating how to perform a stoichiometric calculation starting with the mass of nitric acid (HNO_3) and ending with the mass of calcium hydroxide ($Ca(OH)_2$) required for a chemical reaction. It is arranged as a horizontal sequence of colored rectangular boxes connected by labeled arrows, each representing a calculation step.

The first red box on the left is labeled “mass of HNO_3.” An arrow pointing to the right from this box is labeled “molar mass,” leading to a purple box labeled “moles of HNO_3.” Underneath the arrow is a conversion factor written as “1 mol HNO_3 / 63.02 g HNO_3,” which indicates the use of HNO_3's molar mass to convert grams to moles.

Another arrow continues from the purple box to the right, labeled “stoichiometric ratio,” pointing to a dark blue box labeled “moles $Ca(OH)_2$.” The conversion factor beneath this arrow reads “1 mol $Ca(OH)_2$ / 2 mol HNO_3,” which reflects the balanced chemical reaction between nitric acid and calcium hydroxide, where one mole of calcium hydroxide reacts with two moles of nitric acid.

From the dark blue box, another arrow points to the final green box labeled “mass $Ca(OH)_2$,” with the arrow labeled “molar mass.” The conversion factor under this arrow states “74.10 g $Ca(OH)_2$ / 1 mol $Ca(OH)_2$,” which shows the molar mass used to convert moles of calcium hydroxide into grams.

Overall, the diagram provides a clear path for solving stoichiometry problems involving neutralization reactions, starting from a given mass of nitric acid and ending with the required mass of calcium hydroxide.

Figure 11.6.4: Concept Map of the Calculations

The image is a horizontal flowchart illustrating how to calculate the volume of nitric acid (HNO_3) solution required to react with a given volume of calcium hydroxide ($Ca(OH)_2$) solution, using molarity and stoichiometric relationships. The sequence consists of five colored rectangular boxes, each representing a different step in the calculation. Arrows between the boxes are labeled with conversion factors and chemical relationships used in each step.

The first box on the far left is light brown and labeled “Liters $Ca(OH)_2$ solution.” An arrow points to the right from this box and is labeled “Molarity conversion factor.” This arrow leads to a purple box labeled “moles $Ca(OH)_2$.” Under this arrow is the conversion factor “0.672 mol $Ca(OH)_2$ / 1 L $Ca(OH)_2$ solution,” which indicates that 1 liter of the calcium hydroxide solution contains 0.672 moles of $Ca(OH)_2$.

The next arrow points from the purple box to a dark blue box labeled “moles HNO_3.” This arrow is labeled “Stoichiometric ratio” and includes the conversion factor “2 mol HNO_3 / 1 mol $Ca(OH)_2$.” This represents the balanced chemical equation, showing that two moles of nitric acid are needed to react with one mole of calcium hydroxide.

Continuing to the right, an arrow labeled “Molarity conversion factor” leads from the dark blue box to the final orange box labeled “Liters HNO_3 solution.” The conversion factor under this arrow reads “1 L HNO_3 solution / 0.432 mol HNO_3,” meaning that one liter of the nitric acid solution contains 0.432 moles of HNO_3.

Altogether, the image demonstrates how to go from a known volume of $Ca(OH)_2$ solution to the volume of HNO_3 solution needed for complete neutralization, passing through steps of mole calculation and stoichiometric ratios.

Figure 11.6.5: Acid-Base Titration Set-UP

The image shows a three-step illustration of an acid-base titration process using burettes and Erlenmeyer flasks, progressing from left to right. In the first stage, the burette is positioned above a flask containing a light blue solution, which is being titrated. A few drops are being added from the burette. In the second stage, the burette continues to dispense liquid into the flask, which now has a slightly lighter hue, suggesting that the titration is approaching the equivalence point but has not yet reached it. In the third stage, the flask contains a dark purple solution, indicating the endpoint of the titration has been reached due to the presence of a pH indicator that changes color when neutralization occurs. The burettes are consistently releasing liquid into the flasks in a controlled, drop wise manner. This series visually demonstrates the color change that occurs in the analyte as it reacts with the titrant and signals completion of the titration reaction.

Figure 11.6.6: Mg Reacting with HCI and H_2 Gas Bubbling Out

The image shows the lower portion of a clear glass test tube containing a colorless liquid. A vigorous chemical reaction is occurring near the surface of a solid substance submerged in the liquid, producing a thick layer of gas bubbles that cling to the sides of the tube and rise to the surface. The reaction appears to generate a large amount of gas, indicating the possible production of a gas such as carbon dioxide. This type of bubbling is commonly observed during acid-base reactions involving carbonates or metals, such as the reaction between hydrochloric acid and a metal carbonate. The background is a plain light brown color, providing contrast to clearly highlight the reaction taking place inside the test tube.

Figure 11.6.7: The reaction of Sodium Carbonate with HCI

The image shows two clear test tubes placed side by side against a warm-toned background. Both contain a white solid reacting with a colorless liquid, illustrating the reaction between sodium carbonate and hydrochloric acid. The left test tube shows the initial stage of the reaction, where a layer of white sodium carbonate sits at the bottom and a dense stream of tiny gas bubbles rapidly rises through the liquid. This indicates vigorous carbon dioxide gas evolution. The right test tube shows a slightly later stage of the reaction, where the tube is filled with a thick cloud of gas bubbles, and foam or froth has formed throughout the liquid. The reaction between sodium carbonate and hydrochloric acid produces carbon dioxide gas, water, and a salt, which is visually represented by the bubbling and effervescence seen in both tubes.

Figure 11.6.8: Bottle of Antacid Tablets

The image shows a plastic bottle of antacid tablets lying on its side on a smooth white surface, with its open cap facing downward. A variety of round, chewable antacid tablets have spilled out of the container and are scattered on the surface. The tablets are pastel-colored in shades of yellow, pink, green, and blue. The label on the bottle identifies the product as "Antacid Tablets" with calcium carbonate as the active ingredient, and shows it is a generic brand equivalent to Tums. The packaging also features images of fruit flavors and indicates the tablets are used for relieving heartburn and acid indigestion. This image visually represents over-the-counter antacids, which chemically react with excess hydrochloric acid in the stomach.

Figure 11.6.9: Illustration of Acid Rain

The image is a labeled diagram illustrating the process of acid rain formation and its environmental impact. On the left side of the diagram, a factory with smoke stacks is shown releasing gray clouds of acidic gases, including sulfur dioxide and nitrogen oxides,

into the atmosphere. These gases are represented by upward-pointing arrows and are labeled as acidic gases released into the atmosphere. The gases are then carried by the wind to another area, indicated by curved arrows pointing toward clouds. In the clouds, the gases dissolve in rainwater, forming acid rain. This is shown with rain falling from the clouds onto a landscape of hills, trees, and a river. The rain is marked as acid rain, which affects the environment by killing plant life, polluting rivers and streams, and eroding stonework. The entire scene visualizes the chain reaction beginning with pollution and ending with environmental damage, emphasizing the harmful effects of industrial emissions on ecosystems. The image shows the consequences of acid-base reactions on a large environmental scale.

Figure 11.6.10: Effects of Acid Rain

The image contains two separate photographs labeled (a) and (b), each illustrating the environmental effects of acid rain. Image (a), on the left, shows a forest filled with tall, dead, or dying trees that have lost their leaves and branches, with only bare trunks and skeletal limbs remaining. The ground below is green, but the forest appears heavily damaged, with most trees stripped of foliage, suggesting long-term exposure to acidic conditions. Image (b), on the right, shows a close-up of a stone statue featuring a child reaching out in distress. The statue is significantly weathered and damaged, with visible surface erosion and loss of detail in the sculpture's features. The background shows green leafy trees. Together, the images highlight how acid rain negatively impacts both natural ecosystems, like forests, and man-made structures, such as stone monuments.

Figure 11.7.1: Buffer Effects 1

The image is a bar graph comparing the changes in pH of pure water and a buffer solution after the addition of hydrochloric acid (HCl) or sodium hydroxide (NaOH). The vertical axis represents pH, ranging from 0 to 14. The horizontal axis labels five different test conditions. The first bar shows that 1 liter of pure water has a pH of about 7. The second bar shows that after adding 0.20 moles of HCl to pure water, the pH drops dramatically to around 1.5. The third bar shows another sample of 1 liter of pure water with a pH of about 7, and the fourth bar shows that after adding 0.20 moles of NaOH to it, the pH increases significantly to about 13. The fifth bar represents a buffer solution composed of 0.5 M acetic acid (CH_3COOH) and 0.5 M sodium acetate ($NaCH_3COO$), which has an initial pH around 4.8. The sixth bar shows that after adding 0.20 moles of HCl to this buffer, the pH decreases only slightly. The seventh bar, another sample of the same buffer, starts again at about pH 4.8, and the eighth bar shows that after adding 0.20 moles of NaOH, the pH increases only slightly. The overall visual demonstrates that the buffer resists drastic changes in pH upon addition of strong acid or base, while pure water experiences significant pH shifts.

Figure 11.7.2: Buffer Effects 2

The image is a bar graph illustrating the effect of adding hydrochloric acid (HCl) and sodium hydroxide (NaOH) to pure water and to a phosphate buffer solution. The vertical axis represents pH, ranging from 0 to 14. The horizontal axis labels four experimental conditions involving pure water and a buffer solution made of 0.10 M sodium dihydrogen phosphate (NaH_2PO_4) and 0.5 M disodium hydrogen phosphate (Na_2HPO_4). The first pair of bars shows that 1 liter of pure water has an initial pH of around 7, and after adding 0.10 mol of HCl, the pH drops sharply to around 1. The second pair of bars shows another sample of 1 liter of pure water with an initial pH of 7, which rises to approximately 12 after adding 0.10 mol of NaOH. The third pair of bars shows the phosphate buffer solution with an initial pH around 7. After adding 0.10 mol of HCl to the buffer, the pH remains nearly unchanged, staying around 6.9. The fourth pair of bars shows another sample of the same phosphate buffer with an initial pH around 7. After adding 0.10 mol of NaOH, the pH again remains nearly constant, just above 7. This graph visually demonstrates the buffering capacity of the phosphate buffer, which resists drastic pH changes when small amounts of strong acid or base are added, unlike pure water which experiences significant shifts in pH.

Figure 11.7.3: The Action of Buffers

The image is a three-part diagram that illustrates how a buffer solution resists changes in pH when small amounts of strong acid or strong base are added. The central panel represents a buffer solution in its initial state, labeled as "Buffer," and shows a balanced presence of weak acid molecules, represented by HAn, and their conjugate base, represented by An minus. The right panel shows the effect of adding hydrogen ions (H^+) from a strong acid to the buffer. After the addition, the solution contains more HAn molecules and fewer An minus ions, demonstrating that the added H^+ reacts with An minus to form HAn, thus minimizing the increase in acidity. The left panel shows the effect of adding hydroxide ions (OH^-) from a strong base. After the addition, the number of An minus ions increases and water molecules, labeled H_2O, appear in the solution, while HAn is slightly reduced. This indicates that the added OH^- reacts with HAn to form An minus and water, minimizing the increase in basicity.

The background of the image is blue, and the overall setup visually communicates how buffers work to maintain pH stability in both acidic and basic conditions.

Figure 11.7.4: Mechanism of Blood Buffer Action

The image is divided into two panels labeled (a) and (b), both illustrating buffer systems. Panel (a), titled "Buffer action," shows how acetic acid (CH_3COOH) acts as a buffer in the presence of strong acids and bases. The reversible central reaction is CH_3COOH in equilibrium with hydrogen ions (H^+) and acetate ions (CH_3COO^-). When hydrochloric acid (HCl) is added, represented by H^+ and Cl^- ions, the extra H^+ is neutralized by CH_3COO^- to form CH_3COOH. When sodium hydroxide (NaOH) is added, represented by Na^+ and OH^- ions, the OH^- neutralizes H^+ to form water, and CH_3COOH dissociates to replenish the lost H^+, keeping the pH stable. Panel (b), titled "Buffers in the blood plasma," illustrates the bicarbonate buffer system in human physiology. Carbon dioxide (CO_2) from the respiratory system combines with water (H_2O) to form carbonic acid (H_2CO_3), which is in equilibrium with hydrogen ions (H^+) and bicarbonate ions (HCO_3^-). When H^+ is removed, the system shifts to produce more H^+ and HCO_3^-. When H^+ is added, it combines with HCO_3^- to form H_2CO_3, reducing free H^+ and maintaining pH balance. The diagram shows organs such as lungs and kidneys that regulate this system, with lungs exchanging CO_2 and O_2 and kidneys helping to regulate bicarbonate ion levels. The diagram also refers to other buffer systems that assist in stabilizing blood pH.

Figure 11.7.5: The Molecular Structure of Aspirin

The image shows the structural formula of acetylsalicylic acid, commonly known as aspirin. The molecule consists of a benzene ring, which is a hexagonal ring made of six carbon atoms with alternating double bonds, and each carbon is bonded to a hydrogen atom or a functional group. Attached to one of the carbon atoms of the benzene ring is a carboxylic acid group, consisting of a carbon double bonded to an oxygen and also bonded to a hydroxyl group, shown as COOH. This group is circled in the image to highlight it. Attached to another carbon of the benzene ring is an ester group. This ester group is made up of an oxygen atom bonded to a carbon that is double bonded to another oxygen and also bonded to a methyl group (CH3). The ester group connects to the ring through an oxygen atom, indicating the molecule is an ester of salicylic acid. This entire structure represents aspirin, with its key functional groups labeled and visually emphasized.

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12: NUCLEAR CHEMISTRY

Introduction

In the chemical reactions we have studied so far, changes occur primarily in the electronic structure of atoms and molecules—how electrons are arranged around the nucleus. The nucleus itself, composed of protons and neutrons, remains unchanged in these chemical processes. Nuclear chemistry, by contrast, focuses on reactions that involve changes within the atomic nucleus. These changes can alter the number of protons, neutrons, or both, leading to the formation of new elements or isotopes and the release or absorption of significant amounts of energy.

The study of nuclear chemistry began with the discovery of radioactivity in 1896 by Antoine Becquerel, and it has since evolved into a critical field with applications spanning energy production, medical diagnostics and treatments, dating of archeological and geological materials, and industrial uses. This chapter will explore the fundamental concepts of nuclear chemistry, including nuclear stability, mass defect, binding energy, and the various modes of radioactive decay. We will also examine nuclear reactions such as fission and fusion, the production of artificial isotopes, and the use of radioactive tracers in research and medicine.

An important aspect of nuclear chemistry is understanding the biological effects of ionizing radiation and the techniques used for detection, measurement, and protection. By studying these principles, we gain insight not only into the underlying science of the atom but also into the practical applications and safety considerations associated with nuclear processes.

Nuclear chemistry plays a crucial role in modern medical imaging and diagnostics. Figure 12.1 shows positron emission tomography (PET) scans that utilize radioactive tracers emitting positrons, allowing clinicians to visualize physiological processes in real time. In the image on the left, PET imaging highlights muscle activity, providing insights into metabolism and function. The central images focus on the brain, revealing differences in dopamine signaling between addicts and non-addicts, which can help in understanding neurological and psychological conditions. The rightmost images demonstrate an oncological application, where PET scans identify lymph node metastasis, aiding in the diagnosis, staging, and treatment planning of cancer. These applications illustrate how nuclear chemistry directly contributes to personalized medicine, early detection of diseases, and monitoring treatment effectiveness.

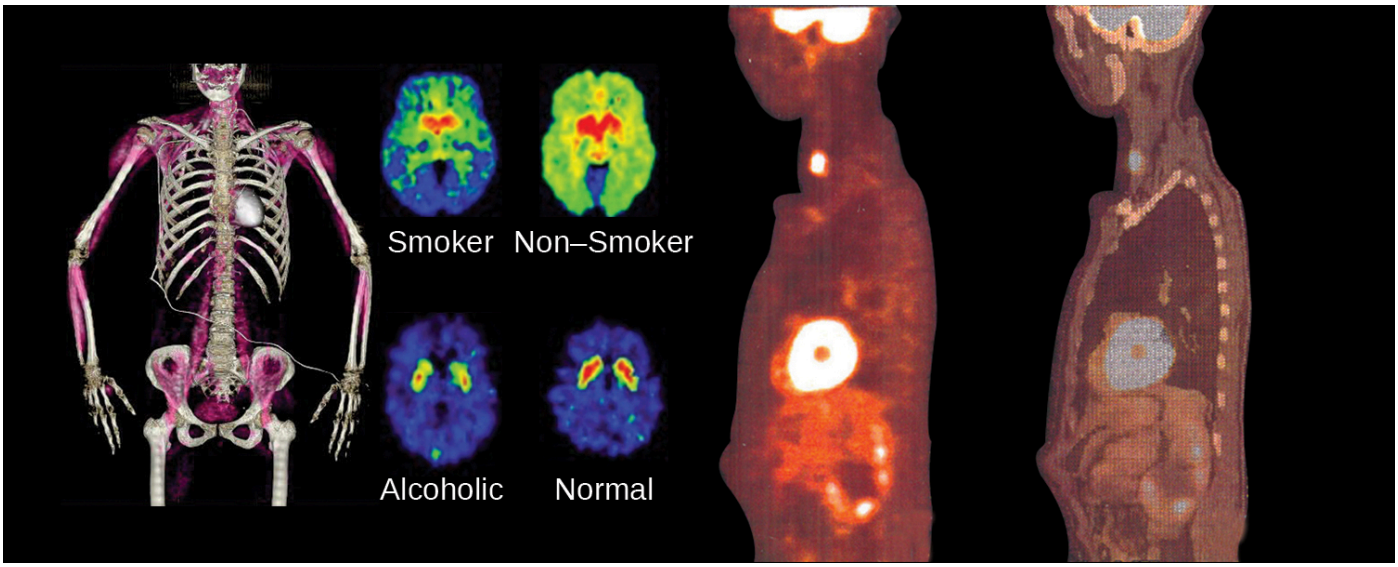


Figure 12.1: Nuclear chemistry provides the basis for many useful diagnostic and therapeutic methods in medicine. (CC BY 4.0; anonymous via LibreTexts).

By the end of this chapter, you will be able to explain how nuclear reactions differ from chemical reactions, understand the factors that govern nuclear stability, describe common types of decay and their applications, and evaluate both the benefits and risks associated with the use of radioactive materials.

Learning Objectives

- Explore the fundamental concepts of radioactivity and types of radiation.
- Understand nuclear reactions and practice writing balanced nuclear equations.
- Learn the units used to measure radioactivity and radiation exposure.
- Apply the concept of half-life to perform decay calculations.
- Examine radioactive decay in dating materials and its limitations.
- Study applications of radioisotopes in medicine and industry, and radiation safety.

12.1: Prelude to Radioactivity
12.2: Nuclear Reactions
12.3: Half-Life
12.4: Applications of Radioactive Isotopes
12.5: Biological Effects of Radiation
12.6: Summary
12.7: Key Terms
12.8: Alternative Text for Chapter 12

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12.1: Prelude to Radioactivity

Introduction

Atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of the atom, while electrons circle the nucleus in shells and subshells and account for an atom's size. Remember, the notation for succinctly representing an isotope of a particular atom:

$$^{12}_{6}\mathrm{C} \tag{12.1.1}$$

The element in this example, represented by the symbol *C*, is carbon. Its atomic number, 6, is the lower left subscript on the symbol and is the number of protons in the atom. The mass number, the superscript to the upper left of the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 12, which means that the number of neutrons in the atom is 12 – 6 = 6 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neutrons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. The two isotopes of hydrogen, ^{2}H and ^{3}H, are given their own names: deuterium (D) and tritium (T), respectively. Another way of expressing a particular isotope is to list the mass number after the element name, like carbon-12 or hydrogen-3.

Atomic theory in the 19th century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source as the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an **alpha particle**, which is symbolized by the Greek letter α. An alpha particle is composed of two protons and two neutrons, and so it is the same as a helium nucleus. It has a 2+ charge.

The second major type of radioactive emission is called a **beta particle**, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus, not from the shells of electrons about the nucleus! Because a beta particle is an electron, it has a 1– charge and we can represent it as $_{-1}^{0}\mathrm{e}$.

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called **gamma rays**, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*).

Penetration Power of Radioactive Emissions

Alpha, beta, and gamma emissions have different abilities to penetrate matter (Figure 12.1.1). The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter.

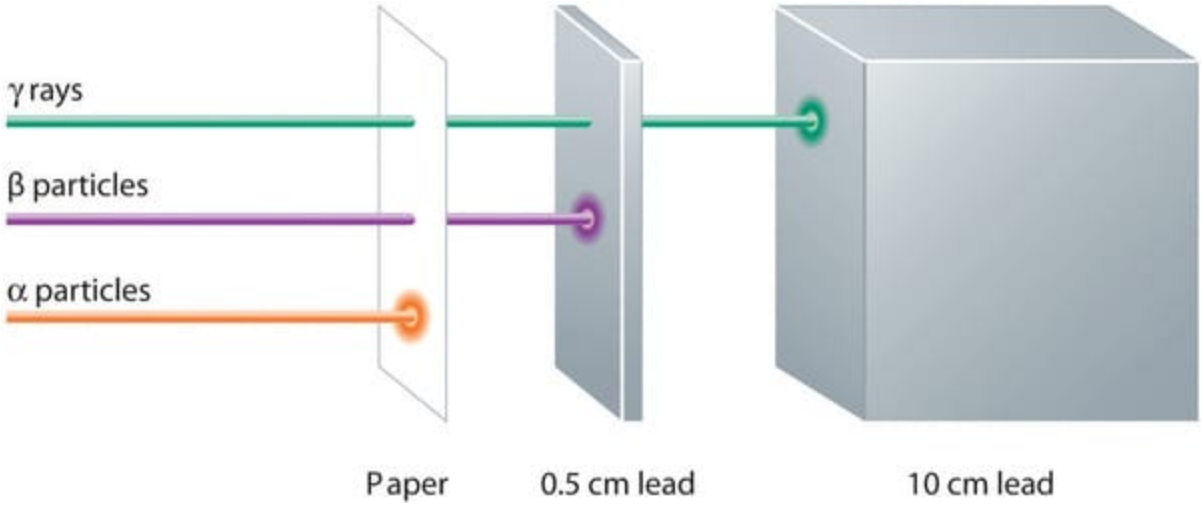


Figure 12.1.1: Alpha particles are stopped by paper, beta particles by thin lead, while gamma rays require thick lead shielding to be absorbed. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 12.1.1.

Table 12.1.1 summarizes the properties of the these main types of radioactive emissions by showing their symbols, particle identity, electrical charge, relative mass number, and typical penetration power. Alpha particles are helium nuclei with a +2 charge and relatively large mass, making them the least penetrating and easily stopped by paper or skin. Beta particles are high-energy electrons or positrons with a –1 charge and negligible mass number, allowing them to penetrate more deeply than alpha particles, though they can still be stopped by materials such as aluminum foil. Gamma rays, unlike the other two, are high-energy electromagnetic radiation with no mass or charge, giving them extremely high penetrating power, requiring dense materials such as lead or thick concrete to provide effective shielding.

Table 12.1.1: The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, $^{4}_{2}\mathrm{He}$	β, $^{0}_{-1}\mathrm{e}$	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1–	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called **spontaneous fission** (or fission). Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\mathrm{U} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 2{}^{1}_{0}\mathrm{n} \tag{12.1.2}$$

where $^{1}_{0}\mathrm{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and the mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208.

Key Takeaway

- Radioactivity is the emission of energy or particles from unstable nuclei.
- Atoms are made of protons, neutrons, and electrons; isotopes differ by their number of neutrons.
- Radioactivity, first discovered by Henri Becquerel in 1896, is the spontaneous emission of particles or energy from the atomic nucleus.
- The three main emissions—alpha, beta, and gamma—differ in charge, mass, and penetration power, with alpha least

penetrating and gamma most penetrating.

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12.2: Nuclear Reactions

Introduction

As discussed earlier, radioactivity is the spontaneous emission of particles or energy from unstable nuclei, with the three main types being alpha, beta, and gamma emissions. In this section, we will look more closely at these processes by examining their nuclear reactions, the nuclides that undergo them, and the corresponding equations that represent each type of decay.

Alpha Particle

The first is called an alpha particle, which is symbolized by the Greek letter α. An alpha particle is composed of two protons and two neutrons, and so it is the same as a helium nucleus. (We often use $^{4}_{2}\mathrm{He}$ to represent an alpha particle.) It has a 2+ charge. When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by two (because of the loss of two protons), and its mass number decreases by four (because of the loss of four nuclear particles). We can represent the emission of an alpha particle with a chemical equation—for example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{231}_{90}\mathrm{Th} \tag{12.2.1}$$

How do we know that a product of the reaction is $^{231}_{90}\mathrm{Th}$? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means we must have the same number of protons and neutrons on both sides of the chemical equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose 4 nuclear particles of the original 235, there are 231 remaining. Thus, we use subtraction to identify the isotope of the thorium atom—in this case, $^{231}_{90}\mathrm{Th}$.

Chemists often use the names **parent isotope** and **daughter isotope** to represent the original atom and the product other than the alpha particle. In the previous example, $^{235}_{92}\mathrm{U}$ is the parent isotope, and $^{231}_{90}\mathrm{Th}$ is the daughter isotope. When one element changes into another in this manner, it undergoes **radioactive decay**.

Example 12.2.1

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as $^{222}_{86}\mathrm{Rn}$. We represent the alpha particle as $^{4}_{2}\mathrm{He}$ and use subtraction (222 – 4 = 218 and 86 – 2 = 84) to identify the daughter isotope as an isotope of polonium, $^{218}_{84}\mathrm{Po}$:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{218}_{84}\mathrm{Po}$$

Example 12.2.2

Write the nuclear equation that represents the radioactive decay of polonium-209 by alpha particle emission and identify the daughter isotope.

Solution

$$^{209}_{84}\mathrm{Po} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{205}_{82}\mathrm{Pb}$$

Beta Particle

The second major type of radioactive emission is called a beta particle, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1– charge. We can also represent a beta particle as ${}^{0}_{-1}\mathrm{e}$ or β^{-}. The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$ {}^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e} \tag{12.2.2} $$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the electron is assigned an “atomic number” of 1–, equal to its charge.)

Gamma Rays

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called gamma rays, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed ionizing radiation). For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$ {}^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{Po} + {}^{4}_{2}\mathrm{He} + \gamma \tag{12.2.3} $$

This may not seem like much energy, but if 1 mol of radon atoms were to decay, the gamma ray energy would be 49 million kJ!

Example 12.2.3

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is ${}^{12}_{5}\mathrm{B}$ while one of the products is an electron, ${}^{0}_{-1}\mathrm{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus, the complete nuclear equation is as follows:

$$ {}^{12}_{5}\mathrm{B} \rightarrow {}^{12}_{6}\mathrm{C} + {}^{0}_{-1}\mathrm{e} + \gamma $$

The daughter isotope is ${}^{12}_{6}\mathrm{C}$.

Example 12.2.4

Write the nuclear equation that represents the radioactive decay of iodine-131 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is ${}^{131}_{53}\mathrm{I}$ while one of the products is an electron, ${}^{0}_{-1}\mathrm{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 131, and its atomic number must be 54. The element having an atomic number of 54 is xenon. Thus, the complete nuclear equation is as follows:

$$ {}^{131}_{53}\mathrm{I} \rightarrow {}^{131}_{54}\mathrm{Xe} + {}^{0}_{-1}\mathrm{e} + \gamma $$

The daughter isotope is ${}^{131}_{54}\mathrm{Xe}$.

Key Takeaway

- Alpha particles have low penetration; stopped by paper or skin.
- Beta particles have moderate penetration; blocked by plastic or aluminum.
- Gamma rays have high penetration; require lead or concrete shielding.

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12.3: Half-Life

Introduction

As discussed earlier, radioactivity is the spontaneous emission of particles or energy from unstable nuclei, most commonly in the form of alpha, beta, or gamma emissions. Each of these processes can be described by nuclear reactions that show how one nuclide transforms into another in order to achieve greater stability. While these reactions explain how unstable nuclei decay, they do not occur all at once but rather at characteristic rates unique to each isotope. This leads to the important concept of **half-life**, which is the time required for half of the atoms in a sample of a radioactive nuclide to undergo decay. In the next section, we will explore this idea in detail, examining how half-life provides a measure of nuclear stability and a practical tool for understanding and applying radioactivity in science and technology.

Consider the following example. Suppose we have 100.0 g of ^{3}H (tritium, a radioactive isotope of hydrogen). It has a half-life of 12.3 y. After 12.3 y, half of the sample will have decayed to ^{3}He by emitting a beta particle, so that only 50.0 g of the original ^{3}H remains. After another 12.3 y—making a total of 24.6 y—another half of the remaining ^{3}H will have decayed, leaving 25.0 g of ^{3}H. After another 12.3 y—now a total of 36.9 y—another half of the remaining ^{3}H will have decayed, leaving 12.5 g of ^{3}H. This sequence of events is illustrated in Figure 12.3.1.

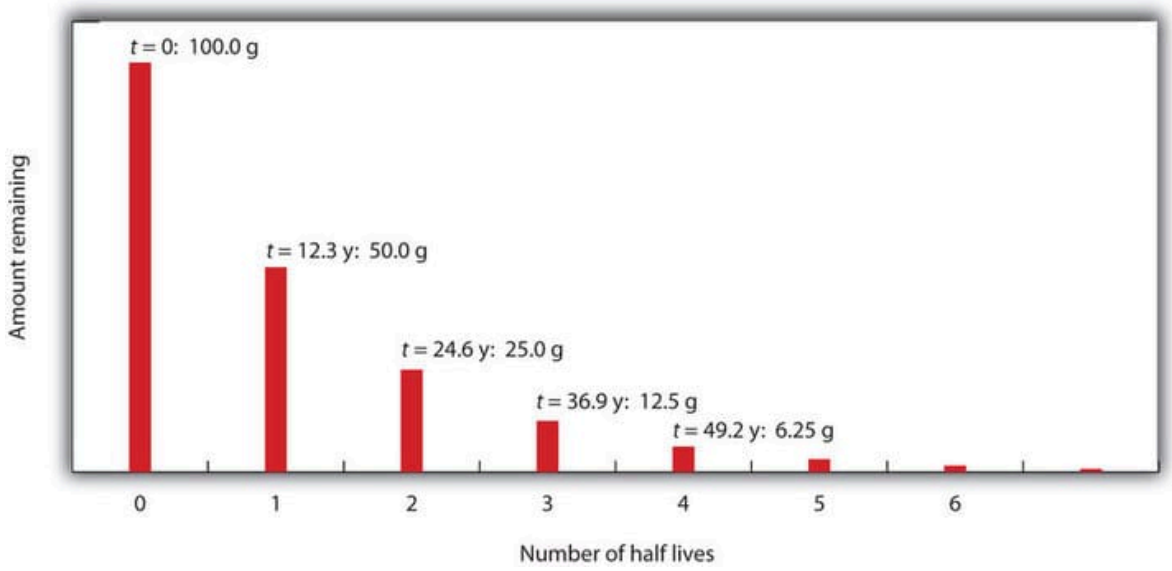


Figure 12.3.1: During each successive half-life, half of the initial amount will radioactively decay. (CC BY-NC-SA 3.0; anonymous via LibreTexts).
Detailed Descrription of Figure 12.3.1.

We can determine the amount of a radioactive isotope remaining after a given number half-lives by using the following expression:

$$\text{amount remaining} = \text{initial amount} \times \left(\frac{1}{2}\right)^{n} \tag{12.3.1}$$

where *n* is the number of half-lives. This expression works even if the number of half-lives is not a whole number.

Example 12.3.1

The half-life of ^{20}F is 11.0 s. If a sample initially contains 5.00 g of ^{20}F, how much ^{20}F remains after 44.0 s?

Solution

If we compare the time that has passed to the isotope's half-life, we note that 44.0 s is exactly 4 half-lives, so we can use Equation 12.3.1 with $n = 4$. Substituting and solving results in the following:

$$\begin{aligned}\text{amount remaining} &= 5.00\ g \times \left(\frac{1}{2}\right)^4 \\ &= 5.00\ g \times \frac{1}{16} \\ &= 0.313\ g\end{aligned}$$

Less than one-third of a gram of ^{20}F remains.

Example 12.3.2

The half-life of ^{44}Ti is 60.0 y. A sample initially contains 0.600 g of ^{44}Ti. How much ^{44}Ti remains after 180.0 y?

Solution

If we compare the time that has passed to the isotope's half-life, we note that 180.0 y is exactly 3 half-lives, so we can use Equation 12.3.1 with $n = 3$. Substituting and solving results in the following:

$$\begin{aligned}\text{amount remaining} &= 0.600\ g \times \left(\frac{1}{2}\right)^3 \\ &= 0.600\ g \times \frac{1}{8} \\ &= 0.075\ g\end{aligned}$$

0.075 g of ^{44}Ti remains after 180.0 y.

Half-lives of isotopes range from fractions of a microsecond to billions of years. Table 12.3.1 lists the half-lives of some isotopes.

Table 12.3.1: Half-Lives of Various Isotopes

Isotope	Half-Life
^{3}H	12.3 y
^{14}C	5,730 y
^{40}K	1.26×10^9 y
^{51}Cr	27.70 d
^{90}Sr	29.1 y
^{131}I	8.04 d
^{222}Rn	3.823 d
^{235}U	7.04×10^8 y
^{238}U	4.47×10^9 y
^{241}Am	432.7 y
^{248}Bk	23.7 h

Isotope	Half-Life
^{260}Sg	4 ms

Example 12.3.3

The isotope I-125 is used in certain laboratory procedures and has a half-life of 59.4 days. If the initial activity of a sample of I-125 is $32,000$ counts per minute (cpm), how much activity will be present in 178.2 days?

Solution

We begin by determining how many half-lives are represented by 178.2 days:

$$\frac{178.2 \text{ days}}{59.4 \text{ days/half-life}} = 3 \text{ half-lives}$$

Then we simply count activity:

$$\begin{aligned} \text{initial activity } (t_0) &= 32,000 \text{ cpm} \\ \text{after one half-life} &= 16,000 \text{ cpm} \\ \text{after two half-lives} &= 8,000 \text{ cpm} \\ \text{after three half-lives} &= 4,000 \text{ cpm} \end{aligned}$$

Be sure to keep in mind that the initial count is at time zero (t_0) and we subtract from that count at the first half-life. The second half-life has an activity of half the previous count (not the initial count).

Equation 12.3.1 can be used to calculate the amount of radioactivity remaining after a given time:

$$N_t = N_0 \times (0.5)^{\text{number of half-lives}}$$

where $N_t =$ activity at time t and $N_0 =$ initial activity at time $t = 0$.

If we have an initial activity of $32,000$ cpm, what will the activity be after four half-lives?

$$\begin{aligned} N_t &= N_0 \, (0.5)^4 \\ &= (32,000) \, (0.5) \, (0.5) \, (0.5) \, (0.5) \\ &= 2000 \text{ cpm} \end{aligned}$$

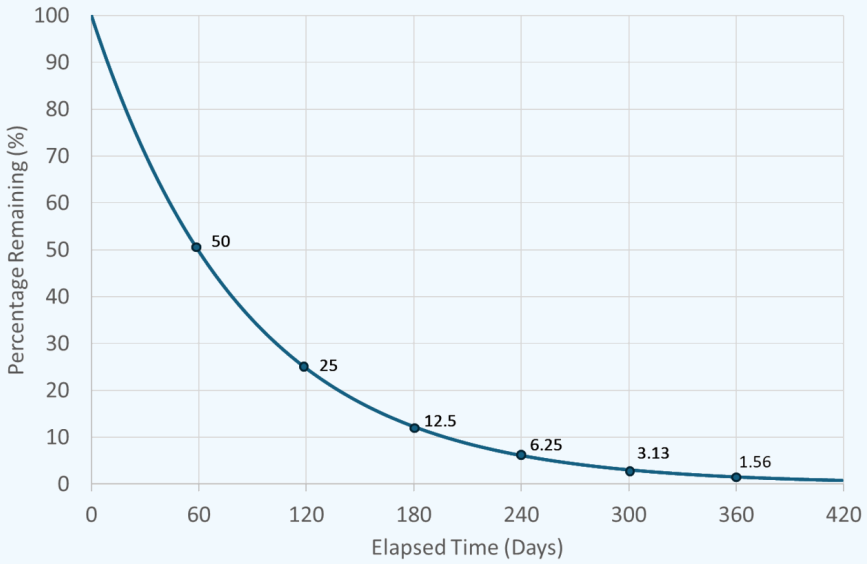


Figure 12.3.2: Typical Radioactive Decay: The graph above illustrates a typical decay curve for $I-125$. The activity decreases by one-half during each succeeding half-life. (CC BY-NC-SA 3.0; anonymous via LibreTexts). Detailed Description of Figure 12.3.2.

Looking Closer: Half-Lives of Radioactive Elements

Many people think that the half-life of a radioactive element represents the amount of time an element is radioactive. In fact, it is the time required for half—not all—of the element to decay radioactively. Occasionally, however, the daughter element is also radioactive, so its radioactivity must also be considered.

The expected working life of an ionization-type smoke detector (described in the opening essay) is about 10 years. In that time, americum-241, which has a half-life of about 432 y, loses less than 4% of its radioactivity. A half-life of 432 y may seem long to us, but it is not very long as half-lives go. Uranium-238, the most common isotope of uranium, has a half-life of about 4.5×10^9 y, while thorium-232 has a half-life of 14×10^9 y.

On the other hand, some nuclei have extremely short half-lives, presenting challenges to the scientists who study them. The longest-lived isotope of lawrencium, ^{262}Lr, has a half-life of 3.6 h, while the shortest-lived isotope of lawrencium, ^{252}Lr, has a half-life of 0.36 s. As of this writing, the largest atom ever detected has atomic number 118, mass number 293, and a half-life of 120 ns. Can you imagine how quickly an experiment must be done to determine the properties of elements that exist for so short a time?

Key Takeaways

- Natural radioactive processes are characterized by a half-life, the time it takes for half of the material to decay radioactively.
- The amount of material left over after a certain number of half-lives can be easily calculated.

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- 11.2: Half-Life by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

12.4: Applications of Radioactive Isotopes

Radioactive Dating Using Carbon-14

Radioactive isotopes are powerful tools for determining the ages of objects, organisms, and geological formations because of their characteristic half-lives and predictable decay rates. This process, called **radiometric dating**, is widely used in archaeology, geology, and paleontology to establish timelines and study the history of Earth and life. For relatively recent materials—up to about 50,000 years old—**carbon-14** is the primary isotope used, making it ideal for dating archeological artifacts, plant remains, and other carbon-containing substances. For much older materials, such as rocks and minerals, isotopes with longer half-lives like **uranium-238**, which decays into lead-206 over billions of years, are employed to estimate the age of geological formations and even the Earth itself. In this section, we will explore the principles of radioactive dating, focusing on C-14 dating for recent materials while also introducing the use of U-238 for dating ancient rocks and determining the age of the planet.

The radioactivity of carbon-14 provides a method for dating objects that were a part of a living organism. This method of radiometric dating, which is also called **radiocarbon dating** or carbon-14 dating, is accurate for dating carbon-containing substances that are up to about 30,000 years old, and can provide reasonably accurate dates up to a maximum of about 50,000 years old.

Naturally occurring carbon consists of three isotopes: ${}^{12}_{6}\mathrm{C}$ which constitutes about 99% of the carbon on earth; ${}^{13}_{6}\mathrm{C}$, about 1% of the total; and trace amounts of ${}^{14}_{6}\mathrm{C}$. Carbon-14 forms in the upper atmosphere by the reaction of nitrogen atoms with neutrons from cosmic rays in space:

$$ {}^{14}_{7}\mathrm{N} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{14}_{6}\mathrm{C} + {}^{1}_{1}\mathrm{H} $$

All isotopes of carbon react with oxygen to produce $\mathrm{CO_2}$ molecules. The ratio of ${}^{14}_{6}\mathrm{CO_2}$ to ${}^{12}_{6}\mathrm{CO_2}$ depends on the ratio of ${}^{14}_{6}\mathrm{CO}$ to ${}^{12}_{6}\mathrm{CO}$ in the atmosphere. The natural abundance of ${}^{14}_{6}\mathrm{CO}$ in the atmosphere is approximately 1 part per trillion; until recently, this has generally been constant over time, as seen is gas samples found trapped in ice. The incorporation of ${}^{14}_{6}\mathrm{CO_2}$ and ${}^{14}_{6}\mathrm{CO_2}$ into plants is a regular part of the photosynthesis process, which means that the ${}^{14}_{6}\mathrm{C} : {}^{12}_{6}\mathrm{C}$ ratio found in a living plant is the **same** as the ${}^{14}_{6}\mathrm{C} : {}^{12}_{6}\mathrm{C}$ ratio in the atmosphere. But when the plant dies, it no longer traps carbon through photosynthesis. Because ${}^{12}_{6}\mathrm{C}$ is a stable isotope and does not undergo radioactive decay, its concentration in the plant does not change. However, carbon-14 decays by β emission with a half-life of 5730 years:

$$ {}^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}_{-1}{}^{0}\,\mathrm{e} $$

Thus, the ratio gradually decreases after the plant dies. The decrease in the ratio with time provides a measure of the time that has elapsed since the death of the plant (or other organism that ate the plant). Figure $12.4.1$: visually depicts this process.

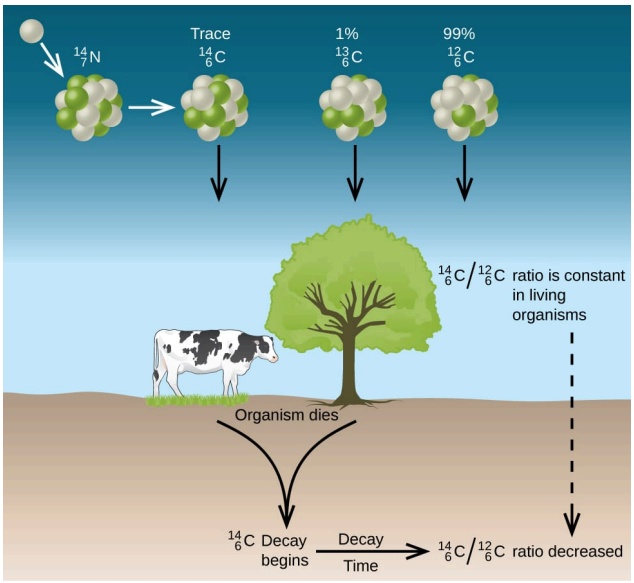


Figure 12.4.1: Along with stable carbon-12, radioactive carbon-14 is taken in by plants and animals, and remains at a constant level within them while they are alive. After death, the C-14 decays and the C-14:C-12 ratio in the remains decreases. Comparing this ratio to the C-14:C-12 ratio in living organisms allows us to determine how long ago the organism lived (and died). (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.4.1.

For example, with the half-life of being 5730 years, if the ratio in a wooden object found in an archaeological dig is half what it is in a living tree, this indicates that the wooden object is 5730 years old. Highly accurate determinations of ratios can be obtained from very small samples (as little as a milligram) by the use of a mass spectrometer.

Example 12.4.1: Radiocarbon Dating

A tiny piece of paper (produced from formerly living plant matter) taken from the Dead Sea Scrolls has an activity of 10.8 disintegrations per minute per gram of carbon. If the initial C-14 activity was 13.6 disintegrations/min/g of C, estimate the age of the Dead Sea Scrolls.

Solution

The rate of decay (number of disintegrations/minute/gram of carbon) is proportional to the amount of radioactive C-14 left in the paper, so we can substitute the rates for the amounts, N, in the relationship:

$$t = -\frac{1}{\lambda}\ln\left(\frac{N_t}{N_0}\right) \longrightarrow t = -\frac{1}{\lambda}\ln\left(\frac{\text{Rate}_{\,t}}{\text{Rate}_{\,0}}\right)$$

where the subscript 0 represents the time when the plants were cut to make the paper, and the subscript t represents the current time.

The decay constant can be determined from the half-life of C-14, 5730 years:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{5730\,\text{y}} = 1.21 \times 10^{-4}\,\text{y}^{-1}$$

Substituting and solving, we have:

$$t = -\frac{1}{\lambda}\ln\left(\frac{\text{Rate}_{\,t}}{\text{Rate}_{\,0}}\right) = -\frac{1}{1.21 \times 10^{-4}\,\text{y}^{-1}}\ln\left(\frac{10.8\ \cancel{\text{dis/min/g C}}}{13.6\ \cancel{\text{dis/min/g C}}}\right) = 1910\,\text{y}$$

Therefore, the Dead Sea Scrolls are approximately 1900 years old (Figure 12.4.2).



Figure 12.4.2: Carbon-14 analysis shows these scrolls were written or copied on plant-based material dating between 100 BC and AD 50. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.4.2.

Radioactive Dating Using Uranium-238

Radioactive dating can also use other radioactive nuclides with longer half-lives to date older events. For example, uranium-238 (which decays in a series of steps into lead-206) can be used for establishing the age of rocks (and the approximate age of the oldest rocks on earth). Since U-238 has a half-life of 4.5 billion years, it takes that amount of time for half of the original U-238 to decay into Pb-206. In a sample of rock that does not contain appreciable amounts of Pb-208, the most abundant isotope of lead, we can assume that lead was not present when the rock was formed. Therefore, by measuring and analyzing the ratio of U-238:Pb-206, we can determine the age of the rock. This assumes that all of the lead-206 present came from the decay of uranium-238. If there is additional lead-206 present, which is indicated by the presence of other lead isotopes in the sample, it is necessary to make an adjustment.

Example 12.4.2: Radioactive Dating of Rocks

An igneous rock contains 9.58×10^{-5} g of U-238 and 2.51×10^{-5} g of Pb-206, and much, much smaller amounts of Pb-208. Determine the approximate time at which the rock formed.

Solution

The sample of rock contains very little Pb-208, the most common isotope of lead, so we can safely assume that all the Pb-206 in the rock was produced by the radioactive decay of U-238. When the rock formed, it contained all of the U-238 currently in it, plus some U-238 that has since undergone radioactive decay.

The amount of U-238 currently in the rock is:

$$9.58 \times 10^{-5} \cancel{\text{g U}} \times \left(\frac{1 \text{ mol U}}{238 \cancel{\text{g U}}} \right) = 4.03 \times 10^{-7} \text{ mol U}$$

Because when one mole of U-238 decays, it produces one mole of Pb-206, the amount of U-238 that has undergone radioactive decay since the rock was formed is:

$$2.51 \times 10^{-5} \cancel{\text{g Pb}} \times \left(\frac{1 \cancel{\text{ mol Pb}}}{206 \cancel{\text{ g Pb}}} \right) \times \left(\frac{1 \text{ mol U}}{1 \cancel{\text{ mol Pb}}} \right) = 1.22 \times 10^{-7} \text{ mol U}$$

The total amount of U-238 originally present in the rock is therefore:

$$4.03 \times 10^{-7} \text{ mol U} + 1.22 \times 10^{-7} \text{ mol U} = 5.25 \times 10^{-7} \text{ mol U}$$

The amount of time that has passed since the formation of the rock is given by:

$$t = -\frac{1}{\lambda} \ln \left(\frac{N_t}{N_0} \right)$$

with N_0 representing the original amount of U-238 and N_t representing the present amount of U-238.

U-238 decays into Pb-206 with a half-life of 4.5×10^9 y, so the decay constant λ is:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{4.5 \times 10^9 \text{ y}} = 1.54 \times 10^{-10} \text{ y}^{-1}$$

Substituting and solving, we have:

$$t = -\frac{1}{1.54 \times 10^{-10} \text{ y}^{-1}} \ln \left(\frac{4.03 \times 10^{-7} \cancel{\text{ mol U}}}{5.25 \times 10^{-7} \cancel{\text{ mol U}}} \right) = 1.7 \times 10^9 \text{ y}$$

Therefore, the rock is approximately 1.7 billion years old.

Key Takeaways

- Radioactive decay acts as a natural clock, allowing age determination of materials.
- C-14 dating is used for recent organic remains (up to ~50,000 years).
- U-238 and other long-lived isotopes date ancient rocks and geological formations.
- Accurate dating depends on initial isotope ratios and may require corrections for environmental changes.

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12.5: Biological Effects of Radiation

Introduction

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All **radioactive nuclides** emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause heating, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure 12.5.1).

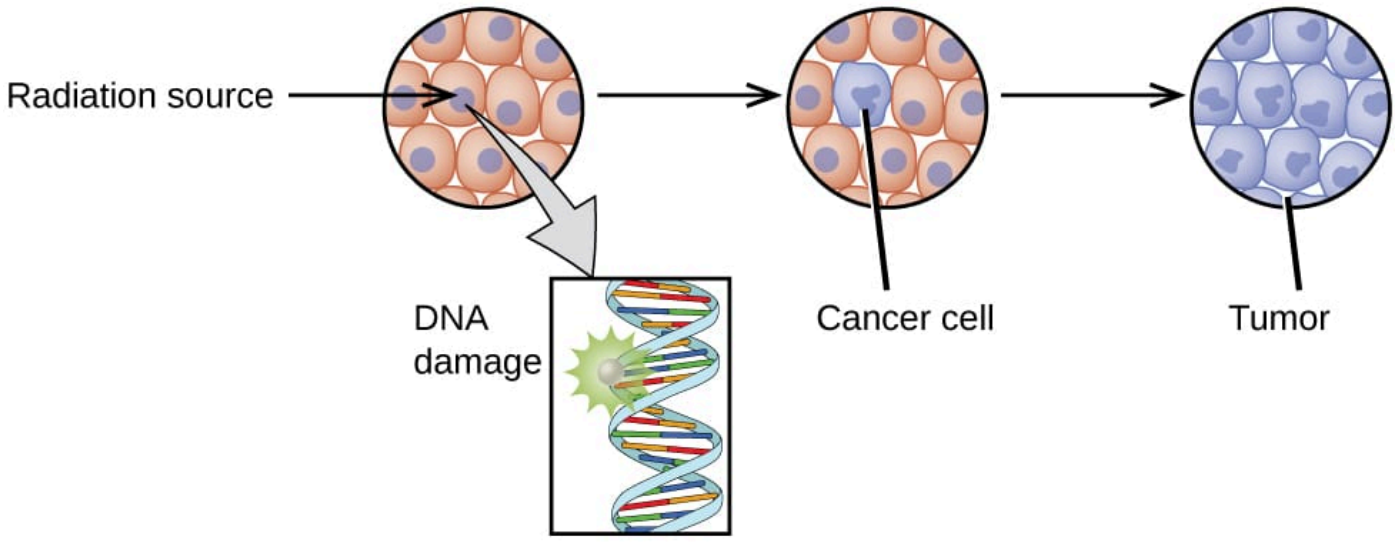


Figure 12.5.1: If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.1.

Ionizing and Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules (for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation), (Figure 12.5.2).

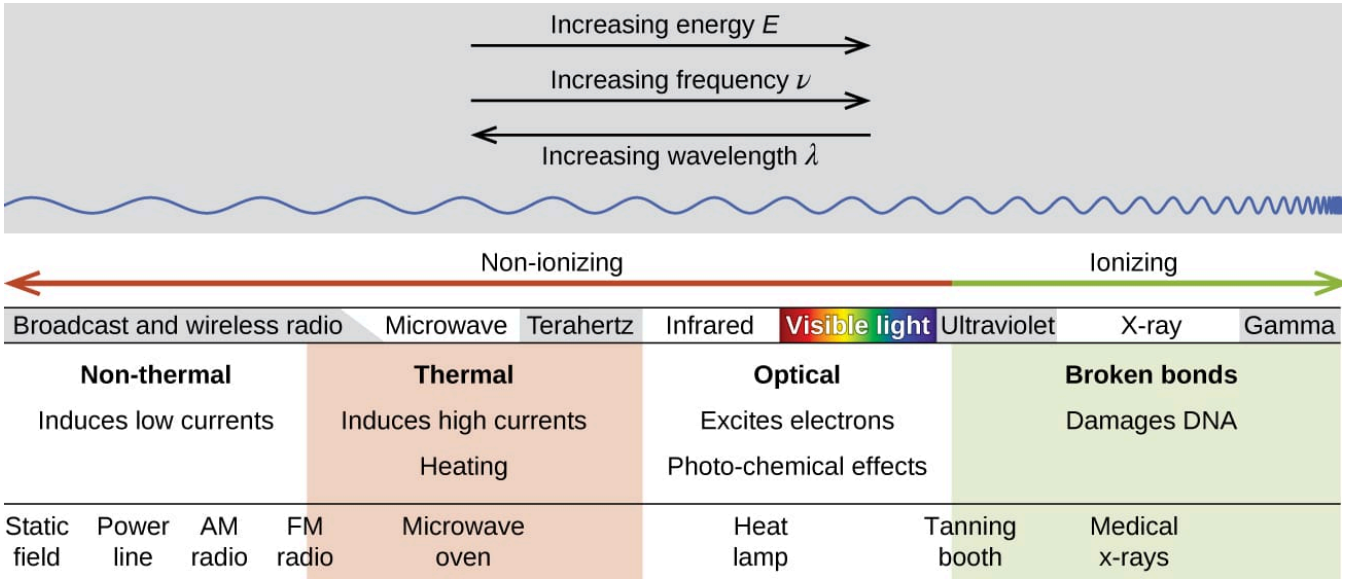
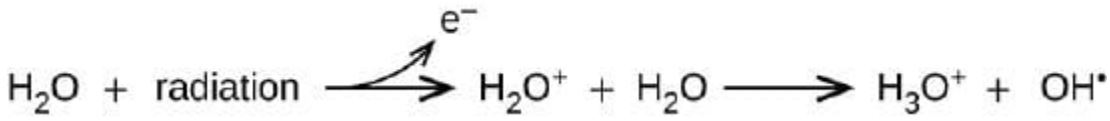


Figure 12.5.2: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.2.

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day at the beach in the sun), a large amount of nonionizing radiation is necessary before dangerous levels are reached. **Ionizing radiation**, however, may cause much more severe damage by breaking bonds or removing electrons in biological molecules, disrupting their structure and function. The damage can also be done indirectly, by first ionizing H_2O (the most abundant molecule in living organisms), which forms a H_2O^+ ion that reacts with water, forming a hydronium ion and a hydroxyl radical:



Because the hydroxyl radical has an unpaired electron, it is highly reactive. (This is true of any substance with unpaired electrons, known as a free radical.) This hydroxyl radical can react with all kinds of biological molecules (DNA, proteins, enzymes, and so on), causing damage to the molecules and disrupting physiological processes. Examples of direct and indirect damage are shown in Figure 12.5.3.

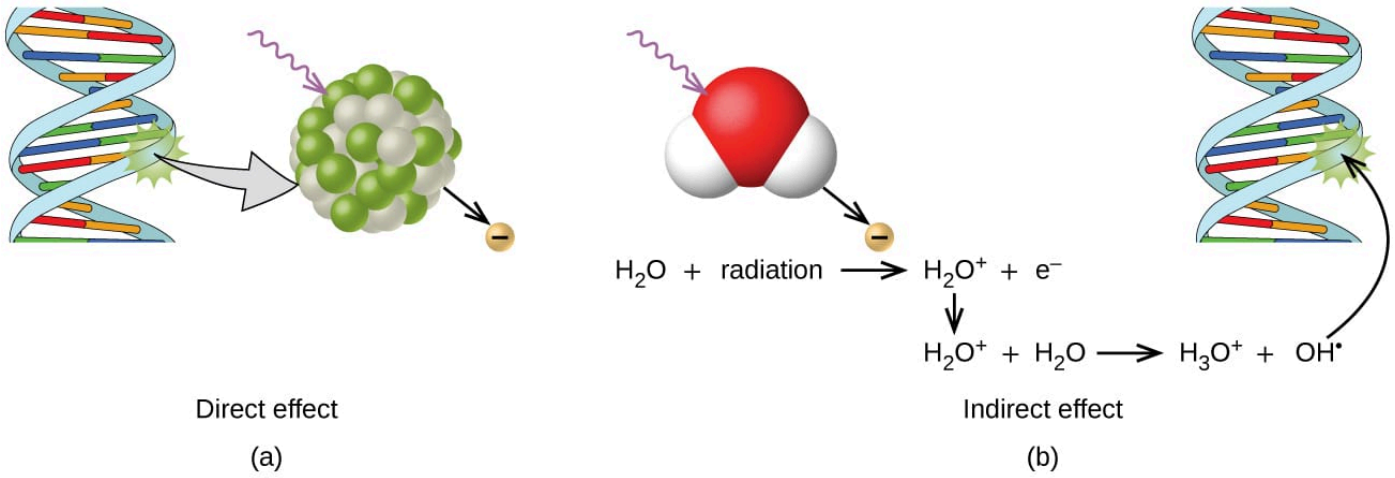


Figure 12.5.3: Ionizing radiation can (a) directly damage a biomolecule by ionizing it or breaking its bonds, or (b) create an H_2O^+ ion, which reacts with H_2O to form a hydroxyl radical, which in turn reacts with the biomolecule, causing damage indirectly. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.3.

Biological Effects of Exposure to Radiation

Radiation can harm either the whole body (**somatic damage**) or eggs and sperm (**genetic damage**). Its effects are more pronounced in cells that reproduce rapidly, such as the stomach lining, hair follicles, bone marrow, and embryos. This is why patients undergoing radiation therapy often feel nauseous or sick to their stomach, lose hair, have bone aches, and so on, and why particular care must be taken when undergoing radiation therapy during pregnancy.

Different types of radiation have differing abilities to pass through material (Figure 12.5.4). A very thin barrier, such as a sheet or two of paper, or the top layer of skin cells, usually stops alpha particles. Because of this, alpha particle sources are usually not dangerous if outside the body, but are quite hazardous if ingested or inhaled (see the Chemistry in Everyday Life feature on Radon Exposure). Beta particles will pass through a hand, or a thin layer of material like paper or wood, but are stopped by a thin layer of metal. Gamma radiation is very penetrating and can pass through a thick layer of most materials. Some high-energy gamma radiation is able to pass through a few feet of concrete. Certain dense, high atomic number elements (such as lead) can effectively attenuate gamma radiation with thinner material and are used for shielding. The ability of various kinds of emissions to cause ionization varies greatly, and some particles have almost no tendency to produce ionization. Alpha particles have about twice the ionizing power of fast-moving neutrons, about 10 times that of β particles, and about 20 times that of γ rays and X-rays.

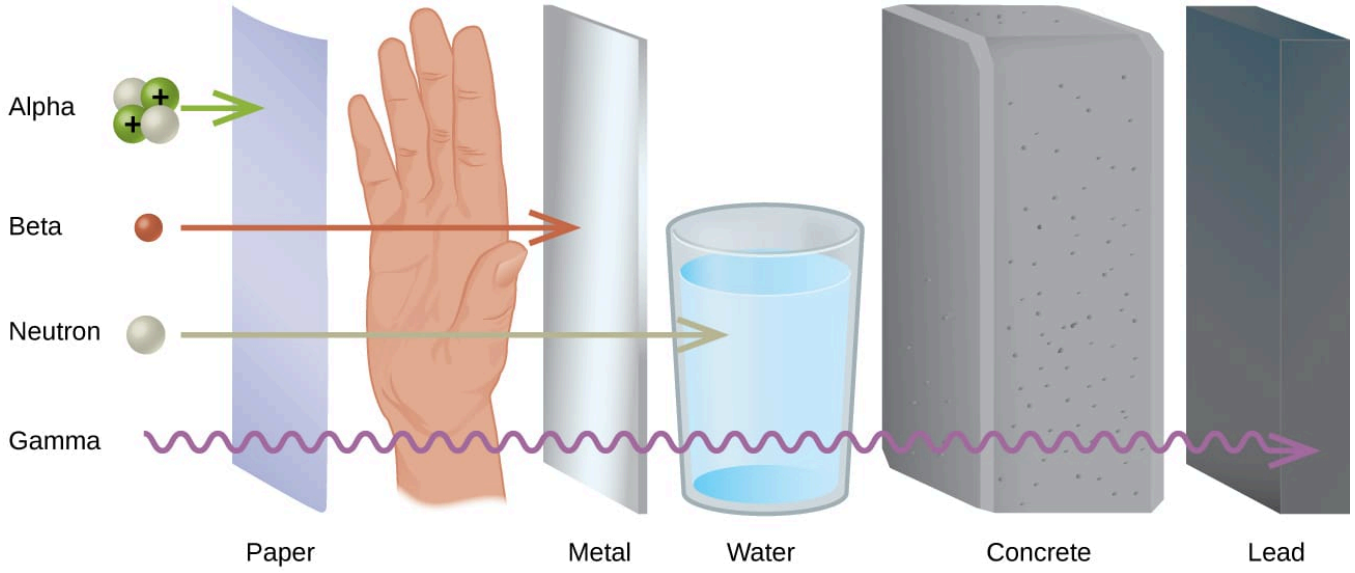


Figure 12.5.4: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.4.

Chemistry in Everyday Life: Radon Exposure

For many people, one of the largest sources of exposure to radiation is from radon gas (Rn-222). Radon-222 is an α emitter with a half–life of 3.82 days. It is one of the products of the radioactive decay series of U-238 (Figure 21.9), which is found in trace amounts in soil and rocks. The radon gas that is produced slowly escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 12.5.5).

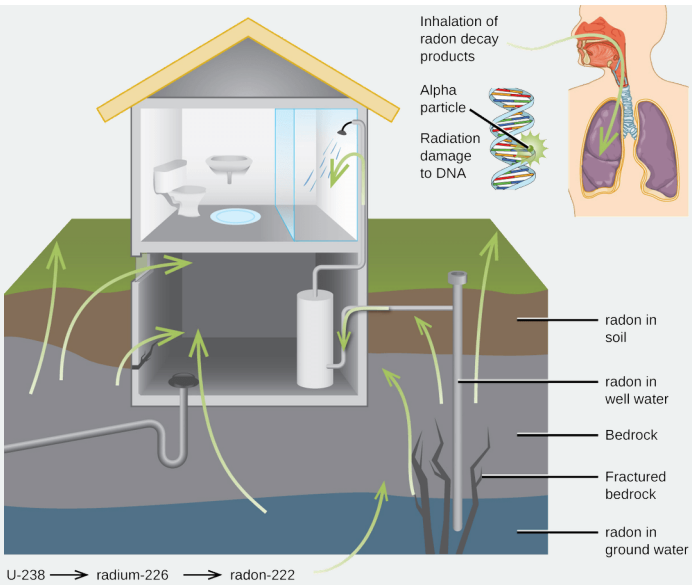


Figure 12.5.5: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinderblock foundations, and openings for water and gas pipes. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.5.

Radon is found in buildings across the country, with amounts depending on where you live. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the levels found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

Measuring Radiation Exposure

Several different devices are used to detect and measure radiation, including **Geiger counters**, **scintillation counters** (scintillators), and **radiation dosimeters** (Figure 12.5.6). Probably the best-known radiation instrument, the Geiger counter (also called the Geiger-Müller counter) detects and measures radiation. Radiation causes the ionization of the gas in a Geiger-Müller tube. The rate of ionization is proportional to the amount of radiation. A scintillation counter contains a scintillator—a material that emits light (luminesces) when excited by ionizing radiation—and a sensor that converts the light into an electric signal. Radiation dosimeters also measure ionizing radiation and are often used to determine personal radiation exposure. Commonly used types are electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

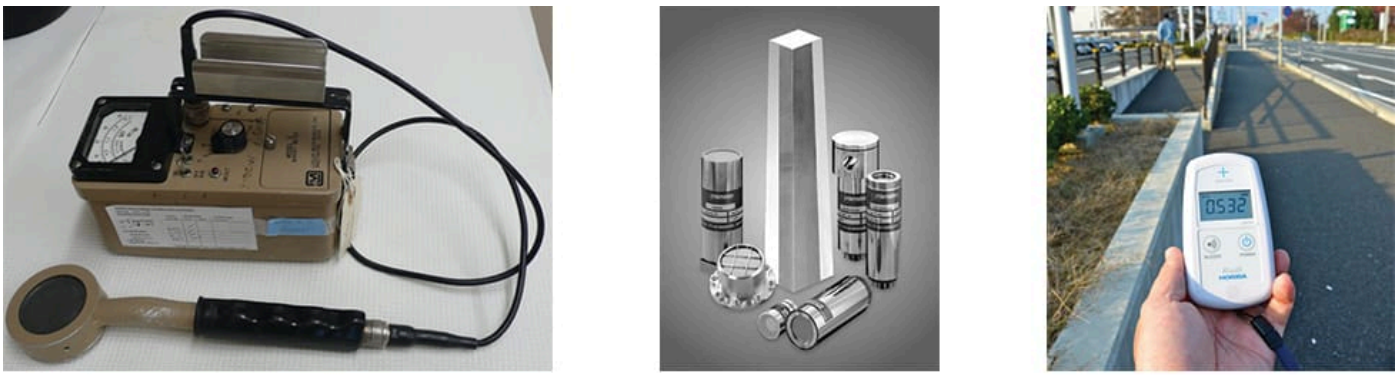


Figure 12.5.6: Devices such as (a) Geiger counters, (b) scintillators, and (c) dosimeters can be used to measure radiation. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.6.

A variety of units are used to measure various aspects of radiation (Figure 12.5.7). The SI unit for rate of radioactive decay is the becquerel (Bq), with 1 Bq = 1 disintegration per second. The curie (Ci) and millicurie (mCi) are much larger units and are frequently used in medicine (1 curie = 1 Ci = 3.7 10^{10} disintegrations per second). The SI unit for measuring radiation dose is the gray (Gy), with 1 Gy = 1 J of energy absorbed per kilogram of tissue. In medical applications, the radiation absorbed dose (rad) is more often used (1 rad = 0.01 Gy; 1 rad results in the absorption of 0.01 J/kg of tissue). The SI unit measuring tissue damage caused by radiation is the sievert (Sv). This takes into account both the energy and the biological effects of the type of radiation involved in the radiation dose. The roentgen equivalent for man (rem) is the unit for radiation damage that is used most frequently in medicine (100 rem = 1 Sv). Note that the tissue damage units (rem or Sv) includes the energy of the radiation dose (rad or Gy) along with a biological factor referred to as the RBE (for relative biological effectiveness) that is an approximate measure of the relative damage done by the radiation. These are related by:

$$\text{number of rems } = \text{RBE} \times \text{ number of rads} \tag{12.5.1}$$

with RBE approximately 10 for α radiation, 2(+) for protons and neutrons, and 1 for β and γ radiation.

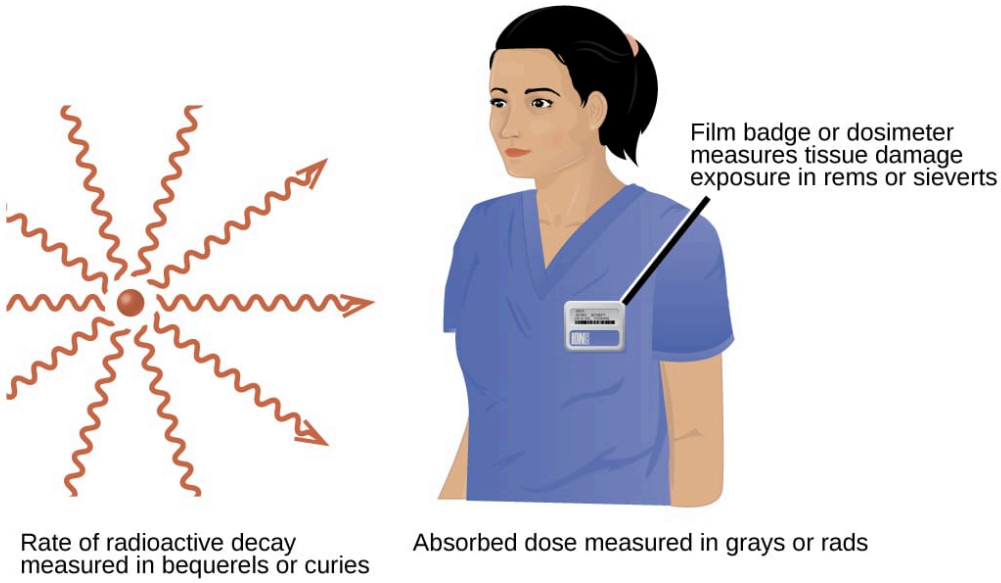


Figure 12.5.7: The energy that is absorbed from the source, and the amount of damage the absorbed radiation does. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.7.

Units of Radiation Measurement

Radiation is measured using specific units depending on what is being quantified—such as the activity of a radioactive source, the energy absorbed, or the potential biological effects. Below is a Table 12.5.1 summarizing key units used in radiation measurement.

Table 12.5.1: Units Used for Measuring Radiation.

Measurement Purpose	Unit	Quantity Measured	Description
activity of source	becquerel (Bq)	radioactive decays or emissions	amount of sample that undergoes 1 decay/second
	curie (Ci)		amount of sample that undergoes 3.7 10^{10} decays/second
absorbed dose	gray (Gy)	energy absorbed per kg of tissue	1 Gy = 1 J/kg tissue
	radiation absorbed dose (rad)		1 rad = 0.01 J/kg tissue

Measurement Purpose	Unit	Quantity Measured	Description
biologically effective dose	sievert (Sv)	tissue damage	Sv = RBE Gy
	roentgen equivalent for man (rem)		Rem = RBE rad

Example 12.5.1: Amount of Radiation

Cobalt-60 ($t_{1/2} = 5.26\ y$) is used in cancer therapy since the γ rays it emits can be focused in small areas where the cancer is located. A 5.00-g sample of Co-60 is available for cancer treatment.

a. What is its activity in Bq?
b. What is its activity in Ci?

Solution

The activity is given by:

$$\text{Activity } = \lambda N = \left(\frac{\ln 2}{t_{1/2}}\right) N = \left(\frac{\ln 2}{5.26\text{y}}\right) \times 5.00\text{ g} = 0.659\frac{\text{g}}{\text{y}} \text{ of Co } -60 \text{ that decay}$$

And to convert this to decays per second:

$$0.659\frac{\text{g}}{\text{y}} \times \frac{1\text{y}}{365\text{ d}} \times \frac{1\text{ d}}{24\text{ h}} \times \frac{1\text{ h}}{3600\text{ s}} \times \frac{1\text{ mol}}{59.9\text{ g}} \times \frac{6.02 \times 10^{23}\text{ atoms}}{1\text{ mol}} \times \frac{1\text{ decay}}{1\text{ atom}}$$
$$= 2.10 \times 10^{14}\frac{\text{decay}}{\text{s}}$$

a. Since $1\text{ Bq} = \frac{1\text{ decay}}{\text{s}}$ the activity in Becquerel (Bq) is:

$$2.10 \times 10^{14}\frac{\text{decay}}{\text{s}} \times \left(\frac{1\text{ Bq}}{1\frac{\text{decay}}{\text{s}}}\right) = 2.10 \times 10^{14}\text{ Bq}$$

b. Since $1\text{Ci} = \frac{3.7\times10^{11}\text{ decay}}{\text{s}}$ the activity in curie (Ci) is:

$$2.10 \times 10^{14}\frac{\text{decay}}{\text{s}} \times \left(\frac{1\text{Ci}}{\frac{3.7\times10^{11}\text{ decay}}{\text{s}}}\right) = 5.7 \times 10^{2}\text{Ci}$$

Effects of Long-term Radiation Exposure on the Human Body

The effects of radiation depend on the type, energy, and location of the radiation source, and the length of exposure. As illustrated in Figure 12.5.8 , the average person is exposed to background radiation, including cosmic rays from the sun and radon from uranium in the ground (reviee the Chemistry in Everyday Life feature on Radon Exposure); radiation from medical exposure, including CAT scans, radioisotope tests, X-rays, and so on; and small amounts of radiation from other human activities, such as airplane flights (which are bombarded by increased numbers of cosmic rays in the upper atmosphere), radioactivity from consumer products, and a variety of radionuclides that enter our bodies when we breathe (for example, carbon-14) or through the food chain (for example, potassium-40, strontium-90, and iodine-131).

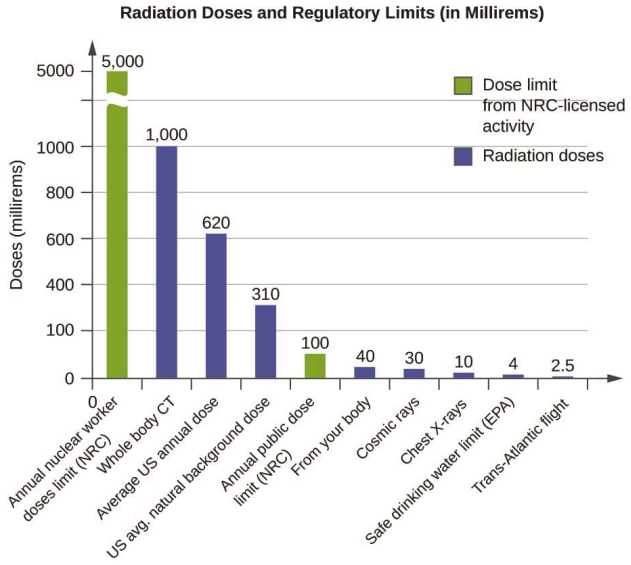


Figure 12.5.8: The various sources and their relative amounts are shown in this bar graph. (CC BY 4.0; via OpenStax). Detailed Description of Figure 12.5.8.

A short-term, sudden dose of a large amount of radiation can cause a wide range of health effects, from changes in blood chemistry to death. Short-term exposure to tens of rems of radiation will likely cause very noticeable symptoms or illness; a dose of about 500 rems is estimated to have a 50% probability of causing the death of the victim within 30 days of exposure. Exposure to radioactive emissions has a cumulative effect on the body during a person's lifetime, which is another reason why it is important to avoid any unnecessary exposure to radiation. Health effects of short-term exposure to radiation are shown in Table 12.5.2.

Table 12.5.2: Health Effects of Radiation.

Exposure (rem)	Health Effect	Time to Onset (without treatment)
5–10	changes in blood chemistry	—
50	nausea	hours
55	fatigue	—
70	vomiting	—
75	hair loss	2–3 weeks
90	diarrhea	—
100	hemorrhage	—
400	possible death	within 2 months
1000	destruction of intestinal lining	—
	internal bleeding	—
	death	1–2 weeks
2000	damage to central nervous system	—
	loss of consciousness;	minutes
	death	hours to days

It is impossible to avoid some exposure to ionizing radiation. We are constantly exposed to background radiation from a variety of natural sources, including cosmic radiation, rocks, medical procedures, consumer products, and even our own atoms. We can minimize our exposure by blocking or shielding the radiation, moving farther from the source, and limiting the time of exposure.

Key Takeaways

- Ionizing radiation can damage or kill cells.
- Somatic damage affects the body; genetic damage affects reproductive cells.
- Rapidly dividing cells (bone marrow, hair, embryos) are most sensitive.
- Shielding and exposure control reduce risks.

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12.6: Summary

Radioactivity is the spontaneous emission of particles or energy from unstable atomic nuclei, including alpha, beta, and gamma radiation, each with distinct penetration and ionizing abilities. These emissions result from nuclear reactions, which transform one nuclide into another, and can be represented by nuclear equations. The half-life of a radioactive isotope—the time for half of a sample to decay—provides a measure of nuclear stability and allows calculation of decay over time. Radioactive isotopes have many applications, including radiometric dating (e.g., carbon-14 for recent organic materials and uranium-238 for ancient rocks), medical diagnostics and treatment, and industrial tracing. Exposure to ionizing radiation can damage living cells, causing somatic or genetic effects, with rapidly dividing cells being most sensitive; protective measures such as shielding help reduce risk. Radioactivity is quantified using units such as becquerels (Bq) or curies (Ci) for activity, gray (Gy) or rad for absorbed dose, and sievert (Sv) or rem for biological effects.

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12.7: Key Terms

- **Radioactivity:** The spontaneous emission of particles or energy from an unstable atomic nucleus.
- **Alpha particle (α):** A particle made of 2 protons and 2 neutrons; identical to a helium nucleus.
- **Parent isotope:** The original unstable atom that undergoes radioactive decay.
- **Daughter isotope:** The new atom formed as a result of radioactive decay
- **Beta particle (β):** A high-speed electron (or positron) emitted from a nucleus during radioactive decay.
- **Gamma ray (γ):** High-energy electromagnetic radiation released from an atomic nucleus.
- **Radioactive decay:** The process by which an unstable atomic nucleus transforms into a more stable one by emitting radiation.
- **Spontaneous fission:** A type of nuclear decay in which a heavy nucleus splits into two or more lighter nuclei along with the release of neutrons and a large amount of energy, occurring without the nucleus being struck by a particle.
- **Half-life:** The time required for half of the nuclei in a radioactive sample to decay.
- **Radiometric dating:** A technique that determines the age of a material by measuring the relative amounts of a radioactive isotope and its decay products.
- **Carbon-14**: A radioactive isotope of carbon used to date recent organic materials up to ~50,000 years old.
- **Uranium-238:** A long-lived radioactive isotope used to date ancient rocks and geological formations.
- **Radiocarbon dating:** A method of dating formerly living organisms by measuring the decay of C-14 relative to stable carbon isotopes.
- **Radioactive nuclides:** Unstable atoms that emit high-energy particles or electromagnetic waves.
- **Ionization radiation:** The process by which radiation removes electrons from atoms or molecules, creating ions.
- **Somatic damage:** Biological damage affecting the body's cells and tissues.
- **Genetic damage:** Biological damage affecting reproductive cells (eggs and sperm), potentially causing mutations.
- **Geiger counter:** A device that detects and measures radiation by the ionization of gas in a Geiger-Müller tube; the rate of ionization is proportional to radiation intensity.
- **Scintillation counter:** An instrument that uses a material that emits light when struck by ionizing radiation; the light is then converted into an electrical signal for measurement.
- **Radiation dosimeter:** A device used to measure personal exposure to ionizing radiation; types include electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

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12.8: Alternative Text for Chapter 12

12.1.1: Different Emissions Exhibit Different Penetration Powers

The image shows alpha (orange), beta (purple), and gamma (green) radiation particles moving horizontally from left to right toward three vertically held barriers: a sheet of paper, a 0.5 cm thick lead block, and a 10 cm thick lead block. Alpha particles at the bottom are completely stopped by the paper. Beta particles in the middle pass through the paper but are stopped by the 0.5 cm lead block. Gamma rays at the top penetrate both the paper and the 0.5 cm lead block but are partially absorbed by the 10 cm lead block.

12.3.1: Radioactive Decay Showing Successive half-lives

The image is a bar graph illustrating the radioactive decay of 100.0 g of tritium (^{3}H) over successive half-lives. The x-axis ranges from 0 to 7, representing the number of half-lives, and the y-axis shows the amount of tritium remaining in grams. Seven vertical red bars correspond to the seven half-lives. Half-lives 1 through 4 are labeled on the x-axis, while only the first four bars have the time in years (t) and remaining mass in grams (Y) indicated on top: after 1 half-life (t = 12.3 y), Y = 50.0 g; after 2 half-lives (t = 24.6 y), Y = 25.0 g; after 3 half-lives (t = 36.9 y), Y = 12.5 g; after 4 half-lives (t = 49.2 y), Y = 6.25 g. The remaining bars continue the halving pattern without labels. The height of each bar decreases progressively, visually representing the exponential decay of the tritium sample over time.

12.3.2: Typical Decay Curve for ^{125}I

The graph illustrates a typical decay curve for ^{125}I, showing how its activity decreases by one-half during each successive half-life. The y-axis represents the percentage of activity remaining, ranging from 0 to 100 % in intervals of 10 %, and the x-axis shows elapsed time in days from 0 to 420 in intervals of 60 days. The elapsed time points are explicitly indicated on the graph: at 60 days, 50 % of the activity remains; at 120 days, it decreases to 25 %; at 180 days, 12.5 %; at 240 days, 6.25 %; at 300 days, 3.13 %; and at 360 days, 1.56 %. The curve slopes downward, visually showing the progressive halving of activity over time as the isotope decays.

12.4.1: Formation and Decay of Carbon-14 in Radiocarbon Dating

The image illustrates the formation and use of carbon-14 for radiocarbon dating. A white diagonal arrow shows a nitrogen-14 nucleus (7 protons in green, 7 neutrons in gray) being struck by a cosmic-ray neutron. A white horizontal arrow indicates the conversion of nitrogen-14 into the three carbon isotopes: carbon-12, carbon-13, and carbon-14. From each of the three isotopes, a black downward arrow points toward a tree and a cow grazing beneath it, representing the incorporation of carbon into living organisms. On the right, it is indicated that the carbon-14/carbon-12 ratio remains constant while the organism is alive. Below the cows, a label reads “organisms die,” and two curves converge downward toward the carbon-14 decay equation. A horizontal arrow labeled “decay time” leads to the carbon-14/carbon-12 ratio on the right. A dashed downward arrow starting from the original ratio indicates the decrease in the ratio over time after the organism dies, illustrating the principle of radiocarbon dating.

12.4.2: Dating Ancient Manuscripts Using Carbon-14

The image shows six pages of the Dead Sea Scrolls arranged from left to right. Carbon-14 dating has determined that the plant material used to make the paper of these pages came from organisms that died between 100 BC and AD 50. This dating demonstrates the approximate historical age of the scrolls and provides context for their creation or copying within that time frame.

12.5.1: Biological Effects of Radiation on DNA and Cells

The image illustrates the biological effects of radiation on DNA. On the left, a radiation source emits high-energy particles or

waves toward a group of several cells arranged within a spherical cluster. Beneath this, an enlarged depiction of a DNA double helix is shown being struck by the radiation, with the beam exaggerated and shattered to make the impact visible. Two horizontal arrows connect the middle cell sphere to the final cell sphere, indicating the progression of damage and the consequences of radiation exposure. On the right, one cell with a single damaged nucleus is labeled "cancer cells", while another cell with multiple damaged nuclei is labeled "tumor cells". The image emphasizes how radiation can affect DNA, potentially causing cellular malfunction, uncontrolled division, and cancer.

12.5.2: Electromagnetic Spectrum: From Nonionizing to Ionizing Radiation

The image illustrates the relationship between electromagnetic radiation and its biological effects. At the top, three horizontal arrows (from top to bottom) indicate increasing energy and increasing frequency from left to right, while an opposite horizontal arrow indicates decreasing wavelength in the same direction. A single horizontal line extends across the spectrum, with the left portion in green labeled nonionizing radiation (the majority of the spectrum) and the right portion in red labeled ionizing radiation (a smaller region near gamma rays). This visual emphasizes that lower-frequency, lower-energy radiation such as light and microwaves is nonionizing, whereas higher-frequency, higher-energy radiation such as alpha and beta particles, gamma rays, X-rays, and high-energy ultraviolet radiation is ionizing.

12.5.3: Direct and Indirect Damage of Biomolecules by Ionizing Radiation

The image illustrates the direct and indirect effects of ionizing radiation on biomolecules. On the left side (a), ionizing radiation strikes a biomolecule directly, causing damage by ionizing the molecule or breaking its chemical bonds. On the right side (b), the radiation first interacts with a water molecule, creating an $H_2O^{\diamond}$ ion, which then reacts with another water molecule to produce a hydroxyl radical (•OH). This highly reactive free radical subsequently reacts with a nearby biomolecule, causing indirect damage. The figure emphasizes that ionizing radiation can harm biological molecules both directly, by interacting with the molecule itself, and indirectly, by generating reactive species such as hydroxyl radicals that then attack biomolecules.

12.5.4: Penetration of Different Types of Radiation Through Materials

The image illustrates the penetration abilities of different types of radiation through various materials. On the left, four types of radiation are shown vertically from top to bottom: alpha particles, beta particles, neutrons, and gamma rays. On the right, several materials are depicted horizontally from left to right as barriers: paper, a hand, metal, water, concrete, and lead. The diagram shows which types of radiation can penetrate each material: alpha particles are stopped by paper or the outer layer of skin; beta particles can pass through paper or a hand but are stopped by thin metal; neutrons have moderate penetration; and gamma rays are highly penetrating, passing through most materials, with dense elements like lead providing effective shielding. This visual emphasizes the relative penetrating power of radiation, from least to most penetrating: alpha < beta < neutron < gamma.

12.5.5: Radon-222 Entry into Buildings from Uranium-Containing Soil

The image illustrates how radon-222 gas seeps into houses and other buildings. Radon-222, an alpha-emitting radioactive gas produced from the decay of uranium-238 in rocks and soil, gradually escapes from the ground and accumulates in low-lying areas. The diagram shows a house built on uranium-containing soil and rocks, with radon gas rising from the ground. Arrows indicate the pathways by which radon enters the building, including cracks in concrete foundations and basement floors, porous stone or cinderblock walls, and openings for water and gas pipes. At the bottom of the image, the decay series is shown: U-238 decays to Ra-226 and then to Rn-222, using two arrows to indicate the sequential conversion. The image emphasizes that radon, being about eight times denser than air, tends to accumulate in basements and lower floors before slowly diffusing throughout the structure.

12.5.6: Devices for Detecting and Measuring Radiation

The image illustrates several devices used to detect and measure radiation. On the left (a) is a Geiger counter, also called a Geiger-Müller counter, which detects radiation by measuring the ionization of gas in a Geiger-Müller tube; the rate of ionization is proportional to the radiation intensity. In the center (b) is a scintillation counter, which contains a scintillator that emits light when excited by ionizing radiation; a sensor converts this light into an electrical signal for measurement. On the right (c) are radiation dosimeters, which measure cumulative ionizing radiation exposure and are commonly used to monitor personal exposure. Types

of dosimeters include electronic, film badge, thermoluminescent, and quartz fiber devices. The figure emphasizes the different ways radiation can be detected and quantified using these instruments.

12.5.7: Units for Measuring Radioactive Emission, Absorbed Dose, and Tissue Damage

The image illustrates the different units used to measure radiation and their applications. On the left, a diagram shows the decay or emission of radiation, represented by arrows, with the rate of radioactive decay indicated at the bottom and measured in becquerels (Bq) or curies (Ci). On the right, a human figure is shown being exposed to radiation while working in a medical laboratory. Beneath the figure, the absorbed dose is indicated, measured in grays (Gy) or rads. Above the upper right of the body, a film badge or dosimeter is shown, representing the measurement of tissue damage, which is recorded in rems or sieverts (Sv). The figure emphasizes the distinction between measuring the rate of emission, the energy absorbed, and the biological effects of radiation on the human body.

12.5.8: Annual Radiation Exposure from Various Sources in the U.S.

The image is a bar graph illustrating the total annual radiation exposure for an average person in the U.S., which is approximately 620 millirems (mrem). The x-axis represents radiation dose in millirems, while the y-axis lists the various sources of exposure. The bars show the relative contributions from different sources, including background radiation (such as cosmic rays from the sun and radon from uranium in the ground), medical exposures (including CAT scans, radioisotope tests, and X-rays), and other human-related sources (such as airplane flights, consumer products, and radionuclides entering the body through air or the food chain, e.g., carbon-14, potassium-40, strontium-90, and iodine-131). The graph visually compares the magnitude of each source's contribution to the total annual dose.

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GLOSSARY

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