CVTC BASIC CHEMISTRY



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CVTC Basic Chemistry

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Textbooks used in Compilation:

The Basics of GOB Chemistry (Ball et al.)

Introductory Chemistry (CK-12)

Chemistry for Changing Times (Hill and McCreary)

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TABLE OF CONTENTS

Licensing

1: Matter, Atoms, and Elements

- 1.1: Prelude to Chemistry, Matter, and Measurement
- 1.2: What is Chemistry?
- 1.3: The Classification of Matter
- 1.4: Measurements
- 1.5: Temperature Units and Conversions
- 1.6: The Elements
- 1.7: Atomic Theory
- 1.8: The Structure of Atoms
- 1.9: Nuclei of Atoms
- 1.10: Uses of Radioactive Isotopes
- 1.11: Atomic Masses
- 1.12: The Periodic Table
- 1.13: Chapter 1 Exercises
  - 1.13.1: Chemistry, Matter, and Measurement (Exercises)
  - 1.13.2: Elements, Atoms, and the Periodic Table (Exercises)

2: Chemical Bonding and Nomenclature

- 2.1: Prelude to Ionic Bonding and Simple Ionic Compounds
- 2.2: Two Types of Bonding
- 2.3: Cations
- 2.4: Anions
- 2.5: Transition Metal Ions
- 2.6: Naming Binary Ionic Compounds
- 2.7: Stock System Naming
- 2.8: Formulas for Binary Ionic Compounds
- 2.9: Polyatomic Ions
- 2.10: Names and Formulas of Ternary Ionic Compounds
- 2.11: Naming Binary Molecular Compounds
- 2.12: Naming Acids
- 2.13: Names and Formulas of Bases
- 2.14: Chapter 2 Exercises
  - 2.14.1: Ions, Molecules, and Compounds (Exercises)
  - 2.14.2: Answers Ions, Molecules, and Compounds (Exercises)
  - 2.14.3: More Practice with Nomenclature (Exercises)

3: Measurement and Calculations

- 3.1: SI Base Units
- 3.2: Metric Prefixes
- 3.3: Scientific Notation in Chemistry
- 3.4: SI Length and Volume Units
- 3.5: SI Mass and Weight Units

- 3.6: SI Kinetic Energy Units
- 3.7: Scientific Dimensional Analysis
- 3.8: Metric Unit Conversions
- 3.9: Derived Units
- 3.10: Dosage Calculations
- 3.11: Density
- 3.12: Chapter 3 Excercises
  - 3.12.E: Chemistry, Matter, and Measurement (Exercises)

4: Chemical Quantities

- 4.1: Avogadro's Number
- 4.2: Conversions Between Moles and Atoms
- 4.3: Molar Mass
- 4.4: Conversions Between Moles and Mass
- 4.5: Conversions Between Mass and Number of Particles
- 4.6: Percent Composition
- 4.7: Chapter 4 Exercises
  - 4.7.1: Mole Practice with Guidence (Exercises)
  - 4.7.2: Chemical Quantities (Exercises)
  - 4.7.3: Answers to Chemical Quantities (Exercises)

5: Chemical Reactions and Quantities

- 5.1: Prelude to Introduction to Chemical Reactions
- 5.2: The Law of Conservation of Matter
- 5.3: Chemical Equations
- 5.4: Quantitative Relationships Based on Chemical Equations
- 5.5: Mole-Mole Relationships in Chemical Reactions
- 5.6: Mole-Mass and Mass-Mass Problems

6: Chemical Energy and Reaction Types

- 6.1: Chemical Potential Energy
- 6.2: Heat
- 6.3: Exothermic and Endothermic Processes
- 6.4: Chemical Reaction Rate
- 6.5: Activation Energy
- 6.6: Potential Energy Diagrams
- 6.7: Factors Affecting Reaction Rate
- 6.8: Catalysts
- 6.9: Some Types of Chemical Reactions
- 6.10: Oxidation-Reduction (Redox) Reactions
- 6.11: Redox Reactions in Organic Chemistry and Biochemistry
- 6.12: Single Replacement Reactions
- 6.13: Double Replacement Reactions
- 6.14: Chapter 6 Exercises
  - 6.14.1: Chemical Energy (Exercises)
  - 6.14.2: Types of Reactions (Exercises)
  - 6.14.3: Answers Types of Reactions (Exercises)
  - 6.14.4: More Practice with Reactions (Exercises)

7: Solutions

- 7.1: Structure of Water
- 7.2: Structure of Ice
- 7.3: Physical Properties of Water
- 7.4: Solute and Solvent
- 7.5: Dissolving Process
- 7.6: Liquid-Liquid Solutions
- 7.7: Electrolytes and Nonelectrolytes
- 7.8: Dissociation
- 7.9: Strong and Weak Electrolytes
- 7.10: Solutions
- 7.11: Rate of Dissolving
- 7.12: Saturated and Unsaturated Solutions
- 7.13: How Temperature Influences Solubility
- 7.14: Henry's Law
- 7.15: Concentration
- 7.16: Molecular and Ionic Equations
- 7.17: Net Ionic Equations
- 7.18: Predicting Precipitates Using Solubility Rules
- 7.19: Chapter 7 Exercises
  - 7.19.1: Solutions (Exercises)
  - 7.19.2: Answers to Solutions (Exercises)

8: Acids and Bases

- 8.1: Prelude to Acids and Bases
- 8.2: Arrhenius Definition of Acids and Bases
- 8.3: Brønsted-Lowry Definition of Acids and Bases
- 8.4: Water - Both an Acid and a Base
- 8.5: The Strengths of Acids and Bases
- 8.6: Buffers
- 8.7: Chapter 8 Exercises
  - 8.7.1: Acids and Bases (Exercises)
  - 8.7.2: Acids and Bases (Summary)

Index

Glossary

Detailed Licensing

Licensing

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

CHAPTER OVERVIEW

1: Matter, Atoms, and Elements

1.1: Prelude to Chemistry, Matter, and Measurement
1.2: What is Chemistry?
1.3: The Classification of Matter
1.4: Measurements
1.5: Temperature Units and Conversions
1.6: The Elements
1.7: Atomic Theory
1.8: The Structure of Atoms
1.9: Nuclei of Atoms
1.10: Uses of Radioactive Isotopes
1.11: Atomic Masses
1.12: The Periodic Table
1.13: Chapter 1 Exercises
1.13.1: Chemistry, Matter, and Measurement (Exercises)
1.13.2: Elements, Atoms, and the Periodic Table (Exercises)

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1.1: Prelude to Chemistry, Matter, and Measurement

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?

Learning Objectives

1. Define chemistry in relation to other sciences.
2. Identify the general steps in the scientific method.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. 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.

Chemistry is one branch of science. 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 (Figure **1.2.1**). Thus, chemistry is the study of matter, 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.

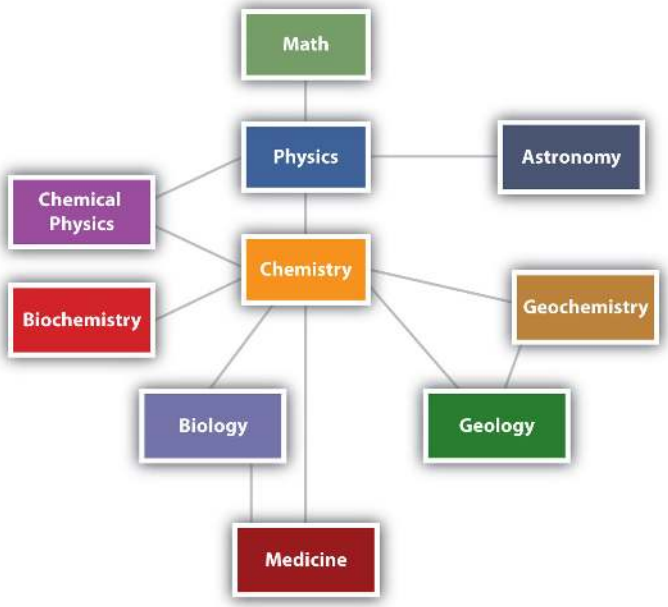


Figure **1.2.1**: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure **1.2.1** shows how many of the individual fields of science are related.

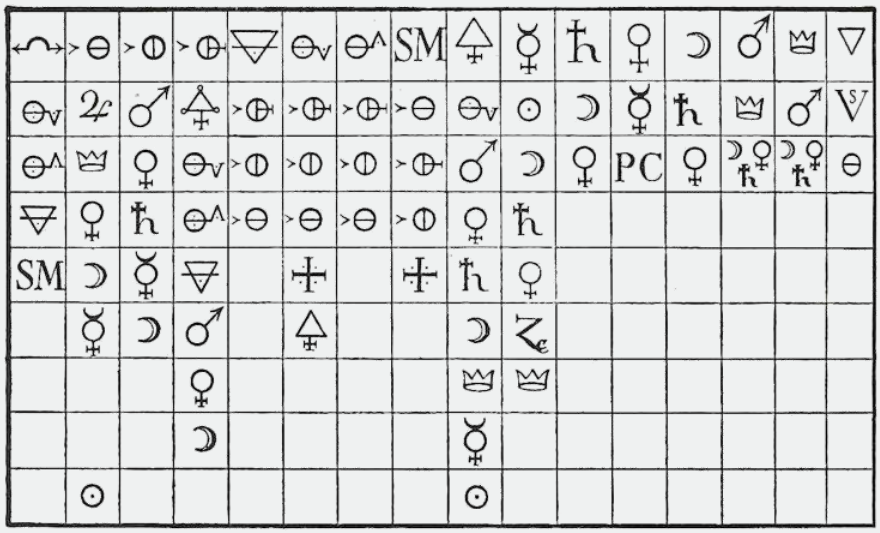
> *There are many other fields of science in addition to the ones listed here.*

Alchemy

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. 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.



The first affinity table. Table of different relations observed in chemistry between different substances; Memoirs of the Royal Academy of Sciences, p. 202-212. Alchemists used symbols like these to represent substances.

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.

Exercise 1.2.1

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Answer a

Sculpture is not considered a science because it is not a study of some aspect of the natural universe.

Answer b

Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

Exercise 1.2.2

Which fields of study are branches of science?

a. physiology (the study of the function of an animal's or a plant's body)
b. geophysics
c. agriculture
d. politics

Answer

a, b and c only

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. 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 **1.2.2**:

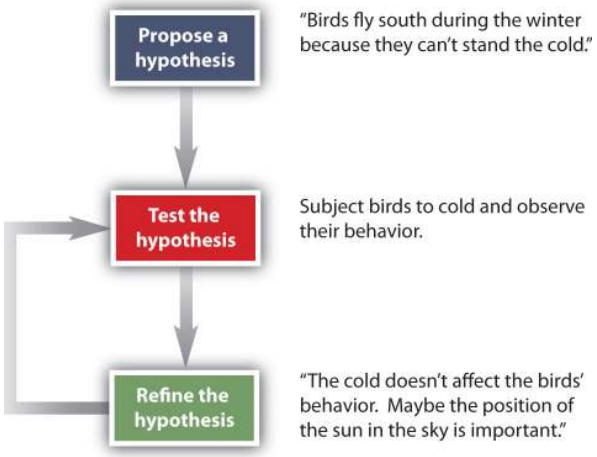


Figure **1.2.2**: The General Steps of the Scientific Method. After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.

The steps may not be as clear-cut in real life as described here, but most scientific work follows this general outline.

1. **Propose a hypothesis.** A scientist generates a testable idea, or hypothesis, to try to answer a question or explain an observation about how the natural universe works. 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.
2. **Test the hypothesis.** A scientist evaluates the hypothesis by devising and carrying out experiments to test it. 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.
3. **Refine the hypothesis if necessary.** 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.

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

1.2.3 Exercise

Define science and chemistry.

Answer

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

1.2.4 Exercise

Name the steps of the scientific method.

Answer

After identifying the problem or making an observation, propose a hypothesis, test the hypothesis, and refine the hypothesis, if necessary

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.

1. Based on what you know, which fields are branches of science?

    a. meteorology (the study of weather)
    b. astrophysics (the physics of planets and stars)
    c. economics (the study of money and monetary systems)
    d. astrology (the prediction of human events based on planetary and star positions)
    e. political science (the study of politics)

2. Based on what you know, which fields are a branches of science?

    a. history (the study of past events)
    b. ornithology (the study of birds)
    c. paleontology (the study of fossils)
    d. zoology (the study of animals)
    e. phrenology (using the shape of the head to determine personal characteristics)

3. Which of the following are examples of matter?

    a. a baby
    b. an idea
    c. the Empire State Building
    d. an emotion
    e. the air
    f. Alpha Centauri, the closest known star (excluding the sun) to our solar system

4. Which of the following are examples of matter?

    a. your textbook
    b. brain cells
    c. love
    d. a can of soda
    e. breakfast cereal

5. Suggest a name for the science that studies the physics of rocks and the earth.
6. Suggest a name for the study of the physics of living organisms.
7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.
8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.
9. When someone says, “I have a theory that excess salt causes high blood pressure,” does that person really have a theory? If it is not a theory, what is it?
10. When a person says, “My hypothesis is that excess calcium in the diet causes kidney stones,” what does the person need to do to determine if the hypothesis is correct?
11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?
12. Most students take multiple English classes in school. Does the study of English use the scientific method?

Answers

1.
   a. science
   b. science
   c. not science
   d. not science
   e. not science

2.
   a. not science
   b. science
   c. science
   d. science
   e. not science

3.
   a. matter
   b. not matter
   c. matter
   d. not matter
   e. matter
   f. matter

4.
   a. matter
   b. matter
   c. not matter
   d. matter
   e. matter

5. geophysics
6. biophysics
7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.
8. Engineering
9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

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1.3: The Classification of Matter

Learning Objectives

- Use physical and chemical properties, including phase, to describe matter.
- Identify a sample of matter as an element, a compound, or a mixture.

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.

Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. **Physical properties** are characteristics that describe matter. They include characteristics such as size, shape, color, and mass. These characteristics can be observed or measured without changing the *identity* of the matter in question. **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.



Figure **1.3.1**: (left) Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical as the 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. (CC-SA-BY-3.0; Andrikkos)

Elements and Compounds

Any sample of matter that has the *same physical and chemical properties throughout* the sample is called a **substance**. There are two types of substances. A substance that cannot be broken down into chemically simpler components is called an **element**. Aluminum, which is used in soda cans and is represented by the symbol Al, is an element. A substance that can be broken down into chemically simpler components (because it consists of more than one element) is called a **compound**. 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. 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*. We will see examples of both macroscopic and microscopic viewpoints throughout this book (Figure **1.3.2**).

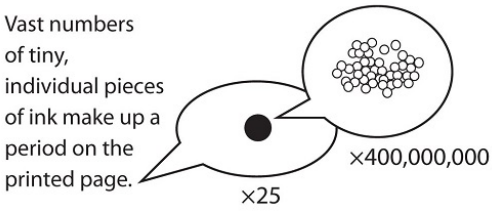


Figure **1.3.2**: How Many Particles Are Needed for a Period in a Sentence? Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The ×25 and ×400,000,000 indicate the number of times the image is magnified.

Mixtures

A material composed of two or more substances is a **mixture**. In a mixture, the individual substances maintain their chemical identities. Many mixtures are obvious combinations of two or more substances, such as a mixture of sand and water. Such mixtures are called **heterogeneous mixtures**. In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Mixtures with a consistent or uniform composition throughout are called **homogeneous mixtures** (or solutions). For example, when sugar is dissolved in water to form a liquid solution, the individual properties of the components cannot be distinguished. Other examples or homogenous mixtures include solid solutions, like the metal alloy steel, and gaseous solutions, like air which is a mixture of mainly nitrogen and oxygen.

✔ Example 1.3.1

How would a chemist categorize each example of matter?

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

Answer a

Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a *homogeneous mixture*, or a solution.

Answer b

Soil is composed of small pieces of a variety of materials, so it is a ***h**eterogeneous mixture*.

Answer c

Water is a *substance*; more specifically, because water is composed of hydrogen and oxygen, it is a *compound*.

Answer d

Oxygen, a *substance*, is an *element*.

? Exercise 1.3.2

How would a chemist categorize each example of matter?

a. breakfast coffee
b. hydrogen
c. an egg

Answer a

homogeneous mixture or solution

Answer b

element

Answer c

heterogeneous mixture

Phases or Physical States of Matter

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. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers.

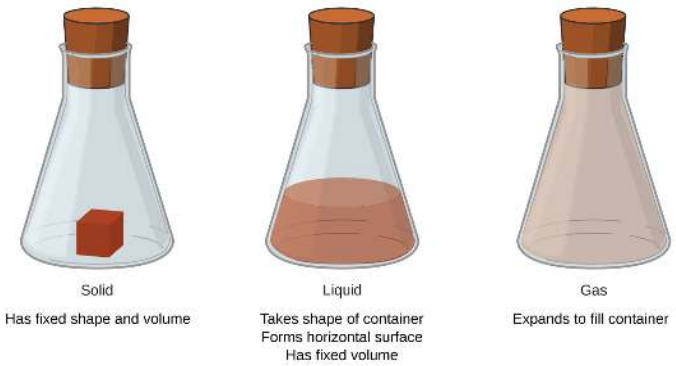


Figure 1.3.3: The three most common states or phases of matter are solid, liquid, and gas. (CC BY-4.0; OpenStax)

A beaker labeled solid contains a cube of red matter and says has fixed shape and volume. A beaker labeled liquid contains a brownish-red colored liquid. This beaker says takes shape of container, forms horizontal surfaces, has fixed volume. The beaker labeled gas is filled with a light brown gas. This beaker says expands to fill container.

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 1.3.2).



Figure 1.3.5: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. (CC BY-SA 3.0 Unported; Markus Schweiss via Wikipedia)

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 1.3.5). Phase changes are identified by particular names depending on what phases are involved, as summarized in Table 1.3.1.

Table 1.3.1: Phase Changes

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

Figure **1.3.3** illustrates the relationships between the different ways matter can be classified.

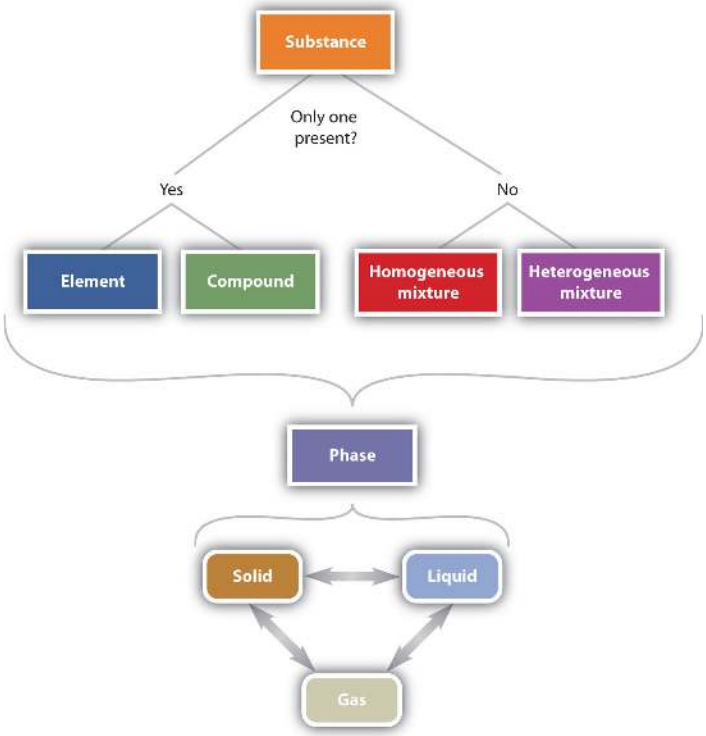


Figure **1.3.6**: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties.
This table starts with a substance. If there is only one present, it can either be an element or a compound. If there is more than one present, it can either be a homogeneous mixture 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.

Concept Review Exercises

1. Explain the differences between the physical properties of matter and the chemical properties of matter.
2. What is the difference between a heterogeneous mixture and a homogeneous mixture? Give an example of each.
3. Give at least two examples of a phase change and state the phases involved in each.

Answers

1. Physical properties describe the existence of matter, and chemical properties describe how substances change into other substances.
2. A heterogeneous mixture is obviously a mixture, such as dirt; a homogeneous mixture behaves like a single substance, such as saltwater.
3. solid to liquid (melting) and liquid to gas (boiling) (answers will vary)

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

Exercise 1.3.3

Does each statement refer to a chemical property or a physical property?

1. Balsa is a very light wood.
2. If held in a flame, magnesium metal burns in air.
3. Mercury has a density of 13.6 g/mL.
4. Human blood is red.

Answer

1. physical property
2. chemical property
3. physical property
4. physical property

Exercise 1.3.4

Does each statement refer to a chemical property or a physical property?

1. The elements sodium and chlorine can combine to make table salt.
2. The metal tungsten does not melt until its temperature exceeds 3,000°C.
3. The ingestion of ethyl alcohol can lead to disorientation and confusion.
4. The boiling point of isopropyl alcohol, which is used to sterilize cuts and scrapes, is lower than the boiling point of water

Answer

1. chemical property
2. physical property
3. chemical property
4. physical property

Exercise 1.3.5

Define *element*. How does it differ from a compound?

Answer

An element is a substance that cannot be broken down into chemically simpler components. Compounds can be broken down into simpler substances.

Exercise 1.3.6

Define *compound*. How does it differ from an element?

Answer

A compound is composed of two or more elements combined in a fixed ratio. An element is the simplest chemical substance.

Exercise 1.3.7

Give two examples of a heterogeneous mixture.

Answer

a salt and pepper mix and a bowl of cereal (answers will vary)

Exercise 1.3.8

Give two examples of a homogeneous mixture.

Answer

vinegar and rubbing alcohol (answers will vary)

Exercise 1.3.9

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

1. xenon, a substance that cannot be broken down into chemically simpler components
2. blood, a substance composed of several types of cells suspended in a salty solution called plasma
3. water, a substance composed of hydrogen and oxygen

Answer

1. element
2. heterogeneous mixture
3. compound

Exercise 1.3.10

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

1. sugar, a substance composed of carbon, hydrogen, and oxygen
2. hydrogen, the simplest chemical substance
3. dirt, a combination of rocks and decaying plant matter

Answer

1. compound
2. element
3. heterogeneous mixture

Exercise 1.3.11

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

1. air, primarily a mixture of nitrogen and oxygen
2. ringer's lactate, a standard fluid used in medicine that contains salt, potassium, and lactate compounds all dissolved in sterile water
3. tartaric acid, a substance composed of carbon, hydrogen, and oxygen

Answer

1. heterogeneous mixture
2. solution
3. compound

? Exercise 1.3.12

What word describes each phase change?

1. solid to liquid
2. liquid to gas
3. solid to gas

Answer

1. melting or fusion
2. boiling or evaporation
3. sublimation

? Exercise 1.3.13

1. What word describes each phase change?
    1. liquid to solid
    2. gas to liquid
    3. gas to solid

Answer

1. freezing
2. condensation
3. deposition

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1.4: Measurements

Learning Objectives

- Express quantities properly using a number and a unit.

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**: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GNU Free Documentation License; Pia von Lützau via Wikipedia).

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.

Exercise 1.4.1

Identify the ***number*** and the ***unit*** in each quantity.

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

Answer a

The number is one, and the unit is dozen.

Answer b

The number is 2.54, and the unit is centimeter.

Answer c

The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.

Answer d

The number is 88, and the unit is meters per second. Note that in this case the unit is actually a combination of two units: meters and seconds.

? Exercise 1.4.2

Identify the ***number*** and the ***unit*** in each quantity.

a. 99 bottles of soda
b. 60 miles per hour
c. 32 fluid ounces
d. 98.6 degrees Fahrenheit

Answer a

The number is 99, and the unit is bottles of soda.

Answer b

The number is 60, and the unit is miles per hour.

Answer c

The number 32, and the unit is fluid ounces

Answer d

The number is 98.6, and the unit is degrees Fahrenheit

? Exercise 1.4.2

What are the two necessary parts of a quantity?

Answer

The two necessary parts are the number and the unit.

Key Takeaway

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

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1.5: Temperature Units and Conversions

Learning Objectives

- Identify the difference between temperature and heat.
- Recognize the different scales used to measure temperature.
- Perform temperature conversions from one unit to another.

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



Figure 1.5.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

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



Figure 1.5.2 Daniel Gabriel Fahrenheit (top), Anders Celsius (center), and Lord Kelvin (bottom).

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724, and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32°F$ and the boiling point as $212°F$. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

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

The Kelvin Scale

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

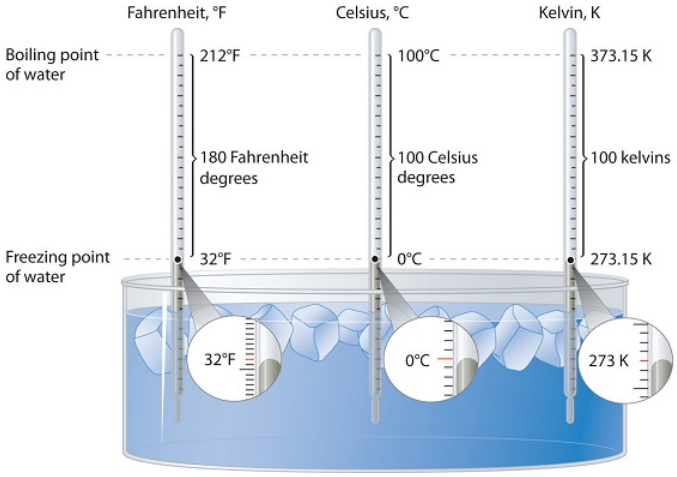


Figure 1.5.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F).

Converting between Scales

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

$$T\,(\text{in °C}) + 273.15 = T\,(\text{in K}) \tag{1.5.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in °C}) \tag{1.5.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows:

$$°C = \frac{5}{9} \times (°F - 32) \tag{1.5.3}$$

$$°F = \frac{9}{5} \times (°C) + 32 \tag{1.5.4}$$

There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F.

Example 1.5.1: Temperature Conversions

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

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 1.5.3:

$$\begin{aligned} °C &= \frac{5}{9} \times (103.5°F - 32) \\ &= 39.7\,°C \end{aligned}$$

Converting from Celsius to Kelvin requires the use of Equation 1.5.1:

$$\begin{aligned} K &= 39.7\,°C + 273.15 \\ &= 312.9\,K \end{aligned}$$

Exercise 1.5.1

Convert each temperature to °C and °F.

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

Answer (a)

5527 °C, 9980 °F

Answer (b)

2807 °C, 5084 °F

Answer (c)

-195.8 °C, -320.4 °F

Summary

- Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).
- The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K.
- Both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit.

Contributors and Attributions

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1.6: The Elements

Learning Objectives

- Define a chemical element and give examples of the abundance of different elements.
- Represent a chemical element with a chemical symbol.

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 **1.6.1** shows some of the chemical elements.



Figure **1.6.1**: Samples of Elements. Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature. © Thinkstock

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.

Table **1.6.1**: Elemental Composition of Earth

Earth's Crust		Earth (overall)	
Element	**Percentage**	**Element**	**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
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*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 14–17.

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 **1.6.1** lists the relative abundances of elements on Earth as a whole and in Earth's crust. Table **1.6.2** lists the relative abundances of elements in the human body. If you compare Table **1.6.1** and Table **1.6.2**, 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 **1.6.1**; carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table **1.6.2**: 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*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 7–24.

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.

LOOKING CLOSER: PHOSPHOROUS, 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, an international organization that makes recommendations concerning all kinds of chemical terminology.

> *Today, new elements are usually named after famous scientists.*

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 **1.6.3** lists the names and symbols of some of the most familiar elements.

Table **1.6.3**: Element Names and 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*
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

*The symbol comes from the Latin name of element.

†The symbol for tungsten comes from its German name—*wolfram*.

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 1.6.1

Write the chemical symbol for each element without consulting Table **1.6.3** "Element Names and Symbols".

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

Answer a

Br

Answer b

B

Answer c

C

Answer d

Ca

Answer e

Au

Exercise 1.6.1

Write the chemical symbol for each element without consulting Table 1.6.3.

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

Answer a

Mn

Answer b

Mg

Answer c

Ne

Answer d

N

Answer e

Ag

Example 1.6.2

What element is represented by each chemical symbol?

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

Answer a

sodium

Answer b

mercury

Answer c

phosphorus

Answer d

potassium

Answer e

iodine

? Exercise 1.6.2

What element is represented by each chemical symbol?

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

Answer a

lead

Answer b

tin

Answer c

uranium

Answer d

oxygen

Answer e

fluorine

Concept Review Exercises

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?

Answers

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.

Key Takeaways

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

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1.7: Atomic Theory

Learning Objectives

- Explain how all matter is composed of atoms.
- Describe the modern atomic theory.
- Recognize which elements exist as diatomic molecules.

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, proposed about 1803 by the English chemist John Dalton (Figure **1.7.1**), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

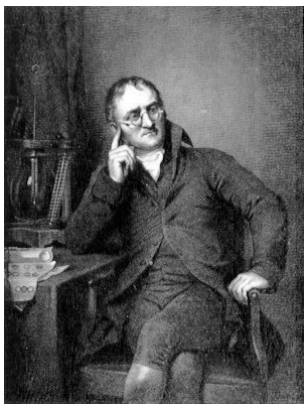


Figure **1.7.1** John Dalton was an English scientist who enunciated the modern atomic theory.

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. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called **diatomic molecules**. In representing a diatomic molecule, we use the symbol of the element and include the ***subscript*** "2" to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2), fluorine (F_2), oxygen (O_2), iodine (I_2), chlorine (Cl_2) and bromine (Br_2). A few other elements can exist as 3-atom molecules like ozone (O_3) and 4-atom molecules like phosphorus (P_4). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S_8 (Figure **1.7.2**).

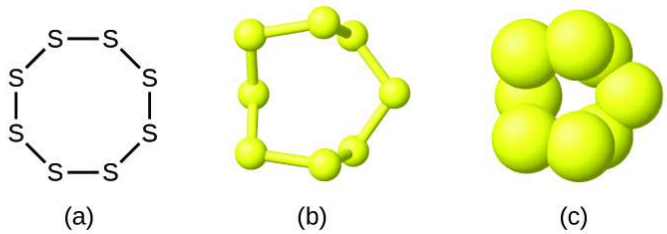


Figure **1.7.2**: A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and 2H represent distinctly different species. H_2 is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression 2H, on the other hand, indicates two separate hydrogen atoms that are **not** combined as a unit. The expression $2H_2$ represents two molecules of diatomic hydrogen (Figure **1.7.3**).

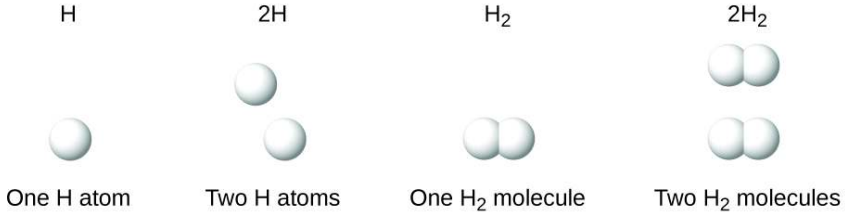


Figure **1.7.3**: The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

Example 1.7.1

Write the chemical formula of the following elements:

a. oxygen
b. carbon
c. potassium

Solution

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

Exercise 1.7.1

Write the chemical formula of the following elements:

a. hydrogen
b. nitrogen
c. neon

Answer a

Hydrogen is one of the seven diatomic molecular elements. Its chemical formula is H_2.

Answer b

Nitrogen is one of the seven diatomic molecular elements. Its chemical formula is N_2.

Answer c

Neon is a monatomic element, hence its formula is Ne.

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.

Concept Review Exercises

1. What is the modern atomic theory?
2. What are atoms?

Answers

1. The modern atomic theory states that all matter is composed of atoms.
2. Atoms are the smallest parts of an element that maintain the identity of that element.

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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1.8: The Structure of Atoms

Learning Objectives

- Describe the three main subatomic particles.
- State how the subatomic particles are arranged in atoms.

There have been several minor but important modifications to Dalton's atomic theory. 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 particle, *subatomic particles*, with their own properties that are different from the chemical properties of atoms.

Subatomic Particles

The first subatomic particle was identified in 1897 and called the electron. It is an extremely tiny particle, with a mass of about 9.109×10^{-31} kg. Experiments with magnetic fields showed that the electron has a negative electrical charge. By 1920, experimental evidence indicated the existence of a second particle. 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).

Finally, additional experiments pointed to the existence of a third particle. Evidence produced in 1932 established the existence of the neutron, a particle with about the same mass as a proton but with no electrical charge, it is *neutral*. We understand now that all atoms can be broken down into subatomic particles: protons, neutrons, and electrons. Table **1.8.1** lists some of their important characteristics and the symbols used to represent each particle.

Table **1.8.1**: 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

The Nucleus

How are these subatomic particles arranged? 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 **1.8.1** shows a diagram of the experimental setup.

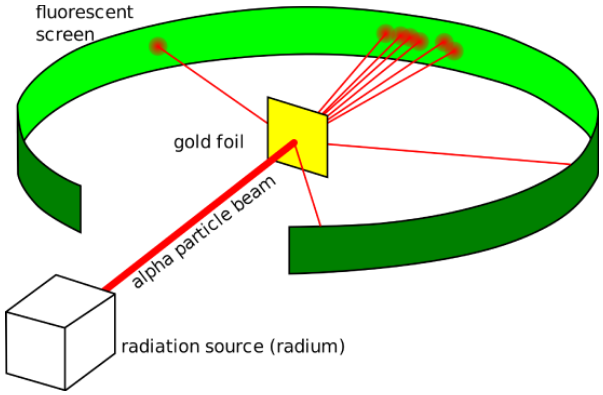


Figure 1.8.1: The Geiger-Marsden Experimental Setup. Experiments using this setup were used to investigate the structure of atoms. In this experiment, most of the particles traveled straight through the foil, but some alpha particles were deflected off to one side. Some were even deflected back toward the source. This was unexpected. Rutherford once said, "It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." (CC SA-BY 3.0; Kurzon).

Rutherford proposed the following model to explain these experimental results. 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 1.8.2 shows how this model explains the experimental results.

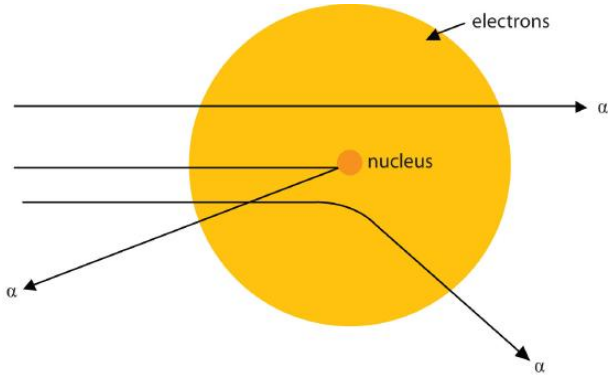


Figure 1.8.2: Rutherford's Metal Foil Experiments. Rutherford explained the results of the metal foil experiments by proposing that most of the mass and the positive charge of an atom are located in its nucleus, while the relatively low-mass electrons orbit about the nucleus. Most alpha particles go straight through the empty space, a few particles are deflected, and fewer still ricochet back toward the source. The nucleus is much smaller proportionately than depicted here.

> *The planetary model of the atom replaced the plum pudding model, which had electrons floating around aimlessly like plums in a "pudding" of positive charge.*

Rutherford's model is essentially the same model that we use today to describe atoms but with one important modification. The planetary model suggests that electrons occupy certain specific, circular orbits about the nucleus. We know now that this model is overly simplistic. A better description is that electrons form fuzzy clouds around nuclei. Figure 1.8.3 shows a more modern version of our understanding of atomic structure.

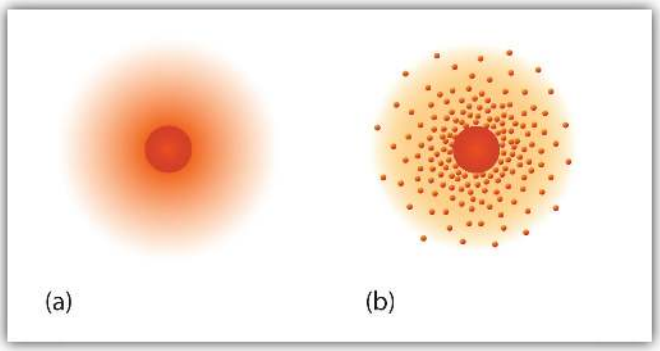


Figure **1.8.3**: A Modern Depiction of Atomic Structure. A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. (a) The darker the color, the higher the probability that an electron will be at that point. (b) In a two-dimensional cross section of the electron in a hydrogen atom, the more crowded the dots, the higher the probability that an electron will be at that point. In both (a) and (b), the nucleus is in the center of the diagram.

Concept Review Exercises

1. What are the charges and the relative masses of the three subatomic particles?
2. Describe the structure of an atom in terms of its protons, neutrons, and electrons.

Answers

1. proton: +1, large; neutron: 0, large; electron: −1, small
2. Protons and neutrons are located in a central nucleus, while electrons orbit about the nucleus.

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.

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1.9: Nuclei of Atoms

Learning Objectives

- Define and differentiate between the atomic number and the mass number of an element.
- Explain how isotopes differ from one another.

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

In the 1910s, experiments with x-rays led to this useful conclusion: the magnitude of the positive charge in the nucleus of every atom of a particular element is the same. In other words, all atoms of the same element have the same number of protons. Furthermore, different elements have a different number of protons in their nuclei, so the number of protons in the nucleus of an atom is characteristic of a particular element. This discovery was so important to our understanding of atoms that the number of protons in the nucleus of an atom is called the atomic number (Z).

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. Table **1.9.1** lists some common elements and their atomic numbers. Based on its **atomic number**, you can determine the **number of protons in the nucleus** of an atom. The largest atoms have over 100 protons in their nuclei.

Example 1.9.1

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

a. aluminum
b. iron
c. carbon

Answer a

According to Table 2.4.1, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.

Answer b

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

Answer c

Carbon has an atomic number of 6. Therefore, every carbon atom has 6 protons in its nucleus.

Exercise 1.9.1

What is the number of protons in the nucleus of each element? Use Table 2.4.1.

a. sodium
b. oxygen
c. chlorine

Answer a

Sodium has 11 protons in its nucleus.

Answer b

Oxygen has 8 protons in its nucleus.

Answer c

Chlorine has 17 protons in its nucleus

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. (Later we will find that some elements may gain or lose electrons from their atoms, so those atoms will no longer be electrically neutral. Thus we will need a way to differentiate the number of electrons for those elements.)

Example 1.9.2

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

a. sulfur
b. tungsten
c. argon

Answer a

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

Answer b

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

Answer c

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

Exercise 1.9.2

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

a. magnesium
b. potassium
c. iodine

Answer a

Mg has 12 electrons.

Answer b

K has 19 electrons.

Answer c

I has 53 electrons.

Table **1.9.1**: Some Common Elements and Their Atomic Numbers

Element	Atomic Number	Element	Atomic Nmbers
aluminum (Al)	13	magnesium (Mg)	12
americium (Am)	95	manganese (Mn)	25
argon (Ar)	18	mercury (Hg)	80
barium (Ba)	56	neon (Ne)	10
beryllium (Be)	4	nickel (Ni)	28
bromine (Br)	35	nitrogen (N)	7
calcium (Ca)	20	oxygen (O)	8
carbon (C)	6	phosphorus (P)	15
chlorine (Cl)	17	platinum (Pt)	78
chromium (Cr)	24	potassium (K)	19
cesium (Cs)	55	radon (Rn)	86
fluorine (F)	9	silver (Ag)	47
gallium (Ga)	31	sodium (Na)	11
gold (Au)	79	strontium (Sr)	38
helium (He)	2	sulfur (S)	16
hydrogen (H)	1	titanium (Ti)	22
iron (Fe)	26	tungsten (W)	74
iodine (I)	53	uranium (U)	92
lead (Pb)	82	zinc (Zn)	30
lithium (Li)	3	zirconium (Zr)	40

Isotopes

How many neutrons are in atoms of a particular element? At first it was thought that the number of neutrons in a nucleus was also characteristic of an element. However, it was found that atoms of the same element can have *different* numbers of neutrons. Atoms of the same element (i.e., same atomic number, Z) that have different numbers of neutrons 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.

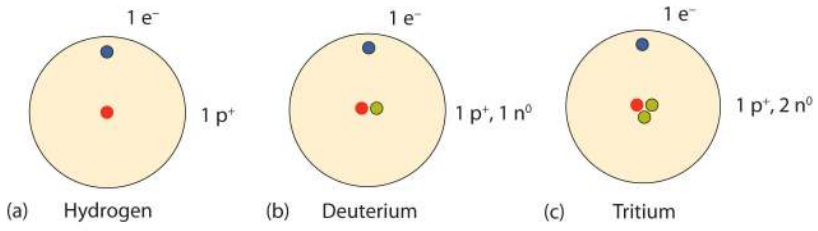


Figure **1.9.1**: Isotopes of Hydrogen. Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

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. Figure **1.9.1** compares the three isotopes of hydrogen.

> The discovery of isotopes required a minor change in Dalton's atomic theory. Dalton thought that all atoms of the same element were exactly the same.

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. The *mass number* (A) 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. Thus, we might see

$$\begin{matrix}\text{mass number}\longrightarrow 56\\ \text{atomic number}\longrightarrow 26\end{matrix}\mathrm{Fe} \tag{1.9.1}$$

which indicates a particular isotope of iron. The 26 is the atomic number (which is the same for all iron atoms), while the 56 is the mass number of the isotope. To determine the number of neutrons in this isotope, we subtract 26 from 56: 56 − 26 = 30, so there are 30 neutrons in this atom.

✔ Example 1.9.3

How many protons and neutrons are in each atom?

a. $^{35}_{17}\mathrm{Cl}$
b. $^{127}_{53}\mathrm{I}$

Answer a

In $^{35}_{17}\mathrm{Cl}$ there are 17 protons, and 35 − 17 = 18 neutrons in each nucleus.

Answer b

In $^{127}_{53}\mathrm{I}$ there are 53 protons, and 127 − 53 = 74 neutrons in each nucleus.

? Exercise 1.9.3

How many protons and neutrons are in each atom?

a. $^{197}_{79}\mathrm{Au}$
b. $^{23}_{11}\mathrm{Na}$

Answer a

In $^{197}_{79}\mathrm{Au}$ there are 79 protons, and 197 − 79 = 118 neutrons in each nucleus.

Answer b

In $^{23}_{11}\mathrm{Na}$ there are 11 protons, and 23 − 11 = 12 neutrons in each nucleus.

It is not absolutely necessary to indicate the atomic number as a subscript because each element has its own unique atomic number. Many isotopes are indicated with a superscript only, such as $^{13}\mathrm{C}$ or $^{235}\mathrm{U}$. You may also see isotopes represented in print as, for example, carbon-13 or uranium-235.

Summary

The atom consists of discrete particles that govern its chemical and physical behavior. Each atom of an element contains the same number of protons, which is the **atomic number (Z)**. Neutral atoms have the same number of electrons and protons. Atoms of an element that contain different numbers of neutrons are called **isotopes**. Each isotope of a given element has the same atomic number but a different **mass number (A)**, which is the **sum of the numbers of protons and neutrons**.

Almost all of the mass of an atom is from the total protons and neutrons contained within a tiny (and therefore very dense) nucleus. The majority of the volume of an atom is the surrounding space in which the electrons reside. A representation of a carbon-12 atom is shown below in Figure **1.9.2**.

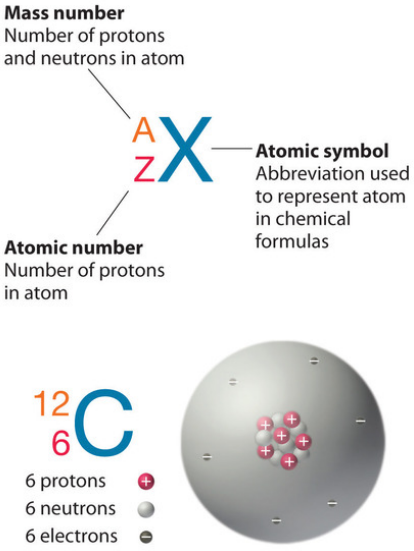


Figure **1.9.2**: Formalism used for identifying specific nuclide (any particular kind of nucleus)

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.

Concept Review Exercises

1. Why is the atomic number so important to the identity of an atom?
2. What is the relationship between the number of protons and the number of electrons in an atom?
3. How do isotopes of an element differ from each other?
4. What is the mass number of an element?

Answers

1. The atomic number defines the identity of an element. It describes the number of protons in the nucleus.
2. In an electrically neutral atom, the number of protons equals the number of electrons.
3. Isotopes of an element have the same number of protons but have different numbers of neutrons in their nuclei.
4. The mass number is the sum of the numbers of protons and neutrons in the nucleus of an atom.

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.

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1.10: Uses of Radioactive Isotopes

Learning Objective

- Learn some applications of radioactivity.

Radioactive isotopes have a variety of applications. Generally, however, they are useful because either we can detect their radioactivity or we can use the energy they release.

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A **tracer** is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some tritium-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium is a radioactive isotope of hydrogen.)

Tracers can also be used to follow the steps of a complex chemical reaction. After incorporating radioactive atoms into reactant molecules, scientists can track where the atoms go by following their radioactivity. One excellent example of this is the use of carbon-14 to determine the steps involved in photosynthesis in plants. We know these steps because researchers followed the progress of carbon-14 throughout the process.

Radioactive Dating

Radioactive isotopes are useful for establishing the ages of various objects. The half-life of radioactive isotopes is unaffected by any environmental factors, so the isotope acts like an internal clock. For example, if a rock is analyzed and is found to contain a certain amount of uranium-235 and a certain amount of its daughter isotope, we can conclude that a certain fraction of the original uranium-235 has radioactively decayed. If half of the uranium has decayed, then the rock has an age of one half-life of uranium-235, or about $4.5 \times 10^9 y$. Many analyses like this, using a wide variety of isotopes, have indicated that age of the earth itself is over $4 \times 10^9 y$.

In another interesting example of radioactive dating, hydrogen-3 dating has been used to verify the stated vintages of some old fine wines.

One isotope, carbon-14, is particularly useful in determining the age of once-living artifacts. A tiny amount of carbon-14 is produced naturally in the upper reaches of the atmosphere, and living things incorporate some of it into their tissues, building up to a constant, albeit very low, level. Once a living thing dies, it no longer acquires carbon-14; as time passes the carbon-14 that was in the tissues decays. (The half-life of carbon-14 is 5,370 y.) If a once-living artifact is discovered and analyzed many years after its death and the remaining carbon-14 is compared to the known constant level, an approximate age of the artifact can be determined. Using such methods, scientists determined that the age of the Shroud of Turin (Figure **1.10.1** purported by some to be the burial cloth of Jesus Christ and composed of flax fibers, a type of plant) is about 600–700 y, not 2,000 y as claimed by some. Scientists were also able to use radiocarbon dating to show that the age of a mummified body found in the ice of the Alps was 5,300 y.

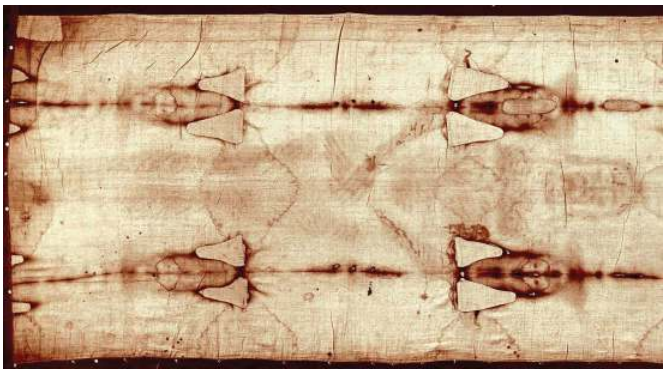


Figure **1.10.1**: Shroud of Turin. In 1989, several groups of scientists used carbon-14 dating to demonstrate that the Shroud of Turin was only 600–700 y. Many people still cling to a different notion, despite the scientific evidence.

Irradiation of Food

The radiation emitted by some radioactive substances can be used to kill microorganisms on a variety of foodstuffs, extending the shelf life of these products. Produce such as tomatoes, mushrooms, sprouts, and berries are irradiated with the emissions from coba . This exposure kills a lot of the bacteria that cause spoilage, so the produce stays fresh longer. Eggs and some meat, such as beef, pork, and poultry, can also be irradiated. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.

Medical Applications

Radioactive isotopes have numerous medical applications in diagnosing and treating illness and diseases. One example of a diagnostic application is using radioactive iodine-131 to test for thyroid activity (Figure **1.10.1**). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of ^{131}I is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

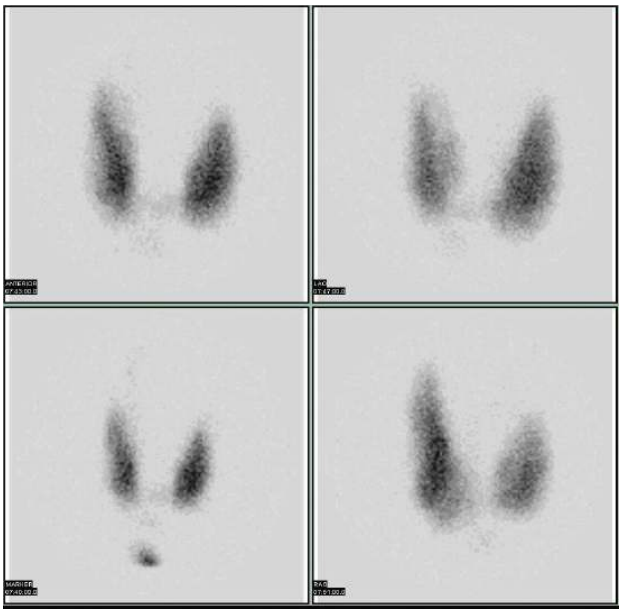


Figure **1.10.2**: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, en.Wikipedia.org/wiki/File:Thyroid_scan.jpg.

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect. However, therapeutic applications usually require much larger doses because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor were detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less than 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect, but also to ease the pain of bone cancers. Table **1.10.1** lists several radioactive isotopes and their medical uses.

Table **1.10.1**: Some Radioactive Isotopes with Medical Applications

Isotope	Use
^{32}P	cancer detection and treatment, especially in eyes and skin
^{59}Fe	anemia diagnosis
^{60}Co	gamma ray irradiation of tumors
$^{99m}Tc^{*}$	brain, thyroid, liver, bone marrow, lung, heart, and intestinal scanning; blood volume determination
^{131}I	diagnosis and treatment of thyroid function
^{133}Xe	lung imaging
^{198}Au	liver disease diagnosis
*The "m" means that it is a metastable form of this isotope of technetium.	

In addition to the direct application of radioactive isotopes to diseased tissue, the gamma ray emissions of some isotopes can be directed toward the tissue to be destroyed. Cobalt-60 is a useful isotope for this kind of procedure.

Food and Drink Application: Radioactivity in Wines

Wine lovers put some stock in *vintages*, or the years in which the wine grapes were grown before they were turned into wine. Wine can differ in quality depending on the vintage. Some wine lovers willingly pay much more for a bottle of wine with a certain vintage. But how does one verify that a bottle of wine was in fact part of a certain vintage? Is the label a fake? Is that stash of wine found in the basement of a French chateau really from the 1940s, or was it made in 2009?

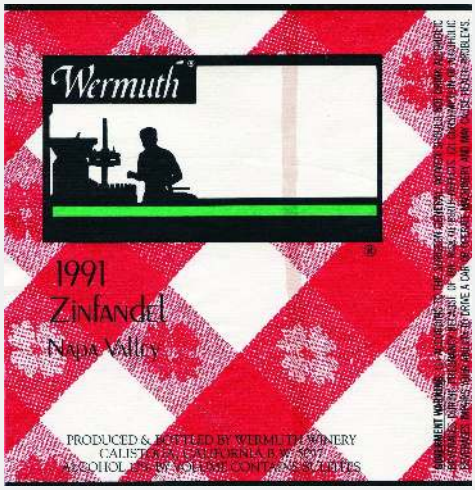


Figure **1.10.3**: A wine label. This wine label from a bottle of wine claims a vintage of 1991. Is the wine really from this vintage, or is it a fake? Radioactivity can help determine the answer. Source: Used by permission of Ralph E. Wermuth.

Cesium-137 is a radioactive isotope that has a half-life of 30.1 y. It was introduced into the atmosphere in the 1940s and 1950s by the atmospheric testing of nuclear weapons by several countries after World War II. A significant amount of cesium-137 was released during the Chernobyl nuclear disaster in 1986. As a result of this atmospheric contamination, scientists have precise measurements of the amount of cesium-137 available in the environment since 1950. Some of the isotope of cesium is taken up by living plants, including grape vines. Using known vintages, oenologists (wine scientists) can construct a detailed analysis of the cesium-137 of various wines through the years.

The verification of a wine's vintage requires the measurement of the activity of cesium-137 in the wine. By measuring the current activity of cesium-137 in a sample of wine (the gamma rays from the radioactive decay pass through glass wine bottles easily, so there's no need to open the bottle), comparing it to the known amount of cesium-137 from the vintage, and taking into account the passage of time, researchers can collect evidence for or against a claimed wine vintage.

> Before about 1950, the amount of cesium-137 in the environment was negligible, so if a wine dated before 1950 shows any measurable cesium-137 activity, it is almost surely a fake, so do not shell out lots of money for it! It may be a good wine, but it is almost definitely not over 60 years old.

Summary

Radioactivity has several practical applications, including tracers, medical applications, dating once-living objects, and preservation of food.

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1.11: Atomic Masses

Learning Objectives

- Define atomic mass and atomic mass unit.
- Calculate atomic mass.

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.

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:

$$1\ \text{u} = \frac{1}{12}\ \text{the mass of}\ ^{12}\text{C atom} \tag{1.11.1}$$

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 1.11.1) are the naturally occurring isotopes of selected elements along with the percent natural abundance of each.

Table 1.11.1: 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 1.11.1

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.

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

1.11.1 Exercise

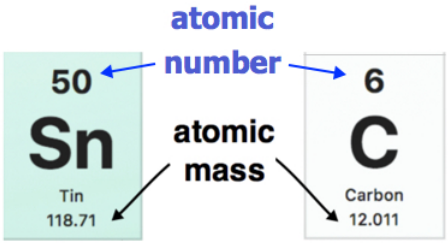
Calculate the atomic mass of copper. Copper exists as a mixture of 2 isotopes. Their respective masses and natural abundance are shown below.

- ^{63}Cu: 62.930 u (69.17%)
- ^{65}Cu: 64.928 u (30.83%)

Answer

63.546 u

The **atomic mass** of each element is found under the element symbol in the **periodic table.** 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.



Atomic mass indicated on entries of the Periodic Table. (public Domain; Pubchem)

The **periodic table** is found in this link:

https://pubchem.ncbi.nlm.nih.gov/periodic-table/png/Periodic_Table_of_Elements_w_Atomic_Mass_PubChem.png

Example 1.11.2: 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:

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

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

? Exercise 1.11.2: Mass of Tin

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

Answer

$$118.71 \cancel{u} \times \frac{1.661 \times 10^{-24}\ \text{g}}{1\ \cancel{u}} = 1.972 \times 10^{-22}\ \text{g}$$

Concept Review Exercises

1. Define atomic mass. Why is it considered a weighted average?
2. What is an atomic mass unit?

Answers

1. The atomic mass is an average of an element's atomic masses, weighted by the natural abundance of each isotope of that element. It is a weighted average because different isotopes have different masses.
2. An atomic mass unit is 1/12th of the mass of a ^{12}C atom.

Key Takeaway

- 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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1.12: The Periodic Table

Learning Objectives

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

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

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

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

The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is shown in Figure **1.12.1**. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

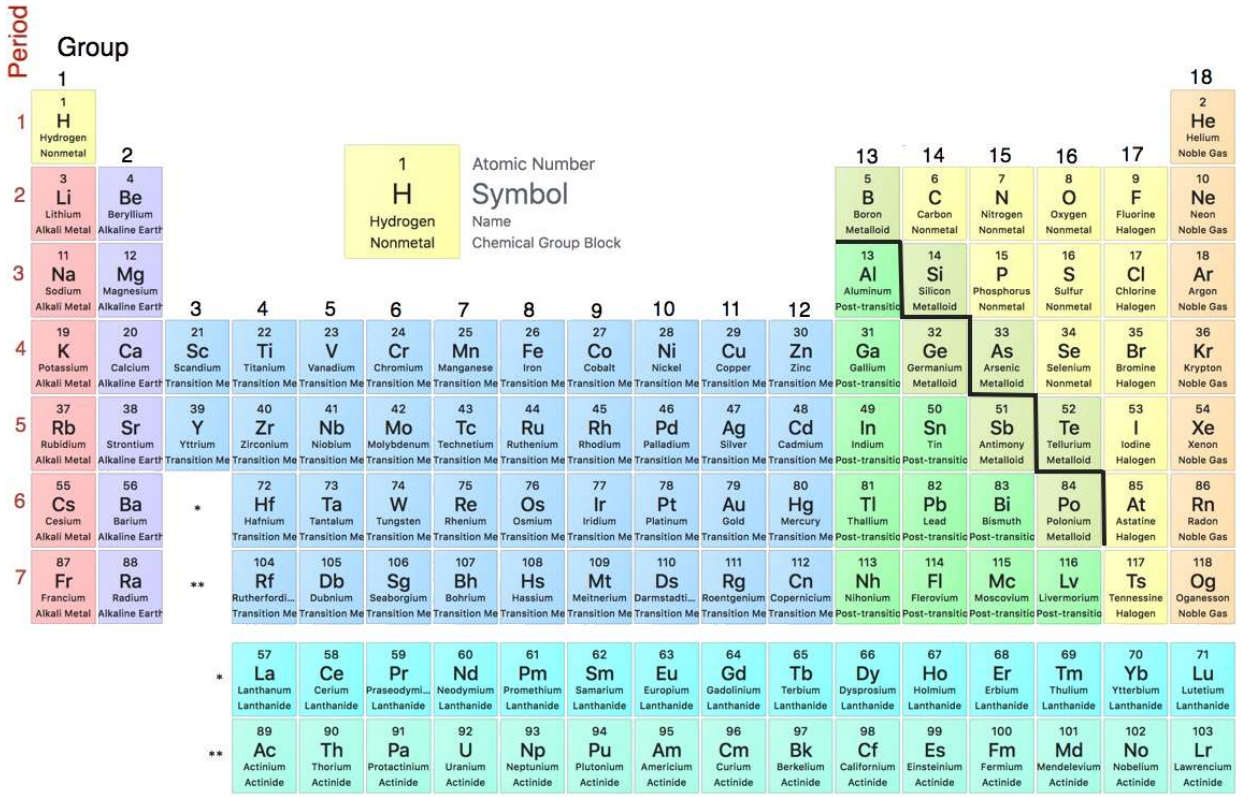


Figure **1.12.1**: Modern Periodic Table. (Public Domain; PubChem modified Leticia Colmenares).

Features of the Periodic Table

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

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 semimetal, as shown in Figure **1.12.2**. A metal is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A nonmetal is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure **1.12.2**, **metals** occupy the left three-fourths of the periodic table, while **nonmetals** (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **semimetals** (or **metalloids**). Elements adjacent to the **bold zigzag** line in the right-hand portion of the periodic table have semimetal properties.

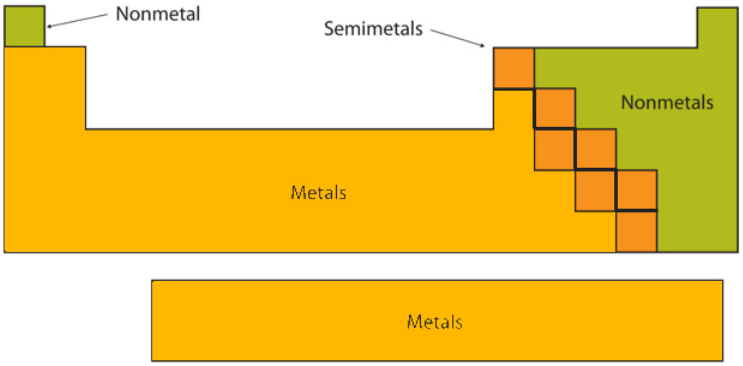


Figure **1.12.2**: Types of Elements. Elements are either metals, nonmetals, or semimetals. Each group is located in a different part of the periodic table.

✔ Example 1.12.1

Based on its position in the periodic table, do you expect selenium (Se) to be a metal, a nonmetal, or a semimetal?

Solution

The atomic number of selenium is 34, which places it in period 4 and group 16. In Figure 2.7.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal. 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.

1.12.1 Exercise

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

Answer

metal

Representative, Transition and Inner-transition

Another way to categorize the elements of the periodic table is shown in Figure **1.12.3**. The first two columns on the left and the last six columns on the right are called the **main group** or **representative** elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the inner transition metals. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

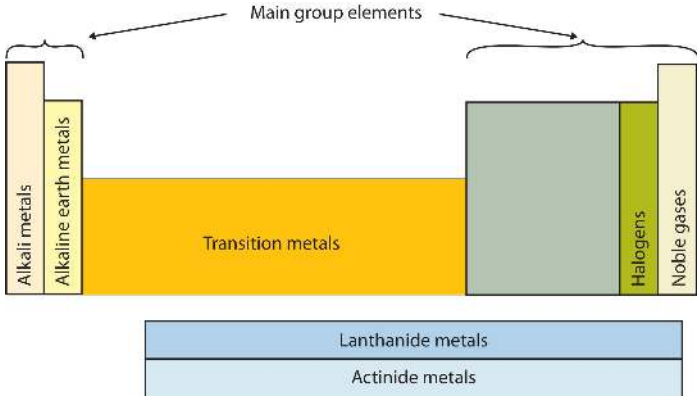


Figure **1.12.3**: Special Names for Sections of the Periodic Table. Some sections of the periodic table have special names. The elements lithium, sodium, potassium, rubidium, cesium, and francium are collectively known as alkali metals.

Alkali metals are the first column. Alkaline earth metals are the second. Halogens are the second to last column. Noble gases are the last column.

To Your Health: Transition Metals in the Body

Most of the elemental composition of the human body consists of main group elements. The first element appearing on the list that is not a 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. Transition metals have interesting chemical properties, partially because some of their electrons are in *d* subshells. 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 not listed explicitly in Table 2.1.2, so they are present in the body in very small quantities. However, even these small quantities are required for the body to function properly.

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** or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

To Your Health: Radon

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

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

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

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 **1.12.4**. 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.**

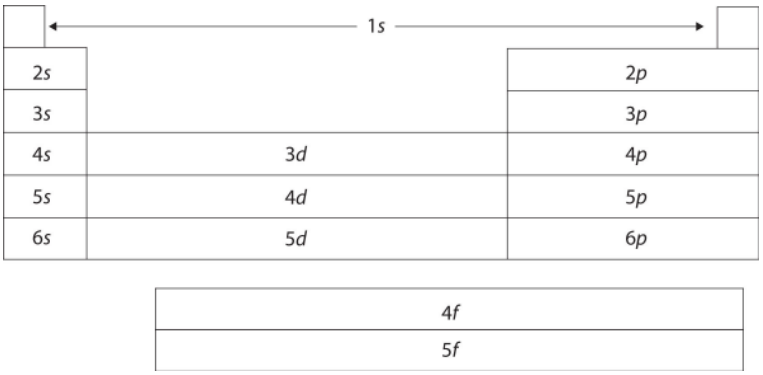


Figure **1.12.4** The Shape of the Periodic Table. The shape of the periodic table reflects the order in which electron shells and subshells fill with electrons.

Example 1.12.1

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

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.

Answer b

The electron configuration of carbon is $1s^22s^22p^2$. Its valence shell electron configuration is $2s^22p^2$. Every element in the same column should have a similar valence shell electron configuration, which we can represent as ns^2np^2.

Exercise 1.12.1

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

a. the halogens
b. the column of elements headed by oxygen

Answer a

The halogens are in the 17th column (or Group 7A) of the periodic table. This column corresponds to the *p* subshell being filled with 5 electrons. Therefore, the valence shell electron configuration is ns^2np^5.

Answer b

The column headed by O is the 16th column (or Group 6A). This column corresponds to the *p* subshell being filled with 4 electrons. Therefore, the valence shell electron configuration is ns^2np^4.

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 1.12.1. 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 1.12.5). 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 $\mathbf{H_2}$ molecule is measured to be $\mathbf{74\ pm}$. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = \mathbf{37\ pm}$.

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 1.12.5

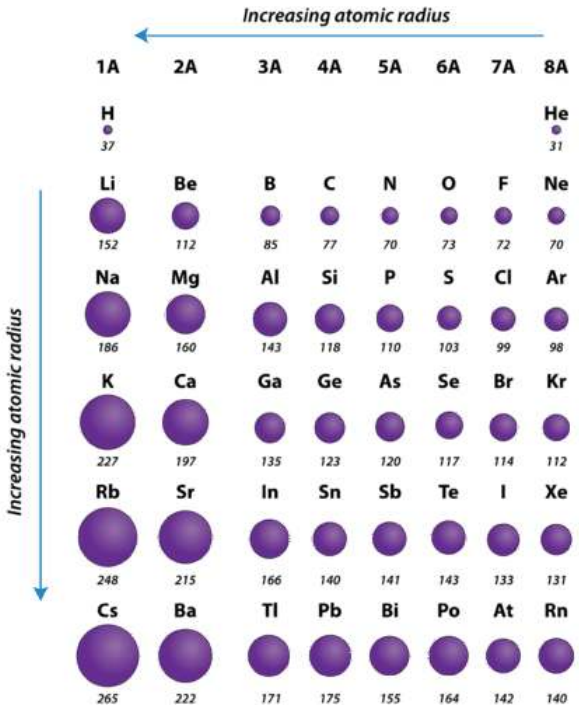


Figure **1.12.5** Trends on the Periodic Table. Atomic radii of the representative elements measured in picometers. The relative sizes of the atoms show several trends with regard to the structure of the periodic table. Atoms become larger going down a group and going from right to left across a period.

✔ Example 1.12.2

Using the periodic table (rather than Figure **1.12.5**), which atom is larger?

a. N or Bi
b. Mg or Cl

Answer a

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.

Answer b

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.

? Exercise 1.12.2

Using the periodic table (rather than Figure **1.12.5**), which atom is larger?

a. Li or F
b. Na or K

Answer a

Li and F are on the same period, but F lies farther to the right. Therefore, we expect Li to be larger than F atoms.

Answer b

K lies below Na in Group 1A, hence has more electron shells, making it larger than Na.

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.

Concept Review Exercises

1. How are the elements organized into the periodic table?
2. Looking at the periodic table, where do the following elements appear?
    a. the metals
    b. the nonmetals
    c. the halogens
    d. the transition metals
    e. the noble gases
3. Describe the trends in atomic radii as related to an element’s position on the periodic table.

Answers

1. Elements are organized in order of increasing atomic number.
2. a. the left three-quarters of the periodic table (to the left of the zigzag band)
    b. the right quarter of the periodic table (to the right of the zigzag band)
    c. the next-to-last column of the periodic table
    d. the middle section of the periodic table
    e. the last column of the periodic table
3. As you go across the periodic table, atomic radii decrease; as you go down the periodic table, atomic radii increase.

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.

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1.13: Chapter 1 Exercises

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1.13.1: Chemistry, Matter, and Measurement (Exercises)

These are homework exercises taken from Chapter 1 of the Ball et al. "The Basics of GOB Chemistry" Textmap.

What Is Chemistry?

Exercises

1. Based on what you know, which fields are branches of science?
    a. meteorology (the study of weather)
    b. astrophysics (the physics of planets and stars)
    c. economics (the study of money and monetary systems)
    d. astrology (the prediction of human events based on planetary and star positions)
    e. political science (the study of politics)
2. Based on what you know, which fields are a branches of science?
    a. history (the study of past events)
    b. ornithology (the study of birds)
    c. paleontology (the study of fossils)
    d. zoology (the study of animals)
    e. phrenology (using the shape of the head to determine personal characteristics)
3. Which of the following are examples of matter?
    a. a baby
    b. an idea
    c. the Empire State Building
    d. an emotion
    e. the air
    f. Alpha Centauri, the closest known star (excluding the sun) to our solar system
4. Which of the following are examples of matter?
    a. your textbook
    b. brain cells
    c. love
    d. a can of soda
    e. breakfast cereal
5. Suggest a name for the science that studies the physics of rocks and the earth.
6. Suggest a name for the study of the physics of living organisms.
7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.
8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.
9. When someone says, “I have a theory that excess salt causes high blood pressure,” does that person really have a theory? If it is not a theory, what is it?
10. When a person says, “My hypothesis is that excess calcium in the diet causes kidney stones,” what does the person need to do to determine if the hypothesis is correct?
11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?

12. Most students take multiple English classes in school. Does the study of English use the scientific method?

Answers

1.
    a. science
    b. science
    c. not science
    d. not science
    e. not science
2.
    a. not science
    b. science
    c. science
    d. science
    e. not science
3.
    a. matter
    b. not matter
    c. matter
    d. not matter
    e. matter
    f. matter
4.
    a. matter
    b. matter
    c. not matter
    d. matter
    e. matter
5. geophysics
6. biophysics
7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.
8. Engineering
9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

The Classification of Matter

Exercises

1. Does each statement refer to a chemical property or a physical property?
    a. Balsa is a very light wood.
    b. If held in a flame, magnesium metal burns in air.
    c. Mercury has a density of 13.6 g/mL.
    d. Human blood is red.

2. Does each statement refer to a chemical property or a physical property?
    a. The elements sodium and chlorine can combine to make table salt.
    b. The metal tungsten does not melt until its temperature exceeds 3,000°C.
    c. The ingestion of ethyl alcohol can lead to disorientation and confusion.
    d. The boiling point of isopropyl alcohol, which is used to sterilize cuts and scrapes, is lower than the boiling point of water.
3. Define *element*. How does it differ from a compound?
4. Define *compound*. How does it differ from an element?
5. Give two examples of a heterogeneous mixture.
6. Give two examples of a homogeneous mixture.
7. Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.
    a. xenon, a substance that cannot be broken down into chemically simpler components
    b. blood, a substance composed of several types of cells suspended in a salty solution called plasma
    c. water, a substance composed of hydrogen and oxygen
8. Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.
    a. sugar, a substance composed of carbon, hydrogen, and oxygen
    b. hydrogen, the simplest chemical substance
    c. dirt, a combination of rocks and decaying plant matter
9. Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.
    a. air, primarily a mixture of nitrogen and oxygen
    b. ringer's lactate, a standard fluid used in medicine that contains salt, potassium, and lactate compounds all dissolved in sterile water
    c. tartaric acid, a substance composed of carbon, hydrogen, and oxygen
10. Identify each material as an element, a compound, a heterogeneous mixture, or a solution.
    a. equal portions of salt and sand placed in a beaker and shaken up
    b. a combination of beeswax dissolved in liquid hexane
    c. hydrogen peroxide, a substance composed of hydrogen and oxygen
11. What word describes each phase change?
    a. solid to liquid
    b. liquid to gas
    c. solid to gas
12. What word describes each phase change?
    a. liquid to solid
    b. gas to liquid
    c. gas to solid

Answers

1.
    a. physical property
    b. chemical property
    c. physical property
    d. physical property
2.
    a. chemical property

b. physical property
c. chemical property
d. physical property

3. An element is a substance that cannot be broken down into chemically simpler components. Compounds can be broken down into simpler substances.

4. A compound is composed of two or more elements combined in a fixed ratio. An element is the simplest chemical substance.

5. a salt and pepper mix and a bowl of cereal (answers will vary)

6. vinegar and rubbing alcohol (answers will vary)

7.

a. element
b. heterogeneous mixture
c. compound

8.

a. compound
b. element
c. heterogeneous mixture

9.

a. solution
b. solution
c. compound

10.

a. heterogeneous mixture
b. solution
c. compound

11.

a. melting or fusion
b. boiling or evaporation
c. sublimation

12.

a. freezing
b. condensation
c. deposition

Measurements

Exercises

1. Why are both parts of a quantity important when describing it?
2. Why are measurements an important part of any branch of science, such as chemistry?
3. You ask a classmate how much homework your chemistry professor assigned. Your classmate answers, “twenty.” Is that a proper answer? Why or why not?

4. Identify the number and the unit in each quantity.

a. five grandchildren
b. 16 candles
c. four score and seven years
d. 40 days and 40 nights
e. 12.01 grams
f. 9.8 meters per second squared
g. 55 miles per hour
h. 98.6 degrees Fahrenheit

Answers

1. The number states how much, and the unit states of what. Without the number and the unit, a quantity cannot be properly communicated.

2. Measurements are needed to carry out experiments.

3. No, it is not a proper answer; you do not know whether the professor meant homework problem number 20 or 20 homework problems.

4.

a. The number is 5, and the unit is grandchildren.
b. The number is 16, and the unit is candles.
c. The number is 4 score and 7 (= 87), and the unit is years.
d. The number is 40, and the units are days and nights.
e. The number is 12.01, and the unit is grams.
f. The number is 9.8, and the unit is meters per second squared.
g. The number is 55, and the unit is miles per hour.
h. The number is 98.6, and the unit is degrees Fahrenheit.

Other Units - Temperature

1. Perform the following conversions.

a. 255°F to degrees Celsius
b. −255°F to degrees Celsius
c. 50.0°C to degrees Fahrenheit
d. −50.0°C to degrees Fahrenheit

2. Perform the following conversions.

a. 1,065°C to degrees Fahrenheit
b. −222°C to degrees Fahrenheit
c. 400.0°F to degrees Celsius
d. 200.0°F to degrees Celsius

3. Perform the following conversions.

a. 100.0°C to kelvins
b. −100.0°C to kelvins
c. 100 K to degrees Celsius
d. 300 K to degrees Celsius

4. Perform the following conversions.

a. 1,000.0 K to degrees Celsius
b. 50.0 K to degrees Celsius
c. 37.0°C to kelvins
d. −37.0°C to kelvins

5. Convert 0 K to degrees Celsius. What is the significance of the temperature in degrees Celsius?

6. Convert 0 K to degrees Fahrenheit. What is the significance of the temperature in degrees Fahrenheit?

7. The record for the world's highest air temperature ever recorded was 134°F. This was measured at Greenland Ranch, Death Valley, on July 10, 1913. What is the temperature in degrees Celsius and in kelvins?

8. The coldest temperature ever recorded on the surface of the earth was −128.6°F in Vostok, Antarctica, in 1983. What is the temperature in degrees Celsius and in kelvins?

Answers

1a. 124°C

1b. −159°C

1c. 122°F

1d. −58°F

2a. 1,949°F

2b. -368°F

2c. 204°C

2d. 93°C

3a. 373 K

3b. 173 K

3c. −173°C

3d. 27°C

4a. 727°C

4b. -223°C

4c. 310 K

4d. 236 K

5. −273°C. This is the lowest possible temperature in degrees Celsius.

6. -460°F. This is the lowest possible temperature in degrees Fahrenheit.

7. 57°C, 330K

8. -89.22°C, 183.9K

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1.13.2: Elements, Atoms, and the Periodic Table (Exercises)

These are homework exercises to accompany Chapter 2 of the Ball et al. "The Basics of GOB Chemistry" Textmap.

The Elements

Exercises

1. Which of the following substances are elements?

   a. sodium
   b. milk
   c. gold
   d. water
   e. air
   f. liquefied nitrogen

2. Which of the following substances are elements?

   a. paper
   b. electricity
   c. neon
   d. carbon
   e. wood
   f. concrete

3. Write the chemical symbol for each element.

   a. silver
   b. sulfur
   c. nitrogen
   d. neon

4. Write the chemical symbol for each element.

   a. bromine
   b. oxygen
   c. lithium
   d. boron

5. Explain why it is improper to write CO as the chemical symbol for cobalt.

6. Explain why it is improper to write NO as the chemical symbol for nobelium.

7. Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	
Cr	
C	
P	

8. Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

Answers

1.
   a. element
   b. not an element
   c. element
   d. not an element
   e. not an element
   f. element

2.
   a. not an element
   b. not an element
   c. element
   d. element
   e. not an element
   f. not an element

3.
   a. Ag
   b. S
   c. N
   d. Ne

4.
   a. Br
   b. O
   c. Li
   d. B

5. By convention, the second letter in an element's symbol is always lowercase.

6. By convention, the second letter in an element's symbol is always lowercase. Additionally, NO represents a compound.

7.

Element Symbol	Element Name
F	fluorine
Fe	iron
I	iodine
Cr	chromium
C	carbon
P	phosphorus

8.

Element Symbol	Element Name
Mg	magnesium
Mn	manganese
Ca	calcium
Cl	chlorine
K	potassium
Pt	platinum

Atomic Theory

Exercises

1. Which of the following elements exist as diatomic molecules?

    a. helium
    b. hydrogen
    c. iodine
    d. gold

2. Which of the following elements exist as diatomic molecules?

    a. chlorine
    b. potassium
    c. silver
    d. oxygen

3. Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H?
4. Why is it proper to represent the elemental form of chlorine as Cl_2 but improper to represent the elemental form of calcium as Ca_2?

Answers

1.
   a. no
   b. yes
   c. yes
   d. no
2.
   a. yes
   b. no
   c. no
   d. yes

3. Hydrogen exists as a diatomic molecule in its elemental form; helium does not exist as a diatomic molecule.

4. Chlorine exists as a diatomic molecule in its elemental form; calcium does not exist as a diatomic molecule.

The Structure of Atoms

Exercises

1. Which is smaller—an electron or a helium atom?
2. Which is larger—a proton or an atom of lead?
3. Which subatomic particle has a positive charge? Which subatomic particle has a negative charge?
4. Which subatomic particle is electrically neutral? Does it exist inside or outside the nucleus?
5. Protons are among the (most, least) massive subatomic particles, and they are found (inside, outside) the nucleus.
6. Electrons are among the (most, least) massive subatomic particles, and they are found (inside, outside) the nucleus.
7. Describe why Rutherford used the term *planetary model* to describe his model of atomic structure.
8. Why is the planetary model not an appropriate way to describe the structure of an atom?
9. What happened to most of the alpha particles in Rutherford's experiment? Explain why that happened.
10. Electrons account for the (majority, minority) of the (mass, volume) of an atom.

Answers

1. An electron is smaller.

2. An atom of lead is larger.

3. proton; electron
4. neutron; inside the nucleus
5. most; inside
6. least; outside
7. Electrons are in orbit about the nucleus.
8. Electrons do not have specific circular orbits about the nucleus.
9. Most of the alpha particles went through the metal sheet because atoms are mostly empty space.
10. minority; mass

Nuclei of Atoms

Exercises

1. How many protons are in the nucleus of each element?
   a. radon
   b. tungsten
   c. chromium
   d. beryllium
2. How many protons are in the nucleus of each element?
   a. sulfur
   b. uranium
   c. calcium
   d. lithium
3. What are the atomic numbers of the elements in Exercise 1?
4. What are the atomic numbers of the elements in Exercise 2?
5. How many electrons are in neutral atoms of the elements in Exercise 1?
6. How many electrons are in neutral atoms of the elements in Exercise 2?
7. Complete the following table.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
80	120		
			$^{55}_{26}\mathrm{Fe}$
	2	hydrogen	

8. Complete the following table.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
			$^{3}_{2}He$
95	153		
	21	potassium	

9. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.
    a. ^{131}I
    b. ^{40}K
    c. ^{201}Hg
    d. ^{19}F
10. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.
    a. ^{3}H
    b. ^{133}Cs
    c. ^{56}Fe
    d. ^{207}Pb
11. What is the mass number of a gallium atom that has 38 neutrons in it?
12. What is the mass number of a uranium atom that has 143 neutrons in it?
13. Complete each sentence.
    a. ^{48}Ti has _____ neutrons.
    b. ^{40}Ar has _____ neutrons.
    c. ^{3}H has _____ neutrons.
14. Complete each sentence.
    a. ^{18}O has _____ neutrons.
    b. ^{60}Ni has _____ neutrons.
    c. ^{127}I has _____ neutrons.

Answers

1.
    a. 86
    b. 74
    c. 24
    d. 4
2.
    a. 16
    b. 92
    c. 20
    d. 3
3. 86, 74, 24, and 4
4. 16, 92, 20, 3

5. 86, 74, 24, and 4

6. 16, 92, 20, 3

7.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
80	120	mercury	$^{200}_{80}\text{Hg}$
26	29	iron	$^{55}_{26}\text{Fe}$
1	2	hydrogen	$^{3}_{1}\text{H}$

8.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
2	1	helium	$^{3}_{2}\text{He}$
95	153	americium	$^{248}_{95}\text{Am}$
19	21	potassium	$^{40}_{19}\text{K}$

9.

a. protons: 53; neutrons: 78; electrons: 53
b. protons: 19; neutrons: 21; electrons: 19
c. protons: 80; neutrons: 121; electrons: 80
d. protons: 9; neutrons: 10; electrons: 9

10.

a. protons: 1; neutrons: 2; electrons: 1
b. protons: 55; neutrons: 78; electrons: 55
c. protons: 26; neutrons: 30; electrons: 26
d. protons: 82; neutrons: 125; electrons: 82

11. 69

12. 235

13.

a. 26
b. 22
c. 2

14.

a. 10
b. 32
c. 74

Atomic Masses

Exercises

1. What is the atomic mass of zinc in atomic mass units?
2. What is the atomic mass of barium in atomic mass units?
3. What is the average mass of a single magnesium atom in grams?
4. What is the average mass of a single calcium atom in grams?
5. What is the mass of 1.000×10^{24} aluminum atoms in grams?
6. What is the mass of 5.000×10^{23} carbon atoms in grams?
7. Which has more mass—1 tungsten atom or 11 oxygen atoms?
8. Which has more mass—1 magnesium atom or 6 helium atoms?
9. Determine the atomic mass of lithium, given the isotopic composition: 92.4% lithium-7 (mass 7.016 u) and 7.60% lithium-6 (mass 6.015 u).
10. Determine the atomic mass of neon, given the isotopic composition: 90.48% neon-20 (mass 19.992 u), 0.27% neon-21 (mass 20.994 u), and 9.25% neon-22 (mass 21.991 u).
11. Determine the atomic mass of magnesium, given the isotopic composition: 78.70% magnesium-24 (mass 23.98 u), 10.13% magnesium-25 (mass 24.99 u) and 11.17% magnesium-26 (mass 25.98).
12. Determine the atomic mass of bromine, given the isotopic composition: 50.69% bromine-79 (mass 78.9183 u), 49.31% and bromine-81 (mass 80.9163 u).

Answers

1. 65.4 u
2. 137.33 u
3. 4.038×10^{-23} g
4. 6.657×10^{-23} g
5. 44.81 g
6. 9.974 g
7. 1 tungsten atom
8. 1 magnesium atom
9. 6.94 u
10. 20.18 u
11. 24.30 u

12. 79.90 u

The Periodic Table

Exercises

1. Which elements have chemical properties similar to those of magnesium?
    a. sodium
    b. fluorine
    c. calcium
    d. barium
    e. selenium
2. Which elements have chemical properties similar to those of lithium?
    a. sodium
    b. calcium
    c. beryllium
    d. barium
    e. potassium
3. Which elements have chemical properties similar to those of chlorine?
    a. sodium
    b. fluorine
    c. calcium
    d. iodine
    e. sulfur
4. Which elements have chemical properties similar to those of carbon?
    a. silicon
    b. oxygen
    c. germanium
    d. barium
    e. argon
5. Which elements are alkali metals?
    a. sodium
    b. magnesium
    c. aluminum
    d. potassium
    e. calcium
6. Which elements are alkaline earth metals?
    a. sodium
    b. magnesium
    c. aluminum
    d. potassium
    e. calcium
7. Which elements are halogens?
    a. oxygen
    b. fluorine
    c. chlorine

d. sulfur
e. carbon

8. Which elements are noble gases?

a. helium
b. hydrogen
c. oxygen
d. neon
e. chlorine

9. Which pairs of elements are located in the same period?

a. H and Li
b. H and He
c. Na and S
d. Na and Rb

10. Which pairs of elements are located in the same period?

a. V and Nb
b. K and Br
c. Na and P
d. Li and Mg

11. In each pair of atoms, which atom has the greater atomic radius?

a. H and Li
b. N and P
c. Cl and Ar
d. Al and Cl

12. In each pair of atoms, which atom has the greater atomic radius?

a. H and He
b. N and F
c. Cl and Br
d. Al and B

13. Scandium is a (metal, nonmetal, semimetal) and is a member of the (main group elements, transition metals).

14. Silicon is a (metal, nonmetal, semimetal) and is a member of the (main group elements, transition metals).

Answers

1.

a. no
b. no
c. yes
d. yes
e. no

2.

a. yes
b. no

c. no
d. no
e. yes

3.

a. no
b. yes
c. no
d. yes
e. no

4.

a. yes
b. no
c. yes
d. no
e. no

5.

a. yes
b. no
c. no
d. yes
e. no

6.

a. no
b. yes
c. no
d. no
e. yes

7.

a. no
b. yes
c. yes
d. no
e. no

8.

a. yes
b. no
c. no
d. yes
e. no

9.

a. no
b. yes
c. yes
d. no

10.

a. no
b. yes
c. yes
d. no

11.

a. Li
b. P
c. Cl
d. Al

12.

a. H
b. N
c. Br
d. Al

13. metal; transition metals

14. semimetal; main group elements

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CHAPTER OVERVIEW

2: Chemical Bonding and Nomenclature

2.1: Prelude to Ionic Bonding and Simple Ionic Compounds
2.2: Two Types of Bonding
2.3: Cations
2.4: Anions
2.5: Transition Metal Ions
2.6: Naming Binary Ionic Compounds
2.7: Stock System Naming
2.8: Formulas for Binary Ionic Compounds
2.9: Polyatomic Ions
2.10: Names and Formulas of Ternary Ionic Compounds
2.11: Naming Binary Molecular Compounds
2.12: Naming Acids
2.13: Names and Formulas of Bases
2.14: Chapter 2 Exercises
2.14.1: Ions, Molecules, and Compounds (Exercises)
2.14.2: Answers Ions, Molecules, and Compounds (Exercises)
2.14.3: More Practice with Nomenclature (Exercises)

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2.1: Prelude to Ionic Bonding and Simple Ionic Compounds

We will see that the word salt has a specific meaning in chemistry, but to most people, this word refers to table salt. This kind of salt is used as a condiment throughout the world, but it was not always so abundant. Two thousand years ago, Roman soldiers received part of their pay as salt, which explains why the words salt and salary come from the same Latin root (salarium). Today, table salt is either mined or obtained from the evaporation of saltwater.

Table salt is sodium chloride (NaCl), which is a simple compound of two elements that are necessary for the human body to function properly. Sodium, for example, is important for nerve conduction and fluid balance. In fact, human blood is about a 0.9% sodium chloride solution, and a solution called normal saline is commonly administered intravenously in hospitals.

Although some salt in our diets is necessary to replenish the sodium and chloride ions that we excrete in urine and sweat, too much is unhealthy, and many people may be ingesting more salt than their bodies need. The RDI of sodium is 2,400 mg—the amount in about 1 teaspoon of salt—but the average intake of sodium in the United States is between 4,000 mg and 5,000 mg, partly because salt is a common additive in many prepared foods. Previously, the high ingestion of salt was thought to be associated with high blood pressure, but current research does not support this link. Even so, some doctors still recommend a low-salt diet (never a “no-salt” diet) for patients with high blood pressure, which may include using a salt substitute. Most salt substitutes use potassium instead of sodium, but some people complain that the potassium imparts a slightly bitter taste.

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2.2: Two Types of Bonding

Learning Objectives

- Define the octet rule.
- Describe how ionic bonds are formed.

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.

> *Of the noble gases, only krypton, xenon, and radon have been found to make compounds.*

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.

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**.

Despite our focus on the octet rule, we must remember that for small atoms, such as hydrogen, helium, and lithium, the first shell is, or becomes, the outermost shell and hold only two electrons. Therefore, these atoms satisfy a "**duet rule**" rather than the octet rule.

Example 2.2.1

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:

Sodium atom		**Sodium ion**	
11 protons	11+	11 protons	11+
11 electrons	11−	10 electrons	10−
	0 overall charge		+1 overall charge

Exercise 2.2.1

A fluorine atom has seven valence electrons. Do you think it is more likely for a fluorine atom to lose seven electrons or gain one electron to obtain an octet? Write the formula of the resulting ion.

Answer

The process that involves less number of electrons is more favorable. Fluorine would gain one electron. The formula of the resulting ion is F^-.

Key Takeaways

- Atoms have a tendency to have eight electrons in their valence shell.
- The attraction of oppositely charged ions is what makes ionic bonds.

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2.3: Cations



Figure **2.3.1** (Public Domain; User:Gold Guru/Wikipedia via Wikipedia)

Have you ever gone digging for gold?

When the prospectors during the California Gold Rush (1848-1855) searched for gold nuggets in the earth, they were able to find the nuggets because gold is an unreactive material that exists in its elemental state in many places. Not everyone was fortunate enough to find significantly large gold nuggets, but a number of miners did become very wealthy (of course, a large number of others went back home broke).

Many of the elements we know about do not exist in their native form. They are so reactive that they are found only in compounds. These non-elemental forms are known as ions. Their properties are very different from those of the elements they come from. The term comes from a Greek word meaning "move" and was first coined by Michael Faraday, who studied the movement of materials in an electrical field.

Cations

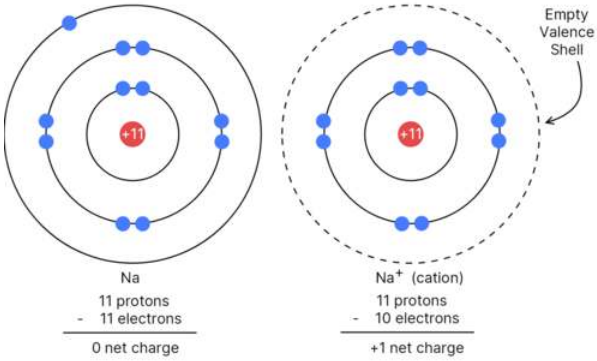


Figure **2.3.2**: Sodium loses an electron to become a cation. (CK-12 Curriculum Materials License; CK-12 Foundation via CK-12 Foundation)

Some elements lose one or more electrons in forming ions. These ions are known as "**cations**" because they are positively charged and migrate toward the negative electrode (**cathode**) in an electrical field. Looking at the periodic table below, we know that the group 1 elements are all characterized by having one ***s*** electron in the outer orbit; group 2 elements have two ***s*** electrons in the outer orbit. These electrons are loosely attached to the atom and can easily be removed, leaving more protons in the atom than there are electrons, so that the resulting ion has a positive charge. Cations can also be formed from electron loss to many of the transition elements.

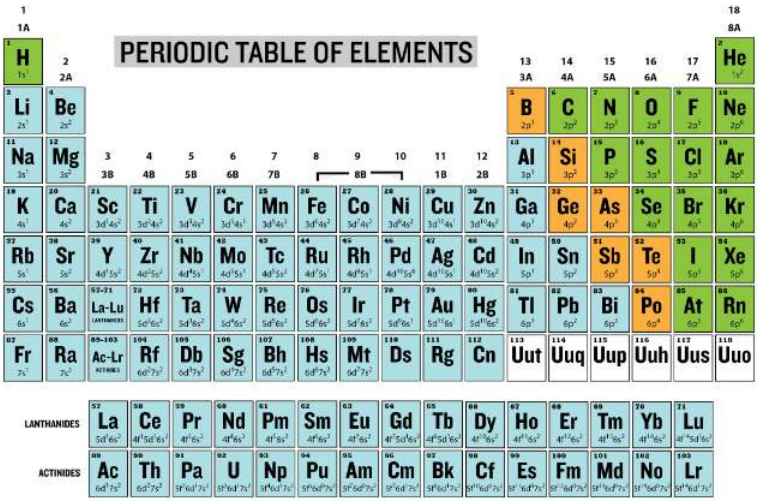
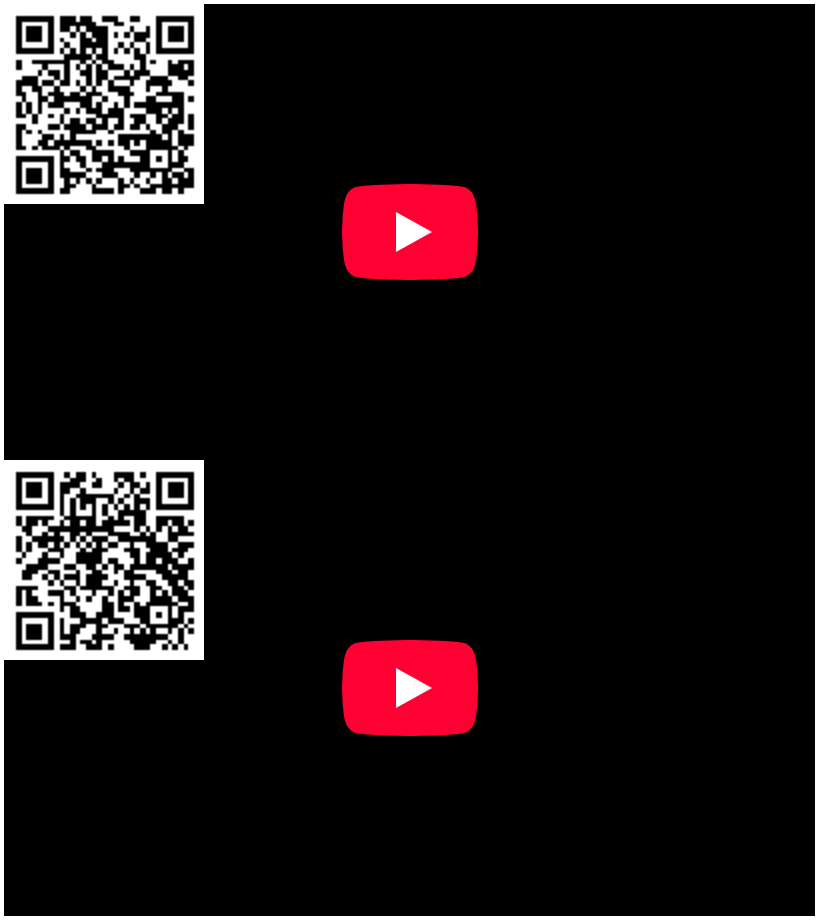


Figure **2.3.3**: Periodic Table of Elements, notated with group numbers. (CC BY-NC 3.0; Christopher Auyeung via CK-12)

The cations are designated by the symbol for the parent element and a plus charge as a superscript after the element symbol—the potassium cation would be indicated as $\mathbf{K^+}$. Note that the charge is placed after the symbol and not before it. The potassium ion is **monovalent**, meaning that it has lost one electron and has a $\mathbf{+1}$ charge. The symbol for the magnesium cation would be $\mathbf{Mg^{2+}}$ or $\mathbf{Mg^{++}}$ to indicate that it has lost two electrons and has a $\mathbf{+2}$ charge, so the magnesium cation would be referred to as a **divalent** cation.

The cations are simply named as the parent element. The sodium cation is still called "sodium". Often, the charge would be attached for clarity, so the sodium cation might be referred to as "sodium one plus".



Applications of Cations

Cations play important roles in our daily lives. Sodium, potassium, and magnesium ions are essential for such processes as blood pressure regulation and muscle contraction. Calcium ions are an important part of bone structure. Sodium ions can be used in water softeners to remove other harmful elements. We put sodium chloride (table salt) on our food and use it as a preservative.

Summary

- Cations are formed by the loss of one or two electrons from an element.
- Groups 1 and 2 elements form cations.
- Cations are named according to the parent element.
- Cation charges are indicated with a superscript following the chemical symbol.

Review

1. What is a cation?
2. Write the symbol for the barium cation.
3. Write the symbol for the cesium cation.
4. List three ways cations are useful.

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2.4: Anions



Figure **2.4.1** (Public Domain; Shane Anderson/NOAA via Commons Wikimedia, Kelp 300 [commons.wikimedia.org])

What does the amount of salt in seaweed tell us?

Before iodized salt was developed, some people experienced a number of developmental difficulties, including problems with thyroid gland function and intellectual disabilities. In the 1920s, we learned that these conditions could usually be treated easily with the addition of iodide anion to the diet. One easy way to increase iodide intake was to add the anion to table salt. This simple step greatly enhanced health and development. Large amounts of iodide ion are also found in seaweed such as kelp (see picture above) and saltwater fish.

When a metal loses an electron, energy is needed to remove that electron. The other part of this process involves the addition of the electron to another element. The electron adds to the outer shell of the new element. Just as the loss of the electron from the metal produces a full shell, when the electron or electrons are added to the new element, it also results in a full shell.

Anions

Anions are negative ions that are formed when a **nonmetal** atom gains one or more electrons. Anions are so named because they are attracted to the **anode** (positive field) in an electrical field. Atoms typically gain electrons so that they will have the electron configuration of a noble gas. All the elements in Group 17 have seven valence electrons due to the outer $ns^2\ np^5$ configuration. Therefore, each of these elements would gain one electron and become an anion with a -1 charge. Likewise, Group 16 elements form ions with a -2 charge, and the Group 15 nonmetals form ions with a -3 charge.

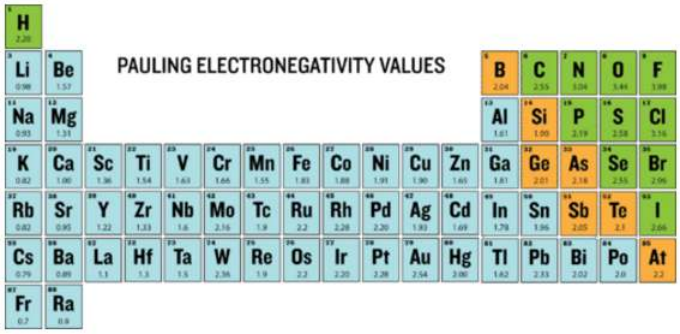
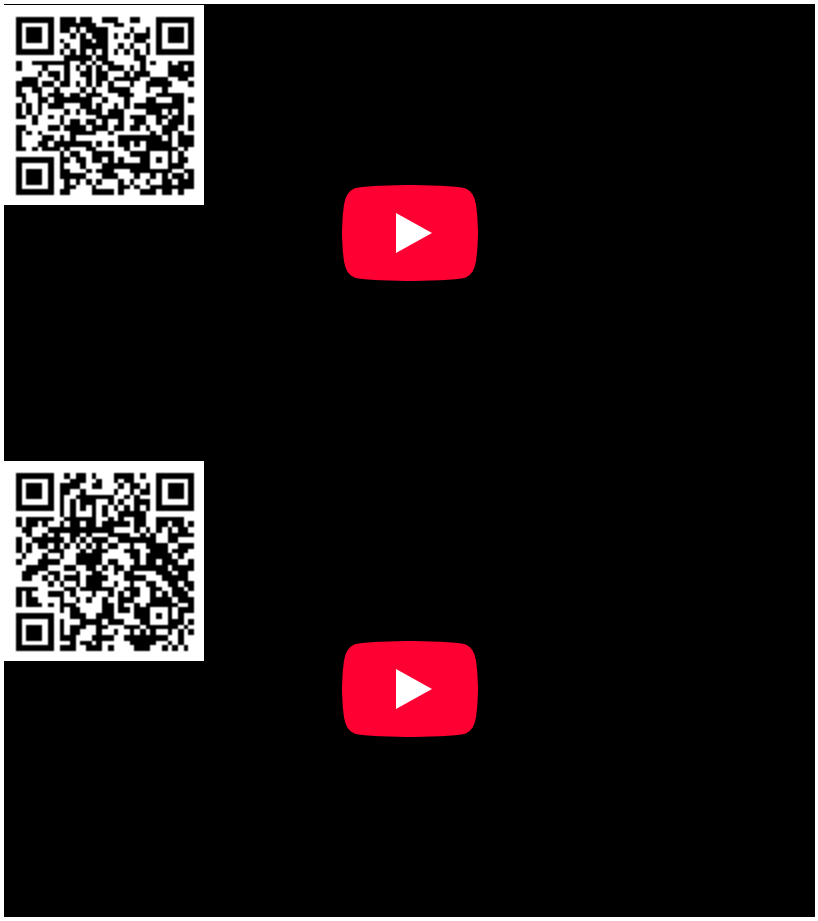


Figure **2.4.2** (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Naming anions is slightly different than naming cations. The ending of the element's name is dropped and replaced with the *-ide* suffix. For example, F^- is the fluoride ion, while O^{2-} is the oxide ion. As is the case with cations, the charge on the anion is indicated by a superscript following the symbol. Common anions are listed in the table below:

Table **2.4.1**

Anion Name	Symbol and Charge
fluoride	F^-
chloride	Cl^-
bromide	Br^-
iodide	I^-
oxide	O^{2-}
sulfide	S^{2-}
nitride	N^{3-}



Uses for Anions

Fluoride ion is widely used in water supplies to help prevent tooth decay. Chloride is an important component in ion balance in blood. Iodide ion is needed by the thyroid gland to make the hormone thyroxine.

Summary

- Anions are formed by the addition of one or more electrons to the outer shell of an atom.
- Group 17 elements add one electron to the outer shell, group 16 elements add two electrons, and group 15 elements add three electrons.
- Anions are named by dropping the ending of the element's name and adding *-ide*.

Review

1. What is an anion?
2. How are anions formed?
3. Why do anions form?
4. How are anions named?
5. List three examples of anions with names, charges, and chemical symbols.
6. List three ways anions are used.

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2.5: Transition Metal Ions



Figure **2.5.1** (Public Domain; the US Mint via Commons Wikimedia, Platinium eagle 101 [commons.wikimedia.org])

What kind of coin is this?

Most of us are familiar with the common coins: penny, nickel, dime, quarter. In some areas (such as Las Vegas), you might see large amounts of silver dollars (these get a little heavy in your pocket). But most of us have probably never seen a platinum eagle—an eagle-depicting coin that is held primarily by collectors. If you were to take a one-ounce platinum eagle into a store to buy one hundred dollars worth of items, the store owner will most likely not believe you when you claim that the coin is worth one hundred dollars. It would also be awkward and annoying if you lost one of these coins out of your pocket. Platinum is just one of several transition metals that is worth a lot of money (gold is another one).

The group 1 and 2 elements form cations through a simple process that involves the loss of one or more outer shell electrons. These electrons come from the ***s*** orbital and are removed very readily.

Transition Metal Ions

Most **transition metals** differ from the metals of Groups 1, 2, and 13 in that they are capable of forming more than one cation with different ionic charges. As an example, iron commonly forms two different ions. It can sometimes lose two electrons to form the Fe^{2+} ion, while at other times it loses three electrons to form the Fe^{3+} ion. Tin and lead, though members of the ***p*** block rather than the ***d*** **block**, also are capable of forming multiple ions.



Figure **2.5.2**: (Public Domain; Cepheus, modified by CK-12 Foundation via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

Ionic formation for transition metals is complicated by the fact that these elements have unfilled inner ***d*** shells. Although the next higher ***s*** orbitals are actually at a lower energy level than the ***d*** level, these ***s*** electrons are the ones that are removed during ionization.

The table below lists the names and formulas of some of the common transition metal ions:

Common Transition Metal Ions

Table 2.5.1: Common Transition Metal Ions

1+	2+	3+	4+
copper (I), Cu^{+}	cadmium, Cd^{2+}	chromium (III), Cr^{3+}	lead (IV), Pb^{4+}
gold (I), Au^{+}	chromium (II), Cr^{2+}	cobalt (III), Co^{3+}	tin (IV), Sn^{4+}
mercury (I), Hg_2^{2+}	cobalt (II), Co^{2+}	gold (III), Au^{3+}	
silver, Ag^{+}	copper (II), Cu^{2+}	iron (III), Fe^{3+}	
	iron (II), Fe^{2+}		
	lead (II), Pb^{2+}		
	manganese (II), Mn^{2+}		
	mercury (II), Hg^{2+}		
	nickel (II), Ni^{2+}		
	platinum (II), Pt^{2+}		
	tin (II), Sn^{2+}		
	zinc, Zn^{2+}		



Uses for Transition Metals

Because there are so many metals in this group, there are a wide variety of uses. Many of the metals are used in electronics, while others (such as gold and silver) are used in monetary systems. Iron is a versatile structural material. Cobalt, nickel, platinum, and other metals are employed as catalysts in a number of chemical reactions. Zinc is a significant component of batteries.

Summary

- Transition metals have unfilled inner ***d*** electron shells.
- Ions form primarily through loss of ***s*** electrons.
- Many transition metals can form more than one ion.
- Transition metals have a wide variety of applications.

Review

1. What is unique about the electron configurations of transition metals?
2. Which electrons of transition metal elements are most likely to be lost during ion formation?
3. How many ions can iron form?
4. Which transition metal forms only one ion?
5. List several uses for transition metals.

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2.6: Naming Binary Ionic Compounds



Figure 2.6.1 (Public Domain; Carl Gustav Calwer, Gustav Jäger, Emil Hochdanz via Wikipedia)

Do you know the proper name for this insect?

Proper naming (or *nomenclature*) is important for identification purposes. Medicine names must be precise so that the correct drug is given—one that will help the patient and not harm them. Biological classification of species requires accurate naming for proper categorization. Insects are categorized by genus and species, so they are uniquely identified. Names must be correct for the tracing of family trees in genealogical studies. Compounds used in chemical reactions must be correctly specified in order for the reaction to occur.

Binary Ionic Compounds

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**.

Naming Binary Ionic Compounds

When examining the formula of a compound in order to name it, you must first decide what kind of compound it is. For a binary ionic compound, a metal will always be the first element in the formula, while a nonmetal will always be the second. The metal cation is named first, followed by the nonmetal anion. Subscripts in the formula do not affect the name. The table below shows three examples.

Table 2.6.1: Naming Binary Ionic Compounds

Formula	Name
KF	potassium fluoride
$\mathrm{Na_3N}$	sodium nitride
$\mathrm{Ca_3P_2}$	calcium phosphide

Notice that in each of the formulas above, the overall charge of the compound is zero. Potassium ion is $\mathrm{K^+}$, while fluoride ion is $\mathrm{F^-}$. Since the magnitude of the charges is equal, the formula contains one of each ion. This would also be the case for a compound such as MgS, in which the ions are $\mathrm{Mg^{2+}}$ and $\mathrm{S^{2-}}$. For sodium nitride, the sodium ion is $\mathrm{Na^+}$, while the nitride ion is $\mathrm{N^{3-}}$. In order to make a neutral compound, three of the $1+$ sodium ions are required in order to balance out the single $3-$ nitride ion. So the Na is given a subscript of 3. For calcium phosphide, the calcium ion is $\mathrm{Ca^{2+}}$, while the phosphide ion is $\mathrm{P^{3-}}$. The least common multiple of 2 and 3 is 6. To make the compound neutral, three calcium ions have a total charge of $6+$, while two phosphide ions have a total charge of $6-$. The Ca is given a subscript of 3, while the P is given a subscript of 2.



Summary

- A binary ionic compound is a compound composed of a monatomic metal cation and a monatomic nonmetal anion.
- The metal part of the compound is named as the element.
- The nonmetallic part of the compound is named by dropping the end of the element and adding *-ide*.
- For binary compounds, it is not necessary to indicate the number of ions in the compound.

Review

1. What is a binary compound?
2. Which of the following is not a binary compound?
    a. NaCl
    b. KH_2PO_4
    c. KBr
3. Name the following compounds:
    a. NaBr
    b. $MgCl_2$
    c. LiI
    d. CaO
    e. $CuBr_2$
    f. FeO

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2.7: Stock System Naming



Figure **2.7.1** (Public Domain; XAVeRY via Wikipedia)

How is an mp3 player designed?

In describing many technological items, it's not enough to simply say what brand or model we have. We talk about details such as how much horsepower is “under the hood” for a car or how fast the chip is for our computer. Even a simple device like an mp3 player has more than one size. We can get an 8 MB player, or a 16 MB player. Designation of the item often is incomplete without other information as to its capabilities. Descriptive nomenclature such as we use in chemistry provides such information for compounds.

Transition metals have more than one possibility for ion formation. In order to name these compounds correctly, we need to be able to indicate which ion is involved in any given compound.

Naming Compounds Using the Stock System

Naming compounds that involve transition metal cations necessitates use of the **Stock system**. Consider the binary ionic compound $\mathbf{FeCl_3}$. To simply name this compound "iron chloride" would be incomplete, because iron is capable of forming two ions with different charges. The name of any iron-containing compound must reflect which iron ion is in the compound. In this case, the subscript in the formula indicates that there are three chloride ions, each with a $1-$ charge. Therefore, the charge of the single iron ion must be $3+$. The correct name of $\mathbf{FeCl_3}$ is iron (III) chloride, with the cation charge written as the Roman numeral. Here are several other examples:

Table **2.7.1**

Formula	Name
$\mathbf{Cu_2O}$	copper (I) oxide
$\mathbf{CuO}$	copper (II) oxide
$\mathbf{SnO_2}$	tin (IV) oxide

The first two examples are both oxides of copper (shown in the figure below). The ratio of copper ions to oxide ions determines the name. Since the oxide ion is O^{2-}, the charges of the copper ion must be $1+$ in the first formula and $2+$ in the second formula. In the third formula, there is one tin ion for every two oxide ions. This means that the tin must carry a $4+$ charge, making the name tin (IV) oxide.



Figure **2.7.2**: Copper (I) oxide, a red solid, and copper (II) oxide, a black solid, are different compounds because of the charge of the copper ion. (Public Domain; Ondřej Mangl via Copper(I) oxide: http://commons.wikimedia.org/wiki/File:Cu2O.png; Copper(II) oxide: http://commons.wikimedia.org/wiki/File:Oxid_m%25C4%259B%25C4%258Fnat%25C3%25BD.PNG)



Summary

- The Stock system allows for the specification of transition metal ionic charge when naming ionic compounds.
- Roman numerals are used to indicate the amount of positive charge on the cation.

Review

1. What is the Stock system?
2. For which group of metal ions would we use the Stock system?
3. What does the Roman numeral stand for?
4. Assign a Roman numeral to each of the following cations:
    a. Sn^{4+}
    b. Fe^{3+}
    c. Co^{2+}
    d. Pb^{4+}

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2.8: Formulas for Binary Ionic Compounds

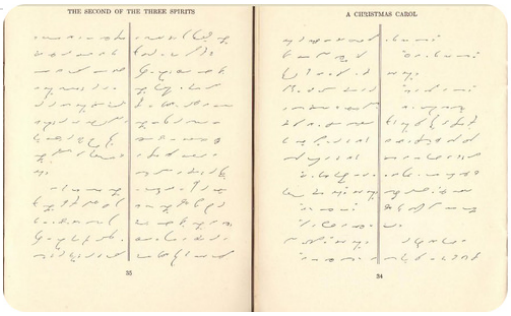


Figure 2.8.1 (Public Domain; Winifred Kenna via Commons Wikimedia, Gregg shorthand A Christmas Carol [commons.wikimedia.org])

How does shorthand work?

Shorthand was a very popular way of recording speech, especially in dictating letters and in court testimony. Instead of trying to write out all the words, the person taking the dictation would use a set of symbols that represented syllables or words. The pages above show a shorthand version of "A Christmas Carol" written by Charles Dickens. Unless you know shorthand, the passage is meaningless. But knowing shorthand allows you to read this classic story. Different professions also use a type of shorthand in communication to save time. Chemists use chemical symbols in combination to indicate specific compounds. There are two advantages to this approach:

1. The compound under discussion is clearly described so that there can be no confusion about its identity.
2. Chemical symbols represent a universal language that all chemists can understand, no matter what their native language is.

Writing Formulas for Binary Ionic Compounds

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out. Consider the compound aluminum nitride. The ions are:

$$Al^{3+} \quad N^{3-}$$

Since the ions have charges that are equal in magnitude, one of each will be the lowest ratio of ions in the formula. The formula for aluminum nitride is AlN.

The ions for the compound lithium oxide are:

$$Li^{+} \quad O^{2-}$$

In this case, two lithium ions are required to balance out the charge of one oxide ion. The formula of lithium oxide is Li_2O.

An alternative way to write a correct formula for an ionic compound is to use the crisscross method. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped. Shown below is the crisscross method for aluminum oxide.



Figure 2.8.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The red arrows indicate that the 3 from the $3+$ charge will cross over to become the subscript of the O. The 2 from the $2-$ charge will cross over to become the subscript of the Al. The formula for aluminum oxide is Al_2O_3.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions. In the case of aluminum nitride, the crisscross method would yield a formula of Al_3N_3, which is not correct. It must be reduced to AlN. Following the crisscross method to write the formula for lead (IV) oxide would involve the following steps:

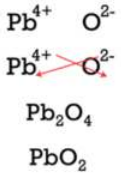
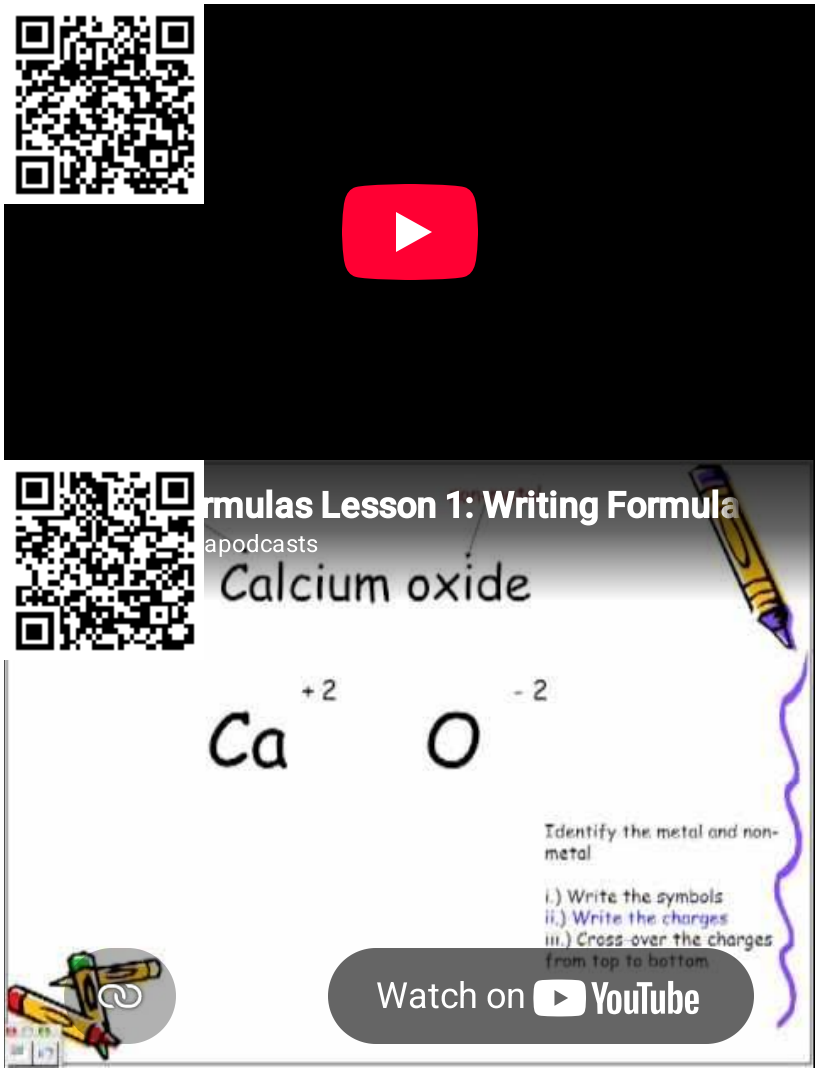


Figure **2.8.3** (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The crisscross first yields Pb_2O_4 for the formula, but that must be reduced to the lower ratio, and PbO_2 is the correct formula.



Summary

- Formulas for binary compounds begin with the metal followed by the nonmetal.
- Positive and negative charges must cancel each other out.
- Ionic compound formulas are written using the lowest ratio of ions.

Review

1. Write formulas for the binary ionic compounds formed between the following pairs of elements:
    a. cesium and fluorine
    b. calcium and sulfur
    c. aluminum and chlorine
    d. zinc and nitrogen

2. Write the formula and give the name for the compound formed by the following ions:
    a. Fe^{3+} and O^{2-}
    b. Ni^{2+} and S^{2-}
    c. Au^{+} and Cl^{-}
    d. Sn^{4+} and I^{-}
3. Give names for the following compounds:
    a. Ag_2S
    b. PdO
    c. $PtCl_4$
    d. V_2O_5

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2.9: Polyatomic Ions



Figure 2.9.1 (Public Domain; Huhu via Wikipedia)

Have you ever read the story of Romeo and Juliet?

Shakespeare wrote, "A rose by any other name would smell as sweet" as part of the romantic balcony scene between Romeo and Juliet in the famous play. The two families were bitter rivals, but Juliet meant this statement to mean that she loved Romeo no matter what his name was. Some names are simple—we know Romeo mainly as Romeo. Most of us have a first name, middle name (often not used), and last name. In some cultures, names will be much more complex. The full name of the famous 20^{th} century artist Pablo Picasso is Pablo Diego José Francisco de Paula Juan Neponuceno María de los Remedios Cipriano de la Santísima Trinidad Martyr Patricio Clito Ruíz y Picasso.

Many materials exist as simply binary compounds composed of a metal cation and a nonmetal anion, with each ion consisting of only one type of atom. Other combinations of atoms also exist, either as larger ionic complexes or complete molecules. Some of the most useful materials we work with contain polyatomic ions.

Polyatomic Ions

A **polyatomic** ion is an ion composed of more than one atom. The ammonium ion consists of one nitrogen atom and four hydrogen atoms. Together, they comprise a single ion with a $1+$ charge and a formula of NH_4^+. The carbonate ion consists of one carbon atom and three oxygen atoms, and carries an overall charge of $2-$. The formula of the carbonate ion is CO_3^{2-}. The atoms of a polyatomic ion are tightly bonded together, and so the entire ion behaves as a single unit. The figures below show several models.

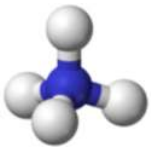


Figure 2.9.2: The ammonium ion (NH_4^+) is a nitrogen atom (blue) bonded to four hydrogen atoms (white). (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)



Figure 2.9.3: The hydroxide ion (OH^-) is an oxygen atom (red) bonded to a hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

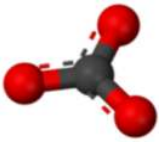


Figure 2.9.4: The carbonate ion (CO_3^{2-}) is a carbon atom (black) bonded to three oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

The table below lists a number of polyatomic ions by name and by formula. The heading for each column indicates the charge on the polyatomic ions in that group. Note that the vast majority of the ions listed are anions—there are very few polyatomic cations.

Table 2.9.1: Common Polyatomic Ions

1−	2−	3−	1+	2+
acetate, CH_3COO^-	carbonate, CO_3^{2-}	arsenate, AsO_3^{3-}	ammonium, NH_4^+	dimercury, Hg_2^{2+}
bromate, BrO_3^-	chromate, CrO_4^{2-}	phosphite, PO_3^{3-}		
chlorate, ClO_3^-	dichromate, $Cr_2O_7^{2-}$	phosphate, PO_4^{3-}		
chlorite, ClO_2^-	hydrogen phosphate, HPO_4^{2-}			
cyanide, CN^-	oxalate, $C_2O_4^{2-}$			
dihydrogen phosphate, $H_2PO_4^-$	peroxide, O_2^{2-}			
hydrogen carbonate, HCO_3^-	silicate, SiO_3^{2-}			
hydrogen sulfate, HSO_4^-	sulfate, SO_4^{2-}			
hydrogen sulfide, HS^-	sulfite, SO_3^{2-}			
hydroxide, OH^-				
hypochlorite, ClO^-				
nitrate, NO_3^-				
nitrite, NO_2^-				
perchlorate, ClO_4^-				
permanganate, MnO_4^-				

The vast majority of polyatomic ions are anions, many of which end in *-ate* or *-ite*. Notice that in some cases, such as nitrate (NO_3^-) and nitrite (NO_2^-), there are multiple anions that consist of the same two elements. In these cases, the difference between the ions is in the number of oxygen atoms present, while the overall charge is the same. As a class, these are called oxoanions. When there are two oxoanions for a particular element, the one with the greater number of oxygen atoms gets the *-ate* suffix, while the one with the fewer number of oxygen atoms gets the *-ite* suffix. The four oxoanions of chlorine are as follows:

- ClO^-, hypochlorite
- ClO_2^-, chlorite
- ClO_3^-, chlorate
- ClO_4^-, perchlorate

In cases such as this, the ion with one more oxygen atom than the *-ate* anion is given a *per-* prefix. The ion with one fewer oxygen atom than the *-ite* anion is given a *hypo-* prefix.



Summary

- Polyatomic ions contain more than one type of atom in the ion.
- The majority of polyatomic ions are anions that are named ending in *-ate* or *-ite*.
- Some anions have multiple forms and are named accordingly.

Review

1. Write the formulas for the following ions:
    a. ammonium
    b. carbonate
    c. sulfate
    d. phosphate
2. Name the following ions:
    1. PO_3^{3-}
    2. SiO_3^{2-}
    3. OH^-
    4. MnO_4^-

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2.10: Names and Formulas of Ternary Ionic Compounds



Figure **2.10.1** (Public Domain; Jon Sullivan via Commons Wikimedia, Pepperoni Pizza [commons.wikimedia.org])

How many varieties of pizza can you think of?

Many people enjoy eating pizza. When someone orders a pizza (either by phone, online, or in a restaurant), they don't just order "pizza". Even the simplest "binary" pizza will have a topping and cheese—maybe pepperoni, maybe something else. However, many pizzas quickly become more complex. One person may want pepperoni and sausage, another may wish to order Canadian bacon and pineapple; then you have the folks that can't decide, so they order half-this and half-that. The combinations may be more complex, but the same basic ideas about pizza are valid.

Ternary Ionic Compounds

Not all ionic compounds are composed of only monatomic ions. A **ternary ionic compound** is an ionic compound composed of three elements. In a typical ternary ionic compound, there is still one type of cation and one type of anion involved. The cation, the anion, or both, are polyatomic ions.

Naming Ternary Ionic Compounds

The process of naming ternary ionic compounds is the same as naming binary ionic compounds. The cation is named first, followed by the anion. Some examples are shown in the table below:

Table **2.10.1**: Examples of Ternary Ionic Compounds

Formula	Name
$NaNO_3$	sodium nitrate
NH_4Cl	ammonium chloride
$Fe(OH)_3$	iron (III) hydroxide

When more than one polyatomic ion is present in a compound, the formula of the ion is placed in parentheses with a subscript outside of the parentheses that indicates how many of those ions are in the compound. In the last example above, there is one Fe^{3+} cation and three OH^- anions.

Writing Formulas for Ternary Ionic Compounds

Writing a formula for a ternary ionic compound also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion. Use the crisscross method to ensure that the final formula is neutral. Calcium nitrate is composed of a calcium cation and a nitrate anion.

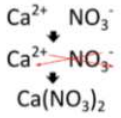


Figure **2.10.2** (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The charge is balanced by the presence of two nitrate ions and one calcium ion. Parentheses are used around the nitrate ion because more than one of the polyatomic ion is needed. If only one polyatomic ion is in a formula, parentheses are not used. As an example,

the formula for calcium carbonate is $CaCO_3$. The carbonate ion carries a $2-$ charge and so exactly balances the $2+$ charge of the calcium ion.

There are two polyatomic ions that produce unusual formulas. The Hg_2^{2+} ion is called either the dimercury ion or, preferably, the mercury (I) ion. When bonded with an anion with a $1-$ charge, such as chloride, the formula is Hg_2Cl_2. Because the cation consists of two Hg atoms bonded together, this formula is not reduced to $HgCl$. Likewise, the peroxide ion, O_2^{2-}, is also a unit that must stay together in its formulas. For example, the formula for potassium peroxide is K_2O_2.

Summary

- Ternary compounds are composed of three or more elements.
- Ternary compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.

Review

1. What is a ternary compound?
2. What is the basic rule for naming ternary compounds?
3. Write the formulas for the following compounds:
    a. mercury(II) nitrate
    b. ammonium phosphate
    c. calcium silicate
    d. lead(II) chromate
4. Name the following compounds:
    a. $KClO_3$
    b. Rb_2SO_4
    c. $Cd(NO_3)_2$
    d. NaCN

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2.11: Naming Binary Molecular Compounds



Figure **2.11.1** (Public Domain; Anonymous via Wikipedia)

Why do so many relatives in royalty share the same name?

Some families name a son (usually the first born) after his father. So, it is somewhat common to meet a John Smith, Jr. named after John Smith the father. Certain families may name a grandson John Smith III. Countries with long histories of royalty take the naming even further. One line of kings named Henry goes up to Henry the Eighth (not the nicest guy in the world—he had six wives and two of them met untimely ends). The use of numbering for names adds clarity to a system—it's easily discernible which Henry is being spoken of.

Inorganic chemical compounds can be broadly classified into two groups: ionic compounds and molecular compounds. The structure of all ionic compounds is an extended three-dimensional array of alternating positive and negative ions. Since ionic compounds do not take the form of individual molecules, they are represented by empirical formulas. Now we will begin to examine the formulas and nomenclature of molecular compounds.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride ($NaCl$). Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **valence electrons** in such a way that a **bond** forms between pairs of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

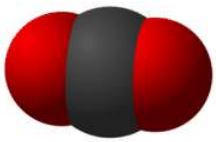


Figure **2.11.2**: Carbon dioxide molecules consist of a central carbon atom bonded to two oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.



Figure **2.11.3**: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines. (Public Domain; User:Greenhorn1/Wikimedia Commons via Wikipedia)

A binary molecular compound is a molecular compound that is composed of two elements. The elements that combine to form binary molecular compounds are both nonmetal atoms. This contrasts with ionic compounds, which are formed from a metal ion and a nonmetal ion. Therefore, binary molecular compounds are different because ionic charges cannot be used to name them or to write their formulas. Another difference is that two nonmetal atoms will frequently combine with one another in a variety of ratios. Consider the elements nitrogen and oxygen. They combine to make several compounds including NO, NO_2, and N_2O. They cannot all be called nitrogen oxide. How would someone know which one you were talking about? Each of the three compounds has very different properties and reactivities. A system to distinguish between compounds such as these is necessary.

Prefixes are used in the names of binary molecular compounds to identify the number of atoms of each element. The table below shows the prefixes up to ten.

Table **2.11.1**: Numerical Prefixes

Number of Atoms	Prefix
1	*mono-*
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*
6	*hexa-*
7	*hepta-*
8	*octa-*
9	*nona-*
10	*deca-*

The rules for using the prefix system of nomenclature of binary compounds can be summarized as follows:

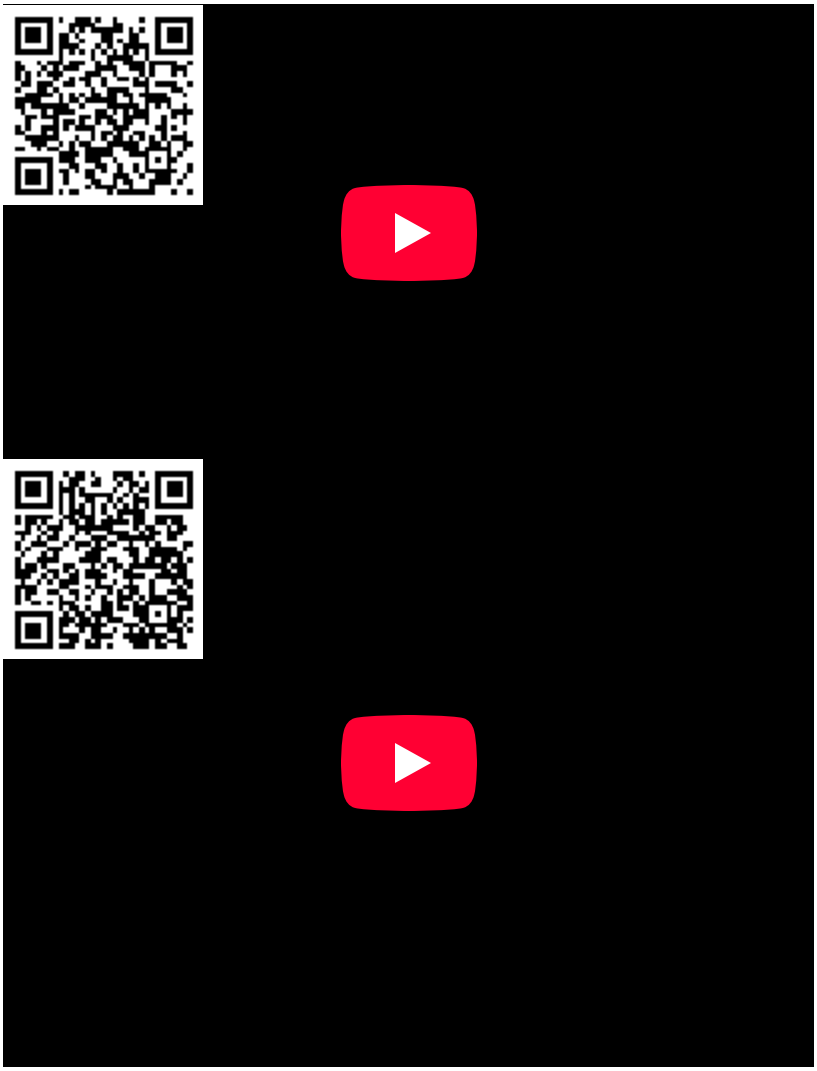
1. Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
2. When naming, the appropriate prefix is used only if there are more than one atom of that element in the formula.
3. The second element is named after the first, but with the ending of the element's name changed to *-ide*. The appropriate prefix is always used for the second element.
4. The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms are tetroxide instead of tetraoxide.

Some examples of molecular compounds are listed in Table **2.11.2**.

Table **2.11.2**

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the *mono-* prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono-* and the *a* of *hepta-* are dropped from the name when paired with oxide.



Summary

- Molecular compounds are inorganic compounds that take the form of discrete molecules.
- The atoms of these compounds are held together by covalent bonds.
- Prefixes are used to indicate the number of atoms of an element that are in the compound.

Review

1. What are molecular compounds?
2. How do they differ from ionic compounds?

3. What holds the atoms of a molecule together?
4. Are the elements of molecular compounds metals or non-metals?
5. What is the prefix for the number of atoms of an element when there are four atoms of that element in the compound?
6. What is the prefix for the number of atoms of an element when there are 7 atoms of that element in the compound?
7. Name the following compounds:
    a. ClF_3
    b. As_2O_5
    c. B_4H_{10}

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2.12: Naming Acids



Figure 2.12.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How is gold tested?

A spot test for gold has been in use for decades. The sample is first treated with nitric acid. Other metals may react or dissolve in this acid, but gold will not. Then the sample is added to a mixture of nitric acid and hydrochloric acid. Gold will only dissolve in this mixture. The term "acid test" arose from the California gold rush in the late 1840s, when this combination was used to test for the presence of real gold. It has since come to mean "tested and approved" in a number of fields.

Acids

An **acid** can be defined in several ways. The most straightforward definition is that an acid is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions (H^+) when dissolved in water.



Figure 2.12.2: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids. (CC BY 2.0; (A) Becky Cortino; (B) Flickr: isox4 via (A) http://www.flickr.com/photos/mediasavvy/8239231530; (B) http://www.flickr.com/photos/isox4/5167980026)

This is a different type of compound than the others we have seen so far. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like $NaCl$. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$HCl \rightarrow H^+ + Cl^-$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here. A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. An **oxoacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic. The name of all monatomic ions ends in *-ide*. The majority of polyatomic ions end in either *-ate* or *-ite*, though there are a few exceptions such as the cyanide ion (CN^-). It is the suffix of the anion that determines how the acid is named, as displayed in the rules and table below.

Table 2.12.1: Naming System for Acids

Anion Suffix	Example	Name of Acid	Example
-ide	chloride (Cl^-)	hydro_____ic acid	hydrochloric acid (HCl)
-ate	sulfate (SO_4^{2-})	_____ic acid	sulfuric acid (H_2SO_4)
-ite	nitrite (NO_2^-)	____ous acid	nitrous acid (HNO_2)

The three different suffixes that are possible for the anions lead to three rules:

1. When the anion ends in *-ide*, the acid name begins with the prefix *hydro-*. The root of the anion name goes in the blank (*chlor* for chloride), followed by the suffix *-ic*. HCl is hydrochloric acid because Cl^- is the chloride ion. HCN is hydrocyanic acid because CN^- is the cyanide ion.
2. When the anion ends in *-ate*, the name of the acid is the root of the anion followed by the suffix *-ic*. There is no prefix. H_2SO_4 is sulfuric acid (not sulfic) because SO_4^{2-} is the sulfate ion.
3. When the anion ends in *-ite*, the name of the acid is the root of the anion followed by the suffix *-ous*. Again, there is no prefix. HNO_2 is nitrous acid because NO_2^- is the nitrite ion.

Note how the root for a sulfur-containing oxoacid is *sulfur-* instead of just *sulf-*. The same is true for a phosphorus-containing oxoacid. The root is *phosphor-* instead of simply *phosph-*.

Many foods and beverages contain citric acid. Vinegar is a dilute solution of acetic acid. Car batteries contain sulfuric acid that helps in the release of electrons to create electricity.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Since H^+ ions carry a single positive charge, the number of H^+ ions in the formula is equal to the quantity of negative charge on the anion. Two examples from the table above illustrate this point. The chloride ion carries a $1-$ charge, so only one H is needed in the formula of the acid (HCl). The sulfate ion carries a $2-$ charge, so two Hs are needed in the formula of the acid (H_2SO_4). Another way to write the correct formula is to utilize the crisscross method, shown below for sulfuric acid.

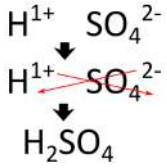
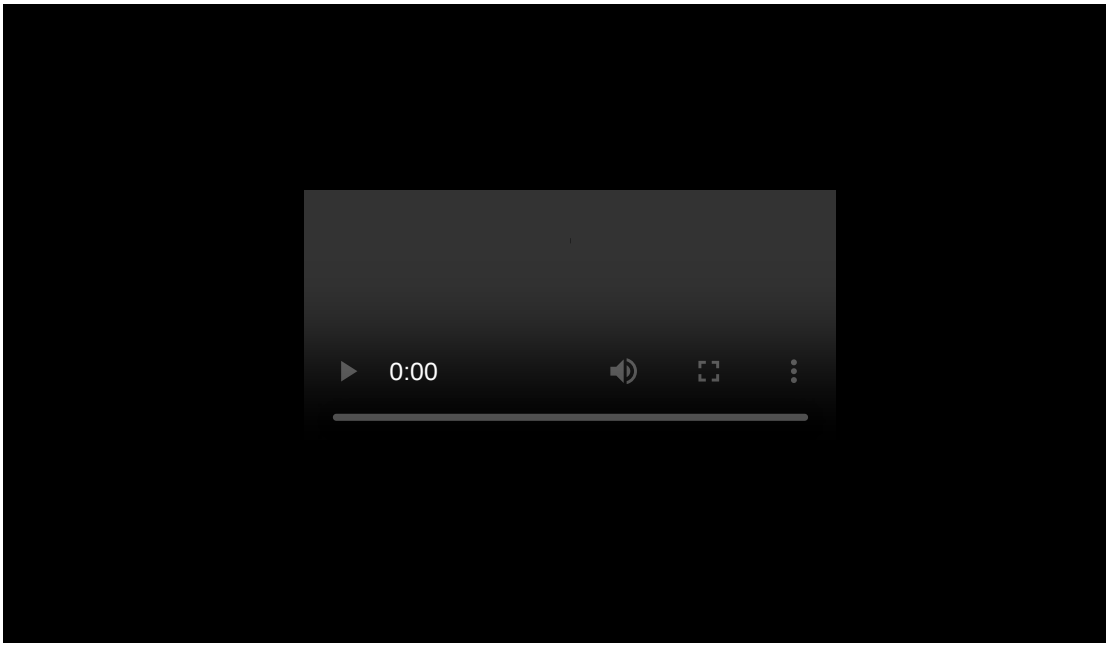


Figure **2.12.3**: Crisscross approach to writing the formula for sulfuric acid. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)



Summary

- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxoacids contain hydrogen, oxygen, and one other element.
- The name of the acid is based on the anion attached to the hydrogen.
- Writing formulas for acids follows the same rules as for binary ionic compounds.

Review

1. Define “acid.”
2. What is a binary acid?
3. What is an oxoacid?

4. Name the following acids:
    a. H_2SO_4
    b. HCN
    c. HCl
    d. H_3PO_4
5. Write formulas for the following acids:
    a. hydrobromic acid
    b. perchloric acid
    c. nitrous acid

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2.13: Names and Formulas of Bases



Figure 2.13.1 (Public Domain; Unknown Author via Wikipedia)

What different things are these workers doing to make soap?

Soap making has a long history. Until recently, soap was made using animal fats and lye from wood ashes. The lye served as a base to break down the fats and help form the soap. Needless to say, unless the soap was washed to remove the lye, it was very harsh on the skin. Many families would make their own soap by boiling the lye and fat in a large kettle over an open fire—a long and hot task.

Bases

A **base** can be simply defined as an ionic compound that produces **hydroxide ions** when dissolved in water. One of the most commonly used bases is sodium hydroxide, illustrated below.

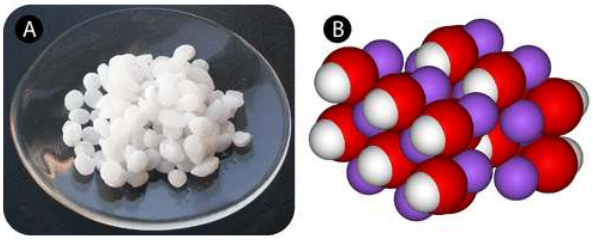


Figure 2.13.2: (A) Sodium hydroxide, a base, is a solid that is typically produced as small white pellets. (B) The structure of sodium hydroxide is an extended three-dimensional network. The purple spheres are the sodium ions (Na^+). The red and white spheres are oxygen and hydrogen atoms, respectively, which are bonded together to form hydroxide ions (OH^-). (Public Domain; (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:SodiumHydroxide.jpg; (B) http://commons.wikimedia.org/wiki/File:Sodium-hydroxide-crystal-3D-vdW.png)

Names and Formulas of Bases

There is no special system for naming bases. Since they all contain the OH^- anion, names of bases end in *hydroxide*. The cation is simply named first. Some examples of names and formulas for bases are shown in the table below.

Names and Formulas of Bases

Table 2.13.1

Formula	Name
$NaOH$	sodium hydroxide
$Ca(OH)_2$	calcium hydroxide
NH_4OH	ammonium hydroxide

Notice that because bases are ionic compounds, the number of hydroxides in the formula does not affect the name. The compound must be neutral, so the charges of the ions are balanced just as for other ionic compounds. The sodium ion (Na^+) requires one OH^-

ion to balance the charge, so the formula is $NaOH$. Calcium (Ca^{2+}) requires two OH^- ions to balance the charge, so the formula is $Ca(OH)_2$. Hydroxide ion is a polyatomic ion, and must be in parentheses when there is more than one in a formula.

Summary

- Bases are ionic compounds that produce hydroxide ions when dissolved in water.
- The cation is named first followed by "*hydroxide*."

Review

1. What is a base?
2. What is the charge on the hydroxide anion?
3. Name the following bases:
    a. LiOH
    b. $Mg(OH)_2$
    c. $Fe(OH)_3$
4. Write the formulas for the following bases:
    a. nickel (II) hydroxide
    b. aluminum hydroxide
    c. silver hydroxide

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2.14: Chapter 2 Exercises

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2.14.1: Ions, Molecules, and Compounds (Exercises)

Two Types of Bonding Exercises

1. What is the octet rule?

2. How are ionic bonds formed?

3. Why is an ionic compound unlikely to consist of two positively charged ions?

4. Why is an ionic compound unlikely to consist of two negatively charged ions?

5. A calcium atom has two valence electrons. Do you think it will lose two electrons or gain six electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.

6. An aluminum atom has three valence electrons. Do you think it will lose three electrons or gain five electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.

7. A selenium atom has six valence electrons. Do you think it will lose six electrons or gain two electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.

8. An iodine atom has seven valence electrons. Do you think it will lose seven electrons or gain one electron to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.

Ions & Bonding Exercises

1. Fill in the following table:

Nuclear Symbol	Atomic Number	Mass Number	Number of Protons	Number of Electrons	Number of Neutrons	Charge
$^{40}_{18}\mathrm{Ar}$						
		39	19	18		
	16				20	-2

2. For each of the following elements, indicate whether it is a main group element (MG), transition metal (TM), or inner transition metal (ITM). If the element is a main group element, indicate the group number and whether it is a metal, a nonmetal or a metalloid. Also indicate the ionic charge that each would most likely create.

a. **Sr** (atomic # 38)
b. **Br** (atomic # 35)
c. **Mo** (atomic # 42)
d. **P** (atomic # 15)
e. **B** (atomic # 5)
f. **U** (atomic # 92)
g. **Sn** (atomic # 50)
h. **Hg** (atomic # 80)

3. Provide the common names of Groups 1, 2, 17 and 18.

4. Give an example of at least one element made up of molecules and one compound made up of molecules.

5. For each of the following atoms, indicate whether it forms a positive or a negative ion, and include the ion charge.

a. **Na**
b. **Ba**
c. **Cl**
d. **S**
e. **Ag**

Nomenclature Exercises

1. Which of the following are ionic compounds? Which are covalent compounds? Name each compound.

a. N_2O
b. K_2O
c. PCl_3
d. $AlPO_4$
e. HCl
f. NH_4F
g. $Pb(NO_2)_2$
h. H_2SO_3

2. Name the following ionic compounds:

a. $CaCO_3$
b. ZnS
c. CuOH
d. $Mg(ClO_4)_2$

3. Give the formulas for the following ionic compounds:

a. potassium phosphate
b. ammonium sulfate
c. cobalt(II) hydroxide
d. iron(III)nitride

4. Provide the formulas for the following covalent compounds:

a. phosphorus triiodide
b. dinitrogen pentoxide
c. chloric acid

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2.14.2: Answers Ions, Molecules, and Compounds (Exercises)

Two Types of Bonding Answers

1. The octet rule is the concept that atoms tend to have eight electrons in their valence electron shell.

2. Ionic bonds are formed by the attraction between oppositely charged ions.

3. Positive charges repel each other, so an ionic compound is not likely between two positively charged ions.

4. Negative charges repel each other also.

5. Ca atom is more likely to lose two electrons. It will become Ca^{2+} ion.

6. An Al atom is more likely to lose three electrons. It will become Al^{3+} ion.

7. Selenium is more likely to gain two electrons. It will become Se^{2-} ion.

8. Iodine is more likely to gain one electron. It will become I^- ion.

Chemical Principles (Zumdahl and Decoste)
Textmap Alternative

Ions & Bonding Answers

1.

Nuclear Symbol	Atomic Number	Mass Number	Number of Protons	Number of Electrons	Number of Neutrons	Charge
$^{40}_{18}Ar$	18	40	18	18	22	0
$^{39}_{19}K^{+}$	19	39	19	18	20	+1
$^{36}_{16}S^{2-}$	16	36	16	18	20	-2

2.

- a. MG, Group 2, metal, +2
- b. MG, Group 17, nonmetal, -1
- c. TM, +6
- d. MG, Group 15, nonmetal, -3
- e. MG, Group 13, metalloid, +3

- f. ITM, +6
- g. MG, Group 14, metal, +4
- h. TM, +2

3. Group 1 = alkali metals; Group 2 = alkaline earth metals; Group 17 = halogens; Group 18 = noble gases.

4. Elements: O_2, N_2, Cl_2, or any other diatomic element. Compounds: CO_2, H_2O, NH_3, or any other covalent compound.

5.

- a. positive, +1
- b. positive, +2
- c. negative, -1
- d. negative, -2
- e. positive, +1

Nomenclature Answers

1.

- a. covalent, dinitrogen oxide or dinitrogen monoxide
- b. ionic, potassium oxide
- c. covalent, phosphorus trichloride
- d. ionic, aluminum phosphate
- e. covalent, hydrochloric acid
- f. ionic, ammonium fluoride
- g. ionic, lead(II) nitrite
- h. covalent, sulfurous acid

2.

- a. calcium carbonate
- b. zinc sulfide
- c. copper(I) hydroxide
- d. magnesium perchlorate

3.

- a. K_3PO_4
- b. $(NH_4)_2SO_4$
- c. $Co(OH)_2$
- d. FeN

4.

- a. PI_3
- b. N_2O_5
- c. $HClO_3$

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2.14.3: More Practice with Nomenclature (Exercises)

Template:HideTOC

These are homework exercises to accompany the Textmap created for Chemistry: A Molecular Approach by Nivaldo Tro. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Q5.29

Classify each compound as ionic or molecular

a. Fe_2O_3
b. SO_4
c. H_2O
d. $NaCl$

Strategy

A. First you need to have good understanding of the difference between Ionic and Molecular.
B. Ionic compound is formed between a metal and a non-metal
C. Molecular compound is formed between two non-metals
D. Then you need to figure out which elements are metals and non-metals by looking at the periodic table

Answers

a. Ionic
b. Molecular
c. Molecular
d. Ionic

Q5.33

Question: How many types of each atom are in each formula?

a. $(NH_4)_2SO_4$
b. $CaCO_3$
c. $Mg(OH)_2$
d. $NaHCO_3$

What We Know: A chemical formula gives the number of atoms in a compound and the element ratio within a compound.

What It's Asking For: Determine the number of each atom in the given compounds.

Strategy: The subscripts after each atom give the number of atoms in the compound. If the subscript is outside of the parenthesis, the subscript is applied to each atom within the parenthesis.

Solution:

a. 2 N, 8 H, 1 S, 4 O
b. 1 Ca, 1 C, 3 O
c. 1 Mg, 2 O, 2 H
d. 1 Na, 1 H, 1 C, 3 O

Q5.48

Write four different ionic compounds that can be formed using the following elements: Sodium, Chlorine, Aluminum and Sulfur.

Strategy:

A. Identify which elements are the metals and which are the nonmetals that will be arranged to form the ionic compounds.

B. Ionic compounds always contain positive and negative ions; therefore identify the positive charge of each metal identified in the above step and the negative charge of each nonmetal identified in the above step using the Periodic Table for reference.

C. Combine each metal with each nonmetal to form the initial foundation for the ionic compound. The sum of the positive charges must equal the sum of the negative charges within the compound which means that the number of each ion may have to be adjusted to achieve the balance of ions. If so, then the formula of the ionic compound will always reflect the smallest whole number ratio of ions. For example: 2:1, 2:3, etc.

Solution

- $NaCl$
- Na_2S
- $AlCl_3$
- Al_2S_3

Q5.49

Write the formula for the compound that forms between Potassium and each of the polyatomic ions listed below:

a. Sulfate
b. Perchlorate
c. Nitrate
d. Phosphate

Strategy

A. Identify the compound listed
B. Write out the chemical equation for each reaction
C. Balance the equations

Hints

- The formula for Sulfate is SO_4^{2-}
- The formula for Perchlorate is ClO_4^{-}
- The formula for Nitrate is NO_3^{-}
- The formula for Phosphate is PO_4^{3-}

Solution

a. $2K^+ + SO_4^{2-} \rightarrow K_2SO_4$
b. $K^+ + ClO_4^- \rightarrow K_2CLO_4$
c. $K^+ + NO_3^- \rightarrow KNO_3$
d. $3K^+ + PO_4^{3-} \rightarrow K_3PO_4$

Q5.60

Write the name from the formula of the formula from the name for each hydrated ionic compound

a. Iron(III) phosphate tetrahydrate
b. $CsBr \cdot 4H_2O$
c. Cobalt(II) chloride hexahydrate
d. $LiI \cdot 3H_2O$

Strategy

A. Name the cation and anion combination in the first part of the equation, or if the name is given, name the element combination with the anion first, then the cation. The anion is capitalized, while the cation remains lowercase.
B. If the compound involves a metal that forms more than one cation, the number of cations it forms goes in between parentheses in roman numeral form. If the compound does not involve a metal, than list cation charges as normal.

C. For each hydrate at the end of the compound, the prefix at the beginning indicates the number of hydrates. Name the appropriate prefix or number needed for each formula or compound name. Make sure the hydrate is lowercase.

Solutions

(1) Iron(III) phosphate tetrahydrate

A. Iron phosphate as an element is written as **$FePO_4$**, since both parts have a charge of 3, they are not needed in the equation.
B. The prefix tetra stands for "4", so therefore tetrahydrate as an element is written as **$4H_2O$**
C. The two parts combined make the final formula $FePO_4 \cdot 4H_2O$

(2) $CsBr \cdot 4H_2O$

A. Cs and Br with 1 charge stands for **Cesium Bromide**
B. 4 as a prefix is Tetra, so the end of the formula is **tetrahydrate**
C. The final name of the compound is **Cesium bromide tetrahydrate**

(3) Cobalt(II) chloride hexahydrate

A. Cobalt Chloride has a compound formula of $CoCl$, and with Cobalt's charge of 2 involved, a subscript of 2 is added to the end to make $CoCl_2$
B. The "hexa" prefix stands for 6, meaning there are 6 hydrates or **$6H_2O$**
C. The two parts put together makes the final compound, $CoCl_2 \cdot 6H_2O$

(4) $LiI \cdot 3H_2O$

A. Li and I both with single charges make up the compound **Lithium iodide**
B. 3 as a prefix i labled as "tri", so with 3 hydrates, the compound is named as **trihydrate**
C. The final compound put together makes **Lithium iodide trihydrate**

Help with naming compounds: http://chemwiki.ucdavis.edu/Wikitext...onic_Compounds

Q5.65

Write the name for each molecular compound.

a. CO
b. H_2S
c. SF_6
d. N_2O_2

Strategy

First, you have to know how to name a molecular compound. To name a compound, its the prefix, name of first element, prefix and name of the second element with the suffix -ide. The prefixes are as follows:

- mono=1
- di=2
- tri=3
- tetra=4
- penta=5
- hexa=6
- hepta=7
- octa=8
- nona=9
- deca=10

After finding the elements and suffixes, just put it together and you have a molecular compound.

Solution

a. CO: Since we only have one of each, the carbon does not have a suffix, so we would just use Carbon and there is only one oxygen so the suffix is mono. So the answer is Carbon Monoxide
b. H_2S: The hydrogen has a coefficient of 2, so the suffix would be di and the sulfur has no coefficients. Therefore, the answer would be dihydrogen monosulfide.
c. SF_6: The sulfur has no coefficients, so it does not change. The fluoride has a coefficient of 6. Therefore, the answer is Sulfur Hexafluoride
d. N_2O_2: The nitrogen has a coefficient of 2, while the oxygen has a coefficient of 3. The compounds name is Dinitrogen trioxide.

Q5.75a

Find the amount of moles contained in each specimen.

a. 7.87 kg H_2O_2
b. 2.34 kg NaCl
c. 12.5 g C_2H_6O
d. 85.72 g NH_3

Strategy

A. First, determine the units that are given for each sample. In order to convert these measurements to moles, each sample should first be written in grams. Two of the given samples are measured in kilograms. Use the following conversion factor to convert kilograms to grams, with x representing the given mass:

$$x\,kg \cdot \frac{1000\,g}{1\,kg} \tag{2.14.3.1}$$

B. Next, find the atomic masses of all the atoms in each compound by using the Periodic Table. Then, add these atomic masses together for each compound. The resulting value will be the number of grams of each sample that make up one mole.

C. Convert the mass in grams of each sample to moles by multiplying it by the following conversion factor, called the molar mass, with *x* representing the mass of the given specimen and y representing the calculated atomic mass found in step B:

$$x\,g \cdot \frac{1\,mole}{y\,g} \tag{2.14.3.2}$$

A Numbers 1 and 2 first need to be converted into grams.

1. $$7.87\,kg\,H_2O_2 \cdot \frac{1000\,g}{1\,kg} = 7,870\,g\,H_2O_2 \tag{2.14.3.3}$$
2. $$2.34\,kg\,NaCl \cdot \frac{1000\,g}{1\,kg} = 2,340\,g\,NaCl \tag{2.14.3.4}$$

B The following are the calculated molar masses of each of the given compounds. The atomic mass of each element has been rounded to 4 significant figures in these calculations.

1. H_2O_2:

$$2\,(1.008\,g\,H) + 2\,(16.00\,g\,O) = \frac{34.02\,g}{1\,mole}\,H_2O_2 \tag{2.14.3.5}$$

2. NaCl:

$$(22.99\,g\,Na) + (35.45\,g\,Cl) = \frac{58.44\,g}{1\,mole}\,NaCl \tag{2.14.3.6}$$

3. C_2H_6O:

$$2\,(12.01\,g\,C) + 6\,(1.008\,g\,H) + (16.00\,g\,O) = \frac{46.07\,g}{1\,mole}\,C_2H_6O \tag{2.14.3.7}$$

4. NH_3:

$$(14.01\,g\,N) + 3\,(1.008\,g\,H) = \frac{17.03\,g}{1\,mole}\,NH_3 \tag{2.14.3.8}$$

C The number of moles in each specimen can now be calculated by multiplying the mass in grams of each sample by its molarity.

1. $$7,870\,g\,H_2O_2 \cdot \frac{1\,mole\,H_2O_2}{34.02\,g\,H_2O_2} = 231\,moles\,H_2O_2 \tag{2.14.3.9}$$
2. $$2,340\,g\,NaCl \cdot \frac{1\,mole\,NaCl}{58.44\,g\,NaCl} = 40.0\,moles\,NaCl \tag{2.14.3.10}$$
3. $$12.5\,g\,C_2H_6O \cdot \frac{1\,mole\,C_2H_6O}{46.07\,g\,C_2H_6O} = 0.271\,moles\,C_2H_6O \tag{2.14.3.11}$$
4. $$85.72\,g\,NH_3 \cdot \frac{1\,mole\,NH_3}{17.03\,g\,NH_3} = 5.033\,moles\,NH_3 \tag{2.14.3.12}$$

Q5.75b

Calculate the number of moles in each of the following examples.

a. 402.5 mg of NO_2
b. 2.7 kg of H_2O
c. 323 g of CBr_4
d. 2.9 kg of CaO

Solution

a)

First, calculate the molar mass of NO_2.

14.007g/mol N + 2 (15.999 g/mol O) = ***46.0055 g/mol***

Now, convert the sample from mg to g.

402.5 ~~mg~~ NO_2	1 g	**0.4025 g NO_2**
	1000 ~~mg~~	

Finally, use the molar mass to determine how many moles your sample size contains.

0.4025 g NO_2	1 mol	**0.0087 moles of NO_2**
	46.0055 g	

This gives us the final answer, **0.0087 moles of NO_2**.

From this, we can deduce that:

$$\text{Moles of substance} = \text{Mass of substance (g)}\,Molar mass of substance \tag{2.14.3.13}$$

b) 150 moles of H_2O

c) 0.974 moles of CBr_4

d) 52 moles of CaO

Q5.80

For each substance, convert the given molecules to mass in grams:

a. 3.2 x 10^{24} Cl_2 molecules
b. 8.25 x 10^{18} CH_2O molecules
c. 1 carbon dioxide molecule

Strategy:

Step 1: Convert the given molecules to moles by dividing by Avogadro's number.

Step 2: Multiply the number of moles by the Molar Mass of the substance to determine the grams.

Step 3: Through correct Dimensional Analysis, the molecules and moles will cancel to leave grams.

$$\frac{molecules}{Avogadro's} \rightarrow moles \rightarrow moles \cdot MM \rightarrow grams \tag{2.14.3.14}$$

Helpful Hints:

- Avogadro's number: 6.022 x 10^{23} mol of a substance
- Molar mass (MM): the weight of a substance in the units g/mol, found by adding the atomic masses of each element.
    - Example: MM of H_2O --> H(1.008 x 2) + O(15.999)= 18 g/mol

Solution

a) 3.8 x 10^{51} grams

b) 4.11 x 10^{45} grams

c) 7.31 x 10^{25} grams

Q5.105

Write the empirical formula and molecular formula for the compound with the given percent composition and molecular weight.

- 40.0% Carbon
- 6.70% Hydrogen
- 53.3% Oxygen

Molecular Weight ≈ 240 g/mol

Solution

Strategy:

A. Calculate amount of grams of each element we have.
B. Convert amount of grams into number of moles of each element.
C. Obtain molecular formula.
D. Find the common factor among the number of atoms each elements has, then obtain the empirical formula.

Solution:

1. Calculate amount of grams in each element:

$$Carbon = .40 \cdot 240g = 96g$$

$$Hydrogen = .067 \cdot 240g = 16.1g$$

$$Oxygen = .533 \cdot 240g = 127.9g$$

2. Convert amount of grams into number of moles of each elements:

Carbon-

$$96g \cdot \frac{1mol}{12.001g} = 8.00mols$$

Hydrogen-

$$16.1g \cdot \frac{1mol}{1.008g} = 16.0mols$$

Oxygen-

$$127.9g \cdot \frac{1mol}{16.0g} = 8mols$$

3. Obtain molecular formula:

$$C_8H_{16}O_8 \tag{2.14.3.15}$$

4. Find common factor and obtain empirical formula:

Common factor is 8

Empirical formula is

$$CH_2O \tag{2.14.3.16}$$

Q5.106

For several compounds both the molar mass and and empirical formulas are listed. What is the molecular formula for each compound.

a. C_2OH_4, 88 g/mol
b. C_2H_8N, 46 g/mol
c. NH_2, 32 g/mol

Strategy

1. A. Find the empirical formula mass of the compound
2. B. Determine the ratio between the empirical formula mass and the molar mass of the compound
3. C.Multiply the empirical formula by the factor found in part B.

Solution

C_2OH_4, 88 g/mol

A.

Carbon: 2(12.0g/mol)

Oxygen: 1(16.0g/mol)

Hydrogen: 4(1.0g/mol)

=44.0 g/mol

B.

$$\frac{88.0g/mol}{44.0g/mol} = 2 \tag{2.14.3.17}$$

C.

2(C_2OH_4)= $C_4O_2H_8$

C_2H_8N, 46 g/mol

A.

Carbon: 2(12.0g/mol)

Hydrogen: 8(1.0g/mol)

Nitrogen: 1(14.0g/mol)

=46.0 g/mol

B.

$$\frac{46.0g/mol}{46.0g/mol} = 1.0 \tag{2.14.3.18}$$

C.

1(C_2H_8N)=C_2H_8N

NH_2

A.

Nitrogen: 1(14.0g/mol)

Hydrogen: 2(1.0g/mol)

=16.0g/mol

B.

$$\frac{32.0g/mol}{16.0g/mol} = 2 \quad (2.14.3.19)$$

C.

2(NH_2)=N_2H_4

Q5.107

After the combustion of a hydrocarbon, we find 66.02g of CO_2 and 27.02g of H_2O. What is the empirical formula of this hydrocarbon?

Solutions

66.02 x 1 mol/ 44.0 g= 1.500 mol CO2= 1.5 mol C

27.02 x 1 mol/ 18g = 1.501 mol H20= 3 mol H

C1.5 H3= C3 H6

You start by taking both, divide by the molar mass, use the mol ratio of the compound and multiple to get rid of decimals.

~~Additional Question~~

~~Additional Questions 1. For the following molecules; write the chemical formula, determine how many atoms are present in one molecule, determine the molecular weight, determine the number of moles in exactly 1 gram, and the number of moles in exactly 10-6 gram. a. carbon dioxide b. iron (II) chloride c. dinitrogen pentoxide d. iron (III) sulfate 2. Name the following compounds, determine the molecular weight, determine the percent composition, and determine how many moles in 8.35 grams of the compound. a. KI b. CaF2 c. Cu2SO4 d. N2O e. LiOH 3. Give the chemical formula (or atomic symbol), molecular (or atomic) weight, and charge for the following ions: a. sulfate b. sulfite c. nitrate d. chloride e. nitride f. acetate g. carbonate 4. Calculate the number of moles in: a. 20.1797 g of Ne b. 45.3594 g of Ne c. 0.198669 g of Ne 5. Calculate the mass of: a. 2.00 mole of Fe b. 4.362 x 10-5 mol of Fe c. 4.362 x 10-5 mol of Li~~

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CHAPTER OVERVIEW

3: Measurement and Calculations

3.1: SI Base Units
3.2: Metric Prefixes
3.3: Scientific Notation in Chemistry
3.4: SI Length and Volume Units
3.5: SI Mass and Weight Units
3.6: SI Kinetic Energy Units
3.7: Scientific Dimensional Analysis
3.8: Metric Unit Conversions
3.9: Derived Units
3.10: Dosage Calculations
3.11: Density
3.12: Chapter 3 Excercises
3.12.E: Chemistry, Matter, and Measurement (Exercises)

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3.1: SI Base Units

How long is a yard?

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

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units, listed in Table **3.1.1**.

Table **3.1.1**: SI Base Units of Measurement

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

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



Figure **3.1.1**: Meter Standard. (Public Domain; US Department of Commerce; National Institute of Standards and Technology via Wikipedia; http://commons.wikimedia.org/wiki/File:Standard_kilogram,_2.jpg)



Figure **3.1.2**: Kilogram standard (Public Domain; National Institute of Standards and Technology via Wikipedia)



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



Figure **3.1.3**: Areas of the world that use the metric system (green). (Public Domain; User:Canuckguy/Wikipedia and User:AzaToth/Wikimedia Commons via Wikipedia)

Summary

- The SI system is based on multiples of ten.
- There are seven basic units in the SI system.
- Five of these seven SI units are commonly used in chemistry.

Review

1. What does SI stand for?
2. When was this system adopted by the international community?
3. Which of the units are commonly used in chemistry?

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3.2: Metric Prefixes



Figure **3.2.1** (Public Domain; Joseph Wright of Derby via Wikipedia)

What do Latin and Greek have to do with modern science?

Isn't it hard enough to learn English terms? For hundreds of years, the languages of the educated class were Latin and Greek. In part, because the literature of philosophy was Latin and Greek. Even the medieval Bibles were written in those two languages – the first English translation was in the late 1380s. Using Latin and Greek allowed scholars from different countries to communicate more easily with one another. Today we still see many Latin phrases in legal communications ("pro bono" meaning to do something "for the good" and not charge legal fees), scientific naming of biological species, and Latin is used for the annual student speech at Harvard University graduations. Not bad for a "dead" language.

Metric Prefixes

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great".

Table **3.2.1** lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table **3.2.1**: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	**G**	1,000,000,000	1 gigameter $(\mathbf{Gm}) = 10^9\ \mathbf{m}$
mega	**M**	1,000,000	1 megameter $(\mathbf{Mm}) = 10^6\ \mathbf{m}$
kilo	**k**	1,000	1 kilometer $(\mathbf{km}) = 1,000\ \mathbf{m}$
hecto	**h**	100	1 hectometer $(\mathbf{hm}) = 100\ \mathbf{m}$
deka	**da**	10	1 dekameter $(\mathbf{dam}) = 10\ \mathbf{m}$
		1	**1 meter** $(\mathbf{m})$
deci	**d**	1/10	1 decimeter $(\mathbf{dm}) = 0.1\ \mathbf{m}$
centi	**c**	1/100	1 centimeter $(\mathbf{cm}) = 0.01\ \mathbf{m}$
milli	**m**	1/1,000	1 millimeter $(\mathbf{mm}) = 0.001\ \mathbf{m}$
micro	μ	1/1,000,000	1 micrometer $(\mu\mathbf{m}) = 10^{-6}\ \mathbf{m}$
nano	**n**	1/1,000,000,000	1 nanometer $(\mathbf{nm}) = 10^{-9}\ \mathbf{m}$
pico	**p**	1/1,000,000,000,000	1 picometer $(\mathbf{pm}) = 10^{-12}\ \mathbf{m}$

There are more prefixes - some of them rarely used. Have you ever heard of a zeptometer?

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

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



Summary

- Metric prefixes derive from Latin or Greek terms.
- Metric prefixes are used to make the units manageable.

Review

1. What is the prefix for “thousand”?
2. What is the prefix for 0.01?
3. How would you write 500 milliliters?
4. How many decimeters in one meter?
5. You have a mass that weighs 1.2 hectograms. How many grams does it weigh?

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3.3: Scientific Notation in Chemistry

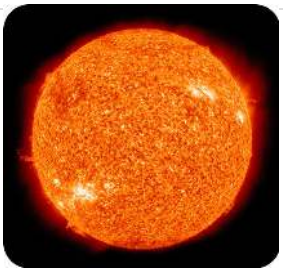


Figure 3.3.1 (Public Domain; NASA via Wikipedia)

How far is the Sun from Earth?

Astronomers are accustomed to very large numbers. While the moon is only $406,697\ \text{km}$ from earth at its maximum distance, the sun is much further away ($150\ \text{million km}$). Proxima Centauri, the star nearest the earth, is $39,900,000,000,000\ \text{km}$ away—and this is just the start of long distances. On the other end of the scale, some biologists deal with very small numbers: a typical fungus could be as small as 30 micrometers ($0.000030\ \text{meters}$) in length, and a virus might only be 0.03 micrometers ($0.00000003\ \text{meters}$) long.

Scientific Notation

Scientific notation is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power. It is a very useful tool for working with numbers that are either very large or very small. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters—a very large distance indeed. In scientific notation, the distance is written as $1.5 \times 10^{11}\ \text{m}$. The coefficient is the 1.5 and must be a number greater than or equal to 1, and less than 10. The power of 10, or exponent, is 11 because you would have to multiply 1.5 by 10^{11} to get the correct number. Scientific notation is sometimes referred as exponential notation. A summary of SI units is given in the table below.

Table 3.3.1: SI Prefixes

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
giga	G	10^9	1,000,000,000	1 gigameter $(\text{Gm}) = 10^9\ \text{m}$
mega	M	10^6	1,000,000	1 megameter $(\text{Mm}) = 10^6\ \text{m}$
kilo	k	10^3	1,000	1 kilometer $(\text{km}) = 1,000\ \text{m}$
hecto	h	10^2	100	1 hectometer $(\text{hm}) = 100\ \text{m}$
deka	da	10^1	10	1 dekameter $(\text{dam}) = 10\ \text{m}$
		10^0	**1**	**1 meter** (m)
deci	d	10^{-1}	1/10	1 decimeter $(\text{dm}) = 0.1\ \text{m}$
centi	c	10^{-2}	1/100	1 centimeter $(\text{cm}) = 0.01\ \text{m}$
milli	m	10^{-3}	1/1,000	1 millimeter $(\text{mm}) = 0.001\ \text{m}$
micro	μ	10^{-6}	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6}\ \text{m}$
nano	n	10^{-9}	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9}\ \text{m}$
pico	p	10^{-12}	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12}\ \text{m}$

When working with small numbers, we use a negative exponent. So 0.1 meters is 1×10^{-1} meters, 0.01 is 1×10^{-2} and so forth. The table above gives examples of smaller units. Note the use of the **leading zero** (the zero to the left of the decimal point). That digit is there to help you see the decimal point more clearly. The figure 0.01 is less likely to be misunderstood than .01, where you may not see the decimal.



Summary

- Scientific notation allows us to express very large or very small numbers in a convenient way.
- Scientific notation uses a coefficient (a number between 1 and 10) and a power of ten sufficient for the actual number.

Review

1. What is scientific notation?
2. What do we use scientific notation for?
3. What is a leading zero?
4. Express 150,000,000 in scientific notation.
5. Express 0.000043 in scientific notation.

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3.4: SI Length and Volume Units



Figure 3.4.1 (Public Domain; User:Żeglarz/Wikimedia Commons via Wikipedia)

How were sailors able to measure the depths of seas?

Back in the days before all the electronic gadgets for measuring depth and locating undersea objects existed, the "fathom" was the unit of measurement for depth. A rope was knotted every six feet and the end was dropped over the side of the ship. You could tell how deep the water was by how many knots went under the water before the rope hit bottom. Today we just turn on an instrument and read the depth to a high level of accuracy.

Length and Volume

Length is the measurement of the extent of something along its greatest dimension. The SI basic unit of length, or linear measure, is the **meter** (m). All measurements of length may be made in meters, though the prefixes listed in various tables will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance, such as the distance between New York City and Chicago, is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 20 micrometers ($\mu\mathrm{m}$).

Volume is the amount of space occupied by a sample of matter. The volume of a regular object can be calculated by multiplying its length by its width and height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter ($\mathrm{m^3}$), which is the volume occupied by a cube that measures $1\ \mathrm{m}$ on each side. This very large volume is not convenient for typical use in a chemistry laboratory. A liter (L) is the volume of a cube that measures $10\ \mathrm{cm}$ ($1\ \mathrm{dm}$) on each side. A liter is thus equal to both $1000\ \mathrm{cm^3}$ ($10\ \mathrm{cm} \times 10\ \mathrm{cm} \times 10\ \mathrm{cm}$), and to $1\ \mathrm{dm^3}$. A smaller unit of volume that is commonly used is the milliliter (mL—note the capital L, which is a standard practice). A milliliter is the volume of a cube that measures $1\ \mathrm{cm}$ on each side. Therefore, a milliliter is equal to a cubic centimeter ($\mathrm{cm^3}$). There are $1000\ \mathrm{mL}$ in $1\ \mathrm{L}$, which is the same as saying that there are $1000\ \mathrm{cm^3}$ in $1\ \mathrm{dm^3}$.



Figure 3.4.2: A typical water bottle is 1 liter in volume. (Public Domain; User:Kenyon/Wikimedia Commons via Wikipedia)

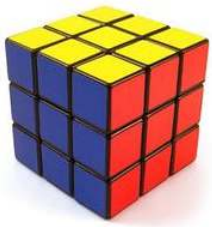


Figure 3.4.3: This Rubik's cube is 5.7 cm on each side and has a volume of 185.2 cm^3 or 185.2 mL. (CC by 2.0; Bram Van Damme (Flickr: Bramus!) via Flickr)



Figure 3.4.4: A graduated cylinder is often used to measure volume in the laboratory, and comes in a variety of sizes. (Public Domain; User:Darrien/Wikipedia via Wikipedia)

Summary

- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of an item.

Review

1. Define length.
2. Define volume.
3. An object measures 6.2 cm×13.7 cm×26.9 cm6.2 cm×13.7 cm×26.9 cm. Which value is the length of the object?
4. How big is a mL?

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3.5: SI Mass and Weight Units



Figure 3.5.1 (Public Domain; NASA via Wikipedia)

How is he floating?

One of the many interesting things about travel in outer space is the idea of weightlessness. If something is not fastened down, it will float in mid-air. Early astronauts learned that weightlessness had negative effects on bone structure. If there was no pressure on the legs, those bones would begin to lose mass. Weight provided by gravity is needed to maintain healthy bones. Specially designed equipment is now a part of every space mission, so the astronauts can maintain good body fitness.

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The base SI unit of mass is the kilogram or kg, which was originally defined as the mass of **1 L** of liquid water at **4°C** (the volume of a liquid changes slightly with temperature). The kilogram is the only SI base unit that contains a prefix. This is because when the first physical standard for mass was created in 1799, the methods to measure the mass of something as small as a gram did not exist. In the laboratory, mass is measured with a balance (figure below), which must be calibrated with a standard mass so that its measurements are accurate.



Figure 3.5.2: An analytical balance makes very sensitive mass measurements in a laboratory, usually in grams. (Public Domain; US Drug Enforcement Administration via Wikipedia)

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are **1000 g** in **1 kg**. A milligram is 1/1000th of a gram, so there are **1000 mg** in **1 g**.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to **gravity** is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale by reading the distance that a certain object pulls down and stretches a spring.



Summary

- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.

Review

1. Define mass.
2. Define weight.
3. If I weigh 180 pounds on Earth, what will I weigh on the moon?

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3.6: SI Kinetic Energy Units



Figure 3.6.1 (Public Domain; User:Dlls/Wikimedia Commons via Wikipedia)

Have you ever watched a cat in action?

When a cat is chasing something, its moves very fast. A person may comment, "that cat has a lot of energy"—which is more correct than they might realize! One form of energy is seen when an object is moving, and this type of energy is the basis for many chemical processes.

SI Kinetic Energy Units

An object's **kinetic energy** is the energy due to motion. Kinetic energy can be defined mathematically as:

$$KE = \frac{1}{2}mv^2$$

where KE is kinetic energy, m is mass, and v is velocity.

Energy is defined as the capacity to do work or to produce heat. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered energy is **potential energy**, a type of energy that is stored in matter and released during a chemical reaction. The **joule** (J) is the SI unit of energy and is named after English physicist James Prescott Joule (1818-1889). If we go back to the equation for kinetic energy written above, we can put units in: kg for mass and $\mathrm{m^2/s^2}$ for velocity squared. Then, in terms of SI base units, a joule is equal to a kilogram times meter squared, divided by a second squared ($\mathrm{kg \cdot m^2/s^2}$). Another common unit of energy that is often used is the **calorie** (cal), which is equivalent to $4.184\ \mathrm{J}$.



Figure 3.6.2: James Prescott Joule. (Public Domain; John Collier via Wikipedia)



Summary

- Energy is the capacity to do work or to produce heat.
- Kinetic energy is energy due to motion.
- Potential energy is energy stored in matter.
- The joule (**J**) is the SI unit of energy and equals $\mathbf{kg \cdot m^2/s^2}$.

Review

1. What is kinetic energy?
2. What is the mathematical equation for kinetic energy?
3. What is potential energy?
4. What is the SI unit for energy?

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3.7: Scientific Dimensional Analysis

Conversion Factors

Many quantities can be expressed in several different ways. The English system measurement of 4 cups is also equal to 2 pints, 1 quart, and $\frac{1}{4}$ of a gallon.

$$4 \text{ cups} = 2 \text{ pints} = 1 \text{ quart} = 0.25 \text{ gallon}$$

Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. This is because the units are different. We can establish the same set of equalities for the metric system:

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 1000 \text{ millimeters}$$

The metric system's use of powers of 10 for all conversions makes this quite simple.

Whenever two quantities are equal, a ratio can be written that is numerically equal to 1. Using the metric examples above:

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

The $\frac{1 \text{ m}}{100 \text{ cm}}$ is called a **conversion factor**. A conversion factor is a ratio of equivalent measurements. Because both 1 m and 100 cm represent the exact same length, the value of the conversion factor is 1. The conversion factor is read as "1 meter per 100 centimeters". Other conversion factors from the cup measurement example can be:

$$\frac{4 \text{ cups}}{2 \text{ pints}} = \frac{2 \text{ pints}}{1 \text{ quart}} = \frac{1 \text{ quart}}{0.25 \text{ gallon}} = 1$$

Since the numerator and denominator represent equal quantities in each case, all are valid conversion factors.

Scientific Dimensional Analysis

Conversion factors are used in solving problems in which a certain measurement must be expressed with different units. When a given measurement is multiplied by an appropriate conversion factor, the numerical value changes, but the actual size of the quantity measured remains the same. **Dimensional analysis** is a technique that uses the units (dimensions) of the measurement in order to correctly solve problems. Dimensional analysis is best illustrated with an example.

Example **3.7.1**

How many seconds are in a day?

Solution

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

Known

- 1 day = 24 hours
- 1 hour = 60 minutes
- 1 minute = 60 seconds

Unknown

The known quantities above represent the conversion factors that we will use. The first conversion factor will have day in the denominator so that the "day" unit will cancel. The second conversion factor will then have hours in the denominator, while the third conversion factor will have minutes in the denominator. As a result, the unit of the last numerator will be seconds and that will be the units for the answer.

Step 2: Calculate.

$$1 \text{ d} \times \frac{24 \text{ hr}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} = 86,400 \text{ s}$$

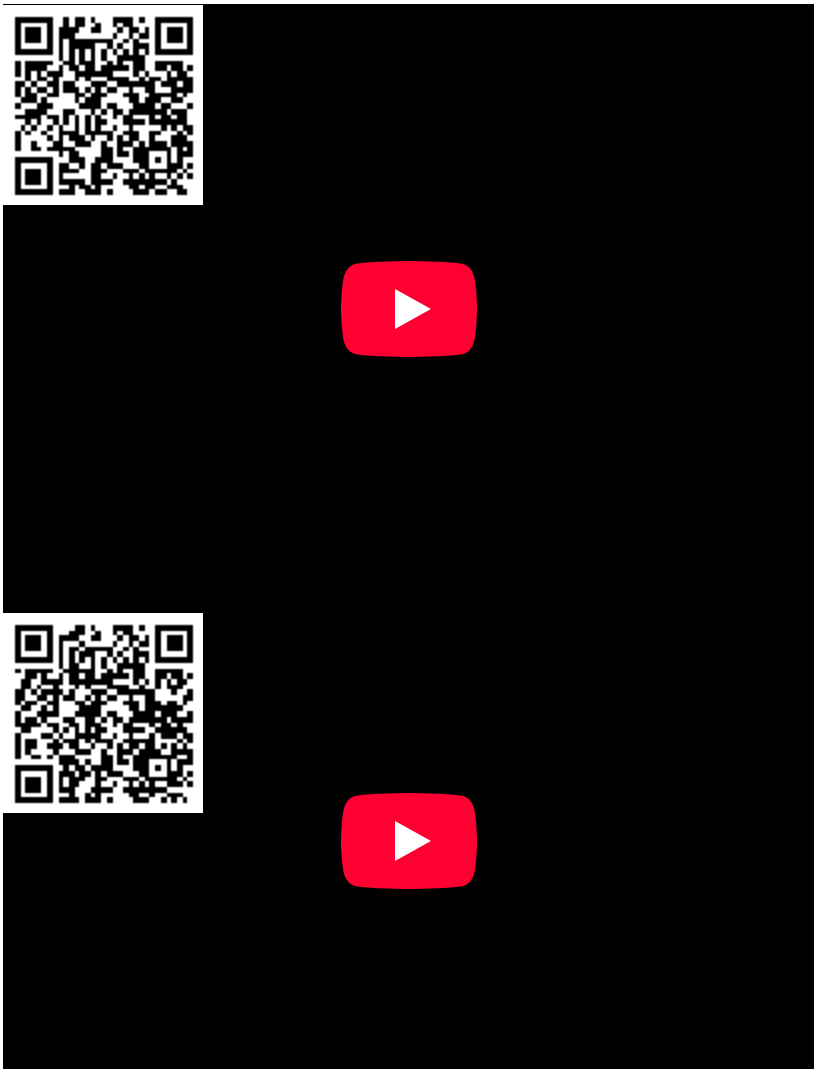
Applying the first conversion factor, the "$\mathbf{d}$" unit cancels and $\mathbf{1 \times 24 = 24}$. Applying the second conversion factor, the "$\mathbf{hr}$" unit cancels and $\mathbf{24 \times 60 = 1440}$. Applying the third conversion factor, the "$\mathbf{min}$" unit cancels and $\mathbf{1440 \times 60 = 86,400}$. The unit that remains is "$\mathbf{s}$" for seconds.

Step 3: Think about your result.

Seconds is a much smaller unit of time than days, so it makes sense that there are a very large number of seconds in one day.

Summary

- A conversion factor is a ratio of equivalent measurements.
- Dimensional analysis is a technique that uses the units (dimensions) of the measurement to solve problems.



Review

1. What is a conversion factor?
2. What is dimensional analysis?
3. How many meters are in 3.7 km?
4. How many kg in 12980 g?

Explore More

Use the link below to answer the following questions**:** http://www.felderbooks.com/papers/units.html

1. What do we always need to express measurements correctly?

2. What does dimensional analysis tell you?
3. How do you know that you have set the problem up incorrectly?
4. How do you know that you have set the problem up correctly?

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3.8: Metric Unit Conversions



Figure 3.8.1 (Public Domain; the US Army via Wikipedia)

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

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

Metric Unit Conversions

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

Example 3.8.1

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

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

Step 3:Think about your result.

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

Two-Step Metric Unit Conversions

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

Example 3.8.2

Convert $4.3\ \text{cm}$ to μm.

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

Step 3: Think about your result.

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



Summary

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

Review

1. Perform the following conversions.
    a. 0.074 km to m
    b. 24,600 µg to g
    c. 4.9 × 107 µg to kg
    d. 84 dm to mm

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3.9: Derived Units



Figure **3.9.1** (Public Domain; User:Joegrimes/Wikipedia via Commons Wikimedia, Farm landly sander [commons.wikimedia.org])

How has farming evolved?

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

Dimensional Analysis and Derived Units

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

Derived SI Units

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

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

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

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

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

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

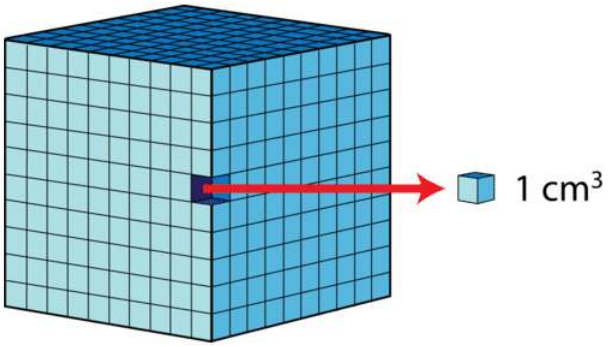


Figure 3.9.2: There are $1000\ \text{cm}^3$ in $1\ \text{dm}^3$. Since a cm^3 is equal to a mL and a dm^3 is equal to a L, we can say that there are $1000\ \text{mL}$ in $1\ \text{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 3.9.1

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

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

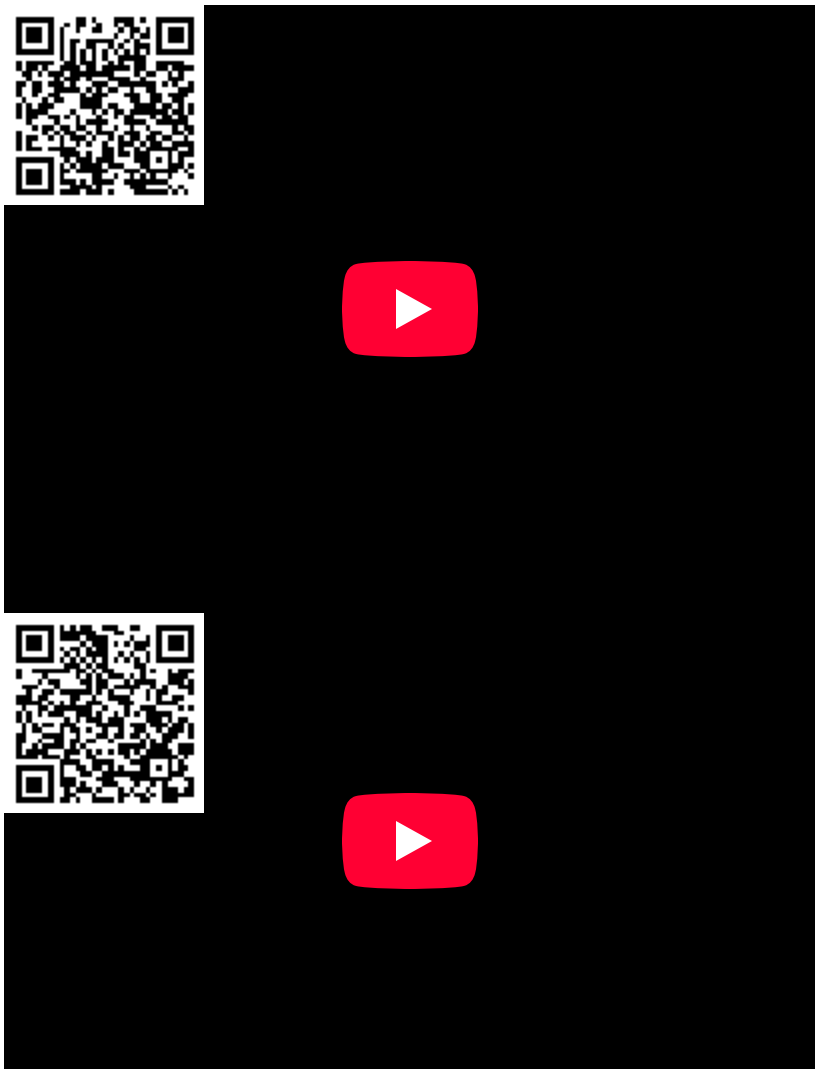
Numerically, the steps are to divide 3.6 by 10^9, followed by multiplying by 10^6. You may find that you can shorten the problem by a step by first determining the conversion factor from mm to cm, and using that instead of first converting to m. There are 10 mm in 1 cm.

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

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

Step 3: Think about your result.

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



Summary

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

Review

1. What is a derived unit?
2. Convert 0.00722 km^3 to m^3
3. Convert 129 cm^3 to L
4. Convert 4.9 × 105 μm^3 to mm^3

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3.10: Dosage Calculations

Learning Objectives

- Calculate drug dosages using conversion factors.

To Your Health: Dosages

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-dependent

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) based on weight. Drug dosage calculation is required if the physician's order is different from what is available.

Example 3.10.1

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.

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

Hence, 1.5 tablets should be taken.

? Exercise 3.10.1

Calculate each of the following.

a. The physician ordered 20 mg of Valium. Valium is available as 10 mg per tablet. How many tablets should the nurse administer?
b. The doctor's order is 1 g of Calcium. What is on hand is Calcium as 500 mg per tablet. How many tablets should be taken?

Answer a

Start with 20 mg of Valium. For the conversion factor, we know that 10 mg Valium = 1 tablet

$$20 \cancel{\text{mg}} \times \frac{1 \text{ tablet}}{10 \cancel{\text{mg}}} = 2 \text{ tablets}$$

Answer b

Start with 1 g (1000 mg) of Calcium. For the conversion factor, we know that 500 mg Calcium = 1 tablet

$$1000 \cancel{\text{mg}} \times \frac{1 \text{ tablet}}{500 \cancel{\text{mg}}} = 2 \text{ tablets}$$

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*. (as in Example 1.8.1).

✓ Example 3.10.2

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.

Exercise 3.10.2

Calculate each of the following.

a. Vancomycin is ordered 15mg/kg for a patient that is 110 lbs. Vancomycin is available as 250 mg per capsule. How many capsules should the nurse administer? (ans. 3 capsules)
b. A doctor prescribes ampicillin 40mg/kg to a patient weighing 55 lbs. Ampicillin is available as 500 mg tablets. How many tablets should the nurse administer? (ans. 2 tablets)

Answer a

$$110 \cancel{\text{ lbs}} \times \frac{1 \cancel{\text{ kg}}}{2.2 \cancel{\text{ lbs}}} \times \frac{15 \cancel{\text{ mg}}}{1 \cancel{\text{ kg}}} \times \frac{1 \text{ tablet}}{250 \cancel{\text{ mg}}} = 3 \text{ tablets}$$

Answer b

$$55 \cancel{\text{ lbs}} \times \frac{1 \cancel{\text{ kg}}}{2.2 \cancel{\text{ lbs}}} \times \frac{40 \cancel{\text{ mg}}}{1 \cancel{\text{ kg}}} \times \frac{1 \text{ tablet}}{500 \cancel{\text{ mg}}} = 2 \text{ tablets}$$

References

1. The Use of Body Surface Area as a Criterion of Drug Dosage in Cancer Chemotherapy(opens in new window) [cancerres.aacrjournals.org] "The Use of Body Surface Area as a Criterion of Drug Dosage in Cancer Chemotherapy" D Pinkel. Cancer Research 1958

Contributors and Attributions

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3.11: Density



Figure 3.11.1 (CC BY 2.0; Tony Hisgett (Flickr:ahisgett) via Flickr)

How do logs stay afloat in water?

After trees are cut, logging companies often move the raw material down a river to a sawmill where it can be shaped into building materials or other products. The logs float on the water because they are less dense than the water they are in. Knowledge of density is important in the characterization and separation of materials. Information about density allows us to make predictions about the behavior of matter.

Density

A golf ball and a table tennis ball are about the same size. However, the golf ball is much heavier than the table tennis ball. Now imagine a similar size ball made out of lead. That would be very heavy indeed! What are we comparing? By comparing the mass of an object relative to its size, we are studying a property called **density**. Density is the ratio of the mass of an object to its volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Density is an intensive property, meaning that it does not depend on the amount of material present in the sample. Water has a density of $1.0\ \text{g/mL}$. That density is the same whether you have a small glass of water or a swimming pool full of water. Density is a property that is constant for the particular identity of the matter being studied.

The SI unit of density is kilograms per cubic meter $\left(\text{kg/m}^3\right)$, since the kg and the m are the SI units for mass and length respectively. In everyday usage in a laboratory, this unit is awkwardly large. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter $\left(\text{g/cm}^3\right)$. Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at $20°\text{C}$ are listed in the table below.

Table 3.11.1: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084
Ice (0°C)	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water (4°C)	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

Since most materials expand as temperature increases, the density of a substance is temperature dependent, and usually decreases as temperature increases.

Ice floats in water, and it can be seen from the table that ice is less dense. Corn syrup, being more dense, would sink if placed in water.

Example 3.11.1

An $18.2\ \text{g}$ sample of zinc metal has a volume of $2.55\ \text{cm}^3$. Calculate the density of zinc.

Solution

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

Known

- Mass $= 18.2\ \text{g}$
- Volume $= 2.55\ \text{cm}^3$

Unknown

Use the equation for density, $D = \frac{m}{V}$, to solve the problem.

Step 2: Calculate.

$$D = \frac{m}{V} = \frac{18.2\ \text{g}}{2.55\ \text{cm}^3} = 7.14\ \text{g/cm}^3$$

Step 3: Think about your result.

If $1\ \text{cm}^3$ of zinc has a mass of about 7 grams, then 2 and a half cm^3 will have a mass about 2 and a half times as great. Metals are expected to have a density greater than that of water, and zinc's density falls within the range of the other metals listed above.

Since density values are known for many substances, density can be used to determine an unknown mass or an unknown volume. Dimensional analysis will be used to ensure that units cancel appropriately.

Example 3.11.2

1. What is the mass of $2.49\ \text{cm}^3$ of aluminum?
2. What is the volume of $50.0\ \text{g}$ of aluminum?

Solution

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

Known

- Density $= 2.70\ \text{g/cm}^3$
- 1. Volume $= 2.49\ \text{cm}^3$
- 2. Mass $= 50.0\ \text{g}$

Unknown

- 1. Mass $=?\ \text{g}$
- 2. Volume $=?\ \text{cm}^3$

Use the equation for density, $D = \frac{m}{V}$, and dimensional analysis to solve each problem.

Step 2: Calculate.

$$1.\ 2.49\ \text{cm}^3 \times \frac{2.70\ \text{g}}{1\ \text{cm}^3} = 6.72\ \text{g}$$

$$2.\ 50.0\ \text{g} \times \frac{1\ \text{cm}^3}{2.70\ \text{g}} = 18.5\ \text{cm}^3$$

In problem 1, the mass is equal to the density multiplied by the volume. In problem 2, the volume is equal to the mass divided by the density.

Step 3: Think about your results.

Because a mass of $1\ \text{cm}^3$ of aluminum is $2.70\ \text{g}$, the mass of about $2.5\ \text{cm}^3$ should be about 2.5 times larger. The $50\ \text{g}$ of aluminum is substantially more than its density, so that amount should occupy a relatively large volume.



Summary

- Density is the ratio of the mass of an object to its volume.
- Gases are less dense than either solids or liquids.
- Both liquid and solid materials can have a variety of densities.
- For liquids and gases, the temperature will affect the density to some extent.

Review

1. Define “density.”
2. Are gases more or less dense that liquids or solids at room temperature?
3. How does temperature affect the density of a material?
4. A certain liquid sample has a volume of 14.7 mL and a mass of 22.8 grams. Calculate the density.
5. A material with a density of 2.7 grams/mL occupies 35.6 mL. How many grams of the material are there?
6. A certain material has a density of 19.3 g/mL. What is the material?

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3.12: Chapter 3 Excercises

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3.12.E: Chemistry, Matter, and Measurement (Exercises)

These are homework exercises to accompany Chapter 1 of the Ball et al. "The Basics of GOB Chemistry" Textmap.

The International System of Units

Exercises

1. List four base units.
2. List four derived units.
3. How many meters are in 1 km? How many centimeters are in 1 m?
4. How many grams are in 1 Mg? How many microliters are in 1 L?
5. Complete the following table:

Unit	Abbreviation
centiliter	
	ms
	cm
	kL
micrometer	

6. Complete the following table:

Unit	Abbreviation
microliter	
kilosecond	
	dL
	ns
millimeter	

7. What are some appropriate units for density?
8. A derived unit for velocity, which is the change of position with respect to time, is meters per second (m/s). Give three other derived units for velocity.

Answers

1. second, meter, kilogram, and kelvin (answers will vary)
2. square meter (m^2), cubic meter (m^3), grams per milliliter (g/mL), milliliters per second (mL/s) answers will vary
3. 1,000; 100
4. 1,000,000; 1,000,000

5.

Unit	Abbreviation
centiliter	cL
millisecond	ms
centimeter	cm
kiloliter	kL
micrometer	µm

6.

Unit	Abbreviation
microliter	µL
kilosecond	ks
deciliter	dL
nanosecond	ns
millimeter	mm

7. grams per liter, grams per milliliter, and kilograms per liter (answers will vary)

8. kilometers per hour; meters per minute; millimeters per second (answers will vary)

Expressing Numbers: Scientific Notation

Exercises

1. Why is scientific notation useful in expressing numbers?
2. What is the relationship between the power and the number of places a decimal point is moved when going from standard to scientific notation?
3. Express each number in scientific notation.
   a. 0.00064
   b. 5,230,000
   c. −56,200
   d. 0.000000000220
   e. 1.0
4. Express each number in scientific notation.
   a. 678
   b. −1,061
   c. 0.000560
   d. 0.0000003003
   e. 100,000,000
5. Express each number in standard form.
   a. 6.72×10^{4}
   b. 2.088×10^{-4}
   c. -3×10^{6}

d. 9.98×10^{-7}

6. Express each number in standard form.

a. 9.05×10^{5}
b. 1.0×10^{-3}
c. 6.022×10^{23}
d. 8.834×10^{-12}

7. Complete the following table:

Incorrect Scientific Notation	Correct Scientific Notation
54.7×10^{4}	
0.0066×10^{3}	
$3{,}078 \times 10^{0}$	

8. Complete the following table:

Incorrect Scientific Notation	Correct Scientific Notation
234.0×10^{1}	
36×10^{-4}	
0.993×10^{5}	

Answers

1. Scientific notation is more convenient than listing a large number of zeros.

2. The power is the number of places a decimal point is moved when going from standard to scientific notation. It is positive if the decimal point is moved to the left; negative if moved to the right.

3.

a. 6.4×10^{-4}
b. 5.23×10^{6}
c. -5.62×10^{4}
d. 2.20×10^{-10}
e. 1.0×10^{0}

4.

a. 6.78×10^{2}
b. -1.061×10^{3}
c. 5.60×10^{-4}
d. 3.003×10^{-7}
e. 1×10^{8}

5.

a. 67,200
b. 0.0002088
c. −3,000,000
d. 0.000000998

6. Express each number in standard form.

a. 905,000
b. 0.0010

c. 602,200,000,000,000,000,000,000
d. 0.000000000008834

7.

Incorrect Scientific Notation	Correct Scientific Notation
54.7×10^4	5.47×10^5
0.0066×10^3	6.6×10^0
$3{,}078 \times 10^0$	3.078×10^3

8.

Incorrect Scientific Notation	Correct Scientific Notation
234.0×10^1	2.340×10^3
36×10^{-4}	3.6×10^{-3}
0.993×10^5	9.93×10^4

Converting Units

Exercises

1. Give the two conversion factors you can construct using each pair of units.
   a. meters and kilometers
   b. liters and microliters
   c. seconds and milliseconds
2. Give the two conversion factors you can construct using each pair of units.
   a. grams and centigrams
   b. millimeters and meters
   c. liters and megaliters
3. How many meters are in 56.2 km?
4. How many seconds are in 209.7 ms?
5. How many microliters are in 44.1 L?
6. How many megagrams are in 90.532 g?
7. Convert 109.6 kg into micrograms. Express your final answer in scientific notation.
8. Convert 3.8×10^5 mm into kilometers. Express your final answer in scientific notation.
9. Convert 3.009×10^{-5} ML into centiliters. Express your final answer in scientific notation.
10. Convert 99.04 dm into micrometers. Express your final answer in scientific notation.
11. The density of ethyl alcohol is 0.79 g/mL. What is the mass of 340 mL of ethyl alcohol in kilograms? Do a 2-step conversion.
12. The density of a certain fraction of crude oil is 1.209 g/mL. What is the mass of 13,500 mL of this fraction?
13. The density of ethyl alcohol is 0.79 g/mL. What is the volume of 340 g of ethyl alcohol in liters? Do a 2-step conversion.
14. The density of a certain component of crude oil is 1.209 g/mL. What is the volume of 13,500 g of this component?

Answers

1.
   a. $\frac{1,000\ \text{m}}{1\ \text{km}}$; $\frac{1\ \text{km}}{1,000\ \text{m}}$
   b. $\frac{1,000,000\ \mu\text{L}}{1\ \text{L}}$; $\frac{1\ \text{L}}{1,000,000\ \mu\text{L}}$
   c. $\frac{1,000\ \text{ms}}{1\ \text{s}}$; $\frac{1\ \text{s}}{1,000\ \text{ms}}$
2.
   a. $\frac{1\ \text{g}}{100\ \text{cg}}$; $\frac{100\ \text{cg}}{1\ \text{g}}$
   b. $\frac{1,000\ \text{mm}}{1\ \text{m}}$; $\frac{1\ \text{m}}{1,000\ \text{mm}}$
   c. $\frac{1,000,000\ \text{L}}{1\ \text{ML}}$; $\frac{1\ \text{ML}}{1,000,000\ \text{L}}$
3. 5.62×10^{4} m
4. 2.097×10^{-1} s
5. 4.41×10^{7} µL
6. 9.0532×10^{-5} g
7. 1.096×10^{11} µg
8. 0.38 km; 3.8×10^{-1} km
9. 3.009×10^{3} cL
10. 9.904×10^{6} µm
11. 0.27 kg
12. 16,300 mL
13. 0.43 L
14. 11,200 mL

Other Units - Density

Exercises

1. Calculate the density of a mineral if the mass was found to be 760.15 grams and the volume of 0.25 Liters.
2. A sample of iron has a volume of 48.2 cm^3. Given the density of iron is 7.874 g/cm^3, what is its mass?
3. The volume of hydrogen used by the *Hindenburg*, the German airship that exploded in New Jersey in 1937, was 2.000×10^{8} L. If hydrogen gas has a

   density of 0.0899 g/L, what mass of hydrogen was used by the airship?
4. The volume of an Olympic-sized swimming pool is 2.50×10^{9} cm^3. If the pool is filled with alcohol (d = 0.789 g/cm^3), what mass of alcohol is in the pool?
5. A typical engagement ring has 0.77 cm^3 of gold. What mass of gold is present? (d = 19.3 g/cm^3)
6. A typical mercury thermometer has 0.039 mL of mercury in it. What mass of mercury is in the thermometer? (d = 5.43 g/cm^3)
7. What is the volume of 100.0 g of lead if lead has a density of 11.34 g/cm^3?
8. What is the volume of 255.0 g of uranium if uranium has a density of 19.05 g/cm^3?
9. What is the volume in liters of 222 g of neon if neon has a density of 0.900 g/L?
10. What is the volume in liters of 20.5 g of sulfur hexafluoride if sulfur hexafluoride has a density of 6.164 g/L?

11. Which has the greater volume, 100.0 g of iron (d = 7.87 g/cm^3) or 75.0 g of gold (d = 19.3 g/cm^3)?

12. Which has the greater volume, 100.0 g of hydrogen gas (d = 0.0000899 g/cm^3) or 25.0 g of argon gas (d = 0.00178 g/cm^3)?

Answers

1. 3.04 g/mL

2. 380. g/cm^3

3. 1.80 x 10^7 grams

4. 1.97 x 10^9 grams

5. 15 grams

6. 0.21 grams

7. 5.249 cm^3

8. 13.39 cm^3

9. 247 Liters

10. 3.33 Liters

11. 100.0 g of iron

12. 100.0 g hydrogen gas

Dosage Calculations

Exercises

1. Vitamin C tablets can come in 500 mg tablets. How many of these tablets are needed to obtain 10 g of vitamin C?

2. A tablet of penicillin contains 250 mg of the antibacterial drug. A prescription contains 44 tablets. What is the total mass of penicillin in the prescription?

3. The recommended daily allowance (RDA) for magnesium for 19-30 yrs old+ men is 400 mg. Magnesium supplements come in 200 mg capsules. How many capsules need to be taken to meet RDA?

4. A 175 lb patient is to undergo surgery and will be given an intravenous anesthetic. The safe dosage of anesthetic is up 12 mg/kg of body weight. Determine the maximum dose of anesthetic that should be used. (Hint: 2.2 lbs = 1 kg)

5. The safe dosage of an IV antibiotic for children weighing more than 2.0kg is 60.mg per kilogram of body weight. How many mg should be administered to a child weighing 16 kg?

6. A drug dose of 1.5 mg/kg is ordered for a child weighing 70.4 lbs. How many mg of the drug should be administered? If the drug is available as 60 mg/2 mL. How many mL must the nurse administer?

Answers

1. 20 tablets

2. 11,000 mg; 11 g

3. 2 tablets

4. 955 mg

5. 960 mg

6. 48 mg; 1.6 mL

Chapter Summary

Exercises

1. A sample of urine has a density of 1.105 g/cm^3. What is the mass of 0.255 L of this urine?
2. The hardest bone in the body is tooth enamel, which has a density of 2.91 g/cm^3. What is the volume, in liters, of 75.9 g of tooth enamel?
3. Some brands of aspirin have 81 mg of aspirin in each tablet. If a person takes 8 tablets per day, how many grams of aspirin is that person consuming every day?
4. The US government has a recommended daily intake (RDI) of 5 µg of vitamin D per day. (The name *recommended daily allowance* was changed to RDI in 1997.) If milk contains 1.2 µg per 8 oz glass, how many ounces of milk are needed to supply the RDI of vitamin D?
5. The population of the United States, according to the 2000 census, was 281.4 million people.
    a. What is the unit in this quantity?
    b. Express this quantity in proper scientific notation.
6. The United States produces 34,800,000,000 lb of sugar each year, and much of it is exported to other countries.
    a. What is the unit in this quantity?
    b. Express this quantity in proper scientific notation.
7. Construct a conversion factor that can convert from one unit to the other in each pair of units.
    a. from millimeters to kilometers
    b. from kilograms to micrograms
    c. from centimeters to micrometers
8. Construct a conversion factor that can convert from one unit to the other in each pair of units.
    a. from kilometers to micrometers
    b. from decaliters to milliliters
    c. from megagrams to milligrams
9. What is the density of a dextrose solution if 355 mL of the solution has a mass of 406.9 g?
10. What is the density of a dental amalgam (an alloy used to fill cavities) if 1.005 kg of the material has a volume of 433 mL? Express your final answer in grams per milliliter.

For Exercises 11–16, see the accompanying table for the relationships between English and 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

11. Approximately how many inches are in 4.76 m?
12. Approximately how many liters are in 1 gal, which is exactly 4 qt?

13. Approximately how many kilograms are in a person who weighs 170 lb?

14. The average distance between Earth and the sun is 9.3×10^7 mi. How many kilometers is that?

15. Show mathematically that 1 L equals 1 dm^3.

16. Show mathematically that 1 L equals 1,000 cm^3.

Answers

1. 282 g

2. 26.1 cm^3; 0.0261 L

3. 650 mg

4. 30 oz

5.

   a. people
   b. 2.841×10^8 people

6.

   a. pound (lb)
   b. 3.48×10^{10} lbs

7

   a. $\frac{1\ \text{km}}{10^6\ \text{mm}}$
   b. $\frac{10^9\ \mu\text{g}}{1\ \text{kg}}$
   c. $\frac{10^4\ \mu\text{m}}{1\ \text{cm}}$

8. Construct a conversion factor that can convert from one unit to the other in each pair of units.

   a. $\frac{10^9\ \mu\text{m}}{1\ \text{km}}$
   b. $\frac{100\ \text{mL}}{1\ \text{dL}}$
   c. $\frac{10^9\ \text{mg}}{1\ \text{Mg}}$

9. 1.15 g/mL

10. 2.321 g/mL

11. 187 in.

12. 3.784 L

13. 77 kg

14. 1.5×10^8 km

15. $1\ \text{L} = 0.001\ \text{m}^3 \times \left(\frac{1\ \text{dm}}{0.1\ \text{m}}\right)^3 = 1\ \text{dm}^3$

16. $1\ \text{L} = 0.001\ \text{m}^3 \times \left(\frac{1\ \text{cm}}{0.01\ \text{m}}\right)^3 = 1000\ \text{cm}^3$

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CHAPTER OVERVIEW

4: Chemical Quantities

4.1: Avogadro's Number
4.2: Conversions Between Moles and Atoms
4.3: Molar Mass
4.4: Conversions Between Moles and Mass
4.5: Conversions Between Mass and Number of Particles
4.6: Percent Composition
4.7: Chapter 4 Exercises
4.7.1: Mole Practice with Guidence (Exercises)
4.7.2: Chemical Quantities (Exercises)
4.7.3: Answers to Chemical Quantities (Exercises)

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4.1: Avogadro's Number



Figure 4.1.1 (Public Domain; User:OSX/Wikimedia Commons via Wikipedia)

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

It certainly is easy to count objects such as bananas, or something as large as elephants (as long as you stay out of their way). However, counting grains of sugar from a sugar canister would take a long, long time. Atoms and molecules are extremely small—far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years, without stopping, you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists of the past needed a name that could stand for a very large number of items. Amadeo Avogadro (1776-1856), an Italian scientist, provided such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.



Figure 4.1.2: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry. (Public Domain; C. Sentier via Wikipedia)

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as $NaCl$ and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.



Figure **4.1.3**: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year. ((left; Left: Michael David Hill, 2005 (Mikiwikipikidikipedia); Right: chrisbb@prodigy.net;(left) Michael David Hill, 2005 (Mikiwikipikidikipedia); (right) chrisbb@prodigy.net via Left: http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; Right: http://www.flickr.com/photos/chrisbrenschmidt/436990097;(left) http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; (right) http://www.flickr.com/photos/chrisbrenschmidt/436990097) CC-BY 2.5; (right) CC-BY 2.0)



Summary

- A mole of any substance contains Avogadro's number (6.02×10^{23}) of representative particles.

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

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4.2: Conversions Between Moles and Atoms



Figure 4.2.1 (Public Domain; Diane A. Reid/National Cancer Institute via Wikipedia)

Big numbers or little numbers?

Do you hate to type subscripts and superscripts? Even with a good word-processing program, having to click on an icon to get a superscript and then remembering to click off after you type the number can be a real hassle. If we did not know about moles and just knew about numbers of atoms or molecules (those big numbers that require lots of superscripts), life would be much more complicated and we would make many more typing errors.

Conversions Between Moles and Atoms

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Example 4.2.1: Converting Number of Particles to Moles

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

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

Known

- number of C atoms $= 4.72 \times 10^{24}$
- 1 mole $= 6.02 \times 10^{23}$ atoms

Unknown

- 4.72 x 10^{24} = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures.

Suppose that you want to know how many hydrogen atoms are in a mole of water molecules. First, you need to know the chemical formula for water, which is H_2O. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen are in two water molecules? There are $2 \times 2 = 4$ hydrogen atoms. How about in a dozen? In that case, a dozen is 12; so $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers (4 and 24), you multiply the given number of molecules by two

atoms of hydrogen per molecule. So, to find the number of hydrogen atoms in a mole of water molecules, the problem can be solved using conversion factors:

$$1\ \text{mol}\ H_2O \times \frac{6.02 \times 10^{23}\ \text{molecules}\ H_2O}{1\ \text{mol}\ H_2O} \times \frac{2\ \text{atoms H}}{1\ \text{molecule}\ H_2O} = 1.20 \times 10^{24}\ \text{atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.



Figure 4.2.2: Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Example 4.2.2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is the sample?

Solution:

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

Known

- $4.89 \times 10^{25} = \text{O}$ atoms
- 1 mole $= 6.02 \times 10^{23}$ molecules H_2SO_4

Unknown

- mol of H_2SO_4 molecules

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25}\ \text{atoms O} \times \frac{1\ \text{molecule}\ H_2SO_4}{4\ \text{atoms O}} \times \frac{1\ \text{mol}\ H_2SO_4}{6.02 \times 10^{23}\ \text{molecules}\ H_2SO_4} = 20.3\ \text{mol}\ H_2SO_4$$

Step 3: Think about your result.

The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.



Summary

- Methods are described for conversions between moles, atoms, and molecules.

Review

1. What conversion factor would we need to convert moles of helium to atoms of helium?
2. I want to convert atoms to moles. My friend tells me to multiply the number of atoms by 6.02×10^{23} atoms/mole. Is my friend correct?
3. Why do you need to know the formula for a molecule in order to calculate the number of moles of one of the atoms?
4. How many atoms of fluorine are in 5.6×10^{22} molecules of MgF_2?

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4.3: Molar Mass



Figure 4.3.1 (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

When creating a solution, how do I know how much of each substance to put in?

I want to make a solution that contains 1.8 moles of potassium dichromate. You do not have a balance calibrated in moles, but you do have one calibrated in grams. If you know the relationship between moles and the number of grams in a mole, you can use your balance to measure out the needed amount of material.

Molar Mass

Molar mass is defined as the mass of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of lithium is $6.94\ \text{g}$, the molar mass of zinc is $65.38\ \text{g}$, and the molar mass of gold is $196.97\ \text{g}$. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole, or g/mol.

Molar Masses of Compounds

The molecular formula of the compound carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen:

$$12.01\ \text{amu} + 2\,(16.00\ \text{amu}) = 44.01\ \text{amu}$$

The **molecular mass** of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is $44.01\ \text{amu}$.

The molar mass of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of $44.01\ \text{g}$, while one mole of sodium sulfide formula units has a mass of $78.04\ \text{g}$. The molar masses are $44.01\ \text{g/mol}$ and $78.04\ \text{g/mol}$ respectively. In both cases, that is the mass of 6.02×10^{23} representative particles. The representative particle of CO_2 is the molecule, while for Na_2S it is the formula unit.

Example 4.3.1: Molar Mass of a Compound

Calcium nitrate, $\text{Ca(NO}_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Solution

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

Known

- Formula $= \text{Ca(NO}_3)_2$
- Molar mass $\text{Ca} = 40.08\ \text{g/mol}$
- Molar mass $\text{N} = 14.01\ \text{g/mol}$
- Molar mass $\text{O} = 16.00\ \text{g/mol}$

Unknown

- molar mass $\text{Ca(NO}_3)_2$

First we need to analyze the formula. Since the **Ca** lacks a subscript, there is one **Ca** atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit and each nitrate ion consists of one nitrogen atom and three oxygen atoms per formula unit. Thus, **1 mol** of calcium nitrate contains **1 mol** of **Ca** atoms, **2 mol** of **N** atoms, and **6 mol** of **O** atoms.

Step 2: Calculate

Use the molar masses of each atom together with the number of atoms in the formula and add together.

$$1\ \text{mol Ca} \times \frac{40.08\ \text{g Ca}}{1\ \text{mol Ca}} = 40.08\ \text{g Ca}$$

$$2\ \text{mol N} \times \frac{14.01\ \text{g N}}{1\ \text{mol N}} = 28.02\ \text{g N}$$

$$6\ \text{mol O} \times \frac{16.00\ \text{g O}}{1\ \text{mol O}} = 96.00\ \text{g O}$$

Molar mass of $\text{Ca(NO}_3\text{)}_2 = 40.08\ \text{g} + 28.02\ \text{g} + 96.00\ \text{g} = 164.10\ \text{g/mol}$



Summary

- Calculations are described for the determination of molar mass of an atom or a compound.

Review

1. What is the molar mass of Pb?
2. Where do you find the molar mass of an element?
3. How many moles of Cl are in one mole of the $CaCl_2$?
4. How many moles of H are in one mole of the compound $(NH_4)_3PO_4$?
5. Calculate the molar mass of $CaCl_2$.

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4.4: Conversions Between Moles and Mass



Figure 4.4.1 (Public Domain; User:Estormiz/Wikimedia Commons via Wikipedia)

How can we get more product?

Chemical manufacturing plants are always seeking to improve their processes. One way that improvement comes about is through measuring the amount of material produced in a reaction. By knowing how much is made, the scientists and engineers can try different ways of getting more product at less cost.

Conversions Between Moles and Mass

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment, you need 3.00 moles of calcium chloride ($\mathrm{CaCl_2}$). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. The molar mass of $\mathrm{CaCl_2}$ is $110.98\ \mathrm{g/mol}$. The conversion factor that can be used is then based on the equality that $1\ \mathrm{mol} = 110.98\ \mathrm{g\ CaCl_2}$. Dimensional analysis will allow you to calculate the mass of $\mathrm{CaCl_2}$ that you should measure.

$$3.00\ \mathrm{mol\ CaCl_2} \times \frac{110.98\ \mathrm{g\ CaCl_2}}{1\ \mathrm{mol\ CaCl_2}} = 333\ \mathrm{g\ CaCl_2}$$

When you measure the mass of $333\ \mathrm{g}$ of $\mathrm{CaCl_2}$, you are measuring 3.00 moles of $\mathrm{CaCl_2}$.



Figure 4.4.2: Calcium chloride is used as a drying agent and as a road deicer. (Public Domain; Martin Walker (User:Walkerma/Wikimedia Commons) via Wikipedia)

Example 4.4.1: Converting Moles to Mass

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

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

Known

- Molar mass of $\mathrm{Cr} = 52.00\ \mathrm{g/mol}$
- $0.560\ \mathrm{mol\ Cr}$

Unknown

- $0.560\ \mathrm{mol\ Cr}$ = ? g

One conversion factor will allow us to convert from the moles of Cr to mass.

Step 2: Calculate.

$$0.560\ \mathrm{mol\ Cr} \times \frac{52.00\ \mathrm{g\ Cr}}{1\ \mathrm{mol\ Cr}} = 29.1\ \mathrm{g\ Cr}$$

Step 3: Think about your result.

Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because of the $0.560\ \mathrm{mol}$.

A similar conversion factor utilizing molar mass can be used to convert from the mass of a substance to moles. In a laboratory situation, you may perform a reaction and produce a certain amount of a product which can be massed. It will often then be necessary to determine the number of moles of the product that was formed. The next problem illustrates this situation.

Example $4.4.2$: Converting Mass to Moles

A certain reaction produces $2.81\ \mathrm{g}$ of copper (II) hydroxide, $\mathrm{Cu(OH)_2}$. Determine the number of moles produced in the reaction.

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

Known

- mass = 2.81g

Unknown

- mol $\mathrm{Cu(OH)_2}$

One conversion factor will allow us to convert from mass to moles.

Step 2: Calculate.

First, it is necessary to calculate the molar mass of $\mathrm{Cu(OH)_2}$ from the molar masses of Cu, O, and H. The molar mass is $97.57\ \mathrm{g/mol}$.

$$2.81\ \mathrm{g\ Cu(OH)_2} \times \frac{1\ \mathrm{mol\ Cu(OH)_2}}{97.57\ \mathrm{g\ Cu(OH)_2}} = 0.0288\ \mathrm{mol\ Cu(OH)_2}$$

Step 3: Think about your result.

The relatively small mass of product formed results in a small number of moles.



Summary

- Calculations involving conversions between moles of a material and the mass of that material are described.

Review

1. You have 19.7 grams of a material and wonder how many moles were formed. Your friend tells you to multiply the mass by grams/mole. Is your friend correct?
2. How many grams of MgO are in 3.500 moles?
3. How many moles of H_2O are in 15.2 grams of pure ice?

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4.5: Conversions Between Mass and Number of Particles

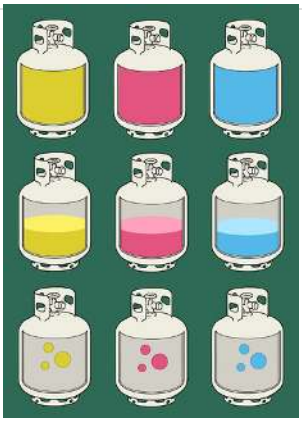


Figure 4.5.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

How much gas is there?

Avogadro was interested in studying gases. He theorized that equal volumes of gases under the same conditions contained the same number of particles. Other researchers studied how many gas particles were in specific volumes of gases. Eventually, scientists were able to develop the relationship between number of particles and mass, using the idea of moles.

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and the mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to grams. In order to convert from mass to number of particles or vice-versa, a conversion to moles is required.



Figure 4.5.2: Conversion from number of particles to mass, or from mass to number of particles requires two steps. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 4.5.1: Converting Mass to Particles

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

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

Known

- Molar mass $\text{Cl}_2 = 70.90\ \text{g/mol}$
- $20.0\ \text{g}\ \text{Cl}_2$

Unknown

- number of molecules of Cl_2

Use two conversion factors. The first converts grams of Cl_2 to moles. The second converts moles of Cl_2 to the number of molecules.

Step 2: Calculate.

$$20.0\ \text{g}\ \text{Cl}_2 \times \frac{1\ \text{mol}\ \text{Cl}_2}{70.90\ \text{g}\ \text{Cl}_2} \times \frac{6.02 \times 10^{23}\ \text{molecules}\ \text{Cl}_2}{1\ \text{mol}\ \text{Cl}_2} = 1.70 \times 10^{23}\ \text{molecules}\ \text{Cl}_2$$

The problem is done using two consecutive conversion factors. There is no need to explicitly calculate the moles of Cl_2.

Step 3: Think about your result.

Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.



Summary

- Calculations are illustrated for conversions between mass and number of particles.

Review

1. Why can’t we convert directly from number of particles to grams?
2. How many atoms of chlorine are present in 1.70×10^{23} molecules Cl_2?
3. How many molecules of BH_3 are in 14.32 grams BH_3?

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4.6: Percent Composition



Figure 4.6.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Is there anything healthy in this jar?

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

Percent Composition

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

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

Percent Composition from Mass Data

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

Example 4.6.1: Percent Composition from Mass

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

Solution

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

Known

- Mass of compound $= 20.00\ \text{g}$
- Mass of $\text{Zn} = 16.07\ \text{g}$

Unknown

- Percent $\text{Zn} = ?\%$
- Percent $\text{O} = ?\%$

Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass.

Step 2: Calculate.

$$\text{Mass of oxygen} = 20.00\ \text{g} - 16.07\ \text{g} = 3.93\ \text{g O}$$

$$\%\ \text{Zn} = \frac{16.07\ \text{g Zn}}{20.00\ \text{g}} \times 100\% = 80.35\%\ \text{Zn}$$

$$\%\ \text{O} = \frac{3.93\ \text{g O}}{20.00\ \text{g}} \times 100\% = 19.65\%\ \text{O}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 4.6.2: Percent Composition from Chemical Formula

Dichlorineheptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorineheptoxide.

Solution

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

Known

- Mass of Cl in $1\text{ mol } Cl_2O_7 = 70.90\text{ g}$
- Mass of O in $1\text{ mol } Cl_2O_7 = 112.00\text{ g}$
- Molar mass of $Cl_2O_7 = 182.90\text{ g/mol}$

Unknown

- Percent Cl $=?\%$
- Percent O $=?\%$

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\%\text{Cl} = \frac{70.90\text{ g Cl}}{182.90\text{ g}} \times 100\% = 38.76\%\text{ Cl}$$

$$\%\text{ O} = \frac{112.00\text{ g O}}{182.90\text{ g}} \times 100\% = 61.24\%\text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is 38.76% Cl and 61.24% O. Suppose that you need to know the masses of chlorine and oxygen present in a 12.50 g sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element:

$$12.50\text{ g } Cl_2O_7 \times \frac{38.76\text{ g Cl}}{100\text{ g } Cl_2O_7} = 4.845\text{ g Cl}$$

$$12.50\text{ g } Cl_2O_7 \times \frac{61.24\text{ g O}}{100\text{ g } Cl_2O_7} = 7.655\text{ g O}$$

The sum of the two masses is 12.50 g, the mass of the sample size.



Science Friday: Stained Glass Conservation

Stained glass from the Middle Ages is often hundreds of years old. Unfortunately, many of these relics are in need of cleaning and maintenance. In this video by Science Friday, conservator Mary Higgins discusses the methods used to protect the stained glass.



Summary

- Processes are described for calculating the percent composition of a material based on mass or on chemical composition.

Review

1. What is the formula for calculating percent composition?
2. What information do you need to calculate percent composition by mass?
3. What do subscripts in a chemical formula tell you?

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4.7: Chapter 4 Exercises

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4.7.1: Mole Practice with Guidence (Exercises)

Template:HideTOC

These are homework exercises to accompany the Textmap created for Chemistry: A Molecular Approach by Nivaldo Tro. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here. In addition to these publicly available questions, access to private problems bank for use in exams and homework is available to faculty only on an individual basis; please contact Delmar Larsen for an account with access permission.

Mole/Mass Conversions

1. Find the amount of moles contained in each specimen.

a. 7.87 kg H_2O_2
b. 2.34 kg NaCl
c. 12.5 g C_2H_6O
d. 85.72 g NH_3

Strategy

A. First, determine the units that are given for each sample. In order to convert these measurements to moles, each sample should first be written in grams. Two of the given samples are measured in kilograms. Use the following conversion factor to convert kilograms to grams, with x representing the given mass:

$$x\,kg \cdot \frac{1000\,g}{1\,kg} \tag{4.7.1.1}$$

B. Next, find the atomic masses of all the atoms in each compound by using the Periodic Table. Then, add these atomic masses together for each compound. The resulting value will be the number of grams of each sample that make up one mole.

C. Convert the mass in grams of each sample to moles by multiplying it by the following conversion factor, called the molar mass, with *x* representing the mass of the given specimen and y representing the calculated atomic mass found in step B:

$$x\,g \cdot \frac{1\,mole}{y\,g} \tag{4.7.1.2}$$

A Numbers 1 and 2 first need to be converted into grams.

1. $$7.87\,kg\,H_2O_2 \cdot \frac{1000\,g}{1\,kg} = 7,870\,g\,H_2O_2 \tag{4.7.1.3}$$

2. $$2.34\,kg\,NaCl \cdot \frac{1000\,g}{1\,kg} = 2,340\,g\,NaCl \tag{4.7.1.4}$$

B The following are the calculated molar masses of each of the given compounds. The atomic mass of each element has been rounded to 4 significant figures in these calculations.

1. H_2O_2:

$$2\,(1.008\,g\,H) + 2\,(16.00\,g\,O) = \frac{34.02\,g}{1\,mole}\,H_2O_2 \tag{4.7.1.5}$$

2. NaCl:

$$(22.99\,g\,Na) + (35.45\,g\,Cl) = \frac{58.44\,g}{1\,mole}\,NaCl \tag{4.7.1.6}$$

3. C_2H_6O:

$$2\,(12.01\,g\,C) + 6\,(1.008\,g\,H) + (16.00\,g\,O) = \frac{46.07\,g}{1\,mole}\,C_2H_6O \tag{4.7.1.7}$$

4. NH_3:

$$(14.01\,g\,N) + 3\,(1.008\,g\,H) = \frac{17.03\,g}{1\,mole}\,NH_3 \tag{4.7.1.8}$$

C The number of moles in each specimen can now be calculated by multiplying the mass in grams of each sample by its molarity.

1. $$7,870\,g\,H_2O_2 \cdot \frac{1\,mole\,H_2O_2}{34.02\,g\,H_2O_2} = 231\,moles\,H_2O_2 \tag{4.7.1.9}$$
2. $$2,340\,g\,NaCl \cdot \frac{1\,mole\,NaCl}{58.44\,g\,NaCl} = 40.0\,moles\,NaCl \tag{4.7.1.10}$$
3. $$12.5\,g\,C_2H_6O \cdot \frac{1\,mole\,C_2H_6O}{46.07\,g\,C_2H_6O} = 0.271\,moles\,C_2H_6O \tag{4.7.1.11}$$
4. $$85.72\,g\,NH_3 \cdot \frac{1\,mole\,NH_3}{17.03\,g\,NH_3} = 5.033\,moles\,NH_3 \tag{4.7.1.12}$$

Mole/Mass Conversions

2. Calculate the number of moles in each of the following examples.

a. 402.5 mg of NO_2
b. 2.7 kg of H_2O
c. 323 g of CBr_4
d. 2.9 kg of CaO

Solution

a)

First, calculate the molar mass of NO_2.

14.007g/mol N + 2 (15.999 g/mol O) = ***46.0055 g/mol***

Now, convert the sample from mg to g.

402.5 ~~mg~~ NO_2	1 g	**0.4025 g NO_2**
	1000 ~~mg~~	

Finally, use the molar mass to determine how many moles your sample size contains.

0.4025 g NO_2	1 mol	**0.0087 moles of NO_2**
	46.0055 g	

This gives us the final answer, **0.0087 moles of NO_2**.

From this, we can deduce that:

$$\text{Moles of substance} = \text{Mass of substance (g)}\,Molar mass of substance \tag{4.7.1.13}$$

b) 150 moles of H_2O

c) 0.974 moles of CBr_4

d) 52 moles of CaO

Mass/Particle Conversions

1. For each substance, convert the given molecules to mass in grams:

a. 3.2 x 10^{24} Cl_2 molecules
b. 8.25 x 10^{18} CH_2O molecules
c. 1 carbon dioxide molecule

Strategy:

Step 1: Convert the given molecules to moles by dividing by Avogadro's number.

Step 2: Multiply the number of moles by the Molar Mass of the substance to determine the grams.

Step 3: Through correct Dimensional Analysis, the molecules and moles will cancel to leave grams.

$$\frac{molecules}{Avogadro's} \rightarrow moles \rightarrow moles \cdot MM \rightarrow grams \tag{4.7.1.14}$$

Helpful Hints:

- Avogadro's number: 6.022 x 10^{23} mol of a substance
- Molar mass (MM): the weight of a substance in the units g/mol, found by adding the atomic masses of each element.
    - Example: MM of H_2O --> H(1.008 x 2) + O(15.999)= 18 g/mol

Solution

a) 3.8 x 10^{51} grams

b) 4.11 x 10^{45} grams

c) 7.31 x 10^{25} grams

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4.7.2: Chemical Quantities (Exercises)

These are homework exercises taken from the following;

Chapter 6 of the Ball et al. "The Basics of GOB Chemistry" Textmap.

Textmap created for "Chemistry" by OpenStax.

- Contributed by Paul Flowers, Klaus Theopold & Richard Langley et al.
- General Chemistry at OpenStax CNX

Textmap created for Chemistry: A Molecular Approach by Nivaldo Tro. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here.

Some questions have been contributed and modified by Kristina Novek

Using Moles & Avogadro's Number

1. How many moles of each element are present in 2.67 mol of each compound?

    a. HCl
    b. H_2SO_4
    c. $Al(NO_3)_3$
    d. $Ga_2(SO_4)_3$

2. How many moles of each element are present in 0.45 mol of each compound?

    a. HCl

    b. H_2SO_4

    c. $Al_2(CO_3)_3$

    d. $Ga_2(SO_4)_3$

3. Calculate the number of atoms or ions of:

    a. 2.00 mole of Fe

    b. 4.36 mol of Si

    c. 1.20×10^{-2} mol of Li^+

    d. 0.875 mol of S^{2-}

4. Calculate the number of molecules or formula units of:

    a. 0.0950 mol of H_2O

    b. 3.20×10^{-3} mol of NaCl

    c. 2.06 mol of $CaBr_2$

    d. 0.14 mol of $C_6H_{12}O_6$ (glucose)

5. Calculate the number of moles from each of the following.

    a. 6.95×10^{23} atoms of Cu

    b. 7.12×10^{21} molecules of CO_2

c. 3.33 x 10^{22} ions of Pb^{4+}

d. 7.00 x 10^{23} formula units of $LiNO_3$

e. 5.16 x 10^{20} molecules of N_2

6. A 55-kg woman has 7.5×10^{-3} mol of hemoglobin in her blood. How many hemoglobin molecules is this?

Molar Mass

1. Calculate the molar mass of each of the following:

(a) P_4

(b) H_2O

(c) $Ca(NO_3)_2$

(d) CH_3CO_2H (acetic acid)

(e) $C_{12}H_{22}O_{11}$ (sucrose, cane sugar).

2.Calculate the molar mass of each of the following:

(a) S_8

(b) C_5H_{12}

(c) $Sc_2(SO_4)_3$

(d) CH_3COCH_3 (acetone)

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

3. If 6.63×10^{-6} mol of a compound has a mass of 2.151 mg, what is the molar mass of the compound?

4. of 8.65 moles of a compound has a mass of 1.47 kilograms, what is the molar mass of the compound?

Mole/Mass Conversions

1. Determine the mass of each of the following:

(a) 0.0146 mol KOH

(b) 10.2 mol ethane, C_2H_6

(c) 1.6×10^{-3} mol $Na_2 SO_4$

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

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

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

3. Determine the mass in grams of each of the following:

(a) 0.600 mol of oxygen atoms

(b) 0.600 mol of oxygen molecules, O_2

(c) 0.600 mol of ozone molecules, O_3

4. Determine the number of moles in each of the following:

(a) 86.5 grams of $MgSO_4$
(b) 2.30 x 10^2 grams of CO
(c) 62.8 grams of KCN
(d) 0.44 grams of Cl_2

5. One of the chemicals that gives lavender its' pleasant smell is called linalool ($C_{10}H_{18}O$). If a sample of lavender contained 9.14 x 10^{-5} grams of linalool, calculate the number of moles of linalool in the sample.

Mass/Particle Conversions

1. Convert the following given masses to number of particles:

(a) 4.2 grams of nickel
(b) 9.30 x 10^{-2} grams of Br^-
(c) 0.846 grams of hydrogen gas
(d) 0.44 grams of H_2O_2 (hydrogen peroxide)

2. Convert the following particles to grams:

(a) 7.90 x 10^{22} formula units of NaOH
(b) 2.03 x 10^{24} molecules of $C_{12}H_{22}O_{11}$ (lactose)
(c) 8.85 x 10^{23} atoms of Fe
(d) 4.34 x 10^{21} ions of NO_2^-

3. Glycerol ($C_3H_8O_3$) also know as glycerin, and is commonly used in soaps, toothpaste, and moisturizers. How many molecules are in one gram of glycerol?

4. Ammonia (NH_3) is mostly used in agriculture as a fertilizer and can be used as a multipurpose cleanser. What is the mass of 9.85 x 10^{22} molecules of ammonia?

5. What is the mass of one tin (Sn) atom in grams?

Percent Composition

1. The chemical formula for aspirin is $C_9H_8O_4$. Calculate the mass percent composition of aspirin.

2. Cannabidiol ($C_{21}H_{26}O_2$) also known as CBD is a chemical found in the cannabis plant. It has been used to treat epilepsy and anxiety. Calculate the mass percent composition of CBD.

3.What is the mass percent of of hydrogen in the common pain reliever acetaminophen ($C_8H_9NO_2$)?

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4.7.3: Answers to Chemical Quantities (Exercises)

Using Moles & Avogadro's Number

1.

a. 2.67 mol of H and 2.67 mol of Cl
b. 5.34 mol of H, 2.67 mol of S, and 10.68 mol of O
c. 2.67 mol of Al, 8.01 mol of N, and 24.03 mol of O
d. 5.34 mol of Ga, 8.01 mol of S, and 32.04 mol of O

2.

a. 0.45 mol of H and 0.45 mol of Cl

b. 0.90 mol of H, 0.45 mol S, and 1.8 mol O

c. 0.90 mol Al, 1.35 mol C, and 4.05 mol O

d. 0.90 mol Ga, 1.35 mol S, 5.4 mol O

3. a. 1.20 x 10^{24} atoms of Fe; b. 2.63 x 10^{24} atoms of Si; c. 7.23 x 10^{21} atoms of Li^{+}; d. 5.27 x 10^{23} S^{2-} ions

4.

a. 5.72 x 10^{22} molecules of water (H_2O)

b. 1.93 x 10^{21} formula units of sodium chloride (NaCl)

c. 1.24 x 10^{24} formula units of calcium bromide ($CaBr_2$)

d. 8.43 x 10^{22} molecules of glucose ($C_6H_{12}O$)

5. a. 1.15 moles Cu; b. 0.0118 moles CO_2; c. 0.0553 mol Pb^{4+}; d. 1.16 mol $LiNO_3$; e. 0.857 mol N_2

6. 4.52 x 10^{21} molecules of hemoglobin

Molar Mass

1. (a) 123.896 g/mol; (b) 18.015 g/mol; (c) 164.086 g/mol; (d) 60.052 g/mol; (e) 342.297 g/mol

2. (a) 256.528 g/mol; (b) 72.150 g/mol; (c) 378.103 g/mol; (d) 58.080 g/mol; (e) 180.158 g/mol

3. 324.43 g/mol

4. 169.94 g/mol

Mole/Mass Conversions

1. (a) 0.819 g; (b) 307 g; (c) 0.23 g; (d) 1.235 × 10^6 g (1235 kg); (e) 765 g

2. 8.38 x 10^{-3} moles of leucine

3. (a) 9.60 g; (b) 19.2 g; (c) 28.8 g

4. (a) 0.719 mol $MgSO_4$; (b) 8.21 mol CO; (c) 0.964 mol KCN; (d) 6.2 x 10^{-3} mol Cl_2

5. 5.93 x 10^{-7} moles of linalool

Mass/Particle Conversions

1. (a) 4.31 x 10^{22} atoms Ni ; (b) 7.01 x 10^{20} ions Br^- ; (c) 2.52 x 10^{23} molecules H_2 ; (d) 7.79 x 10^{21} molecules H_2O_2

2. (a) 5.25 g NaOH; (b) 1.15 x 10^3 g lactose; (c) 82.1 g Fe; (d) 0.332 g NO_2^-

3. 6.54 x 10^{21} molecules of $C_3H_8O_3$

4. 2.79 grams of NH_3

5. 1.97 x 10^{-22} grams of Sn

Percent Composition

1. 60.00% C, 4.48% H, 35.52% O

2. 80.20% C, 8.35% H, 10.18% O

3. 6.01% H

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CHAPTER OVERVIEW

5: Chemical Reactions and Quantities

5.1: Prelude to Introduction to Chemical Reactions
5.2: The Law of Conservation of Matter
5.3: Chemical Equations
5.4: Quantitative Relationships Based on Chemical Equations
5.5: Mole-Mole Relationships in Chemical Reactions
5.6: Mole-Mass and Mass-Mass Problems

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5.1: Prelude to Introduction to Chemical Reactions

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 ($\mathrm{C_6H_{12}O_6}$). They convert glucose into carbon dioxide ($\mathrm{CO_2}$) and ethanol ($\mathrm{C_2H_5OH}$) in a chemical transformation that is represented as follows:

$$\mathrm{C_6H_{12}O_6\,(s) \longrightarrow 2\,CO_2\,(g) + 2\,C_2H_5OH(\ell)}$$

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.

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5.2: The Law of Conservation of Matter

Learning Objectives

- Correctly define a law as it pertains to science.
- State 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*.

With the development of more precise ideas on elements, compounds and mixtures, scientists began to investigate how and why substances react. French chemist A. Lavoisier laid the foundation to the scientific investigation of matter by describing that substances react by following certain laws. These laws are called the laws of chemical combination. These eventually formed the basis of Dalton's Atomic Theory of Matter.

Law of Conservation of Mass

According to this law, during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants.

$$\underbrace{\overbrace{\mathrm{HgO\,(s)}}^{\text{Mercuric oxide}}}_{100\ \mathrm{g}} \rightarrow \underbrace{\overbrace{\mathrm{Hg\,(l)}}^{\text{Mercury}}}_{92.6\ \mathrm{g}} + \underbrace{\overbrace{\mathrm{O_2\,(g)}}^{\text{Oxygen}}}_{7.4\ \mathrm{g}}$$

Another way of stating this is, "In a chemical reaction, matter is neither created nor destroyed." The law of conservation of mass is also known as the "law of indestructibility of matter."

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.



A sample of calcium carbonate (CaCO3). (Public Domain; Walkerma).

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.

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.

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Figure 5.1.1 shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose we had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose we had been able to measure the oxygen used by the fire and the gases produced by the fire. What would we find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

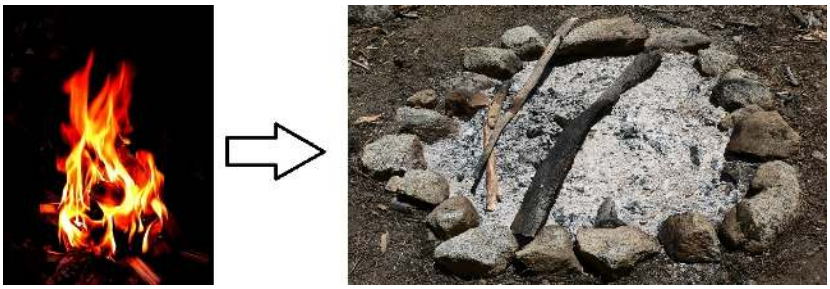


Figure 5.1.1: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). Images used with permission (CC BY-SA 2.5; Einar Helland Berger for fire and for ash).

? Exercise 5.2.1

a. What is the law of conservation of matter?
b. How does the law of conservation of matter apply to chemistry?

Answer a:

The law of conservation of matter states that in any given system that is closed to the transfer of matter, the amount of matter in the system stays constant

Answer b:

The law of conservation of matter says that in chemical reactions, the total mass of the products must equal the total mass of the reactants.

Key Takeaway

The amount of matter in a closed system is conserved.

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5.3: Chemical Equations

Learning Objectives

- Define chemical reaction.
- Use a balanced chemical equation to represent a chemical reaction.

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 **5.3.1** shows a rather dramatic example of this very reaction.

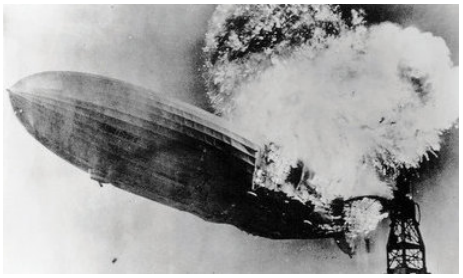


Figure **5.3.1**: The Formation of Water. Hydrogen and oxygen combine to form water. Here, the hydrogen gas in the zeppelin *SS Hindenburg* reacts with oxygen in the air to make water. Source: Photo courtesy of the US Navy. For a video of this see www.youtube.com/watch?v=CgWHbpMVQ1U.

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** (→) represents the chemical change:

$$\mathbf{H_2 + O_2 \rightarrow H_2O} \tag{5.3.1}$$

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 would be as follows:

$$\mathbf{H_2(g) + O_2(g) \rightarrow H_2O(\ell)} \tag{5.3.2}$$

Chemical equations can also be used to describe physical changes. We will see examples of this soon.

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. The equation is **not balanced** because the number of oxygen atoms on each side is not the same (Figure **5.3.2**).

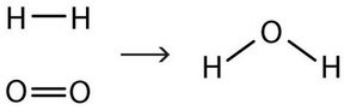


Figure **5.3.2**: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is not balanced as written.

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, perhaps two water molecules are produced, not just one:

$$H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) \tag{5.3.3}$$

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.

> This point is so important that we should repeat it. You **cannot** change the formula of a chemical substance to balance a chemical reaction! You **must** use the proper chemical formula of the substance.

Unfortunately, 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. This can be easily fixed, however, by putting a coefficient of 2 in front of the diatomic hydrogen reactant:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) \tag{5.3.4}$$

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.

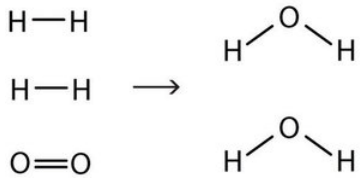


Figure **5.3.3**: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is now balanced.

Proper chemical equations should be balanced. Writing balanced reactions is a chemist's way of acknowledging the law of conservation of matter.

✔ Example 5.3.1

Is each chemical equation balanced?

a. $2Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$
b. $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$
c. $AgNO_3(aq) + 2KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$

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.

? Exercise 5.3.1

Is each chemical equation balanced?

a. $2Hg(l) + O_2(g) \longrightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$
c. $Mg(NO_3)_2(s) + 2Li(s) \longrightarrow Mg(s) + 2LiNO_3(s)$.

Answer a:

balanced

Answer b:

O is not balanced; the 4 atoms of oxygen on the left does not balance with the 6 oxygen atoms on the right

Answer c:

balanced

How does one balance a chemical equation, starting with the correct formulas of the reactants and products? Basically, a **back-and-forth approach** is adopted, counting the number of atoms of one element on one side, checking the number of atoms of that element on the other side, and changing a coefficient if necessary. Then check another element, going back and forth from one side of the equation to another, until each element has the same number of atoms on both sides of the arrow. In many cases, it does not matter which element is balanced first and which is balanced last, as long as all elements have the same number of atoms on each side of the equation.

Below are guidelines for writing and balancing chemical equations.

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

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + HCl}$$

Step 1: Write the skeleton equation with the correct formulas.

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + HCl} \tag{5.3.5}$$

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 ions	1 H ions
2 Cl atom	5 Cl atoms

Step 3: 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:

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{5.3.6}$$

Reactants	Products
1 C atom	1 C atom
4 H ions	4 H ions
2 Cl atom	8 Cl atoms

Now each side has four hydrogen atoms. The product side has a total of eight chlorine atoms (four from the $\mathrm{CCl_4}$ 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 $\mathrm{Cl_2}$ reactant:

$$CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl \tag{5.3.7}$$

Reactants	Products
1 C atom	1 C atom
4 H ions	4 H ions
8 Cl atom	8 Cl atoms

Step 3: 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 5.3.2

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.

$$Pb(NO_3)_2\,(aq) + NaCl\,(aq) \rightarrow NaNO_3\,(aq) + PbCl_2\,(s)$$

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 atom	2 Na atom
2 Cl atom	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.

Exercise 5.3.2

Fermentation is a biochemical process that enables yeast cells to live in the absence of oxygen. Humans have exploited it for centuries to produce wine and beer and make bread rise. In fermentation, sugars such as glucose ($C_6H_{12}O_6$) are converted to ethanol (C_2H_5OH) and carbon dioxide CO_2. Write a balanced chemical reaction for the fermentation of glucose.



Commercial use of fermentation. (a) Microbrewery vats are used to prepare beer. (b) The fermentation of glucose by yeast cells is the reaction that makes beer production possible.

Answer

$$C_6H_{12}O_6\,(s) \longrightarrow 2\,C_2H_5OH\,(l) + 2\,CO_2\,(g)$$

Key Takeaway

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: Quantitative Relationships Based on Chemical Equations

Learning Objectives

- To calculate the amount of one substance that will react with or be produced from a given amount of another substance.

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.

> *The word stoichiometry is pronounced "stow-eh-key-OM-et-tree." It is of mixed Greek and English origins, meaning roughly "measure of an element."*

Looking Closer: Stoichiometry in Cooking

Let us consider a stoichiometry analogy from the kitchen. A recipe that makes 1 dozen biscuits needs 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 we were to write this as a chemical equation, we would write

$$\text{2 c flour + 1 egg + 4 tbsp shortening + 1 tsp salt + 1 tsp baking soda + 1 c milk} \rightarrow \text{12 biscuits}$$

(Unlike true chemical reactions, this one has all 1 coefficients written explicitly—partly because of the many different units here.) This equation gives us 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.

We can use the ratios we derive from the equation for predictive purposes. For instance, if we have 4 cups of flour, how many biscuits can we make if we have enough of the other ingredients? It should be apparent that we can make a double recipe of 24 biscuits.

But how would we find this answer formally, that is, mathematically? We would set up a conversion factor, much like we did in Chapter 1. Because 2 cups of flour make 12 biscuits, we can set up an equivalency ratio:

$$\frac{\text{12 biscuits}}{\text{2 c flour}}$$

We then can use this ratio in a formal conversion of flour to biscuits:

$$\text{4 c flour} \times \frac{\text{12 biscuits}}{\text{2 c flour}} = \text{24 biscuits}$$

Similarly, by constructing similar ratios, we can determine how many biscuits we can make from any amount of ingredient. When you are doubling or halving a recipe, you are doing a type of stoichiometry. Applying these ideas to chemical reactions should not be difficult if you use recipes when you cook.



A recipe shows how much of each ingredient is needed for the proper reaction to take place. (CC BY-SA 3.0 Unported; Peng via Wikipedia)

Consider the following balanced chemical equation:

$$2\,C_2H_2 + 5\,O_2 \rightarrow 4\,CO_2 + 2\,H_2O \tag{5.4.1}$$

The coefficients on the chemical formulas give the ratios in which the reactants combine and the products form. Thus, we can make the following statements and construct the following ratios:

Table uses a statement from the balanced chemical reaction to show ratios and inverse ratios.

Statement from the Balanced Chemical Reaction	Ratio	Inverse Ratio
two C_2H_2 molecules react with five O_2 molecules	$\frac{2C_2H_2}{5O_2}$	$\frac{5O_2}{2C_2H_2}$
two C_2H_2 molecules react to make four CO_2 molecules	$\frac{2C_2H_2}{4CO_2}$	$\frac{4CO_2}{2C_2H_2}$
five O_2 molecules react to make two H_2O molecules	$\frac{5O_2}{2H_2O}$	$\frac{2H_2O}{5O_2}$
four CO_2 molecules are made at the same time as two H_2O molecules	$\frac{2H_2O}{4CO_2}$	$\frac{4CO_2}{2H_2O}$

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

$$26C_2H_2 \times \frac{4CO_2}{2C_2H_2}$$

The molecules of C_2H_2 cancel, and we are left with molecules of CO_2. Multiplying through, we get

$$26C_2H_2 \times \frac{4CO_2}{2C_2H_2} = 52CO_2$$

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 5.4.1

Start with this balanced chemical equation.

$$\mathbf{KMnO_4 + 8\,HCl + 5\,FeCl_2 \longrightarrow 5\,FeCl_3 + MnCl_2 + 4\,H_2O + KCl}$$

a. Verify that the equation is indeed balanced.
b. Give 2 ratios that give the relationship between $\mathbf{HCl}$ and $\mathbf{FeCl_3}$.

Solution a

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.

Solution b

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: $\mathbf{\frac{8HCl}{5FeCl_3}}$ **and** $\mathbf{\frac{5FeCl_3}{8HCl}}$. There are a total of 42 possible ratios. Can you find the other 40 relationships?

Exercise 5.4.1

Start with this balanced chemical equation.

$$\mathbf{2KMnO_4 + 3CH_2{=}CH_2 + 4H_2O \rightarrow 2MnO_2 + 3HOCH_2CH_2OH + 2KOH}$$

a. Verify that the equation is balanced.
b. Give 2 ratios that give the relationship between $KMnO_4$ and $CH_2{=}CH_2$. (A total of 30 relationships can be constructed from this chemical equation. Can you find the other 28?)

Answer a:

Each side has 2 K atoms and 2 Mn atoms. On the reactant side, 3 $CH_2{=}CH_2$ yield 6 C atoms and on the product side, 3 $HOCH_2CH_2OH$ also yield 6 C atoms, so the C atoms are balanced. There are 20 H atoms on the reactants side: 12 H atoms from 3 $CH_2{=}CH_2$ and 8 H atoms from 4 H_2O. On the product side, there are also 20 H atoms: 18 H atoms from 3 $HOCH_2CH_2OH$ and 2 H atoms from 2 KOH. So, the H atoms are balanced. As for O, on the reactant side, there are 8 O atoms from 2 $KMnO_4$ and 4 O atoms from 4 H_2O, for a total of 12 O atoms. On the product side, there are 4 O atoms from the 2 MnO_2 formula units, 6 O atoms from 3 $HOCH_2CH_2OH$, and 2 O atoms from the 2 KOH formula units. This is a total of 12 O atoms in the products, so the O atoms are balanced. All the elements are balanced, so the entire chemical equation is balanced.

Answer b:

Because the balanced chemical equation tells us that 2 $KMnO_4$ formula units react with 3 $CH_2{=}CH_2$ molecules, we have the following 2 ratios: $\mathbf{\frac{2KMnO_4}{3CH_2{=}CH_2}}$ **and** $\mathbf{\frac{3CH_2{=}CH_2}{2KMnO_4}}$. There are a total of 30 possible ratios. Can you find the other 28 relationships?

Key Takeaway

The coefficients in a balanced chemical equation give the ratios in which molecules of substances react and are produced in a chemical reaction.

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5.5: Mole-Mole Relationships in Chemical Reactions

Learning Objectives

- To use a balanced chemical reaction to determine molar relationships between the substances.

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). Here, we will extend the meaning of the coefficients in a chemical equation.

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.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O$$

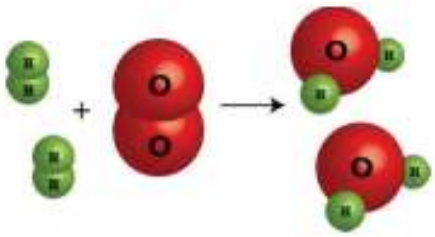
These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ \text{mol}\ \ H_2 + 1\ \ \text{mol}\ \ O_2 \longrightarrow 2\ \text{mol}\ \ H_2O$$

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

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

Now 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."



2 molecules H_2 1 molecule O_2 2 molecules H_2O

2 moles H_2 1 mole O_2 2 moles H_2O

2 x 2.02 g=4.04 g H_2 32.0 g O_2 2 x 18.02 g=36.04 g H_2O

Figure 5.5.1: This representation of the production of water from oxygen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

By the same token, 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 amount of a substance, in moles, 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.**

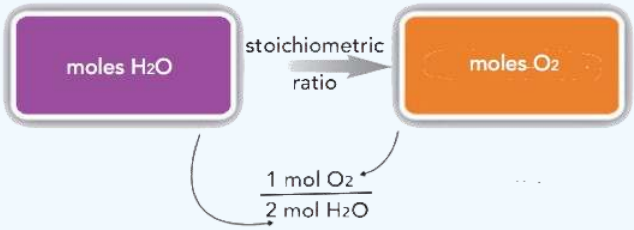
Example 5.5.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O? The balanced equation is as follows:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Solution

Because we are dealing with quantities of H_2O and O_2, we will use the **stoichiometric ratio** that relates those two substances. Because we are given an amount of H_2O and want to determine an amount of O_2, we will use the ratio that has H_2O in the denominator (so it cancels) and O_2 in the numerator (so it is introduced in the answer). Thus,



$$27.6\ \text{mol}\ H_2O \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} = 13.8\ \text{mol}\ O_2$$

To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.

Exercise 5.5.1

Using $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of hydrogen react with 3.07 mol of oxygen to produce H_2O?

Answer

$$3.07\ \text{mol}\ O_2 \times \frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2} = 6.14\ \text{mol}\ H_2$$

Key Takeaway

- The balanced chemical reaction can be used to determine molar relationships between substances.

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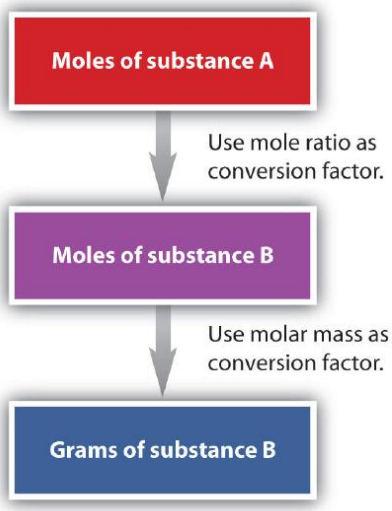
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5.6: Mole-Mass and Mass-Mass Problems

Learning Objectives

- To convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

We have established that a balanced chemical equation is balanced in terms of moles as well as atoms or molecules. We have used balanced equations to set up ratios, now in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions, such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that ability to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations.

As an example, consider the balanced chemical equation

$$\mathrm{Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3}$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of 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 SO_3 needed. Then using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

The first step resembles the exercises we did in Section 6.4. As usual, we start with the quantity we were given:

$$\mathrm{3.59\ mol\ Fe_2O_3 \times \frac{3\ mol\ SO_3}{1\ mol\ Fe_2O_3} = 10.77\ mol\ SO_3}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$\mathrm{10.77\ mol\ SO_3 \times \frac{80.07\ g\ SO_3}{1\ mol\ SO_3} = 862.4\ g\ SO_3}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of 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:



$$3.59 \text{ mol } Fe_2O_3 \times \underbrace{\frac{3 \text{ mol } SO_3}{1 \text{ mol } Fe_2O_3}}_{\text{converts to moles of } SO_3} \times \underbrace{\frac{80.07 \text{ g } SO_3}{1 \text{ mol } SO_3}}_{\text{converts to grams of } SO_3} = 862.4 \text{ g } SO_3$$

We get exactly the same answer when combining all the math steps together as we do when we calculate one step at a time.

Example 5.6.1

How many grams of CO_2 are produced if 2.09 mol of HCl are reacted according to this balanced chemical equation?

$$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

Solution

Our strategy will be to convert from moles of HCl to moles of CO_2 and then from moles of CO_2 to grams of CO_2. We will need the molar mass of CO_2, which is 44.01 g/mol. Performing these two conversions in a single-line gives 46.0 g of CO_2:

$$2.09 \text{ mol } HCl \times \underbrace{\frac{1 \text{ mol } CO_2}{2 \text{ mol } HCl}}_{\text{from the coefficients of the balanced equation}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 46.0 \text{ g } CO_2$$

The molar ratio between CO_2 and HCl comes from the balanced chemical equation.

Exercise

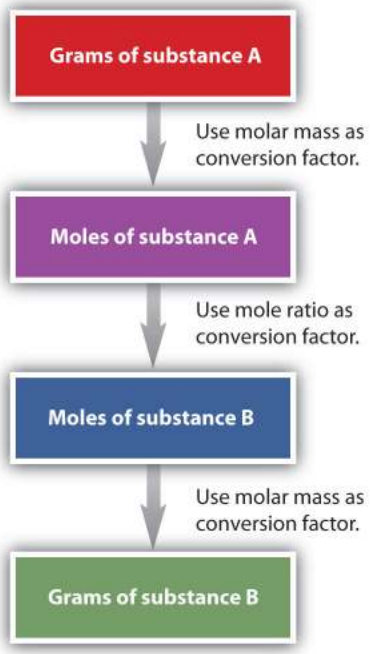
How many grams of glucose ($C_6H_{12}O_6$) are produced if 17.3 mol of H_2O are reacted according to this balanced chemical equation?

$$6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2$$

Answer

$$17.3 \text{ mol } H_2O \times \frac{1 \text{ mol } C_6H_{12}O_6}{6 \text{ mol } H_2O} \times \frac{180.18 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6} = 520 \text{ g } C_6H_{12}O_6$$

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

Example 5.6.2: Chlorination of Carbon

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows:

$$CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl$$

How many grams of HCl are produced by the reaction of 100.0 g of CH_4?

Solution

First, let us work the problem in stepwise fashion. We begin by converting the mass of CH_4 to moles of CH_4, using the molar mass of CH_4 (16.05 g/mol) as the conversion factor:

$$100.0\ g\ CH_4 \times \frac{1\ mol\ CH_4}{16.05\ g\ CH_4} = 6.231\ mol\ CH_4$$

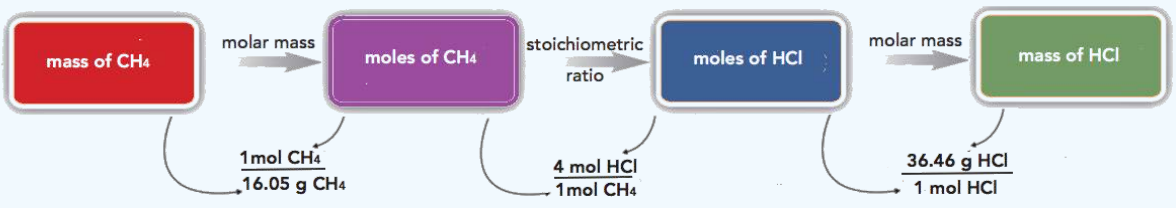
Note that we inverted the molar mass so that the gram units cancel, giving us an answer in moles. Next, we use the balanced chemical equation to determine the ratio of moles CCl_4 and moles HCl and convert our first result into moles of HCl:

$$6.231\ mol\ CH_4 \times \frac{4\ mol\ HCl}{1\ mol\ CH_4} = 24.92\ mol\ HCl$$

Finally, we use the molar mass of HCl (36.46 g/mol) as a conversion factor to calculate the mass of 24.92 mol of HCl:

$$24.92\ mol\ HCl \times \frac{36.46\ g\ HCl}{1\ mol\ HCl} = 908.5\ g\ HCl$$

In each step, we have limited the answer to the proper number of significant figures. If desired, we can do all three conversions on a single line:



$$100.0\ \text{g CH}_4 \times \frac{1\ \text{mol CH}_4}{16.05\ \text{g CH}_4} \times \frac{4\ \text{mol HCl}}{1\ \text{mol CH}_4} \times \frac{36.46\ \text{g HCl}}{1\ \text{mol HCl}} = 908.7\ \text{g HCl}$$

This final answer is slightly different from our first answer because only the final answer is restricted to the proper number of significant figures. In the first answer, we limited each intermediate quantity to the proper number of significant figures. As you can see, both answers are essentially the same.

? Exercise 5.6.2: Oxidation of Propanal

The oxidation of propanal (CH_3CH_2CHO) to propionic acid (CH_3CH_2COOH) has the following chemical equation:

$$CH_3CH_2CHO + 2K_2Cr_2O_7 \rightarrow CH_3CH_2COOH + \text{other products}$$

How many grams of propionic acid are produced by the reaction of 135.8 g of $K_2Cr_2O_7$?

Answer

$$135.8\ \text{g K}_2\text{Cr}_2\text{O}_7 \times \frac{1\ \text{mol K}_2\text{Cr}_2\text{O}_7}{294.20\ \text{g K}_2\text{Cr}_2\text{O}_7} \times \frac{1\ \text{mol CH}_3\text{CH}_2\text{COOH}}{2\ \text{mol K}_2\text{Cr}_2\text{O}_7} \times \frac{74.09\ \text{g CH}_3\text{CH}_2\text{COOH}}{1\ \text{mol CH}_3\text{CH}_2\text{COOH}} = 17.10\ \text{g CH}_3\text{CH}_2\text{COOH}$$

To Your Health: The Synthesis of Taxol

Taxol is a powerful anticancer drug that was originally extracted from the Pacific yew tree (*Taxus brevifolia*). As you can see from the accompanying figure, taxol is a very complicated molecule, with a molecular formula of $\text{C}_{47}\text{H}_{51}\text{NO}_{14}$. Isolating taxol from its natural source presents certain challenges, mainly that the Pacific yew is a slow-growing tree, and the equivalent of six trees must be harvested to provide enough taxol to treat a single patient. Although related species of yew trees also produce taxol in small amounts, there is significant interest in synthesizing this complex molecule in the laboratory.

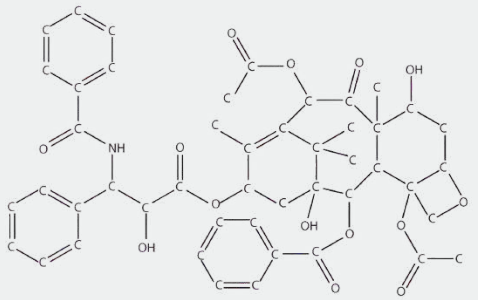


Figure 5.6.1 The Structure of the Cancer Drug Taxol. Because of the complexity of the molecule, hydrogen atoms are not shown, but they are present on every atom to give the atom the correct number of covalent bonds (four bonds for each carbon atom).

After a 20-year effort, two research groups announced the complete laboratory synthesis of taxol in 1994. However, each synthesis required over 30 separate chemical reactions, with an overall efficiency of less than 0.05%. To put this in perspective, to obtain a single 300 mg dose of taxol, you would have to begin with 600 g of starting material. To treat the 26,000 women who are diagnosed with ovarian cancer each year with one dose, almost 16,000 kg (over 17 tons) of starting material must be converted to taxol. Taxol is also used to treat breast cancer, with which 200,000 women in the United States are diagnosed every year. This only increases the amount of starting material needed.

Clearly, there is intense interest in increasing the overall efficiency of the taxol synthesis. An improved synthesis not only will be easier but also will produce less waste materials, which will allow more people to take advantage of this potentially life-saving drug.

Key Takeaway

- A balanced chemical equation can be used to relate masses or moles of different substances in a reaction.

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CHAPTER OVERVIEW

6: Chemical Energy and Reaction Types

6.1: Chemical Potential Energy
6.2: Heat
6.3: Exothermic and Endothermic Processes
6.4: Chemical Reaction Rate
6.5: Activation Energy
6.6: Potential Energy Diagrams
6.7: Factors Affecting Reaction Rate
6.8: Catalysts
6.9: Some Types of Chemical Reactions
6.10: Oxidation-Reduction (Redox) Reactions
6.11: Redox Reactions in Organic Chemistry and Biochemistry
6.12: Single Replacement Reactions
6.13: Double Replacement Reactions
6.14: Chapter 6 Exercises
6.14.1: Chemical Energy (Exercises)
6.14.2: Types of Reactions (Exercises)
6.14.3: Answers Types of Reactions (Exercises)
6.14.4: More Practice with Reactions (Exercises)

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6.1: Chemical Potential Energy

Gunpowder was originally developed by the Chinese in the ninth century AD, primarily for rockets. This material is composed of charcoal, sulfur, and saltpeter (potassium nitrate). The explosive reaction that occurs involves the conversion of the charcoal to carbon dioxide, with the potassium nitrate providing the extra oxygen needed for a rapid reaction. Although sulfur was included to stabilize the product, gunpowder is still highly explosive.

Types of Energy

Two basic types of energy exist: potential energy and kinetic energy. **Potential energy** is stored energy. It has not yet been released, but is ready to go. **Kinetic energy** is the energy of motion. It causes work to be done through movement.

Chemical Potential Energy

Energy is the capacity for doing work or supplying heat. When you fill your car with gasoline, you are providing it with potential energy. **Chemical potential energy** is the energy stored in the chemical bonds of a substance. The various chemicals that make up gasoline contain a large amount of chemical potential energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things: some of the potential energy is transformed into work, which is used to move the car; at the same time, some of the potential energy is converted to heat and makes the car's engine very hot. The energy changes of a system occur as either heat or work, or some combination of both.



Figure **6.1.1**: A dragster is able to accelerate because of the chemical potential energy of its fuel. The burning of the fuel also produces large amounts of heat. (CC BY-NC; CK-12)

Dynamite is another example of chemical potential energy. The major component of dynamite is nitroglycerin, a very unstable material. By mixing it with diatomaceous earth, the stability is increased and it is less likely to explode if it receives a physical shock. When ignited, the nitroglycerin explodes rapidly, releasing large amounts of nitrogen and other gases along with a massive amount of heat.



Figure **6.1.2**: Dynamite explosion. (CC BY-NC; CK-12)

Summary

- Potential energy is stored energy.
- Kinetic energy is the energy of motion.
- Chemical potential energy is energy available in the chemical bonds of a compound.

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6.2: Heat

Blacksmiths heat solid iron in order to shape it into a variety of different objects. Iron is a rigid, solid metal. At room temperature, it is extremely difficult to bend iron. However, when heated to a high enough temperature, iron can be easily worked. The heat energy in the forge is transferred to the metal, making the iron atoms vibrate more and move around more readily.

Heat is energy that is transferred from one object or substance to another because of a difference in temperature between the two. Heat always flows from an object at a higher temperature to an object at a lower temperature (see figure below). The flow of heat will continue until the two objects are at the same temperature.

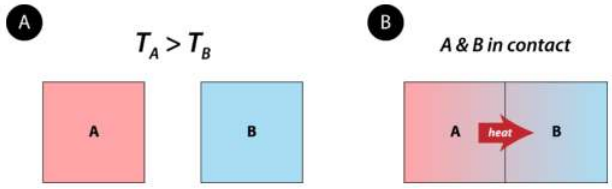


***Figure* 6.2.1:** *Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.*

Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state. When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Summary

- Heat is transferred energy from a site of higher energy to a site of lower energy.
- Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state.

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6.3: Exothermic and Endothermic Processes

A campfire is an example of basic thermochemistry. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, roast marshmallows, or simply to keep warm when it's cold outside.

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe: *the system* and *the surroundings*. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** is everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings is the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings, and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

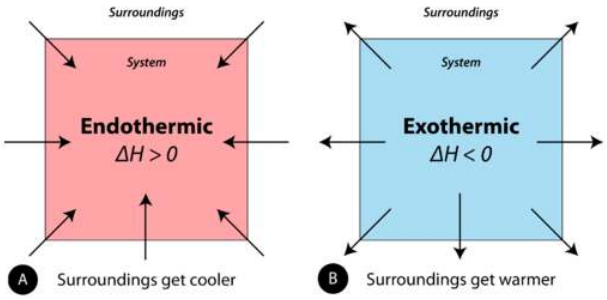


Figure 6.3.1: (A) Endothermic reaction. (B) Exothermic reaction. (CC BY-NC; CK-12)

Units of Heat

Heat flow is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by $1°C$. For example, raising the temperature of $100\ g$ of water from $20°C$ to $22°C$ would require $100 \times 2 = 200\ cal$.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains $85\ kcal$ or $85,000\ cal$. In order to make the distinction, the dietary calorie is written with a capital C.

$$1\ kilocalorie = 1\ Calorie = 1000\ calories$$

To say that the snack "contains" 85 Calories means that $85\ kcal$ of energy are released when that snack is processed by the human body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1\ J = 0.2390\ cal \text{ or } 1\ cal = 4.184\ J$$

We can calculate the amount of heat released in kilojoules when a 400 Calorie hamburger is digested:

$$400.\ \text{Cal} = 400.\ \text{kcal} \times \frac{4.184\ \text{kJ}}{1\ \text{kcal}} = 1.67 \times 10^3\ \text{kJ}$$

Summary

- The law of conservation of energy states that in any physical or chemical process, energy is neither created nor destroyed.
- A specific portion of matter in a given space that is being studied during an experiment or an observation is the *system*.
- The *surroundings* is everything in the universe that is not part of the system.
- A chemical reaction or physical change is endothermic if heat is absorbed by the system from the surroundings.
- A reaction or change is exothermic if heat is released by the system into the surroundings.

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6.4: Chemical Reaction Rate

Drag racing is a sport that involves two cars starting from a dead stop, and driving as fast as they can down a quarter-mile strip. At the end of the strip are timers that determine both elapsed time (how long it took for the cars to cover the quarter mile) and top speed (how fast they were going as they went through the timer chute). Both pieces of data are important. One car may accelerate faster and get ahead that way, while the other car may be slower off the line, but can get up to a higher top speed at the end of the run.

Chemical Reaction Rate

Chemical reactions vary widely in terms of the speed with which they occur. Some reactions occur very quickly. If a lit match is brought in contact with lighter fluid or another flammable liquid, it erupts into flames instantly and burns fast. Other reactions occur very slowly. A container of milk in the refrigerator will be good to drink for weeks before it begins to turn sour. Millions of years were required for dead plants under Earth's surface to accumulate and eventually turn into fossil fuels such as coal and oil.

Chemists need to be concerned with the rates at which chemical reactions occur. Rate is another word for speed. If a sprinter takes $11.0\ \text{s}$ to run a $100\ \text{m}$ dash, his rate is given by the distance traveled divided by the time.

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100\ \text{m}}{11.0\ \text{s}} = 9.09\ \text{m/s}$$

The sprinter's average running rate for the race is $9.09\ \text{m/s}$. We say that it is his average rate because he did not run at that speed for the entire race. At the very beginning of the race, while coming from a standstill, his rate must be slower until he is able to get up to his top speed. His top speed must then be greater than $9.09\ \text{m/s}$ so that, taken over the entire race, the average ends up at $9.09\ \text{m/s}$.



Figure 6.4.1: Runner. (CC BY-NC; CK-12)

Chemical reactions can't be measured in units of meters per second, as that would not make any sense. A **reaction rate** is the change in concentration of a reactant or product with time. Suppose that a simple reaction were to take place in which a $1.00\ \text{M}$ aqueous solution of substance A was converted to substance B.

$$\text{A}\,(aq) \rightarrow \text{B}\,(aq)$$

Suppose that after 20.0 seconds, the concentration of A had dropped from $1.00\ \text{M}$ to $0.72\ \text{M}$ as A was slowly being converted to B. We can express the rate of this reaction as the change in concentration of A divided by time.

$$\text{rate} = -\frac{\Delta\,[\text{A}]}{\Delta t} = -\frac{[\text{A}]_{\text{final}} - [\text{A}]_{\text{initial}}}{\Delta t}$$

A bracket around a symbol or formula means the concentration in molarity of that substance. The change in concentration of A is its final concentration minus its initial concentration. Because the concentration of A is decreasing over time, the negative sign is used. Thus, the rate for the reaction is positive and the units are molarity per second or M/s.

$$\text{rate} = -\frac{0.72\ \text{M} - 1.00\ \text{M}}{20.0\ \text{s}} = -\frac{-0.28\ \text{M}}{20.0\ \text{s}} = 0.014\ \text{M/s}$$

The molarity of A decreases by an average rate of $0.014\ \text{M}$ every second. In summary, the rate of a chemical reaction is measured by the change in concentration over time for a reactant or product. The unit of measurement for a reaction rate is molarity per second (M/s).

Summary

- Chemists need to be concerned with the rates at which chemical reactions occur.
- The reaction rate indicates how fast the reaction proceeds.

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6.5: Activation Energy

The sight of fireworks cascading across the night sky is a hallmark of special occasions. These materials, invented hundreds of years ago, can be dangerous if not handled properly. The chemicals do not react until the fuse burns down and heat is applied to the system. Then, the rocket is launched and explodes high in the sky.

Activation Energy



Figure **6.5.1**: Calcium metal stored in an argon atmosphere. (CC BY-NC; CK-12)

Why do some chemical reactions occur readily while others require input of heat in order to take place? If we mix metallic sodium with water, a reaction occurs immediately, releasing a great deal of heat (enough to ignite the hydrogen gas that is formed). Group II metals, such as calcium, react at a much slower rate. Unlike the extremely vigorous reaction with sodium, the reaction with calcium is slow enough that we can trap the hydrogen gas released.

Supplying reactant particles with energy causes the bonds between the atoms to vibrate with a greater frequency. This increase in vibrational energy makes a chemical bond more likely to break, and a chemical reaction more likely to occur, when those particles collide with other particles. The **activation energy** for a reaction is the minimum energy that colliding particles must have in order to undergo a reaction. Some reactions occur readily at room temperature because the reacting particles already have the requisite activation energy at that temperature. Other reactions only occur when heated because the particles do not have enough energy unless an external source of heat provides the particles with more kinetic energy.

Summary

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6.6: Potential Energy Diagrams

Sisyphus was a mythological being who was a very evil king. As a punishment for his misdoings, he was supposed to roll a large stone up to the top of a long hill. A spell had been placed on the stone so that it would roll back down before reaching the top, never to complete the task. Sisyphus was condemned to an eternity of trying to get to the top of the hill, but never succeeding.

Potential Energy Diagrams

The energy changes that occur during a chemical reaction can be shown in a diagram called a potential energy diagram, or sometimes called a reaction progress curve. A **potential energy diagram** shows the change in potential energy of a system as reactants are converted into products. The figure below shows basic potential energy diagrams for an endothermic (A) and an exothermic (B) reaction. Recall that the enthalpy change (ΔH) is positive for an endothermic reaction and negative for an exothermic reaction. This can be seen in the potential energy diagrams. The total potential energy of the system increases for the endothermic reaction as the system absorbs energy from the surroundings. The total potential energy of the system decreases for the exothermic reaction as the system releases energy to the surroundings.

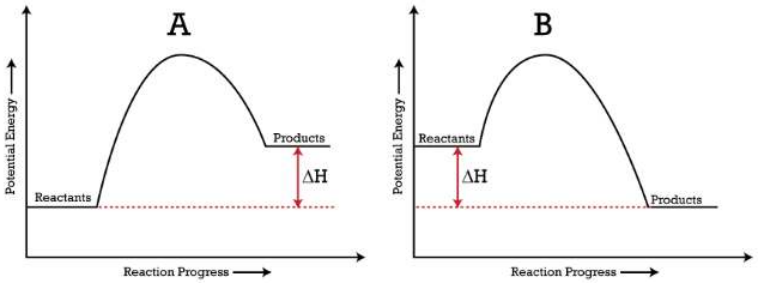


Figure **6.6.1**: A potential energy diagram shows the total potential energy of a reacting system as the reaction proceeds. (A) In an endothermic reaction, the energy of the products is greater than the energy of the reactants and ΔH is positive. (B) In an exothermic reaction, the energy of the products is lower than the energy of the reactants and ΔH is negative. (CC BY-NC; CK-12)

The activation energy for a reaction is illustrated in the potential energy diagram by the height of the hill between the reactants and the products. For this reason, the activation energy of a reaction is sometimes referred to as the activation energy barrier. Reacting particles must have enough energy so that when they collide, they can overcome that barrier (see figure below).

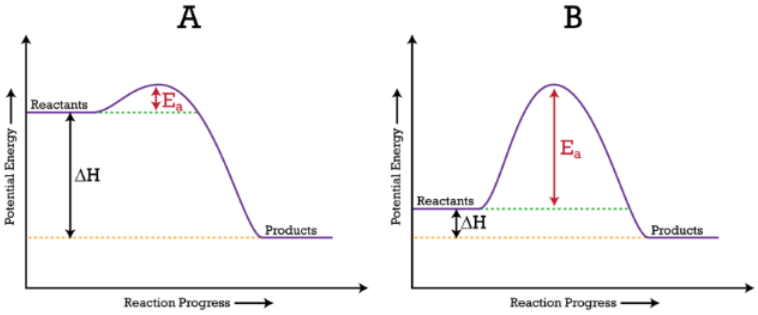


Figure **6.6.2**: The activation energy (E_a) of a reaction is the barrier that must be overcome for the reactants to be able to become products. (A) The activation energy is low, meaning that the reaction is likely to be fast. (B) The activation energy is high, meaning that the reaction is likely to be slow. (CC BY-NC; CK-12)

Summary

- A potential energy diagram shows the change in potential energy of a system as reactants are converted into products.
- Potential energy diagrams for endothermic and exothermic reactions are described.
- Diagrams of activation energy and reaction progress are given.

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6.7: Factors Affecting Reaction Rate

Driving on a crowded freeway can be stressful. Lots of cars, drivers who aren't paying attention, people who speed or who drive too slow—the chances of a collision are rather high. A lot of cars in a particular amount of space means a high car concentration and many opportunities for unwanted connections with other cars.

Factors Affecting Reaction Rates

According to their nature, some reactions occur very quickly, while others are very slow. However, certain changes in the reacting conditions can have an effect on the rate of a given chemical reaction. Collision theory can be utilized to explain these rate effects.

Concentration

An increase in the concentration of one or more reacting substances results in an increase in the rate of reaction. When more particles are present in a given amount of space, a greater number of collisions will naturally occur between those particles. Since the rate of a reaction is dependent on the number of collisions occurring between reactants, the rate increases as the concentration increases.

Pressure

When the pressure of a gas is increased, its particles are forced closer together, decreasing the amount of empty space between the particles. Therefore, an increase in the pressure of a gas is also an increase in the concentration of the gas. For gaseous reactions, an increase in pressure increases the rate of reaction due to a greater number of collisions between reacting particles.

Surface Area

A large log placed in a fire will burn relatively slowly. If the same mass of wood were added to the fire in the form of small twigs, the twigs would burn much more quickly. This is because the twigs provide a greater surface area than the log does. An increase in the surface area of a reactant increases the rate of a reaction. Surface area is larger when a given amount of a solid is present as smaller particles. A powdered reactant has a greater surface area than the same reactant as a solid chunk. In order to increase the surface area of a substance, it may be ground into smaller particles or dissolved into a liquid. In solution, the dissolved particles are separated from each other and will react more quickly with other reactants.

Temperature

Raising the temperature of a chemical reaction usually results in a higher rate of reaction. When the reactant particles are heated, they move faster and faster. This results in a greater frequency of collisions. A more important effect of the temperature increase is that the collisions occur with a greater force, and are thus more likely to surmount the activation energy barrier and go on to form products. Increasing the temperature of a reaction increases the number of effective collisions between reacting particles, so the reaction rate increases.

Summary

- Factors that affect (and generally increase) reaction rate are:
  - Concentration of reactants
  - Pressure (in the case of a gas)
  - Surface area
  - Temperature

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6.8: Catalysts

Gasoline-powered vehicles emit various harmful materials. Nitrogen oxides are formed when atmospheric nitrogen reacts with oxygen at the high temperatures found in a car engine. Carbon monoxide is a by-product of the incomplete combustion of hydrocarbons. Evaporated and unused fuel releases volatile hydrocarbons into the atmosphere to help form smog. The presence of a catalytic converter in the exhaust system of the car causes these materials to react and be changed into less harmful products.

Catalysts

Sometimes a particular substance added to a chemical reaction will cause that reaction to undergo a dramatic increase in rate. Hydrogen peroxide is used as a disinfectant for scrapes and cuts, and is found in many medicine cabinets as a 3% aqueous solution. Hydrogen peroxide naturally decomposes to produce water and oxygen gas, but the reaction is very slow. A bottle of hydrogen peroxide will last for several years before it needs to be replaced. However, the addition of just a small amount of manganese (IV) oxide to hydrogen peroxide will cause it to decompose completely in just a matter of minutes. A **catalyst** is a substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction. After the reaction occurs, a catalyst returns to its original state; so catalysts can be used over and over again. Because it is neither a reactant nor a product, a catalyst is shown in a chemical equation by being written above the yield arrow.

$$\mathrm{H_2O_2\,(aq)} \xrightarrow{\mathrm{MnO_2}} \mathrm{2H_2O\,(l) + O_2\,(g)}$$

A catalyst works by changing the specific way in which the reaction occurs, called its mechanism. The important outcome from the use of a catalyst is that the overall activation energy of the reaction is lowered (see figure below). With a lower activation energy barrier, a greater percentage of reactant molecules are able to have effective collisions, and the reaction rate increases.

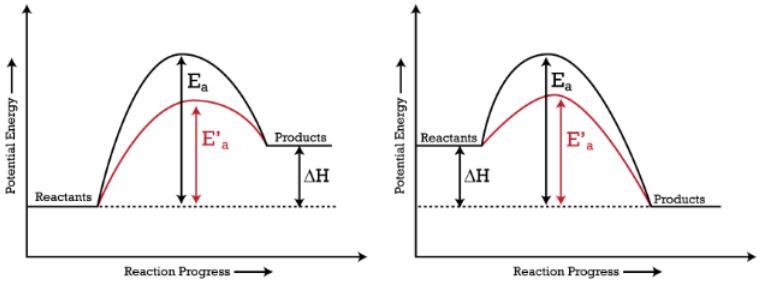


Figure **6.8.1**: The addition of a catalyst to a reaction lowers the activation energy, increasing the rate of the reaction. The activation energy of the uncatalyzed reaction is shown by E_a, while the catalyzed reaction is shown by E'_a. The heat of reaction (ΔH) is unchanged by the presence of the catalyst. (CC BY-NC; CK-12)

Catalysts are an extremely important part of many chemical reactions. Enzymes in your body acts as nature's catalysts, allowing important biochemical reactions to occur at reasonable rates. Chemical companies constantly search for new and better catalysts to make reactions go faster and thus make the company more profitable.

Summary

- A catalyst is a substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction.
- Catalysts are an important part of many chemical reactions.

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6.9: Some Types of Chemical Reactions

Learning Objectives

- To classify a given chemical reaction into a variety of types.

Although there are untold millions of possible chemical reactions, most can be classified into a small number of general reaction types. Classifying reactions has two purposes: it helps us to recognize similarities among them, and it enables us to predict the products of certain reactions. A particular reaction may fall into more than one of the categories that we will define in this book.

Combination (composition) Reactions

A combination (composition) reaction is a chemical reaction that makes a single substance from two or more reactants. There may be more than one molecule of product in the balanced chemical equation, but there is only one substance produced. For example, the equation

$$4\,Fe + 3\,O_2 \rightarrow 2\,Fe_2O_3$$

is a combination reaction that produces Fe_2O_3 from its constituent elements — Fe and O_2. Combination reactions do not have to combine elements, however. The chemical equation

$$Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3$$

shows a combination reaction in which Fe_2O_3 combines with three molecules of SO_3 to make $Fe_2(SO_4)_3$.

Example 6.9.1

Which equations are combination reactions?

a. $Co\,(s) + Cl_2\,(g) \longrightarrow CoCl_2\,(s)$
b. $CO\,(g) + Cl_2\,(g) \longrightarrow COCl_2\,(g)$
c. $N_2H_4(\ell) + O_2\,(g) \longrightarrow N_2\,(g) + 2\,H_2O(\ell)$

Solution

a. This is a combination reaction.
b. This is a combination reaction. (The compound $COCl_2$ is called phosgene and, in the past, was used as a gassing agent in chemical warfare.)
c. This is not a combination reaction.

Exercise 6.9.1

Which equations are combination reactions?

a. $P_4\,(s) + 6\,Cl_2\,(g) \longrightarrow 4\,PCl_3\,(g)$
b. $SO_3(\ell) + H_2O(\ell) \longrightarrow H_2SO_4(\ell)$
c. $NaOH\,(s) + HCl\,(g) \longrightarrow NaCl\,(s) + H_2O(\ell)$

Answer

a and b are combination reactions

Decomposition Reactions

A decomposition reaction is the reverse of a combination reaction. In a decomposition reaction, a single substance is converted into two or more products. There may be more than one molecule of the reactant, but there is only one substance initially. For example, the equation

$$2\,NaHCO_3\,(s) \rightarrow Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(\ell)$$

is a decomposition reaction that occurs when $NaHCO_3$ is exposed to heat. Another example is the decomposition of $KClO_3$:

$$2\,KClO_3\,(s) \rightarrow 2\,KCl\,(s) + 3\,O_2\,(g)$$

This reaction was once commonly used to generate small amounts of oxygen in the chemistry lab.

> The decomposition reaction of $NaHCO_3$ is the reaction that occurs when baking soda is poured on a small kitchen fire. The intent is that the H_2O and CO_2 produced by the decomposition will smother the flames.

Combustion Reactions

A combustion reaction occurs when a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction. Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amounts of heat energy. An example is propane (C_3H_8), a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$C_3H_8\,(g) + 5\,O_2\,(g) \longrightarrow 3\,CO_2\,(g) + 4\,H_2O\,(g)$$

Another example is the burning of acetylene (C_2H_2) in welding torches:

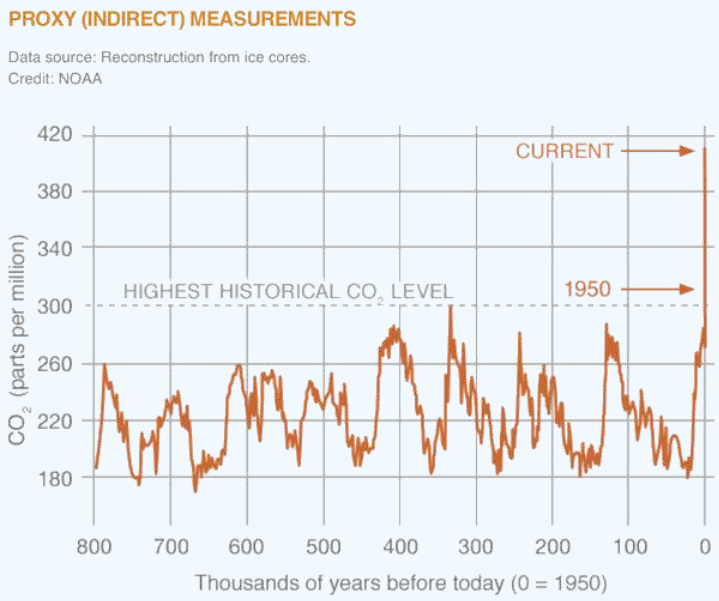
$$2\,C_2H_2 + 5\,O_2 \rightarrow 4\,CO_2 + 2\,H_2O$$

Oxygen (in its elemental form) is a crucial reactant in combustion reactions.

> *Energy in the form of heat is usually given off as a product in a combustion reaction as well.*

Example 6.9.2

Carbon dioxide (CO_2) is an important heat-trapping (greenhouse) gas, which is released through human activities such as deforestation and burning fossil fuels, as well as natural processes such as respiration and volcanic eruptions. The graph above shows CO_2 levels during the last three glacial cycles, as reconstructed from ice cores.



Carbon dioxide is the primary greenhouse gas emitted through human activities. In 2015, CO_2 accounted for about 82.2% of all U.S. greenhouse gas emissions from human activities. Carbon dioxide is naturally present in the atmosphere as part of the Earth's carbon cycle (the natural circulation of carbon among the atmosphere, oceans, soil, plants, and animals). Human activities are altering the carbon cycle–both by adding more CO_2 to the atmosphere and by influencing the ability of natural sinks, like forests, to remove CO_2 from the atmosphere. While CO_2 emissions come from a variety of natural sources, human-related emissions are responsible for the increase that has occurred in the atmosphere since the industrial revolution.

The main human activity that emits CO_2 is the combustion of fossil fuels (coal, natural gas, and oil) for energy and transportation, although certain industrial processes and land-use changes also emit CO_2. An example of how CO_2 can be generated is the **combustion of octane** (C_8H_{10}), a component of gasoline.

Write the balanced equation to represent the combustion of octane.

Solution

$$2\,C_8H_{18}\,(l) + 25\,O_2\,(g) \longrightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)$$

The balanced reaction argues that for every two molecules of octane that are burned, 16 molecules of CO_2 are generated.

Example 6.9.2

Identify each type of reaction.

a. $2\,K\,(s) + S\,(s) + 2\,O_2\,(g) \longrightarrow K_2SO_4\,(s)$
b. $(NH_4)_2Cr_2O_7\,(s) \longrightarrow N_2\,(g) + Cr_2O_3\,(s) + 4\,H_2O(\ell)$
c. $CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O(\ell)$

Solution

a. Multiple reactants are combining to make a single product, so this reaction is a combination reaction.
b. A single substance reacts to make several products, so we have a decomposition reaction.
c. Oxygen reacts with a compound to make carbon dioxide (an oxide of carbon) and water (an oxide of hydrogen). This is a combustion reaction.

Exercise 6.9.2

Identify each type of reaction.

a. $C_2H_5OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$
b. $2\,Mg\,(s) + O_2\,(g) \longrightarrow 2\,MgO\,(s)$
c. $CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(g)$

Answer a

combustion

Answer b

combination (also combustion)

Answer c

decomposition

Key Takeaway

There are several recognizable types of chemical reactions: combination, decomposition, and combustion reactions are examples.

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6.10: Oxidation-Reduction (Redox) Reactions

Learning Objectives

- To identify a chemical reaction as an oxidation-reduction reaction.

When zinc metal is submerged into a quantity of aqueous $\mathbf{HCl}$, the following reaction occurs (Figure $\mathbf{6.10.1}$):

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow H_2\,(g) + ZnCl_2\,(aq)} \tag{6.10.1}$$

This is one example of what is sometimes called a *single replacement reaction* because $\mathbf{Zn}$ replaces $\mathbf{H}$ in combination with $\mathbf{Cl}$.



Figure $\mathbf{6.10.1}$: Zinc Metal plus Hydrochloric Acid. It is fairly obvious that zinc metal reacts with aqueous hydrochloric acid! The bubbles are hydrogen gas (right side of Equation 6.10.1).

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow H_2\,(g) + Zn^{2+}\,(aq) + 2\,Cl^-\,(aq)}$$

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**.

> *Because batteries are used as sources of electricity (that is, of electrons), all batteries are based on redox reactions.*

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, in our earlier equation, now written without the chloride ions,

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)}$$

zinc atoms are oxidized to Zn^{2+}. The **half reaction for the oxidation** reaction, omitting phase labels, is as follows:

$$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-}$$

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. If you check each side of this reaction, you will note that both sides have a zero net charge.

Hydrogen is reduced in the reaction. The balanced **reduction half reaction** is as follows:

$$2\,H^+ + 2\,e^- \longrightarrow H_2$$

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 is simply the combination of the two half reactions and is shown by adding them together.

$$\begin{aligned} &Zn \rightarrow Zn^{2+} + \cancel{2e^-} \\ &\underline{2H^+ + \cancel{2e^-} \rightarrow H_2} \\ &Zn + 2H^+ \rightarrow Zn^{2+} + H_2 \end{aligned}$$

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 6.10.1: 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:

$$Ag^+ + Al \longrightarrow Ag + Al^{3+}$$

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:

$$Ag^+ + e^- \longrightarrow Ag$$

Aluminum is oxidized, losing three electrons to change from Al to Al^{3+}:

Oxidation half-reaction:

$$Al \longrightarrow Al^{3+} + 3\,e^-$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{aligned} &3\left(Ag^+ + e^- \rightarrow Ag\right) \\ &\underline{\qquad Al \rightarrow Al^{3+} + 3e^-} \\ &3Ag^+ + Al \rightarrow 3Ag + Al^{3+} \end{aligned}$$

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: $\mathbf{Ag^+}$

Exercise 6.10.1

Write and balance the redox reaction that has calcium ions and potassium metal as reactants and calcium metal and potassium ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Answer

Reduction: $\mathbf{Ca^{2+} + 2\,e^- \longrightarrow Ca}$

Oxidation: \(2 \times (\ce{K -> K^{+} + e^{-}})\)

Combined: $\mathbf{Ca^{2+} + 2\,K \longrightarrow Ca + 2\,K^+}$

- The **substance oxidized** is the reactant that had undergone oxidation: $\mathbf{K}$
- The **substance reduced** is the reactant that had undergone reduction: $\mathbf{Ca^{2+}}$
- The **reducing agent** is the same as the substance oxidized: $\mathbf{K}$
- The **oxidizing agent** is the same as the substance reduced: $\mathbf{Ca^{2+}}$

Potassium has been used as a reducing agent to obtain various metals in their elemental form.

To Your Health: 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:

$$\mathbf{Cd\,(s) + 2\,NiOOH\,(s) + 2\,H_2O(\ell) \longrightarrow Cd(OH)_2\,(s) + 2\,Ni(OH)_2\,(s)}$$

The cadmium is oxidized, while the nickel atoms in $\mathbf{NiOOH}$ 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:

$$\mathbf{2\,Li\,(s) + I_2\,(s) \longrightarrow 2\,LiI\,(s)}$$

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 **6.10.1**: A small button battery like this is used to power a watch, pacemaker, or calculator. (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).

$$2\,CH_3CHO + O_2 \longrightarrow 2\,CH_3COOH$$

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 shows how each process affects the change in oxygen and change in hydrogen.

Process	Change in oxygen (some reactions)	Change in hydrogen (some reactions)
Oxidation	gain	lose
Reduction	lose	gain

Example 6.10.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $N_2 \longrightarrow NH_3$
b. $CH_3CH_2OHCH_3 \longrightarrow CH_3COCH_3$
c. $HCHO \longrightarrow HCOOH$

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.

Exercise 6.10.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $CH_4 \longrightarrow CO_2 + H_2O$
b. $NO_2 \longrightarrow N_2$
c. $CH_2{=}CH_2 \longrightarrow CH_3CH_3$

Answer a:

Oxygen is being added. Oxidation is occurring.

Answer b:

Oxygen is being removed. Reduction is occurring.

Answer a:

Hydrogen is being added. Reduction is occurring.

Key Takeaway

Chemical reactions in which electrons are transferred are called oxidation-reduction, or redox, reactions. Oxidation is the loss of electrons. Reduction is the gain of electrons. Oxidation and reduction always occur together, even though they can be written as separate chemical equations.

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6.11: Redox Reactions in Organic Chemistry and Biochemistry

Learning Objectives

- To identify oxidation-reduction reactions with organic compounds.

Oxidation-reduction reactions are of central importance in organic chemistry and biochemistry. The burning of fuels that provides the energy to maintain our civilization and the metabolism of foods that furnish the energy that keeps us alive both involve redox reactions.



Figure **6.11.1**: The Burning of Natural Gas. 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-SA 2.0 Generic; Simon Law via Flickr).

All combustion reactions are also redox reactions. A typical combustion reaction is the burning of methane, the principal component of natural gas (Figure **6.11.1**).

$$\mathrm{CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O} \tag{6.11.1}$$

In respiration, the biochemical process by which the oxygen we inhale in air oxidizes 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:

$$\mathrm{C_6H_{12}O_6 + 6\,O_2 \rightarrow 6\,CO_2 + 6\,H_2O} \tag{6.11.2}$$

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:

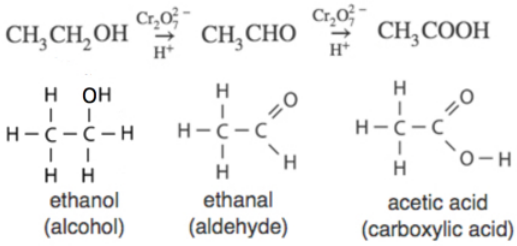
$$\mathrm{3\,CH_3CH_2OH + Cr_2O_7^{2-} + 8\,H^+ \rightarrow 3\,CH_3CHO + 2\,Cr^{3+} + 7\,H_2O} \tag{6.11.3}$$

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:

$$\mathrm{3\,CH_3CH_2OH + 2\,Cr_2O_7^{2-} + 16\,H^+ \rightarrow 3\,CH_3COOH + 4\,Cr^{3+} + 11\,H_2O} \tag{6.11.4}$$

In this reaction, the chromium atom is reduced because it lost 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 acetaldehayde (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.

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.



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 ($\mathbf{CH_3CHOHCH_3}$). It is the common solvent acetone [$\mathbf{(CH_3)_2CO}$], which is used in varnishes, lacquers, rubber cement, and nail polish remover. Acetone can be formed by the following redox reaction:

$$\mathbf{3\,CH_3CHOHCH_3 + Cr_2O_7^{2-} + 8\,H^+ \rightarrow 3\,(CH_3)_2CO + 2\,Cr^{3+} + 7\,H_2O} \tag{6.11.5}$$

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 ($\mathbf{CH_3COCOOH}$) is reduced to lactic acid ($\mathbf{CH_3CHOHCOOH}$) in the muscles.

$$\mathbf{CH_3COCOOH \rightarrow CH_3CHOHCOOH} \tag{6.11.6}$$

(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; $\mathbf{C_6H_8O_6}$) is thought to retard potentially damaging oxidation of living cells. In the process, it is oxidized to dehydroascorbic acid ($\mathbf{C_6H_6O_6}$). In the stomach, ascorbic acid reduces the nitrite ion ($\mathbf{NO_2^-}$) to nitric oxide ($\mathbf{NO}$):

$$\mathbf{C_6H_8O_6 + 2\,H^+ + 2\,NO_2^- \rightarrow C_6H_6O_6 + 2\,H_2O + 2\,NO} \tag{6.11.7}$$

> *If reaction in Equation 6.11.7 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 **6.11.2**).



Figure **6.11.2**: Citrus Fruits. Citrus fruits, such as oranges, lemons, and limes, are good sources of vitamin C, which is an antioxidant. (Pexels License; Scott Webb via Pexels)

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 ($\mathbf{C_6H_{12}O_6}$). The synthesis

of glucose requires a variety of proteins called enzymes and a green pigment called chlorophyll that converts sunlight into chemical energy (Figure **6.11.3**). The overall change that occurs is as follows:

$$6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2 \tag{6.11.8}$$

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 **6.11.3**: 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).

Summary

Redox reactions are common in organic and biological chemistry, including the combustion of organic chemicals, respiration, and photosynthesis.

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6.12: Single Replacement Reactions



Figure 6.12.1 (Pixabay License; wisteriax via Pixabay)

Why is the silver dark?

The platter and pitcher shown above provides an example of tarnish, a chemical reaction caued when silver metal reacts with hydrogen sulfide gas produced by some industrial processes or as a result of decaying animal or plant materials:

$$2Ag + H_2S \rightarrow Ag_2S + H_2$$

The tarnish can be removed using a number of polishes, but the process also removes a small amount of silver along with the tarnish.

Single-Replacement Reactions

A **single-replacement reaction** is a reaction in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$A + BC \rightarrow AC + B$$

In this general reaction, element A is a metal and replaces element B (also a metal) in the compound. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$Y + XZ \rightarrow XY + Z$$

(Where Y is a nonmetal and replaces the nonmetal Z in the compound with X.)

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$Mg\,(s) + Cu(NO_3)_2\,(aq) \rightarrow Mg(NO_3)_2\,(aq) + Cu\,(s)$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

Hydrogen Replacement

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$Zn\,(s) + 2HCl\,(aq) \rightarrow ZnCl_2\,(aq) + H_2\,(g)$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal.



Figure **6.12.2**: Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Public Domain; User: Chemicalinterest/Wikimedia Commons via Wikipedia)

Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$\mathrm{2Na}\,(s) + \mathrm{2H_2O}\,(l) \rightarrow \mathrm{2NaOH}\,(aq) + \mathrm{H_2}\,(g)$$



Figure **6.12.3**: Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite. (Public Domain; User:Ajhalls/Wikimedia Commons via Wikipedia)

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine:

$$\mathrm{Cl_2}\,(g) + \mathrm{2NaBr}\,(aq) \rightarrow \mathrm{2NaCl}\,(aq) + \mathrm{Br_2}\,(l)$$

The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.



Summary

- The activity series describes the relative reactivities of metals and halogens.

Review

1. What is a metal replacement reaction?
2. Will a non-metal replace a metal?
3. What is the most reactive halogen?
4. What products will I get if I add potassium metal to water?

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6.13: Double Replacement Reactions



Figure **6.13.1** (Public Domain; F. S. Church, published in Harper's Weekly, January 17, 1874, p. 61 via Wikipedia)

Want to trade?

The practice of barter (trading one thing for another) has been in existence for a very long time. In the illustration above, items like chickens were bartered for newspapers. You have something I want, and I have something you want. So we trade and we each have something new. Some chemical reactions are like that. Compounds swap parts and you have new materials.

Double-Replacement Reactions

A **double-replacement reaction** is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement (also called double-displacement) reaction is:

$$\mathrm{AB + CD \rightarrow AD + CB}$$

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:

$$\mathrm{2KI\,(\mathit{aq}) + Pb(NO_3)_2\,(\mathit{aq}) \rightarrow 2KNO_3\,(\mathit{aq}) + PbI_2\,(\mathit{s})}$$

There are very strong attractive forces that occur between $\mathrm{Pb^{2+}}$ and $\mathrm{I^-}$ ions and the result is a brilliant yellow precipitate (see figure below). The other product of the reaction, potassium nitrate, remains soluble.



Figure **6.13.2**: Formation of lead iodide precipitate. (CC-BY 2.0; Paige Powers - "Lead Iodide" via Flickr)

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas.

$$\mathrm{Na_2S\,(\mathit{aq}) + 2HCl\,(\mathit{aq}) \rightarrow 2NaCl\,(\mathit{aq}) + H_2S\,(\mathit{g})}$$

Formation of a Molecular Compound

Another kind of double-replacement reaction is one that produces a molecular compound as one of its products. Many examples in this category are reactions that produce water. When aqueous hydrochloric acid is reacted with aqueous sodium hydroxide, the

products are aqueous sodium chloride and water:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Example 6.13.1: Double-Replacement Reactions

Write a complete and balanced chemical equation for the following double-replacement reactions. One product is indicated as a guide.

a. $\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow$ (hydrogen cyanide gas is formed)
b. $(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow$ (a precipitate of barium sulfate forms)

Solution

Step 1: Plan the problem.

In A, the production of a gas drives the reaction. In B, the production of a precipitate drives the reaction. In both cases, use the ionic charges of both reactants to construct the correct formulas of the products.

Step 2: Solve.

A. The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partners, the balanced equation is:

$$\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow \mathrm{NaBr}\,(aq) + \mathrm{HCN}\,(g)$$

B. Ammonium ion and nitrate ion are $1+$ and $1-$ respectively, while barium and sulfate are $2+$ and $2-$. This must be taken into account when exchanging partners and writing the new formulas. Then, the equation is balanced.

$$(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow 2\mathrm{NH_4NO_3}\,(aq) + \mathrm{BaSO_4}\,(s)$$

Step 3: Think about your result.

Both are double replacement reactions. All formulas are correct and the equations are balanced.

Occasionally, a reaction will produce both a gas and a molecular compound. The reaction of a sodium carbonate solution with hydrochloric acid produces aqueous sodium chloride, carbon dioxide gas, and water.

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\ \mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$



Summary

- The double-replacement reaction generally takes the form of AB + CD → AD + CB where A and C are positively-charged cations, while B and D are negatively-charged anions.
- In a double replacement reactions, typically one of the products is a precipitate, a gas, or a molecular compound.

Review

1. In a double-replacement reaction, what type of compounds are usually the reactants? A double-replacement reaction occurs between sodium sulfide and hydrogen chloride. Write the rest of the chemical equation and balance it.
2. $Na_2S + HCl \rightarrow$
3. In double-replacement reaction, one of three possible types of products usually form. What are the three types?

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6.14: Chapter 6 Exercises

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6.14.1: Chemical Energy (Exercises)

Energy

1. Identify each example as either potential or kinetic energy.

a. A sled at the top of a hill

b. pushing a toy car

c. an airplane sitting on a runway

d. a bullet being shot from a gun

2. Chemical energy is a specific type of ____ energy.

3. Chemical energy is best described as...

a. energy stored in chemical bonds and atoms

b. the movement of electrons

c. ionic bonds that are broken when a compound is dissolved in water

d. the neutron proton ratio that exists in isotopes

4. What is the law of conservation of energy?

5. Identify each example as either endothermic or exothermic.

a. to activate a cold pack to help with inflammation you must shake it to mix the contents which in turn cools the pack.

b. striking a match

c. burning charcoal

d. baking a cake

6. A mixture of hydrogen and oxygen gases is stable at room temperature unless a match or other heat source is ignited in the room. This ignition source will produce an immediate explosion. Which term best describes the change that occurs in this scenario?

a. chemical collisions

b. kinetic energy

c. activation energy

d. chemical decomposition

7. Imagine you are burning a log (this is a chemical reaction!). Describe what changes are needed to increase the rate of reaction for each of the following factors that affect reaction rates.

a. Concentration

b. Surface Area

c. Temperature

8. True or False, the presents of a catalyst will speed up a reaction.

9. What does a catalyst do to the activation energy of a chemical reaction?

10. What are the most abundant catalyst found in the human body?

a. carbohydrates

b. enzymes

c. DNA

d. Water

Answers

1. a. potential b. kinetic c. potential d. kinetic

2. potential

3. a

4. energy cannot be created or destroyed, however it can be converted from one form to another

5. a. endo b. exo c. exo d. endo

6. c

7.

a. Concentration - add more oxygen to fuel the combustion

b. Surface Area - cut the log into smaller pieces will allow the log to burn faster

c. Temperature - if the burning log is placed in an insulated chamber then the temperature will increase, therefore, increasing the rate of reaction

Pressure is also a contributing factor to reaction rate but was not used in this example.

8. true

9. catalyst lower the activation energy

10. b

Questions provided by Kristina Novek

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6.14.2: Types of Reactions (Exercises)

These are homework exercises to accompany the Textmap created for "Chemistry" by OpenStax. Some have been modified and added by Kristina Novek.

Types of Chemical Reactions Exercises

1. Write a balanced molecular equation describing each of the following chemical reactions.

   a. Solid calcium carbonate is heated and decomposes to solid calcium oxide and carbon dioxide gas.

   b. Gaseous butane, C_4H_{10}, reacts with diatomic oxygen gas to yield gaseous carbon dioxide and water vapor.

   c. Aqueous solutions of magnesium chloride and sodium hydroxide react to produce solid magnesium hydroxide and aqueous sodium chloride.

   d. Water vapor reacts with sodium metal to produce solid sodium hydroxide and hydrogen gas.

2. Colorful fireworks often involve the decomposition of barium nitrate and potassium chlorate and the reaction of the metals magnesium, aluminum, and iron with oxygen.

   a. Write the formulas of barium nitrate and potassium chlorate.

   b. The decomposition of solid potassium chlorate leads to the formation of solid potassium chloride and diatomic oxygen gas. Write an equation for the reaction.

   c. The decomposition of solid barium nitrate leads to the formation of solid barium oxide, diatomic nitrogen gas, and diatomic oxygen gas. Write an equation for the reaction.

3. Aqueous hydrogen fluoride (hydrofluoric acid) is used to etch glass and to analyze minerals for their silicon content. Hydrogen fluoride will also react with sand (silicon dioxide).

   a. Write an equation for the reaction of solid silicon dioxide with hydrofluoric acid to yield gaseous silicon tetrafluoride and liquid water.

   b. Identify the type of reaction

4. Use the following equations to answer the questions.

I. $H_2O(s) \rightarrow H_2O(l)$

II. $CH_3OH(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

III. $2H_2O(l)$ à $2H_2(g) \rightarrow O_2(g)$

a. Which equation describes a physical change?

b. Which equation identifies the reactants and products of a combustion reaction?

c. Which equation is not balanced?

d. Which equation is a decomposition reaction?

5. Indicate what type, or types, of reaction each of the following represents:

a. $Ca(s) + Br_2(l) \rightarrow CaBr_2(s)$

b. $Ca(OH)_2(aq) + 2\ HBr(aq) \rightarrow CaBr_2(aq) + 2\ H_2O(l)$

c. $C_6H_{12}(l) + 9\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$

d. $Ca(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

6. Indicate what type, or types, of reaction each of the following represents:

a. $H_2O(g) + C(s) \rightarrow CO(g) + H_2(g)$

b. $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$

c. $Fe_2O_3 + 3\ SO_3 \rightarrow Fe_2(SO_4)_3$

d. $Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2\ HNO_3(aq)$

7. Indicate what type, or types, of reaction each of the following represents:

a. $3\ Mg(s) + 2\ AlPO_4(aq) \rightarrow 2\ Al(s) + Mg_3(PO_4)_2(aq)$

b. $CuCl_2(aq) + 2\ AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\ AgCl(s)$

c. $CaI_2(s) + Cl_2(g) \rightarrow CaCl_2(s) + 2\ I_2(s)$

d. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

e. $KBr(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgBr(s)$

8. Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing compounds are peroxides or superoxides.

a. H_3PO_4
b. $Al(OH)_3$
c. SeO_2
d. KNO_2
e. In_2S_3
f. P_4O_6

9. Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing compounds are peroxides or superoxides.

a. H_2SO_4
b. $Ca(OH)_2$
c. $BrOH$
d. $ClNO_2$
e. $TiCl_4$
f. NaH

10. In the following oxidation-reduction reactions, identify the species that are oxidized and reduced and the change in oxidation state for each:

a. $2\ Na(s) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2(g)$
b. $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
c. $K_3P(s) + 2\ O_2(g) \rightarrow K_3PO_4(s)$
d. $C_2H_4(g) + 3\ O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

11. In the following oxidation-reduction reactions, identify the species that are oxidized and reduced and the change in oxidation state for each:

a. $Mg(s) + NiCl_2(aq) \rightarrow MgCl_2(aq) + Ni(s)$
b. $PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$
c. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

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6.14.3: Answers Types of Reactions (Exercises)

Answers

1.

a. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

b. $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

c. $MgCl_2(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + 2NaCl(aq)$

d. $2H_2O(g) + 2Na(s) \rightarrow 2NaOH(s) + H_2(g)$

2.

a. $Ba(NO_3)_2$, $KClO_3$;

b. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

c. $2Ba(NO_3)_2(s) \rightarrow 2BaO(s) + 2N_2(g) + 5O_2(g)$

3.

a. $4HF(aq) + SiO_2(s) \rightarrow SiF_4(g) + 2H_2O(l)$

b. double replacement

4. a. I; b. II; c. II; d. III

5.

a. Redox, Combination

b. Double Replacement

c. Redox, Combustion

d. Redox, Single Replacement

6.

a.Redox, Single Replacement

b. Redox, Decomposition

c. Combination

d. Double Replacement

7.

a. Redox, Single Replacement

b. Double Replacement

c. Redox, Single Replacement

d. Decomposition

e. Double Replacement

8.

a. H +1, P +5, O -2

b. Al +3, H +1, O -2

c. Se +4, O -2

d. K +1, N +3, O -2

e. In +3, S -2

f. P +3, O -2

9.

a. H^{1+}, O^{2-}, S^{6+}

b. H^{1+}, O^{2-}, Ca^{+2}

c. H^{1+}, O^{2-}, Br^{1+}

d. O^{2-}, Cl^{1-}, N^{5+}

e. Cl^{1-}, Ti^{4+}

f. H^{1+}, Na^{1-}

10.

a. Na is oxidized, H^{+} is reduced

b. Mg is oxidized, Cl_2 is reduced

c. P^{3-} is oxidized, O_2 is reduced

d. C^{+2} is oxidized, O_2 is reduced

11.

a. Mg is oxidized, Ni^{+2} is reduced

b. P^{3+} is oxidized, Cl_2 is reduced

c. Zn is oxidized, H^{+} is reduced

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6.14.4: More Practice with Reactions (Exercises)

5.1: The Law of Conservation of Matter

1. What is the law of conservation of matter?
2. How does the law of conservation of matter apply to chemistry?

ANSWERS

1. The law of conservation of matter states that in any given system that is closed to the transfer of matter, the amount of matter in the system stays constant
2. The law of conservation of matter says that in chemical reactions, the total mass of the products must equal the total mass of the reactants.

Exercises

1. Express the law of conservation of matter in your own words.

2. Explain why the concept of conservation of matter is considered a scientific law.

3. Potassium hydroxide ($\mathbf{KOH}$) readily reacts with carbon dioxide ($\mathbf{CO_2}$) to produce potassium carbonate ($\mathbf{K_2CO_3}$) and water ($\mathbf{H_2O}$). How many grams of potassium carbonate is produced if 224.4 g of $\mathbf{KOH}$ reacted with 88.0 g of $\mathbf{CO_2}$. The reaction also produced 36.0 g of water.

Answers

1. Matter may not be created or destroyed.

2. The concept is a scientific law because it is based on experimentation.

3. 276.4 g

5.2: Chemical Equations

Concept Review Exercises

a. What are the parts of a chemical equation?
b. Explain why chemical equations need to be balanced.

Answers

a. reactants and products

b. Chemical equations need to be balanced to satisfy the law of conservation of matter.

Exercises

1. Write a chemical equation to express the fact that hydrogen gas and solid iodine react to make gaseous hydrogen iodide. Make sure the equation satisfies the law of conservation of matter.
2. Write a chemical equation to express the fact that sodium metal and chlorine gas react to make solid sodium chloride. Make sure the equation satisfies the law of conservation of matter.
3. Write an equation expressing the fact that hydrogen gas and fluorine gas react to make gaseous hydrogen fluoride. Make sure the equation satisfies the law of conservation of matter.
4. Write an equation expressing the fact that solid potassium and fluorine gas react to make solid potassium fluoride. Make sure the equation satisfies the law of conservation of matter.
5. Mercury reacts with oxygen to make mercury(II) oxide. Write a balanced chemical equation that summarizes this reaction.

6. Octane (C_8H_{18}) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.
7. Propyl alcohol (C_3H_7OH) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.
8. Sulfuric acid (H_2SO_4) reacts with iron metal to make iron(III) sulfate and hydrogen gas. Write a balanced chemical equation that summarizes this reaction.
9. Balance each equation.
    a. $MgCl_2 + K \rightarrow KCl + Mg$
    b. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
    c. $NaN_3 \rightarrow Na + N_2$ (This is the reaction used to inflate airbags in cars.)
10. Balance each equation.
    a. $NH_4NO_3 \rightarrow N_2O + H_2O$
    b. $TiBr_4 + H_2O \rightarrow TiO_2 + HBr$
    c. $C_3H_5N_3O_9 \rightarrow CO_2 + N_2 + O_2 + H_2O$ (This reaction represents the decomposition of nitroglycerine.)
11. Balance each equation.
    a. $NH_3 + O_2 \rightarrow NO + H_2O$
    b. $Li + N_2 \rightarrow Li_3N$
    c. $AuCl \rightarrow Au + AuCl_3$
12. Balance each equation.
    a. $NaOH + H_3PO_4 \rightarrow Na_3PO_4 + H_2O$
    b. $N_2H_4 + Cl_2 \rightarrow N_2 + HCl$
    c. $Na_2S + H_2S \rightarrow NaSH$
13. Chromium(III) oxide reacts with carbon tetrachloride to make chromium(III) chloride and phosgene ($COCl_2$). Write the balanced chemical equation for this reaction.
14. The reaction that occurs when an Alka-Seltzer tablet is dropped into a glass of water has sodium bicarbonate reacting with citric acid ($H_3C_6H_5O_7$) to make carbon dioxide, water, and sodium citrate ($Na_3C_6H_5O_7$). Write the balanced chemical equation for this reaction.
15. When sodium hydrogen carbonate is used to extinguish a kitchen fire, it decomposes into sodium carbonate, carbon dioxide and water. Write a balanced chemical equation for this reaction.
16. Elemental bromine gas can be generated by reacting sodium bromide with elemental chlorine. The other product is sodium chloride. Write a balanced chemical equation for this reaction.

Answers

1. $H_2(g) + I_2(s) \rightarrow 2HI(g)$
2. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
3. $H_2(g) + F_2(g) \rightarrow 2HF(g)$
4. $2K(s) + F_2(g) \rightarrow 2KF(s)$
5. $2Hg + O_2 \rightarrow 2HgO$
6. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
7. $2C_3H_7OH + 9O_2 \rightarrow 6CO_2 + 8H_2O$
8. $3H_2SO_4 + 2Fe \rightarrow Fe_2(SO_4)_3 + 3H_2$
9. a. $MgCl_2 + 2K \rightarrow 2KCl + Mg$

b. $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$
c. $2NaN_3 \rightarrow 2Na + 3N_2$

10.

a. $NH_4NO_3 \rightarrow N_2O + 2H_2O$
b. $TiBr_4 + 2H_2O \rightarrow TiO_2 + 4HBr$
c. $4C_3H_5N_3O_9 \rightarrow 12CO_2 + 6N_2 + O_2 + 10H_2O$

11.

a. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
b. $6Li + N_2 \rightarrow 2Li_3N$
c. $3AuCl \rightarrow 2Au + AuCl_3$

12.

a. $3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$
b. $N_2H_4 + 2Cl_2 \rightarrow N_2 + 4HCl$
c. $Na_2S + H_2S \rightarrow 2NaSH$

13. $Cr_2O_3 + 3CCl_4 \rightarrow 2CrCl_3 + 3COCl_2$

14. $3NaHCO_3 + H_3C_6H_5O_7 \rightarrow 3CO_2 + 3H_2O + Na_3C_6H_5O_7$

15. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

16. $2NaBr + Cl_2 \rightarrow Br_2 + 2NaCl$

5.3: Quantitative Relationships Based on Chemical Equations

Concept Review Exercises

1. Explain how stoichiometric ratios are constructed from a chemical equation.
2. Why is it necessary for a chemical equation to be balanced before it can be used to construct conversion factors?

Answers

1. Stoichiometric ratios are made using the coefficients of the substances in the balanced chemical equation.
2. A balanced chemical equation is necessary so one can construct the proper stoichiometric ratios.

Exercises

1. Balance this equation and write every stoichiometric ratio you can from it.

$$NH_4NO_3 \rightarrow N_2O + H_2O$$

2. Balance this equation.

$$N_2 + H_2 \rightarrow NH_3$$

3. Balance this equation and write every stoichiometric ratio you can from it.

$$Fe_2O_3 + C \rightarrow Fe + CO_2$$

4. Balance this equation.

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

5. Balance this equation and determine how many molecules of CO_2 are formed if 15 molecules of C_6H_6 are reacted.

$$C_6H_6 + O_2 \rightarrow CO_2 + H_2O$$

6. Balance this equation and determine how many formula units of $Ag_2CO_3(s)$ are produced if 20 formula units of Na_2CO_3 are reacted.

$$Na_2CO_3(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag_2CO_3(s)$$

7. Copper metal reacts with nitric acid according to this equation:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Cu atoms will react if 488 molecules of aqueous HNO_3 are reacted?

8. Gold metal reacts with a combination of nitric acid and hydrochloric acid according to this equation:

$$Au(s) + 3HNO_3(aq) + 4HCl(aq) \rightarrow HAuCl_4(aq) + 3NO_2(g) + 3H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Au atoms react with 639 molecules of aqueous HNO_3?

9. Sulfur can be formed by reacting sulfur dioxide with hydrogen sulfide at high temperatures according to this equation:

$$SO_2(g) + 2H_2S(g) \rightarrow 3S(g) + 2H_2O(g)$$

   a. Verify that this equation is balanced.
   b. How many S atoms will be formed from by reacting 1,078 molecules of H_2S?

10. Nitric acid is made by reacting nitrogen dioxide with water:

$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

   a. Verify that this equation is balanced.
   b. How many molecules of NO will be formed by reacting 2,268 molecules of NO_2?

Answers

1. $NH_4NO_3 \rightarrow N_2O + 2H_2O$; the stoichiometric ratios are $\frac{1NH_4NO_3}{1N_2O}$, $\frac{1NH_4NO_3}{2H_2O}$, $\frac{1N_2O}{2H_2O}$, and their reciprocals.
2. $N_2 + 3H_2 \rightarrow 2NH_3$
3. $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$; the stoichiometric ratios are $\frac{2Fe_2O_3}{3C}$, $\frac{2Fe_2O_3}{4Fe}$, $\frac{2Fe_2O_3}{3CO_2}$, $\frac{3C}{4Fe}$, $\frac{3C}{3CO_2}$, $\frac{4Fe}{3CO_2}$, and their reciprocals.
4. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
5. $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$; 90 molecules
6. $Na_2CO_3(aq) + 2AgNO_3(aq) \rightarrow 2NaNO_3(aq) + Ag_2CO_3(s)$; 20 formula units
7.
   a. It is balanced.
   b. 183 atoms
8.
   a. It is balanced.
   b. 213 atoms
9.
   a. It is balanced.
   b. 1,617 atoms
10.
   a. It is balanced.
   b. 756 molecules

5.4: Some Types of Chemical Reactions

Concept Review Exercises

1. What is the difference between a combination reaction and a combustion reaction?
2. Give the distinguishing characteristic(s) of a decomposition reaction.
3. How do we recognize a combustion reaction?

Answers

1. A combination reaction produces a certain substance; a combustion reaction is a vigorous reaction, usually a combination with oxygen, that is accompanied by the production of light and/or heat.
2. In a decomposition reaction, a single substance reacts to make multiple substances as products.
3. A combustion reaction is typically a vigorous reaction accompanied by light and/or heat, usually because of reaction with oxygen.

Exercises

1. Identify each type of reaction.

   a. $C_6H_5CH_3 + 9O_2 \rightarrow 7CO_2 + 4H_2O$
   b. $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
   c. $C + 2H_2 \rightarrow CH_4$

2. Identify each type of reaction.

   a. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$
   b. $FeO + SO_3 \rightarrow FeSO_4$
   c. $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

3. Identify each type of reaction.

   a. $2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$
   b. $Hg(\ell) + ½O_2\ (g) \rightarrow HgO(s)$
   c. $CH_2CH_2(g) + Br_2(\ell) \rightarrow CH_2BrCH_2Br$

4. Identify each type of reaction.

   a. $Ti(s) + O_2(g) \rightarrow TiO_2(s)$
   b. $H_2SO_3(aq) \rightarrow H_2O(\ell) + SO_2(g)$
   c. $3O_2(g) \rightarrow 2O_3(g)$

Answers

1.
   a. combustion
   b. decomposition
   c. combination

2.
   a. combination
   b. combination
   c. decomposition

3.
   a. decomposition
   b. combustion (also combination)
   c. combination

4.

a. combination

b. decomposition

c. combination

5.5: Oxidation-Reduction (Redox) Reactions

Concept Review Exercises

1. Give two different definitions for oxidation and reduction.
2. Give an example of each definition of oxidation and reduction.

Answers

1. Oxidation is the loss of electrons or the addition of oxygen; reduction is the gain of electrons or the addition of hydrogen.
2. $Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation); $C_2H_4 + H_2 \rightarrow C_2H_6$ (reduction) (answers will vary)

Exercises

1. Which reactions are redox reactions? For those that are redox reactions, identify the oxidizing and reducing agents.

   a. $NaOH + HCl \rightarrow H_2O + NaCl$
   b. $3Mg + 2AlCl_3 \rightarrow 2Al + 3MgCl_2$
   c. $H_2O_2 + H_2 \rightarrow 2H_2O$
   d. $KCl + AgNO_3 \rightarrow AgCl + KNO_3$

2. Which reactions are redox reactions? For those that are redox reactions, identify the oxidizing and reducing agents.

   a. $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$
   b. $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
   c. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
   d. $2K + 2H_2O \rightarrow 2KOH + H_2$

3. Balance each redox reaction by writing appropriate half reactions and combining them to cancel the electrons.

   a. $Ca(s) + H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   b. $I^-(aq) + Br_2(\ell) \rightarrow Br^-(aq) + I_2(s)$

4. Balance each redox reaction by writing appropriate half reactions and combining them to cancel the electrons.

   a. $Fe(s) + Sn^{4+}(aq) \rightarrow Fe^{3+}(aq) + Sn^{2+}(aq)$
   b. $Pb(s) + Pb^{4+}(aq) \rightarrow Pb^{2+}(aq)$ (Hint: both half reactions will yield the same product.)

Answers

1.

   a. no

   b. yes; oxidizing agent: $AlCl_3$; reducing agent: Mg

   c. yes; oxidizing agent: H_2O_2; reducing agent: H_2

   d. no

2.

   a. yes; oxidizing agent: HNO_3; reducing agent: Cu
   b. yes; oxidizing agent: O_2; reducing agent: C_2H_6
   c. no

d. yes; oxidizing agent: H_2O; reducing agent: K

3.

a. Combined: $Ca + 2H^+ \rightarrow Ca^{2+} + H_2$
b. Combined: $Br_2 + 2I^- \rightarrow 2Br^- + I_2$

4.

a. $(Fe \rightarrow Fe^{3+} + 3e^-) \times 2$

$(Sn^{4+} + 2e^- \rightarrow Sn^{2+}) \times 3$

Combined: $2Fe + 3Sn^{4+} \rightarrow 2Fe^{3+} + 3Sn^{2+}$

b. $Pb \rightarrow Pb^{2+} + 2e^-$

$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$

Combined: $Pb + Pb^{4+} \rightarrow 2Pb^{2+}$

5.6: Redox Reactions in Organic Chemistry and Biochemistry

Concept Review Exercise

1. Give some biochemical examples of oxidation and reduction reactions.

Answer

1. photosynthesis and antioxidants in foods (answers will vary)

Exercises

1. A typical respiratory reaction discussed in the text is the oxidation of glucose ($C_6H_{12}O_6$):

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

Is this a redox reaction? If so, what are the oxidizing and reducing agents?

2. The major net reaction in photosynthesis is as follows:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Is this a redox reaction? If so, what are the oxidizing and reducing agents?

3. What would be the ultimate organic product if $CH_3CH_2CH_2OH$ were to react with a solution of $K_2Cr_2O_7$?
4. What would be the ultimate organic product if $CH_3CH_2CH_2CH_2OH$ were to react with a solution of $K_2Cr_2O_7$?
5. What would be the final organic product if $CH_3CH_2CHOHCH_3$ were to react with a solution of $K_2Cr_2O_7$?
6. What would be the major organic product if $CH_3CH_2CHOHCH_2CH_3$ were to react with a solution of $K_2Cr_2O_7$?
7. What alcohol is produced in the reduction of acetone [$(CH_3)_2CO$]?
8. What alcohol is produced in the reduction of propanal (CH_3CH_2CHO)?

Answers

1. yes; oxidizing agent: O_2; reducing agent: $C_6H_{12}O_6$
2. yes; oxidizing agent: CO_2; reducing agent: H_2O
3. CH_3CH_2COOH
4. $CH_3CH_2CH_2COOH$
5. $CH_3CH_2C(O)CH_3$, where the carbon is double bonded to the oxygen
6. $CH_3CH_2C(O)CH_2CH_3$, carbon #3 is double bonded to the oxygen

7. $CH_3CHOHCH_3$, or isopropyl alcohol

8. $CH_3CH_2CH_2OH$

5.

Additional Exercises

1. Isooctane (C_8H_{18}) is used as a standard for comparing gasoline performance. Write a balanced chemical equation for the combustion of isooctane.

2. Heptane (C_7H_{16}), like isooctane (see Exercise 1), is also used as a standard for determining gasoline performance. Write a balanced chemical equation for the combustion of heptane.

3. What is the difference between a combination reaction and a redox reaction? Are all combination reactions also redox reactions? Are all redox reactions also combination reactions?

4. Are combustion reactions always redox reactions as well? Explain.

5. A friend argues that the equation

$$Fe^{2+} + Na \rightarrow Fe + Na^{+}$$

is balanced because each side has one iron atom and one sodium atom. Explain why your friend is incorrect.

6. Some antacids contain aluminum hydroxide [$Al(OH)_3$]. This compound reacts with excess hydrochloric acid (HCl) in the stomach to neutralize it. If the products of this reaction are water and aluminum chloride, what is the balanced chemical equation for this reaction? What is the stoichiometric ratio between the number of HCl molecules made to the number of H_2O molecules made?

7. Sulfuric acid is made in a three-step process: (1) the combustion of elemental sulfur to produce sulfur dioxide, (2) the continued reaction of sulfur dioxide with oxygen to produce sulfur trioxide, and (3) the reaction of sulfur trioxide with water to make sulfuric acid (H_2SO_4). Write balanced chemical equations for all three reactions.

8. If the products of glucose metabolism are carbon dioxide and water, what is the balanced chemical equation for the overall process? What is the stoichiometric ratio between the number of CO_2 molecules made to the number of H_2O molecules made?

9. Historically, the first true battery was the Leclanché cell, named after its discoverer, Georges Leclanché. It was based on the following reaction:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Identify what is being oxidized, what is being reduced, and the respective reducing and oxidizing agents.

Answers

1. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

2. $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

3. A combination reaction makes a new substance from more than one reactant; a redox reaction rearranges electrons. Most (not all) combination reactions are redox reactions. Not all redox reactions are combination reactions.

4. All combustion reactions are redox reactions. In combustion a chemical is combining with oxygen and that chemical is being oxidized. Oxygen, on the other hand, is being reduced.

5. Your friend is incorrect because the number of electrons transferring is not balanced. A balanced equation must not only have the same number of atoms of each element on each side of the equation but must also have the same charge on both sides.

6. $Al(OH)_3 + 3HCl \rightarrow 3H_2O + AlCl_3$; 1:1

7. (1) $S + O_2 \rightarrow SO_2$; (2) $2SO_2 + O_2 \rightarrow 2SO_3$; (3) $SO_3 + H_2O \rightarrow H_2SO_4$

8. $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$; 1:1

9. oxidized and reducing agent: Zn; reduced and oxidizing agent: Cu^{2+}

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CHAPTER OVERVIEW

7: Solutions

7.1: Structure of Water
7.2: Structure of Ice
7.3: Physical Properties of Water
7.4: Solute and Solvent
7.5: Dissolving Process
7.6: Liquid-Liquid Solutions
7.7: Electrolytes and Nonelectrolytes
7.8: Dissociation
7.9: Strong and Weak Electrolytes
7.10: Solutions
7.11: Rate of Dissolving
7.12: Saturated and Unsaturated Solutions
7.13: How Temperature Influences Solubility
7.14: Henry's Law
7.15: Concentration
7.16: Molecular and Ionic Equations
7.17: Net Ionic Equations
7.18: Predicting Precipitates Using Solubility Rules
7.19: Chapter 7 Exercises
7.19.1: Solutions (Exercises)
7.19.2: Answers to Solutions (Exercises)

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7.1: Structure of Water

In his well-known poem "The Rime of the Ancient Mariner", Samuel Coleridge wrote "water, water everywhere, nor any drop to drink." Coleridge was talking about being out on the ocean, but not having any water because he had killed an albatross (apparently bringing bad luck to everyone on the ship). About 75% of the Earth's surface is water. The major constituent of the human body (over 60%) is water. This simple molecule plays important roles in all kinds of processes.

Structure of Water

Water is a simple molecule consisting of one oxygen atom bonded to two different hydrogen atoms. Because of the higher **electronegativity** of the oxygen atom, the bonds are polar covalent (**polar bonds**). The oxygen atom attracts the shared electrons of the covalent bonds to a significantly greater extent than the hydrogen atoms. As a result, the oxygen atom acquires a partial negative charge ($\delta-$), while the hydrogen atoms each acquire a partial positive charge ($\delta+$). The molecule adopts a bent structure because of the two lone pairs of electrons on the oxygen atom. The $\mathrm{H-O-H}$ bond angle is about $105°$, slightly smaller than the ideal $109.5°$ of an sp^3 hybridized atomic orbital.

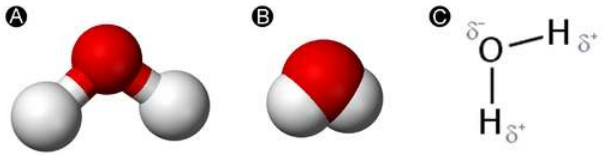


Figure 7.1.1: The water molecule, visualized three different ways: ball-and-stick model, space-filling model, and structural formula with partial charges.

The bent shape of the water molecule is critical because the polar $\mathrm{O-H}$ bonds do not cancel one another and the molecule as a whole is polar. The figure below illustrates the net polarity of the water molecule. The oxygen is the negative end of the molecule, while the area between the hydrogen atoms is the positive end of the molecule.

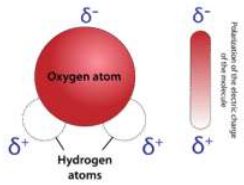


Figure 7.1.2: Water is a polar molecule, as greater electron density is found around the more electronegative oxygen atom.

Polar molecules attract one another by dipole-dipole forces, as the positive end of one molecule is attracted to the negative end of the nearby molecule. In the case of water, the highly polar $\mathrm{O-H}$ bonds results in very little electron density around the hydrogen atoms. Each hydrogen atom is strongly attracted to the lone-pair electrons on an adjacent oxygen atom. These are called hydrogen bonds and are stronger than conventional dipole-dipole forces.

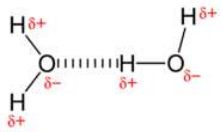


Figure 7.1.3: A hydrogen bond is the attraction between a lone pair of electrons on the oxygen atom of one molecule and the electron-deficient hydrogen atom of a nearby molecule.

Because each oxygen atom has two lone pairs, it can make hydrogen bonds to the hydrogen atoms of two separate other molecules. The figure below shows the result—an approximately tetrahedral geometry around each oxygen atom, consisting of two covalent bonds and two hydrogen bonds.

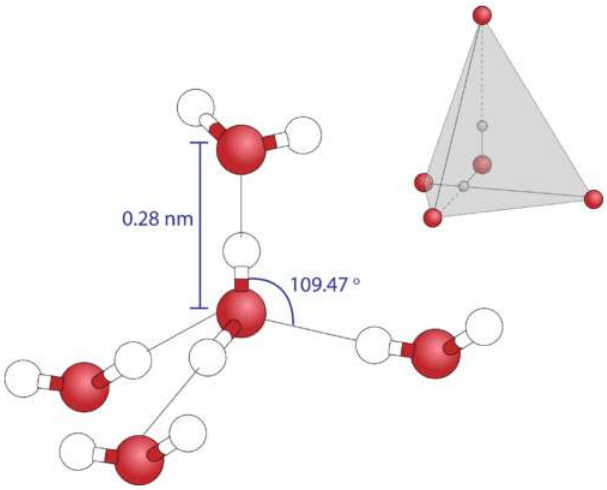


Figure **7.1.4**: As a result of two covalent bonds and two hydrogen bonds, the geometry around each oxygen atom is approximately tetrahedral.

Summary

- Water is a molecular compound consisting of polar molecules that have a bent shape.
- The oxygen atom acquires a partial negative charge, while the hydrogen atom acquires a partial positive charge.

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7.2: Structure of Ice

Ice is an interesting and useful material. It can be used to cool food and keep it fresh. Ice can provide recreation, such as in the case of ice-skating. It can do great damage when it freezes—roads can buckle, houses can be damaged, water pipes can burst. All of this happens because of a unique property of water: when water freezes, it expands in volume as ice is formed.

Structure of Ice

Liquid water is a fluid. The hydrogen bonds in liquid water constantly break and reform as the water molecules tumble past one another. As water cools, its molecular motion slows and the molecules move gradually closer to one another. The **density** of any liquid increases as its temperature decreases. For most liquids, this continues as the liquid freezes; the solid state is denser than the liquid state. However, water behaves differently. It actually reaches its highest density at about $4°C$.

Table 7.2.1: Density of Water and Ice

Temperature (°C)	Density $\left(g/cm^3\right)$
100 (liquid)	0.9584
50	0.9881
25	0.9971
10	0.9997
4	1.0000
0 (liquid)	0.9998
0 (solid)	0.9168

Between $4°C$ and $0°C$, the density gradually decreases as the hydrogen bonds begin to form a network characterized by a generally **hexagonal** structure with open spaces in the middle of the hexagons (see figure below).

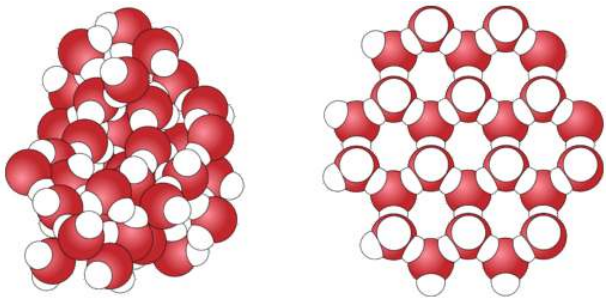


Figure 7.2.1: The structure of liquid water (left) consists of molecules connected by short-lived hydrogen bonds because water is a fluid. In ice (right), the hydrogen bonds become permanent, resulting in an interconnected hexagonally-shaped framework of molecules.

Ice is less dense than liquid water and so it floats. Ponds or lakes begin to freeze at the surface, closer to the cold air. A layer of ice forms, but does not sink as it would if water did not have this unique structure dictated by its shape, polarity, and hydrogen bonding. If the ice were to sink as it froze, entire lakes would freeze solid. Since the ice does not sink, liquid water remains under the ice all winter long. This is important, as fish and other organisms are capable of surviving through winter. Ice is one of few solids that is less dense than its liquid form.

Summary

- The density of any liquid increases as its temperature decreases. As an exception, ice is less dense than liquid water.
- The intermolecular structure of ice has spaces that are not present in liquid water.

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7.3: Physical Properties of Water

Water loss to the atmosphere is a significant problem in many parts of the world. When water supplies are low, anything that can be done to decrease water loss is important for farmers. An evaporation pan can be used to measure how fast water evaporates in a given location. This information can be used as part of project development to cut down on evaporation, and increase the amount of usable water in a region.

Properties of Water

Compared to other molecular compounds of relatively low molar mass, ice melts at a very high temperature. A great deal of energy is required to break apart the hydrogen-bonded network of ice and return it to the liquid state. Likewise, the boiling point of water is very high. Most molecular compounds of similar molar mass are gases at room temperature.

Surface Tension

Water has a high **surface tension** (attraction between molecules at the surface of a liquid) because of its hydrogen bonding. Liquids that cannot hydrogen bond do not exhibit nearly as much surface tension. Surface tension can be seen by the curved meniscus that forms when water is in a thin column, such as in a graduated cylinder or a buret.



Figure **7.3.1**: The meniscus of water in a graduated cylinder forms because of water's hydrogen bonding.

Vapor Pressure

The **vapor pressure** of a liquid is the pressure of the vapor produced by evaporation of a liquid or solid above the liquid or solid in a closed container. The hydrogen bonding between liquid water molecules explains why water has an unusually low vapor pressure. Relatively few molecules of water are capable of escaping the surface of the liquid and entering the vapor phase. Evaporation is slow, and thus the vapor exerts a low pressure in a closed container. Low vapor pressure is an important physical property of water, since lakes, oceans, and other large bodies of water would otherwise evaporate much more quickly.

Vapor pressure is influenced by temperature. As the temperature increases, more molecules are released from the surface of the liquid. This increases movement above the liquid surface, increasing the pressure in the vapor stage. The image below illustrates the effect of temperature on vapor pressure.

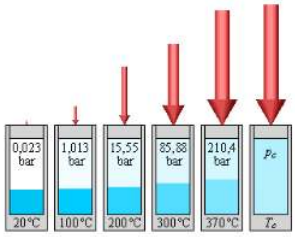


Figure **7.3.2**: Vapor pressure of water at different temperatures.

Summary

- Water has high surface tension because of extensive hydrogen bonding.
- The vapor pressure of water is low due to hydrogen bonding.
- Vapor pressure increases as temperature increases.

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7.4: Solute and Solvent

In the winter, the temperature often gets well below the freezing point of water. This condition can create problems in car radiators. If the water freezes, water hoses will break, the engine block can crack, and significant damage can be done to the car.

Solute and Solvent

When one substance dissolves into another, a **solution** is formed. A solution is a homogeneous mixture consisting of a **solute** dissolved into a **solvent**. The solute is the substance that is being dissolved, while the solvent is the dissolving medium. Solutions can be formed with many different types and forms of solutes and solvents.

We know of many types of solutions. Check out a few examples in the table below.

Table **7.4.1**: Types of Solutions

Type	Solvent	Solute	Example
gas/gas	nitrogen	oxygen	air
gas/liquid	water	carbon dioxide	soda pop
liquid/liquid	water	ethylene glycol	antifreeze
solid/liquid	water	salts	seawater

This section's focus is on solutions where the solvent is water. An aqueous solution is water that contains one or more dissolved substances. The dissolved substances in an aqueous solution many be solids, gases, or other liquids. Some examples are listed in the table above. Other examples include vinegar (acetic acid in water), alcoholic beverages (ethanol in water), and liquid cough medicines (various drugs in water).



Figure **7.4.1**: Typical solution: clear and stable.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will be unchanged. The dissolved sugar particles will pass through the filter along with the water. This is because the dissolved particles in a solution are very small, usually less than $1\ \text{nm}$ in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved.

Summary

- A solution is a homogeneous mixture of a solute in a solvent.
- A solute is the material present in the smaller amount in the solution.
- A solvent is the material present in the larger amount in the solution.

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7.5: Dissolving Process

Many people start their day with a cup of coffee (others drink coffee all day long). The coffee typically consumed is a type of solution—sometimes a very complex one. The coffee itself has been brewed so that material from the coffee bean will dissolve in hot water. Some people add sugar, while others add milk or cream. In some areas of the country, lattés are popular; one or more special flavors might be added along with the milk. Whatever the situation, the end result is an enjoyable solution to drink.

The Process of Dissolution

Water typically dissolves many ionic compounds and polar molecules. Nonpolar molecules such as those found in grease or oil do not dissolve in water. We will first examine the process that occurs when an ionic compound such as table salt (sodium chloride) dissolves in water.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative ions. Water is attracted to the sodium chloride crystal because water is polar and has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see image below).

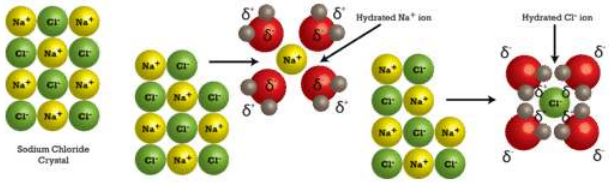


Figure **7.5.1**: The dissolving of sodium chloride in water.

After coming apart from the crystal, the individual ions are then surrounded by solvent particles in a process called **solvation**. Note that the individual $\mathbf{Na^+}$ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. **Hydration** is the process of solute particles being surrounded by water molecules arranged in a specific manner. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is sucrose ($\mathbf{C_{12}H_{22}O_{11}}$), and is an example of a molecular compound. Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but does not break the covalent bonds between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated, but without as distinct an orientation to the water molecules as in the case of the ions. The sugar molecules contain many $\mathbf{-OH}$ groups that can form hydrogen bonds with the water molecules, helping to form the sucrose solution.

Summary

- Motion of water molecules helps break up interactions between solid ions or molecules.
- Solvation involves surrounding ions with solvent particles.
- Ionic solute molecules are hydrated (surrounded by water molecules in a specific orientation).

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7.6: Liquid-Liquid Solutions

In 2010, a major oil spill occurred when an explosion on a drilling rig in the Gulf of Mexico released millions of gallons of crude oil into the Gulf. Oil is primarily a mixture of hydrocarbons (organic compounds composed of only carbon and hydrogen atoms). Because of its composition, oil does not dissolve in water. As a result, much of the Gulf of Mexico was contaminated, as was a great deal of shoreline in the affected area.

Liquid-Liquid Solutions

Nonpolar compounds do not dissolve in water. The attractive forces that operate between the particles in a nonpolar compound are weak dispersion forces. However, the nonpolar molecules are more attracted to themselves than they are to the **polar** water molecules. When a nonpolar liquid such as oil is mixed with water, two separate layers form because the liquids will not dissolve into each other (figure below). When another polar liquid such as ethanol is mixed with water, they completely blend and dissolve into one another. Liquids that dissolve in one another in all proportions are said to be **miscible**. Liquids that do not dissolve in one another are deemed **immiscible**. The general rule for deciding if one substance is capable of dissolving another is "like dissolves like". A nonpolar solid such as iodine will dissolve in nonpolar lighter fluid, but will not dissolve in polar water.



Figure **7.6.1**: Water and oil form separate layers when they are mixed because the nonpolar oil will not dissolve into the polar water. The oil forms the top layer because it is less dense than the water.

For molecular compounds, the major factor that contributes to the material dissolving in water is the ability to form hydrogen bonds with the water solvent. Small compounds such as methanol, ethanol, acetic acid, and acetone have polar groups that can interact with the polar **H** of water. However, as the nonpolar portion of the molecule gets larger, solubility with water drops off. The nonpolar portion of the molecule increasingly repels the water and eventually overrides the interaction of the polar component with water.

Summary

- Liquids that dissolve in one another in all proportions are miscible; liquids that do not dissolve in one another are immiscible.
- The general rule for deciding if one substance is capable of dissolving another is "like dissolves like".
- Nonpolar molecules are usually insoluble in water.
- A non-ionized molecule must be relatively polar to interact with water molecules.

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7.7: Electrolytes and Nonelectrolytes

People around the world jog for exercise. For the most part, jogging can be a healthy way to stay fit. However, problems can develop for those who jog in the heat. Excessive sweating can lead to electrolyte loss, which can be life-threatening. Early symptoms of electrolyte deficiency can include nausea, fatigue, and dizziness. If not treated, individuals can experience muscle weakness and increased heart rate (which could lead to a heart attack). Sports drinks can be consumed to restore electrolytes quickly in the body.

Electrolytes and Nonelectrolytes

An **electrolyte** is a compound that conducts an electric current when it is in an aqueous solution or melted. In order to conduct a current, a substance must contain mobile ions that can move from one electrode to the other. All ionic compounds are electrolytes. When ionic compounds dissolve, they break apart into ions which are then able to conduct a current (**conductivity**). Even insoluble ionic compounds such as $CaCO_3$ are electrolytes because they can conduct a current in the molten (melted) state.

A **nonelectrolyte** is a compound that does not conduct an electric current in either aqueous solution or in the molten state. Many molecular compounds, such as sugar or ethanol, are nonelectrolytes. When these compounds dissolve in water, they do not produce ions. The figure below illustrates the difference between an electrolyte and a nonelectrolyte.

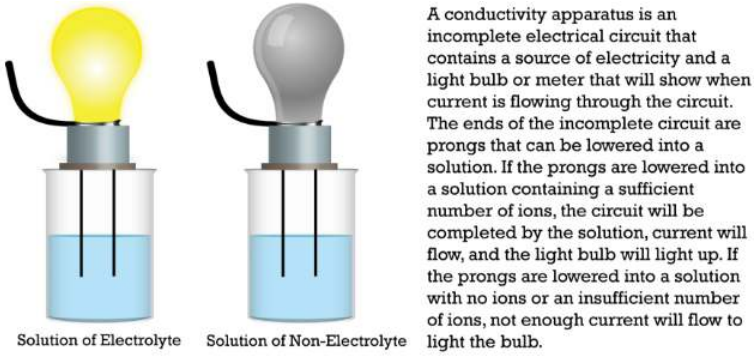


Figure **7.7.1**: Conductivity apparatus.

Roles of Electrolytes in the Body

Several electrolytes play important roles in the body. Here are a few significant electrolytes:

1. Calcium - in bones and teeth. Also important for muscle contraction, blood clotting, and nerve function.

2. Sodium - found outside the cell. Mainly involved in water balance and nerve signaling.

3. Potassium - major cation inside the cell. Important for proper functioning of heart, muscles, kidneys, and nerves.

4. Magnesium - in bone and cells. Involved in muscle, bone, nervous system, and takes part in many biochemical reactions.

Summary

- Electrolytes conduct electric current when in solution or melted.
- Nonelectrolytes do not conduct electric current when in solution or melted.
- Some electrolytes play important roles in the body.

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7.8: Dissociation

In many areas, winter ice on the streets and sidewalks represents a serious walking and driving hazard. One common approach to melting the ice is to put some form of deicing salt on the surface. Materials such as sodium chloride or calcium chloride are frequently employed for this purpose. In order to be effective, the solid material must first dissolve and break up into the ions that make up the compound.

Dissociation

An ionic crystal lattice breaks apart when it is dissolved in water. **Dissociation** is the separation of ions that occurs when a solid ionic compound dissolves. It is important to be able to write dissociation equations. Simply undo the crisscross method that you learned when writing chemical formulas of ionic compounds. The subscripts for the ions in the chemical formulas become the coefficients of the respective ions on the product side of the equation. Shown below are dissociation equations for NaCl, $\mathrm{Ca(NO_3)_2}$, and $\mathrm{(NH_4)_3PO_4}$.

$$\mathrm{NaCl}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

$$\mathrm{Ca(NO_3)_2}\,(s) \rightarrow \mathrm{Ca^{2+}}\,(aq) + 2\mathrm{NO_3^-}\,(aq)$$

$$\mathrm{(NH_4)_3PO_4}\,(s) \rightarrow 3\mathrm{NH_4}\,(aq) + \mathrm{PO_4}\,(aq)$$

The formula unit of sodium chloride dissociates into one sodium ion and one chloride ion. The calcium nitrate formula unit dissociates into one calcium ion and two nitrate ions. This is because of the $2+$ charge of the calcium ion. Two nitrate ions, each with a $1-$ charge are required to make the equation balance electrically. The ammonium phosphate formula unit dissociates into three ammonium ions and one phosphate ion. Note that the polyatomic ions themselves do not dissociate further, but remain intact.

Do not confuse the subscripts of the atoms within the polyatomic ion for the subscripts that result from the crisscrossing of the charges that make up the original compound neutral. The 3 subscript of the nitrate ion and the 4 subscript of the ammonium ion are part of the polyatomic ion and simply remain as part of its formula after the compound dissociates. Notice that the compounds are solids (s) which then become ions in aqueous solution (aq).



Figure 7.8.1: Calcium nitrate is a typical ionic compound. In an aqueous solution, it dissociates into calcium ions and nitrate ions.

Nonelectrolytes do not dissociate when forming an aqueous solution. An equation can still be written that simply shows the solid going into solution. For the dissolution of sucrose:

$$\mathrm{C_{12}H_{22}O_{11}}\,(s) \rightarrow \mathrm{C_{12}H_{22}O_{11}}\,(aq)$$

Summary

- Dissociation is the separation of ions that occurs when a solid ionic compound dissolves.
- Nonionic compounds do not dissociate in water.

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7.9: Strong and Weak Electrolytes

Car batteries are used around the world to provide the power to start car engines. One essential component of car batteries is the strong electrolyte sulfuric acid. In the battery, this material ionizes into hydrogen ions and sulfate ions. As the battery is used, the concentrations of these ions decreases. Older batteries had openings in the top where new sulfuric acid could be added to replenish the supply. Today, batteries are sealed to prevent leakage of hazardous sulfuric acid.

Strong and Weak Electrolytes

Some polar molecular compounds are nonelectrolytes when they are in their pure state, but become electrolytes when they are dissolved in water. Hydrogen chloride (HCl) is a gas in its pure molecular state and is a nonelectrolyte. However, when HCl is dissolved in water, it conducts a current well because the HCl molecule ionizes into hydrogen and chloride ions.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

HCl dissolved into water is called hydrochloric acid. A **strong electrolyte** is a solution in which a large fraction of the dissolved solute exists as ions. Ionic compounds, and some polar compounds, are completely broken apart into ions and thus conduct a current very well—this makes them strong electrolytes.

Some other polar molecular compounds become electrolytes upon being dissolved into water, but do not ionize to very great extent. Gaseous nitrous acid ionizes in solution to hydrogen ions and nitrite ions, but does so very weakly. Aqueous nitrous acid is composed of only about 5% ions and 95% intact nitrous acid molecules. A **weak electrolyte** is a solution in which only a small fraction of the dissolved solute exists as ions. The equation showing the ionization of a weak electrolyte utilizes a double arrow indicating an equilibrium between the reactants and products.

$$\mathrm{HNO_2}\,(g) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{NO_2^-}\,(aq)$$

Summary

- A strong electrolyte exists mainly as ions in solution.
- A weak electrolyte has only a small amount of ionization in solution.

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7.10: Solutions

Learning Objectives

- To understand what causes solutions to form.

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 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. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table **7.10.1** lists some common types of solutions, with examples of each.

Table **7.10.1**: 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 (recall Section 4.6) 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). However, water does not dissolve nonpolar solutes, such as many oils and greases (Figure **7.10.1**).



Figure **7.10.1**: A beaker holds water with blue food dye (upper liquid layer) and a much more dense perfluoroheptane (a fluorocarbon) lower liquid layer. The two fluids cannot mix and the dye cannot dissolve in fluorocarbon. A goldfish and a crab have been introduced into the water. The goldfish cannot penetrate the dense fluorocarbon. The crab floats at the liquid boundary with only parts of his legs penetrating the fluorocarbon fluid, unable to sink to the bottom of the beaker. Quarter coins rest on the bottom of the beaker. Animals were rescued from their predicament after the photo was taken. Figure used with permission from Wikipedia (Sbharris (Steven B. Harris)).

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.

✔ Example 7.10.1

Water is considered a polar solvent. Which substances should dissolve in water?

1. methanol (CH_3OH)
2. sodium sulfate (Na_2SO_4)
3. octane (C_8H_{18})

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

1. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
2. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
3. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

? Exercise 7.10.1

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer

Octane only.

Example 7.10.2

Predict which of the following compounds will be most soluble in water:

a. $\underset{\text{Ethanol}}{CH_3CH_2OH}$

b. $\underset{\text{Hexanol}}{CH_3CH_2CH_2CH_2CH_2CH_2OH}$

Solution

Since ethanol contains an **OH** group, it can hydrogen bond to water. Although the same is true of hexanol, the OH group is found only at one end of a fairly large molecule. The rest of the molecule can be expected to behave much as though it were a nonpolar alkane. This substance should thus be much less soluble than the first. Experimentally we find that ethanol is completely miscible with water, while only 0.6 g hexanol dissolves in 100 g water.

7.10.2 Exercise

Would I_2 be more soluble in CCl_4 or H_2O?

Answer

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

Key Takeaway

- Solutions form because a solute and a solvent experience similar intermolecular interactions.

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7.11: Rate of Dissolving

Many people enjoy a cold glass of iced tea on a hot summer day. Some like it unsweetened, while others like to put sugar in it. How sugar dissolves in the tea depends on two factors: how much sugar was put into the tea, and how cold it is. Tea usually has to be stirred for a while to get all the sugar dissolved.

Rate of Dissolution

Dissolution is the process by which a solute dissolves into a solvent and forms a solution. We know that the dissolution of a solid by water depends upon the collisions that occur between the solvent molecules and the particles in the solid crystal. Anything that can be done to increase the frequency of those collisions and/or to give those collisions more energy will increase the rate of dissolution. Imagine that you were trying to dissolve some sugar in a glassful of tea. A packet of granulated sugar would dissolve faster than a cube of sugar. The rate of dissolution would be increased by stirring, or agitating the solution. Finally, the sugar would dissolve faster in hot tea than it would in cold tea.

Surface Area

The rate at which a solute dissolves depends upon the size of the solute particles. Dissolution is a surface phenomenon, since it depends on solvent molecules colliding with the outer surface of the solute. A given quantity of solute dissolves faster when it is ground into small particles, rather than in the form of large pieces, because more **surface area** is exposed. A packet of granulated sugar exposes far more surface area to the solvent and dissolves more quickly than a sugar cube.

Agitation of the Solution

Dissolving sugar in water will occur more quickly if the water is stirred. The stirring allows fresh solvent molecules to continually be in contact with the solute. If it is not stirred, then the water right at the surface of the solute becomes saturated with dissolved sugar molecules, meaning that it is more difficult for additional solute to dissolve. The sugar cube would eventually dissolve because random motions of the water molecules would bring enough fresh solvent into contact with the sugar, but the process would take much longer. It is important to realize that neither stirring nor breaking up a solute affect the overall amount of solute that dissolves—these actions only affect the rate of dissolution.

Temperature

Heating up a solvent gives the molecules more kinetic energy. The increased rapid motion means that the solvent molecules collide with the solute with greater frequency, and that the collisions occur with more force. Both factors increase the rate at which the solute dissolves. As we will see in the next section, a temperature change not only affects the rate of dissolution, but also affects the amount of solute that dissolves.

Summary

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7.12: Saturated and Unsaturated Solutions

When compounds are synthesized, contaminating materials are often mixed in with them. The process of recrystallization can be used to remove these impurities. The crystals are dissolved in a hot solvent, forming a solution. When the solvent is cooled, the compound is no longer as soluble and will precipitate out of solution, leaving contaminating materials still dissolved.

Saturated and Unsaturated Solutions

Table salt (NaCl) readily dissolves in water. Suppose that you have a beaker of water to which you add some salt, stirring until it dissolves. You add more and that dissolves. You keep adding more and more salt, eventually reaching a point at which no more of the salt will dissolve—no matter how long or how vigorously you stir it. Why? On the molecular level, we know that action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by water molecules. What also happens is that some of the dissolved ions collide back again with the crystal and remain there. **Recrystallization** is the process of dissolved solute returning to the solid state. At some point, the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged. **Solution equilibrium** is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate. The solution equilibrium for the dissolution of sodium chloride can be represented by one of two equations:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{NaCl}\,(aq)$$

While the above equation shows the change of state back and forth between solid and aqueous solution, the preferred equation also shows the dissociation that occurs as an ionic solid dissolves:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

When the solution equilibrium point is reached and no more solute will dissolve, the solution is said to be saturated. A **saturated solution** is a solution that contains the maximum amount of solute that is capable of dissolving. At $20°\mathrm{C}$, the maximum amount of NaCl that will dissolve in \100. \: \text{g}\) of water is $36.0\ \mathrm{g}$. If any more NaCl is added past this point, it will not dissolve, because the solution is saturated. What if more water is added to the solution instead? In that case, more NaCl would be capable of dissolving in the additional solvent. An **unsaturated solution** is a solution that contains less than the maximum amount of solute that is capable of being dissolved. The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

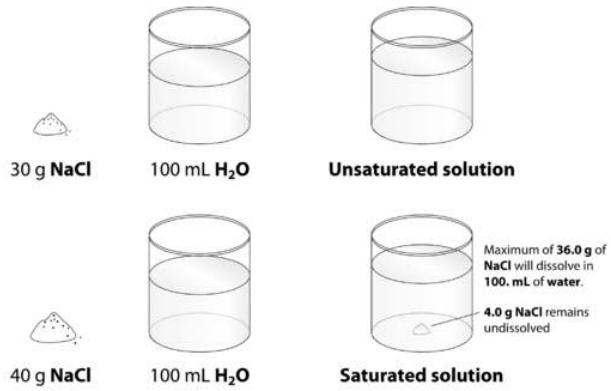


Figure 7.12.1: When $30.0\ \mathrm{g}$ of NaCl is added to $100\ \mathrm{mL}$, it all dissolves, forming an unsaturated solution. When $40.0\ \mathrm{g}$ is added, $36.0\ \mathrm{g}$ dissolves and $4.0\ \mathrm{g}$ remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

Summary

- Recrystallization is the process of dissolved solute returning to the solid state.
- A saturated solution is a solution that contains the maximum amount of solute that is capable of dissolving.
- An unsaturated solution is a solution that contains less than the maximum amount of solute that is capable of being dissolved.
- Solution equilibrium exists when the rate of dissolution equals the rate of recrystallization.

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7.13: How Temperature Influences Solubility

Nuclear power plants require large amounts of water to generate steam for turbines and to cool equipment. They are usually situated near bodies of water to use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per $100\ \text{g}$ of solvent. The solubility of sodium chloride in water is $36.0\ \text{g}$ per $100\ \text{g}$ water at $20°\text{C}$. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. In this section, we will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (see figure below).

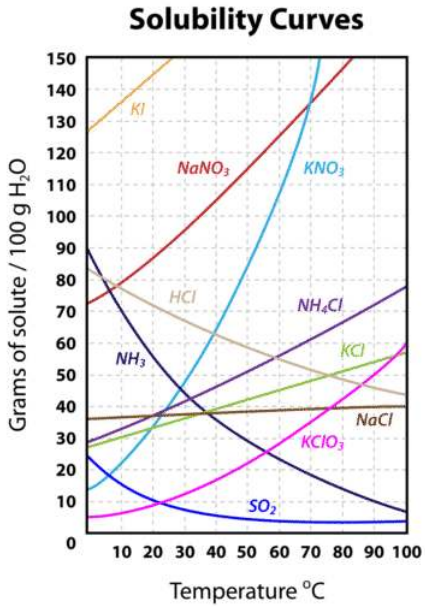


Figure **7.13.1**: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep, and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, NH_3, and SO_2—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ \text{g}$ of KNO_3 is added to $100\ \text{g}$ of water at $30°\text{C}$. According to the solubility curve, approximately $48\ \text{g}$ of KNO_3 will dissolve at $30°\text{C}$. This means that the solution will be saturated since $48\ \text{g}$ is less than $80\ \text{g}$. We can also determine that there will be $80 - 48 = 32\ \text{g}$ of undissolved KNO_3 remaining at the bottom of the container. In a second scenario, suppose that this saturated solution is heated to $60°\text{C}$. According to the curve, the solubility of KNO_3 at $60°\text{C}$ is about $107\ \text{g}$. The solution, in this case, is unsaturated since it contains only the original $80\ \text{g}$ of dissolved solute. Suppose in a third case, that the solution is cooled all the way down to $0°\text{C}$. The solubility at $0°\text{C}$ is about $14\ \text{g}$, meaning that $80 - 14 = 66\ \text{g}$ of the KNO_3 will recrystallize.

Summary

- The solubility of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
- A solubility curve is a graph of the solubility vs. temperature
- The solubility of a solid in water increases with an increase in temperature.
- Gas solubility decreases as the temperature increases.

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7.14: Henry's Law

Having a soft drink in outer space poses some special problems. Under microgravity, the carbonation can quickly dissipate if not kept under pressure. You can't open the can, or you will lose carbonation. So, a special pressurized container has been developed to get around the problem of gas loss at low gravity.

Henry's Law

Pressure has very little effect on the solubility of solids or liquids, but has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the $\mathrm{CO_2}$ molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more $\mathrm{CO_2}$ is added to the space above the container, causing a pressure increase. In this case, more $\mathrm{CO_2}$ molecules are in contact with the water and so more of them dissolve. Thus, the solubility increases as the pressure increases. As with a solid, the $\mathrm{CO_2}$ that is undissolved reaches an equilibrium with the dissolved $\mathrm{CO_2}$, represented by the equation:

$$\mathrm{CO_2}\,(g) \rightleftharpoons \mathrm{CO_2}\,(aq)$$

At equilibrium, the rate of gaseous $\mathrm{CO_2}$ dissolution is equal to the rate of dissolved $\mathrm{CO_2}$ coming out of the solution.

When carbonated beverages are packaged, they are done so under high $\mathrm{CO_2}$ pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is open, the equilibrium is disrupted because the $\mathrm{CO_2}$ pressure above the liquid decreases. Immediately, bubbles of $\mathrm{CO_2}$ rapidly exit the solution and escape out of the top of the open bottle. The amount of dissolved $\mathrm{CO_2}$ decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more $\mathrm{CO_2}$ comes out of the liquid.

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774-1836). **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. Henry's law can be written as follows:

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 and P_1 are the solubility and the pressure at an initial set of conditions; S_2 and P_2 are the solubility and pressure at another changed set of conditions. The solubility of a gas is typically reported in $\mathrm{g/L}$.

Example 7.14.1

The solubility of a certain gas in water is $0.745\ \mathrm{g/L}$ at standard pressure. What is its solubility when the pressure above the solution is raised to $4.50\ \mathrm{atm}$? The temperature is constant at $20^\circ\mathrm{C}$.

Solution

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

Known

- $S_1 = 0.745\ \mathrm{g/L}$
- $P_1 = 1.00\ \mathrm{atm}$
- $P_2 = 4.50\ \mathrm{atm}$

Unknown

Substitute into Henry's law and solve for S_2.

Step 2: Solve.

$$S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.745\ \mathrm{g/L} \times 4.50\ \mathrm{atm}}{1.00\ \mathrm{atm}} = 3.35\ \mathrm{g/L}$$

Step 3: Think about your result.

The solubility is increased to 4.5 times its original value, according to the direct relationship.

Summary

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7.15: Concentration

Learning Objectives

- Calculate percentage concentration (m/m, v/v, m/v), ppm and ppb.
- Calculate the molarity of a solution.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms whose meanings depend on various factors.

Solubility

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 **7.15.1** lists the solubilities of various solutes in water. Solubilities vary with temperature, so Table **7.15.1** includes the temperature at which the solubility was determined.

Table **7.15.1**: Solubilities of Various Solutes in Water at 25°C (Except as Noted)

Substance	Solubility (g in 100 mL of H2O)
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 **7.15.1**. 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.

Precipitation from Supersaturated Solutions

Recrystallization of excess solute from a supersaturated solution usually gives off energy as heat. Commercial heat packs containing supersaturated sodium acetate ($NaC_2H_3O_2$) take advantage of this phenomenon. You can probably find them at your local drugstore.



Video* 7.15.1*: Watered-down sodium acetate trihydrate. Needle crystal is truly wonderful structures

Most solutions we encounter are unsaturated, so knowing the solubility of the solute does not accurately express the amount of solute in these solutions. There are several common ways of specifying the concentration of a solution.

Percent Composition

There are several ways of expressing the concentration of a solution by using a percentage. The **mass/mass percent** (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\,\mathrm{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{7.15.1}$$

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 7.15.1

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:

$$\%\,\mathrm{m/m} = \frac{36.5\ \mathrm{g}}{355\ \mathrm{g}} \times 100\% = 10.3\%$$

Exercise 7.15.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer

7.90%

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** (% v/v): the volume of a solute divided by the volume of a solution times 100:

$$\%\,\mathrm{v/v} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{7.15.2}$$

Again, the units of the solute and the solution must be the same. A hybrid concentration unit, **mass/volume percent** (% m/v), is commonly used for intravenous (IV) fluids (Figure 7.15.1). It is defined as the mass in grams of a solute, divided by volume in milliliters of solution times 100:

$$\%\ \text{m/v} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\% \tag{7.15.3}$$



Figure 7.15.1: Mass/Volume Percent.The 0.9% NaCl concentration on this IV bag is mass/volume percent (left). Such solution is used for other purposes and available in bottles (right). Figures used with permission from Wikipedia

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. 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 = 100 mL solution

The two possible conversion factors are written as follows:

$$\frac{\text{5 mL alcohol}}{\text{100 mL solution}} \text{ or } \frac{\text{100 mL solution}}{\text{5 mL alcohol}}$$

Use the first conversion factor to convert from a given amount of solution to amount of solute. The second conversion factor is used to convert from a given amount of solute to amount of solution. Given any two quantities in any percent composition, the third quantity can be calculated, as the following example illustrates.

Example 7.15.2

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:

$$115\ \text{mL solution} \times \frac{45\ \text{mL}\ C_2H_5OH}{100\ \text{mL solution}} = 51.8\ \text{mL}\ C_2H_5OH$$

Exercise 7.15.2

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 $C_6H_{12}O_6$?

Answer

$$50.0\ \text{g}\ C_6H_12O_6 \times \frac{100\ \text{mL solution}}{12.75\ \text{g}\ C_6H_12O_6} = 392\ \text{mL solution}$$

Example 7.15.3

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.

$$\%\ \text{m/v} = \frac{9.0\ \text{g NaCl}}{1,000\ \text{mL solution}} \times 100\% = 0.90\%\ \text{m/v}$$

Exercise 7.15.3

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?

Answer

$$\%\ \text{m/v} = \frac{27.0\ \text{g NaOCl}}{500.0\ \text{mL solution}} \times 100\% = 5.40\%\ \text{m/v}$$

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:

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000 \tag{7.15.4}$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000,000 \tag{7.15.5}$$

> *Similar to parts per million and parts per billion, related units include parts per thousand (ppth) and parts per trillion (ppt).*

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 7.15.4

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.

Exercise 7.15.4

An 85 kg body contains 0.012 g of Ni. What is the concentration of Ni in parts per million?

Answer

0.14 ppm

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:

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{7.15.6}$$

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

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl} \tag{7.15.7}$$

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**.

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 = 1 L solution

Example

Calculate the molarity of a solution prepared by dissolving 4.25 moles of KOH in enough water to make 2.5 liters of solution.

Solution

To calculate the molar concentration of the solution, the number of moles must be divided by the liters of solution.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{4.25 \text{ moles of KOH}}{2.5 \text{ Liters of solution}} = 1.7\ M \text{ KOH}$$

Exercise

What is the concentration of a 1.75 liter solution if it contains 8.16 moles of solute?

Answer

4.66 M

The Human Connection

Many of the fluids found in our bodies are solutions. The solutes range from simple ionic compounds to complex proteins. Table **7.15.2** lists the typical concentrations of some of these solutes.

Table **7.15.2**: 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.

Looking Closer: The Dose Makes the Poison

Why is it that we can drink 1 qt of water when we are thirsty and not be harmed, but if we ingest 0.5 g of arsenic, we might die? There is an old saying: *the dose makes the poison*. This means that what may be dangerous in some amounts may not be dangerous in other amounts.

Take arsenic, for example. Some studies show that arsenic deprivation limits the growth of animals such as chickens, goats, and pigs, suggesting that arsenic is actually an essential trace element in the diet. Humans are constantly exposed to tiny amounts of arsenic from the environment, so studies of completely arsenic-free humans are not available; if arsenic is an essential trace mineral in human diets, it is probably required on the order of 50 ppb or less. A toxic dose of arsenic corresponds to about 7,000 ppb and higher, which is over 140 times the trace amount that may be required by the body. Thus, arsenic is not poisonous in and of itself. Rather, it is the amount that is dangerous: the dose makes the poison.

Similarly, as much as water is needed to keep us alive, too much of it is also risky to our health. Drinking too much water too fast can lead to a condition called water intoxication, which may be fatal. The danger in water intoxication is not that water itself becomes toxic. It is that the ingestion of too much water too fast dilutes sodium ions, potassium ions, and other salts in the bloodstream to concentrations that are not high enough to support brain, muscle, and heart functions. Military personnel, endurance athletes, and even desert hikers are susceptible to water intoxication if they drink water but do not replenish the salts lost in sweat. As this example shows, even the right substances in the wrong amounts can be dangerous!

Concept Review Exercises

1. What are some of the units used to express concentration?
2. Distinguish between the terms *solubility* and *concentration*.

Answers

1. % m/m, % m/v, ppm, ppb, molarity, and Eq/L (answers will vary)
2. Solubility is typically a limit to how much solute can dissolve in a given amount of solvent. Concentration is the quantitative amount of solute dissolved at any concentration in a solvent.

Concentration. (2021, November 4). https://chem.libretexts.org/@go/page/16090

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7.16: Molecular and Ionic Equations

One of the unfortunate byproducts of industrialized society is acid rain. Sulfur dioxide from burning coal and nitrogen oxides from vehicle emissions both form acids. When these acids react with limestone (calcium carbonate), reactions occur that dissolve the limestone and release water and carbon dioxide. Over a period of time, serious damage is caused to the limestone structure.

Molecular and Ionic Equations

When ionic compounds are dissolved into water, the polar water molecules break apart the solid crystal lattice, resulting in the hydrated ions being evenly distributed through the water. This process is called dissociation and is the reason that all ionic compounds are strong electrolytes. When two different ionic compounds that have been dissolved in water are mixed, a chemical reaction may occur between certain pairs of the hydrated ions. Consider the double-replacement reaction that occurs when a solution of sodium chloride is mixed with a solution of silver nitrate:

$$\mathrm{NaCl}\,(aq) + \mathrm{AgNO_3}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{AgCl}\,(s)$$

The driving force behind this reaction is the formation of the silver chloride precipitate (see figure below).

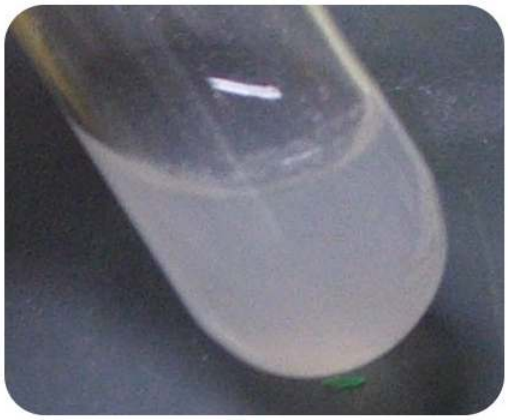


Figure **7.16.1**: The white silver chloride precipitate instantly forms when a solution of silver nitrate is added to a solution of sodium chloride.

This is called a molecular equation. A **molecular equation** is an equation in which the formulas of the compounds are written as though all substances exist as molecules. However, there is a better way to show what is happening in this reaction. All of the aqueous compounds should be written as ions, because they are present in the water as separated ions—a result of their dissociation.

$$\mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) + \mathrm{NO_3^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{NO_3^-}\,(aq) + \mathrm{AgCl}\,(s)$$

This equation is called an **ionic equation**, an equation in which dissolved ionic compounds are shown as free ions.

Some other double-replacement reactions do not produce a precipitate as one of the products. The production of a gas and/or a molecular compound such as water may also drive the reaction. For example, consider the reaction of a solution of sodium carbonate with a solution of hydrochloric acid (HCl). The products of the reaction are aqueous sodium chloride, carbon dioxide, and water. The balanced molecular equation is:

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

The ionic equation is:

$$2\mathrm{Na^+}\,(aq) + \mathrm{CO_3^{2-}}\,(aq) + 2\mathrm{H^+}\,(aq) + 2\mathrm{Cl^-}\,(aq) \rightarrow 2\mathrm{Na^+}\,(aq) + 2\mathrm{Cl^-}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

A single-replacement reaction is one in which an element replaces another element in a compound. An element is in either the solid, liquid, or gas state and is not an ion. The example below shows the reaction of solid magnesium metal with aqueous silver nitrate to form aqueous magnesium nitrate and silver metal.

Balanced molecular equation:

$$\mathrm{Mg}\,(s) + 2\mathrm{AgNO_3}\,(aq) \rightarrow \mathrm{Mg(NO_3)_2}\,(aq) + 2\mathrm{Ag}\,(s)$$

Ionic equation:

$$\mathrm{Mg}\,(s) + 2\mathrm{Ag^+}\,(aq) + 2\mathrm{NO_3^-}\,(aq) \rightarrow \mathrm{Mg^{2+}}\,(aq) + 2\mathrm{NO_3^-}\,(aq) + \mathrm{Ag}\,(s)$$

This type of single-replacement reaction is called a metal replacement. Other common categories of single-replacement reactions are: hydrogen replacement and halogen replacement.

Summary

- A molecular equation is an equation in which the formulas of the compounds are written as though all substances exist as molecules.
- An ionic equation is one in which dissolved ionic compounds are shown as free ions.
- Examples of molecular and ionic equations are shown.

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7.17: Net Ionic Equations

At sports events around the world, a small number of athletes fiercely compete on fields and in stadiums. They get tired, dirty, and sometimes hurt as they try to win the game. Surrounding them are thousands of spectators watching and cheering. Would the game be different without the spectators? Definitely! Spectators provide encouragement to the team and generate enthusiasm. Although the spectators are not playing the game, they are certainly a part of the process.

Net Ionic Equations

We can write a molecular equation for the formation of silver chloride precipitate:

$$\mathrm{NaCl} + \mathrm{AgNO_3} \rightarrow \mathrm{NaNO_3} + \mathrm{AgCl}$$

The corresponding ionic equation is:

$$\mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) + \mathrm{NO_3^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{NO_3^-}\,(aq) + \mathrm{AgCl}\,(s)$$

If you look carefully at the ionic equation, you will notice that the sodium ion and the nitrate ion appear unchanged on both sides of the equation. When the two solutions are mixed, neither the $\mathrm{Na^+}$ nor the $\mathrm{NO_3^-}$ ions participate in the reaction. They can be eliminated from the reaction.

$$\cancel{\mathrm{Na^+}\,(aq)} + \mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) + \cancel{\mathrm{NO_3^-}\,(aq)} \rightarrow \cancel{\mathrm{Na^+}\,(aq)} + \cancel{\mathrm{NO_3^-}\,(aq)} + \mathrm{AgCl}\,(s)$$

A **spectator ion** is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction. In the above reaction, the sodium ion and the nitrate ion are both spectator ions. The equation can now be written without the spectator ions:

$$\mathrm{Ag^+}\,(aq) + \mathrm{Cl^-}\,(aq) \rightarrow \mathrm{AgCl}\,(s)$$

The **net ionic equation** is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction. Notice that in writing the net ionic equation, the positively-charged silver cation was written first on the reactant side, followed by the negatively-charged chloride anion. This is somewhat customary because that is the order in which the ions must be written in the silver chloride product. However, it is not absolutely necessary to order the reactants in this way.

Net ionic equations must be balanced by both mass and charge. Balancing by mass means ensuring that there are equal masses of each element on the product and reactant sides. Balancing by charge means making sure that the overall charge is the same on both sides of the equation. In the above equation, the overall charge is zero, or neutral, on both sides of the equation. As a general rule, if you balance the molecular equation properly, the net ionic equation will end up being balanced by both mass and charge.

Example **7.17.1**

When aqueous solutions of copper (II) chloride and potassium phosphate are mixed, a precipitate of copper (II) phosphate is formed. Write a balanced net ionic equation for this reaction.

Solution

Step 1: Plan the problem.

Write and balance the molecular equation first, making sure that all formulas are correct. Then write the ionic equation, showing all aqueous substances as ions. Carry through any coefficients. Finally, eliminate spectator ions and write the net ionic equation.

Step 2: Solve.

Molecular equation:

$$3\mathrm{CuCl_2}\,(aq) + 2\mathrm{K_3PO_4}\,(aq) \rightarrow 6\mathrm{KCl}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Ionic equation:

$$3\mathrm{Cu^{2+}}\,(aq) + 6\mathrm{Cl^-}\,(aq) + 6\mathrm{K^+}\,(aq) + 2\mathrm{PO_4^{3-}}\,(aq) \rightarrow 6\mathrm{K^+}\,(aq) + 6\mathrm{Cl^-}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Notice that the balance of the equation is carried through when writing the dissociated ions. For example, there are six chloride ions on the reactant side because the coefficient of 3 is multiplied by the subscript of 2 on the copper (II) chloride formula. The spectator ions, K^+ and Cl^-, can be eliminated.

Net ionic equation:

$$3Cu^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Cu_3(PO_4)_2(s)$$

Step 3: Think about your result.

For a precipitation reaction, the net ionic equation always shows the two ions that come together to form the precipitate. The equation is balanced by mass and charge.

Summary

- A spectator ion is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction.
- The net ionic equation is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction.
- An example of writing a net ionic equation is outlined.

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7.18: Predicting Precipitates Using Solubility Rules

Predicting the weather is tricky business. A thorough examination of a large amount of data is needed to make the daily forecast. Wind patterns, historical data, barometric pressure—these and many other data are fed into computers that then use a set of rules to predict what will happen based on past history.

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all.

It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (shown in Table 7.18.1).

Table 7.18.1: *Solubility Rules for Ionic Compounds in Water*

Soluble	Compounds containing the alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+) and ammonium ion (NH_4^+).
Soluble	Compounds containing the nitrate ion (NO_3^-), acetate ion (CH_3COO^-), chlorate ion (ClO_3^-), and bicarbonate ion (HCO_3^-).
Mostly Soluble	Compounds containing the chloride ion (Cl^-), bromide ion (Br^-), and iodide ion (I^-). Exceptions are those of silver (Ag^+), mercury (I) (Hg_2^{2+}), and lead (II) (Pb^{2+}).
Mostly Soluble	Compounds containing the sulfate ion (SO_4^{2-}). Exceptions are those of silver (Ag^+), calcium (Ca^{2+}), strontium (Sr^{2+}), barium (Ba^{2+}), mercury (I) (Hg_2^{2+}), and lead (II) (Pb^{2+}).
Mostly Insoluble	Compounds containing the carbonate ion (CO_3^{2-}), phosphate ion (PO_4^{3-}), chromate ion (CrO_4^{2-}), sulfide ion (S^{2-}), and silicate ion (SiO_3^{2-}). Exceptions are those of the alkali metals and ammonium.
Mostly Insoluble	Compounds containing the hydroxide ion (OH^-). Exceptions are those of the alkali metals and the barium ion (Ba^{2+}).

For practice using the solubility rules, predict if a precipitate will form when solutions of cesium bromide and lead (II) nitrate are mixed.

$$Cs^+(aq) + Br^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow ?$$

The potential precipitates from a double-replacement reaction are cesium nitrate and lead (II) bromide. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, the cesium and nitrate ions are spectator ions and the lead (II) bromide is a precipitate. The balanced net ionic reaction is:

$$Pb^{2+}(aq) + 2Br^-(aq) \rightarrow PbBr_2(s)$$

Summary

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7.19: Chapter 7 Exercises

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7.19.1: Solutions (Exercises)

Attributions

Exercises selected from the following sources. Some questions have been modified and/or added by Kristina Novek

- *Textmap created for Chemistry: A Molecular Approach by Nivaldo Tro.*
- ***Sacramento City College, CHEM 309: General, Organic and Biochemistry Prof. Dianne Bennett***
- *Textmap created for "Principles of Modern Chemistry" by Oxtoby et al.*
- *Chapter 11 Solutions and Colloids (Exercises) Contributed by Paul Flowers, Klaus Theopold & Richard Langley et al. General Chemistry at OpenStax CNX*
- *Chapter 9 Book: The Basics of GOB Chemistry (Ball et al.)*
- *Complementary General Chemistry question banks can be found for other Textmaps and can be accessed here.*

Water, Solutions, & Electrolytes

1. Identify if the statements about the properties of water are true or false.
   a) Ice floats in liquid water because of its lower density
   b) Water contains oxygen atoms with a slight positive charge
   c) Water is considered a nonpolar molecule
   d) Hydrogen bonding creates high surface tension

2. Identify each as either an example of (a) gas in a liquid, (b) gas in a gas, (c) solid in a solid, or (d) liquid in a liquid.
   I. bronze (solution of tin or other metals in copper)
   II. O_2 in N_2 (air)
   III. CO_2 in water

3. Predict whether each of the following substances would be more soluble in water (polar solvent) or in a hydrocarbon such as heptane (C_7H_{16}, nonpolar solvent):
   a) vegetable oil (nonpolar)
   b) isopropyl alcohol (polar)
   c) potassium bromide (ionic)

4. Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules.

5. Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.

6. What is the expected electrical conductivity of the following solutions?
   a) NaOH(*aq*)

b) $HCl(aq)$

c) $C_6H_{12}O_6(aq)$ (glucose)

d) $NH_3(l)$

7. Answer the following True or False statements.

a) In a solution, the solvent is the substance being dissolved

b) An electrolyte is another word for molecule

c) A solution is a pure substance

d) Homogenous is the best way to describe a solution

e) A solute is all that is needed to make a solution

8. Of the following examples of solutions, determine whether the solute is an electrolyte or a molecule.

a) Plain soda (carbonated) water

b) Apple juice from concentrate

c) A bottle of rubbing alcohol (2-propanol)

d) A saline solution of KCl

9. It is important to distinguish between the solute and the solvent. For each example that is a solution, identify which substance is the solute and which is the solvent.

a) carbonated water

b) a saline solution of sodium chloride

c) sugar water for a bird feeder

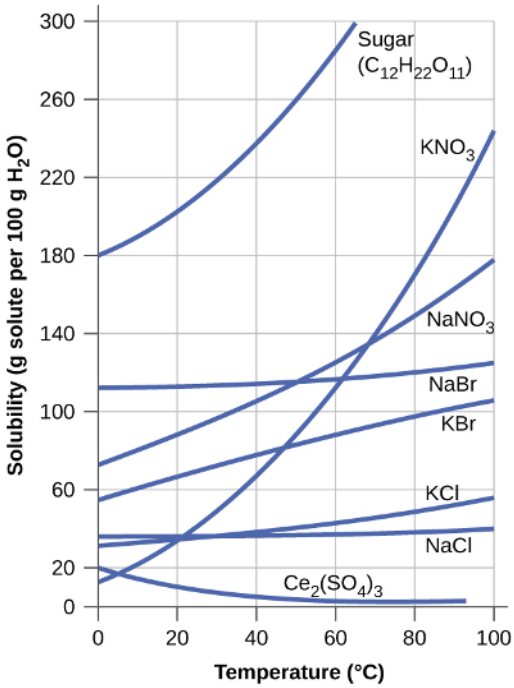
d) a sodium bicarbonate buffer solution

e) oxygen in air

Solubility

1. Define *solubility*. Do all solutes have the same solubility?

Solubility of Solids in Water



2. Use the solubility graph above to answer the following questions.

a) At what temperature is a 60 gram, KBr solution saturated?

b) If 140 grams of NaBr was added to 100 g of water at 40^0C, how much would not dissolve?

c) If 15 grams of KCl was added to 100 g of water at 25^0C would the solution be saturated or unsaturated?

d) A solution of KNO_3 contained 50 grams at 70 ^{0}C. How many additional grams of KNO_3 would have to be added to create a saturated solution?

e) A solution of KBr contains 40 grams of KBr in only 50 grams of water. The temperature of the solution 20^0C. Is the solution saturated or unsaturated?

f) If a solution containing 115 grams of $NaNO_3$ was cooled from 90^0C to 20^0C, how much would precipitate out (assuming a supersaturate solution did not form)?

3. Which of the following would allow a greater amount of carbon dioxide to dissolve in water?

a) Decrease the temperature

b) Increase the temperature

c) Add a catalyst

d) Decrease the pressure

Concentration

1. Explain why the terms *dilute* or *concentrated* are of limited usefulness in describing the concentration of solutions.

Table 7.16.1: Solubilities of Various Solutes in Water at 25°C (Except as Noted)

Substance	Solubility (g in 100 mL of H2O)
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)

Table Retrieved from:*Chapter 9 Section 9.2 Concentration* *Book: The Basics of GOB Chemistry (Ball et al.)*

2. Use the data from the above table to answer the following questions.

   a) Calculate the mass/mass percent of a saturated solution of NaCl. Assume that masses of the solute and the solvent are additive and use the density of H_2O (1.00 g/mL) as a conversion factor.

   b) Calculate the mass/mass percent of a saturated solution of $MgCO_3$. Assume that masses of the solute and the solvent are additive and use the density of H_2O (1.00 g/mL) as a conversion factor.

3. Only 1.36 mL of C_6H_6 will dissolve in 100.000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of benzene in water.

4. Only 35 mL of aniline ($C_6H_5NH_2$) will dissolve in 1,000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of aniline in water.

5. A solution of ethyl alcohol (C_2H_5OH) in water has a concentration of 20.56% v/v. What volume of C_2H_5OH is present in 255 mL of solution?

6. What mass of KCl is present in 475 mL of a 1.09% m/v aqueous solution?

7. The average human body contains 5,830 g of blood. What mass of arsenic is present in the body if the amount in blood is 0.55 ppm?

8. The Occupational Safety and Health Administration has set a limit of 200 ppm as the maximum safe exposure level for carbon monoxide (CO). If an average breath has a mass of 1.286 g, what is the maximum mass of CO that can be inhaled at that maximum safe exposure level?

9. Which concentration is greater—15 ppm or 1,500 ppb?

10. Express the concentration 7,580 ppm in parts per billion.

11. What is the molarity of 0.500 L of a potassium chromate solution containing 0.0650 mol of K_2CrO_4?

12. What is the molarity of 4.50 L of a solution containing 0.206 mol of urea $[(NH_2)_2CO]$?

13. What is the molarity of a 2.66 L aqueous solution containing 0.96 moles of NaBr?

14. If 3.08 moles of $Ca(OH)_2$ is dissolved in enough water to make 0.875 L of solution, what is the molarity of the $Ca(OH)_2$?

Precipitates

1. Rewrite the following balanced equations as net ionic equations.

a) $2Na_3PO_4(aq) + 3CaCl_2(aq) \rightarrow 6NaCl(aq) + Ca3(PO_4)_2(s)$

b) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

c) $2AgNO_3(aq) + ZnCl_2(aq) \rightarrow 2AgCl(s) + Zn(NO_3)_2(aq)$

2. When the following solutions are mixed, does a precipitate form? Write out the total molecular, total ionic, and net ionic equations.

a) lead (II) nitrate and potassium chloride

b) magnesium nitrate and calcium chloride

c) calcium chloride and sodium carbonate

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7.19.2: Answers to Solutions (Exercises)

Water, Solutions, & Electrolytes

1. (a) true; (b) false; (c) false; (d) true

2. (I) c (II) b (III) a

3. (a) heptane; (b) water; (c) water

4. Crystals of NaCl dissolve in water, a polar liquid with a very large dipole moment, and the individual ions become strongly solvated. Hexane is a nonpolar liquid with a dipole moment of zero and, therefore, does not significantly interact with the ions of the NaCl crystals.

5. HBr is an acid and so its molecules react with water molecules to form H_3O^+ and Br^- ions that provide conductivity to the solution. Though HBr is soluble in benzene, it does not react chemically but remains dissolved as neutral HBr molecules. With no ions present in the benzene solution, it is electrically nonconductive.

6.

(a) high conductivity (solute is an ionic compound that will dissociate when dissolved)

(b) high conductivity (solute is a strong acid and will ionize completely when dissolved)

(c) nonconductive (solute is a covalent compound, neither acid nor base, unreactive towards water)

(d) low conductivity (solute is a weak base and will partially ionize when dissolved)

7. (a) true (b) false (c) false (d) true (e) false

8. (a) molecule (b) electrolyte (c) molecule (d) electrolyte

9.

(a) solute: CO_2; solvent: water

(b) solute: NaCl; solvent: water

(c) solute: $C_{12}H_{22}O_{11}$ (sucrose); solvent: water

(d) solute: $NaHCO_3$ (baking soda); solvent: water

(e) solute: O_2; solvent: N_2

Solubility

1. Solubility is the amount of a solute that can dissolve in a given amount of solute, typically 100 mL. The solubility of solutes varies widely.

2.

a. ~10^0C

b. ~ 30 grams of NaBr

c. unsaturated

d. 90 grams of KNO_3

e. The KBr solution is saturated since the amount of water is only 50 grams. Therefore, the saturation point at 20^0C is 33 grams of KBr in 50 grams of water.

f. ~25 grams of $NaNO_3$ would precipitate out.

3. a

Concentration

1.The term *dilute* means relatively less solute and the term *concentrated* implies relatively more solute. Both are of limited usefulness because these are not accurate.

2. a. 26.5% NaCl m/m% solution b. 2.15% $MgCO_3$ m/m% solution

3. 1.34% C_6H_6 v/v% solution

4. 3.4% $C_6H_5NH_2$ v/v% solution

5. 52.4 mL C_2H_5OH

6. 5.18 g

7. 0.00321 g

8. 2.57 x 10^{-4} g

9. 15 ppm

10. 7,580,000 ppb

11. 0.130 M

12. 0.0458 M

13. 0.36 M

14. 3.52 M

Precipitates

1.

a. $2PO_4^{3-}(aq) + 3Ca^{2+}(aq) \rightarrow Ca_3(PO_4)_2(s)$

b. $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

c. $Ag^{+}(aq) + Cl^{-} \rightarrow AgCl(s)$

2.

a. Molecular: $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

Ionic: $Pb^{2+}(aq) + 2NO_3^{-}(aq) + 2K^{+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s) + 2K^{+}(aq) + 2NO_3^{-}(aq)$

Net Ionic: $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

b. no precipitate forms (no reaction)

c. Molecular: $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$

Ionic: $Ca^{+2}(aq) + 2Cl^{-1}(aq) + 2Na^{+1}(aq) + CO_3^{-2}(aq) \rightarrow 2Na^{+1}(aq) + 2Cl^{-1}(aq) + CaCO_3(s)$

Net Ionic: $Ca^{+2}(aq) + CO_3^{-2}(aq) \rightarrow CaCO_3(s)$

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CHAPTER OVERVIEW

8: Acids and Bases

8.1: Prelude to Acids and Bases
8.2: Arrhenius Definition of Acids and Bases
8.3: Brønsted-Lowry Definition of Acids and Bases
8.4: Water - Both an Acid and a Base
8.5: The Strengths of Acids and Bases
8.6: Buffers
8.7: Chapter 8 Exercises
8.7.1: Acids and Bases (Exercises)
8.7.2: Acids and Bases (Summary)

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8.1: Prelude to Acids and Bases

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. 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.



Figure **8.1.1**: Stomach with acidic digestive juices. (CC BY-SA 4.0; BruceBlaus).

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.

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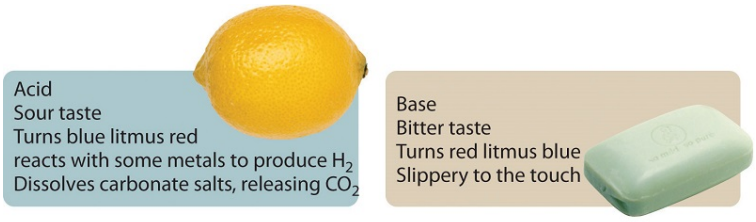
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8.2: Arrhenius Definition of Acids and Bases

Learning Objectives

- To recognize a compound as an Arrhenius acid or an Arrhenius base.
- To describe characteristics of acids and bases.
- To write equations of neutralization reactions.

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.



Although we include their tastes among the common characteristics of acids and bases, we never advocate tasting an unknown chemical!

Chemists prefer, however, to have definitions for 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. Many acids are simple compounds that release a hydrogen cation into solution when they dissolve. Similarly, Arrhenius defined a base as a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution. Many bases are ionic compounds that have the hydroxide ion as their anion, which is released when the base dissolves in water.

Table **8.2.1**: Formulas and Names for Some Acids and Bases

Acids		Bases	
Formula	Name	Formula	Name
HCl(aq)	hydrochloric acid	NaOH(aq)	sodium hydroxide
HBr(aq)	hydrobromic acid	KOH(aq)	potassium hydroxide
HI(aq)	hydriodic acid	$Mg(OH)_2$(aq)	magnesium hydroxide
H_2S(aq)	hydrosulfuric acid	$Ca(OH)_2$(aq)	calcium hydroxide
$HC_2H_3O_2$(aq)	acetic acid	NH_3(aq)	ammonia
HNO_3(aq)	nitric acid	$NaHCO_3$ (aq)	sodium bicarbonate
HNO_2(aq)	nitrous acid	$CaCO_3$ (aq)	calcium carbonate
H_2SO_4(aq)	sulfuric acid		
H_2SO_3(aq)	sulfurous acid		
$HClO_3$(aq)	chloric acid		
$HClO_4$(aq)	perchloric acid		
$HClO_2$(aq)	chlorous acid		
H_3PO_4(aq)	phosphoric acid		
H_3PO_3(aq)	phosphorous acid		
H_2CO_3(aq)	carbonic acid		

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. For example, the base sodium hydroxide (NaOH) is both an ionic compound and an aqueous solution. However, aqueous solutions of acids have their own naming rules. 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. 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. Other prefixes, like *per-* and *hypo-*, also appear in the names for some acids. 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 **8.2.1** lists some acids and bases and their names. 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.

> *The name oxygen comes from the Latin meaning "acid producer" because its discoverer, Antoine Lavoisier, thought it was the essential element in acids. Lavoisier was wrong, but it is too late to change the name now.*

Example 8.2.1

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.

Exercise 8.2.1

Name each substance.

a. $H_2Se(aq)$
b. $Ba(OH)_2(aq)$

Answer

a. hydroselenic acid

b. barium hydroxide

Notice that one base listed in Table **8.2.1**—ammonia—does not have hydroxide as part of its formula. How does this compound increase the amount of hydroxide ion in aqueous solution? Instead of dissociating into hydroxide ions, ammonia molecules react with water molecules by taking a hydrogen ion from the water molecule to produce an ammonium ion and a hydroxide ion:

$$NH_{3(aq)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{8.2.1}$$

Because this reaction of ammonia with water causes an increase in the concentration of hydroxide ions in solution, ammonia satisfies the Arrhenius definition of a base. Many other nitrogen-containing compounds are bases because they too react with water to produce hydroxide ions in aqueous solution.

Neutralization

As we noted previously, acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen ion containing compound and the base is a hydroxide ion containing compound, water is also a product. The general reaction is as follows:

$$\text{acid + base } \rightarrow \text{ water + salt}$$

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:

$$NaOH(aq) + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)} \tag{8.2.2}$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(\ell)} \tag{8.2.3}$$

Once a neutralization reaction is properly balanced, we can use it to perform stoichiometry calculations, such as the ones we practiced earlier.

There are a number of examples of acid-base chemistry in everyday life. 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.

$$HCO^-_3(aq) + H^+(aq) \rightarrow H_2CO_3(aq) \tag{8.2.4}$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l) \tag{8.2.5}$$

Example 8.2.2

Nitric acid [$HNO_3(aq)$] can be neutralized by calcium hydroxide [$Ca(OH)_2(aq)$].

Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Solution

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.

Exercise 8.2.2

Hydrocyanic acid [HCN(aq)] can be neutralized by potassium hydroxide [KOH(aq)].

Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Answer

$KOH(aq) + HCN(aq) \rightarrow KCN(aq) + H_2O(\ell)$ KCN is the salt

Hydrocyanic acid (HCN) is one exception to the acid-naming rules that specify using the prefix *hydro-* for binary acids (acids composed of hydrogen and only one other element).

Stomach Antacids

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$CaCO_3(s) + 2HCl(aq) \rightleftharpoons CaCl_2(aq) + H_2O(l) + CO_2(g) \tag{8.2.6}$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq) \tag{8.2.7}$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that :

$$H_3O^+ + OH^- \rightleftharpoons 2H_2O(l) \tag{8.2.8}$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

Concept Review Exercises

1. Give the Arrhenius definitions of an acid and a base.
2. What is neutralization?

Answers

1. Arrhenius acid: a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution; Arrhenius base: a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.
2. the reaction of an acid and a base

Key Takeaway

- An Arrhenius acid increases the H^+ ion concentration in water, while an Arrhenius base increases the OH^- ion concentration in water.

Exercises

1. Give two examples of Arrhenius acids.
2. Give two examples of Arrhenius bases.
3. List the general properties of acids.
4. List the general properties of bases.
5. Name each compound. (For acids, look up the name in Table 10.1.1. For bases, use the rules for naming ionic compounds from Chapter 3.)
    a. HBr(aq)
    b. $Ca(OH)_2$(aq)
    c. HNO_3(aq)
    d. $Fe(OH)_3$(aq)
6. Name each compound.
    a. HI(aq)
    b. $Cu(OH)_2$(aq)
    c. H_3PO_4(aq)
    d. CsOH(aq)
7. Write a balanced chemical equation for the neutralization of $Ba(OH)_2$(aq) with HNO_3(aq).
8. Write a balanced chemical equation for the neutralization of H_2SO_4(aq) with $Cr(OH)_3$(aq).
9. Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction.
10. Identify the salt produced in each acid-base reaction below. Then, balance the equation.
    a. $2HCl + Sr(OH)_2 \rightarrow 2H_2O$ + ??
    b. KNO_3; $HNO_3 + KOH \rightarrow$?? + H_2O
    c. $HF + Ca(OH)_2$ ---> ?? + H_2O

Answers

1. HCl and HNO_3 (answers will vary)
2. NaOH and $Ca(OH)_2$ (answers will vary)
3. sour taste, react with metals, react with bases, and turn litmus red
4. bitter taste, feels slippery, react with acids and turn litmus blue
5. a. hydrobromic acid
    b. calcium hydroxide

c. nitric acid

d. iron(III) hydroxide

6. a. hydroiodic acid

b. cupric hydroxide

c. phosphoric acid

d. cesium hydroxide

7. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O$

8. $3H_2SO_4(aq) + 2Cr(OH)_3(aq) \rightarrow Cr_2(SO_4)_3(aq) + 6H_2O$

9. $Mg(OH)_2 + 2HCl$ --> $MgCl_2 + 2H_2O$

10. a. $SrCl_2$; $2HCl + Sr(OH)_2 \rightarrow 2H_2O + SrCl_2$

b. KNO_3; $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

c. CaF_2; $2HF + Ca(OH)_2 \rightarrow CaF_2 + 2H_2O$

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8.3: Brønsted-Lowry Definition of Acids and Bases

Learning Objectives

1. Recognize a compound as a Brønsted-Lowry acid or a Brønsted-Lowry base.
2. Illustrate the proton transfer process that defines a Brønsted-Lowry acid-base reaction.

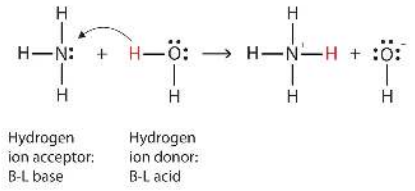
Ammonia (NH_3) increases the hydroxide ion concentration in aqueous solution by reacting with water rather than releasing hydroxide ions directly. In fact, the Arrhenius definitions of an acid and a base focus on hydrogen ions and hydroxide ions. Are there more fundamental definitions for acids and bases?

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.

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:

$$NH_3(aq) + H_2O(\ell) \rightarrow NH_4^+(aq) + OH^-(aq) \tag{8.3.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:



Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

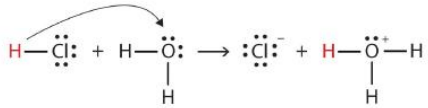
Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

> 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:

$$HCl(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + Cl^-(aq) \tag{8.3.2}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, **HCl acts as a Brønsted-Lowry acid**; as a hydrogen ion acceptor, **H_2O is a Brønsted-Lowry base**. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

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 8.3.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$C_6H_5NH_2(aq) + H_2O(\ell) \rightarrow C_6H_5NH_3^+(aq) + OH^-(aq)$$

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.

? Exercise 8.3.1

Caffeine ($C_8H_{10}N_4O_2$) is a stimulant found in coffees and teas. When dissolved in water, it can accept a proton from a water molecule. Write the chemical equation for this process and identify the Brønsted-Lowry acid and base.

Answer

$$C_8H_{10}N_4O_2(aq) + H_2O(\ell) \rightarrow C_8H_{11}N_4O_2^+(aq) + OH^-(aq)$$

B-L base B-L acid

The Brønsted-Lowry definitions of an acid and a base can be applied to chemical reactions that occur in solvents other than water. The following example illustrates.

✔ Example 8.3.2

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.

? Exercise 8.3.2

Pyridinium chloride (C_5H_5NHCl) dissolves in ethanol (C_2H_5OH) and separates into pyridinium ions ($C_5H_5NH^+$) and chloride ions. The pyridinium ion can transfer a hydrogen ion to a solvent molecule. Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

Answer

$C_5H_5NH^+(solv) + C_2H_5OH(\ell) \rightarrow C_5H_5N(solv) + C_2H_5OH_2^+(solv)$

B-L acid B-L base

Application in Everyday Life

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure **8.3.1**). 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.



$CH_3COOH + NH_2CH_2CH_2CH_2CH_2NH_2 \longrightarrow CH_3COO^- + NH_3^+CH_2CH_2CH_2CH_2NH_2$

Acetic acid + Putrescine → Acetate ion + Putrescinium ion

Figure **8.3.1**: 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. (CC BY 4.0; OpenStax)

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.

To Your Health: 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 used with permission from Wikipedia.

Conjugate Acid-Base Pairs

According to the Bronsted-Lowry theory of acids and bases, 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^-.

$$\mathrm{Acid = H^+ + Conjugate\ base\ of\ Acid^-}$$

$$\mathrm{HF \rightleftharpoons H^+ + F^-}$$

$$\mathrm{H_2O \rightleftharpoons H^+ + OH^-}$$

$$\mathrm{NH_4^+ \rightleftharpoons H^+ + NH_3}$$

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.

$$\mathrm{H^+ + Base = Conjugate\ acid\ of\ Base^+}$$

$$\mathrm{H^+ + F^- \rightleftharpoons HF}$$

$$\mathrm{H^+ + OH^- \rightleftharpoons H_2O}$$

$$\mathrm{H^+ + H_2O \rightleftharpoons H_3O^+}$$

$$\mathrm{H^+ + NH_3 \rightleftharpoons NH_4^+}$$

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 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 **8.3.1**. Conjugate acid-base pairs.

Conjugate Acid	Conjugate Base
H_3O^+	H_2O
H_2O	OH^-
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 8.3.3 : 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.

8.3.3Exercise : Conjugate Pairs

Write the formula of the conjugate base of (a) HNO_2 and (b) H_2CO_3.

Write the formula of the conjugate acid of (c) C_6H5NH_2 and (d) HCO_3^-.

Answer

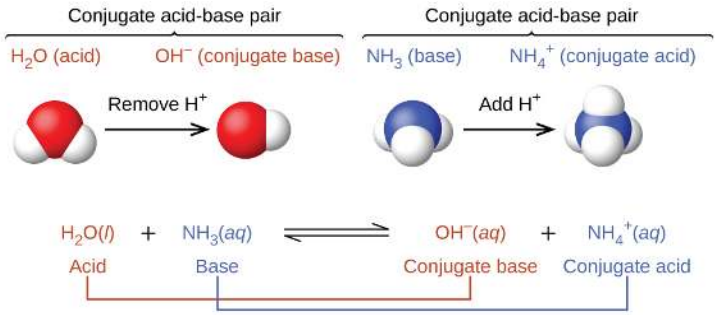
a. NO_2^- is the conjugate base of HNO_2.

b. HCO_3^- is the conjugate base of H_2CO_3

c. $C_6H5NH_3^+$ is the conjugate acid of C_6H5NH_2.

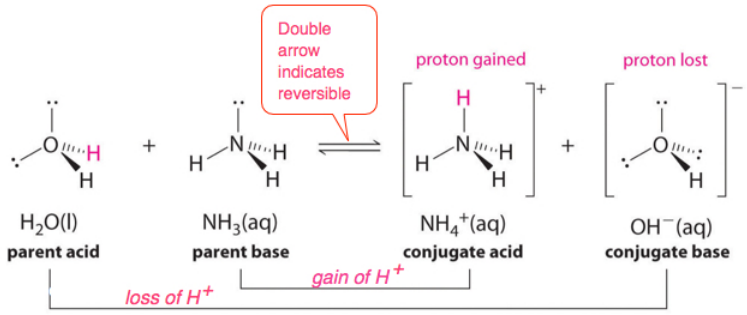
d. H_2CO_3 is the conjugate acid of HCO_3^-

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The **conjugate acid–base pairs** for this reaction are $\boldsymbol{NH_4^+/NH_3}$ and $\boldsymbol{H_2O/OH^-}$. 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.



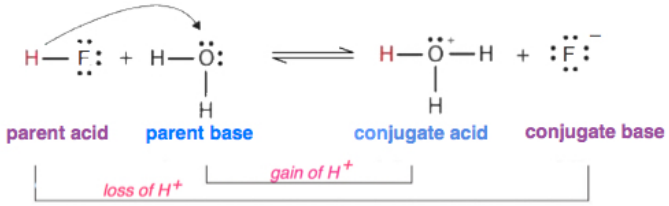
This figure has three parts in two rows. In the first row, two diagrams of acid-base pairs are shown. On the left, a space filling model of H subscript 2 O is shown with a red O atom at the center and two smaller white H atoms attached in a bent shape. Above this model is the label “H subscript 2 O (acid)” in purple. An arrow points right, which is labeled “Remove H superscript plus.” To the right is another space filling model with a single red O atom to which a single smaller white H atom is attached. The label in purple above this model reads, “O H superscript negative (conjugate base).” Above both of these red and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” To the right is a space filling model with a central blue N atom to which three smaller white H atoms are attached in a triangular pyramid arrangement. A label in green above reads “N H subscript 3 (base).” An arrow labeled “Add H superscript plus” points right. To the right of the arrow is another space filling model with a blue central N atom and four smaller white H atoms in a tetrahedral arrangement. The green label above reads “N H subscript 3 superscript plus (conjugate acid).” Above both of these blue and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” The second row of the figure shows the chemical reaction, H subscript 2 O (l) is shown in purple, and is labeled below in purple as “acid,” plus N H subscript 3 (a q) in green, labeled below in green as “base,” followed by a double sided arrow arrow and O H superscript negative (a q) in purple, labeled in purple as “conjugate base,” plus N H subscript 4 superscript plus (a q)” in green, which is labeled in green as “conjugate acid.” The acid on the left side of the equation is connected to the conjugate base on the right with a purple line. Similarly, the base on the left is connected to the conjugate acid on the right side.

In the **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.



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):

$$HF + H_2O \rightleftharpoons H_3O^+ + F^-$$



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Example 8.3.3

Identify the conjugate acid-base pairs in this equilibrium.

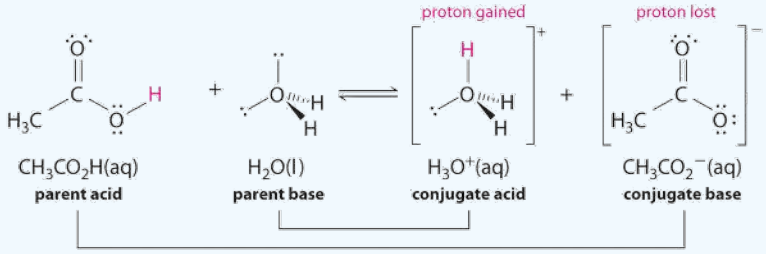
$$CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-$$

Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid–base pairs:

- the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and
- the parent base and its conjugate acid (H_3O^+/H_2O).



Example 8.3.4

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.

Exercise 8.3.4

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer:

H_2O (acid) and OH^- (base); NH_2^- (base) and NH_3 (acid)

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.*

Key Takeaways

- 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.

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8.4: Water - Both an Acid and a Base

Learning Objectives

- To write chemical equations for water acting as an acid and as a base.

Water (H_2O) is an interesting compound in many respects. Here, we will consider its ability to behave as an acid or 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:

$$\mathrm{HCl} + \mathrm{H_2O}(\ell) \rightarrow \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)} \tag{8.4.1}$$

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 (see Example 4 in Section 10.2), a water molecule donates a proton, making ammonia as a product:

$$H_2O(\ell) + NH_2^-(aq) \rightarrow OH^-(aq) + NH_3(aq) \tag{8.4.2}$$

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.

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:

$$H_2O(\ell) + H_2O(\ell) \rightarrow H_3O^+(aq) + OH^-(aq) \tag{8.4.3}$$

This process is called the **autoionization of water** (Figure 8.4.1) 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.)

Note

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.

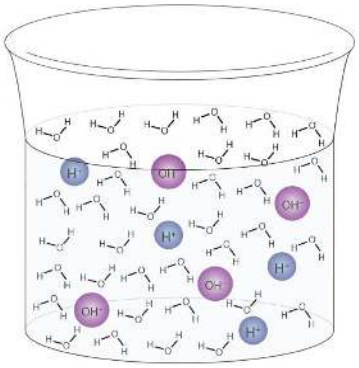


Figure 8.4.1 Autoionization. A small fraction of water molecules—approximately 6 in 100 million—ionize spontaneously into hydronium ions and hydroxide ions. This picture necessarily overrepresents the amount of autoionization that really occurs in pure water.

Example 8.4.1

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $H_2O(\ell) + NO_2^-(aq) \rightarrow HNO_2(aq) + OH^-(aq)$
2. $HC_2H_3O_2(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$

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.

Exercise 8.4.2

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $HCOOH(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + HCOO^-(aq)$
2. $H_2O(\ell) + PO_4^{3-}(aq) \rightarrow OH^-(aq) + HPO_4^{2-}(aq)$

Answer

1. H_2O acts as the proton acceptor (Brønsted-Lowry base)
2. H_2O acts as the proton donor (Brønsted-Lowry acid)

Key Takeaway

- Water molecules can act as both an acid and a base, depending on the conditions.

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8.5: The Strengths of Acids and Bases

Learning Objectives

- Describe the difference between strong and weak acids and bases.
- Describe how a chemical reaction reaches chemical equilibrium.
- Define the pH scale and use it to describe acids and bases.

Acids and bases do not all demonstrate the same degree of chemical activity in solution. Different acids and bases have different strengths.

Strong and Weak Acids

Let us consider the strengths of acids first. A small number of acids ionize completely in aqueous solution. For example, when HCl dissolves in water, every molecule of HCl separates into a hydronium ion and a chloride ion:

$$\mathrm{HCl(g) + H_2O(l) \xrightarrow{\sim 100\,\%} H_3O^+(aq) + Cl^-(aq)} \tag{8.5.1}$$

HCl(aq) is one example of a **strong acid**, which is a compound that is essentially **100% ionized** in aqueous solution. There are very few strong acids. The important ones are listed in Table 8.5.1.

Table 8.5.1: Strong Acids and Bases (All in Aqueous Solution)

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	$Mg(OH)_2$
H_2SO_4	$Ca(OH)_2$
$HClO_4$	

By analogy, a strong base is a compound that is essentially 100% ionized in aqueous solution. As with acids, there are only a few strong bases, which are also listed in Table 8.5.1.

If an acid is **not listed** in Table 8.5.1, it is likely a **weak acid**, which is a compound that is not 100% ionized in aqueous solution. Similarly, a weak base is a compound that is not 100% ionized in aqueous solution. For example, acetic acid ($\mathrm{HC_2H_3O_2}$) is a weak acid. The ionization reaction for acetic acid is as follows:

$$\mathrm{HC_2H_3O_2(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)} \tag{8.5.2}$$

Depending on the concentration of $HC_2H_3O_2$, the ionization reaction may occur only for 1%–5% of the acetic acid molecules.

Looking Closer: 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 (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:

$$\mathrm{NH_3(aq) + H_2O(\ell) \rightarrow NH_4^+(aq) + OH^-(aq)} \tag{8.5.3}$$

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.



Figure 8.5.1: (left) Bottles of alkaline drain cleaners containing sodium hydroxide can dissolve greases and hair. (right) Chemical burns caused by sodium hydroxide solution photographed 44 hours after exposure. Figures used with permission from Wikipedia.

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. Safer, nonlye drain cleaners use peroxide compounds to react on the materials in the clog and clear the drain.

Chemical Equilibrium in Weak Acids and Bases

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 $\mathbf{HC_2H_3O_2\,(aq)}$ is as follows:

$$\mathbf{HC_2H_3O_2\,(aq) + H_2O(\ell) \;\rightarrow H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)} \tag{8.5.4}$$

The reverse process also begins to occur:

$$\mathbf{H_3O^+\,(aq) + C_2H_3O_2^-\,(aq) \;\rightarrow HC_2H_3O_2\,(aq) + H_2O(\ell)} \tag{8.5.5}$$

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:

$$\mathbf{HC_2H_3O_2\,(aq) + H_2O(\ell) \rightleftharpoons H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)} \tag{8.5.6}$$

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 8.5.1: 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. $HNO_2(aq) + H_2O(\ell) \leftrightarrows NO_2^-(aq) + H_3O^+(aq)$
b. $C_5H_5N(aq) + H_2O(\ell) \leftrightarrows C_5H_5NH^+(aq) + OH^-(aq)$

? Exercise 8.5.1

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. $HF_{(aq)}$
b. $AgOH_{(aq)}$
c. $CH_3NH_2(aq)$

Answer

a. $HF(aq) + H_2O(\ell) \leftrightarrows F^-(aq) + H_3O^+(aq)$

b. $AgOH(aq) \leftrightarrows Ag^+(aq) + OH^-(aq)$

c. $CH_3NH_2(aq) + H_2O(\ell) \leftrightarrows CH_3NH_3^+(aq) + OH^-(aq)$

Acid Ionization Constant, K_a

The ionization for a general weak acid, **HA**, can be written as follows:

$$\mathbf{HA}\,(aq) \rightleftharpoons \mathbf{H^+}\,(aq) + \mathbf{A^-}\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_a = \frac{[H^+]\,[A^-]}{[HA]}$$

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 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 8.5.2: 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 ionization. A **strong base** is a base, which ionizes completely in an aqueous solution. The most common strong bases are soluble metal hydroxide compounds such as potassium

hydroxide. Some metal hydroxides are not as strong simply because they are not as soluble. Calcium hydroxide is only slightly soluble in water, but the portion that does dissolve also dissociates into ions.

A **weak base** is a base that ionizes only slightly in an aqueous solution. Recall that a base can be defined as a substance, which accepts a hydrogen ion from another substance. When a weak base such as ammonia is dissolved in water, it accepts an H^+ ion from water, forming the hydroxide ion and the conjugate acid of the base, the ammonium ion.

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

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 ionization 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 8.5.3 is a listing of base ionization constants for several weak bases.

Table 8.5.3: 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}
Urea	$H_2NCONH_2 + H_2O \rightleftharpoons H_2NCONH_3^+ + OH^-$	1.5×10^{-14}

The Ion-Product of Water

As we have already seen, H_2O can act as an acid or 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. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{8.5.7}$$

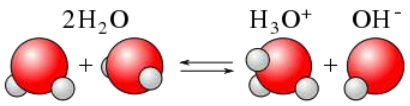


Figure 8.5.2: Autoionization of water, giving hydroxide and hydronium ions.

Similar to a weak acid, the autoionization of water is an equilibrium process, and is more properly written as follows:

$$H_2O(\ell) + H_2O(\ell) \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq) \tag{8.5.8}$$

We often use the simplified form of the reaction:

$$H_2O\,(l) \rightleftharpoons H^+\,(aq) + OH^-\,(aq)$$

The equilibrium constant for the autoionization of water is referred to as the ion-product for water and is given the symbol Kw.

$$K_w = [H^+][OH^-]$$

The **ion-product of water (*Kw*)** is the mathematical product of the concentration of hydrogen ions and hydroxide ions. 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}.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

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.

$$[H^+] = [OH^-] = 1.0\times10^{-7}$$

The product of these two concentrations is 1.0×10^{-14}

$$[H^+]\times[OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

- For acids, the concentration of H^+ or $[H^+]$) is greater than 1.0×10^{-7} M
- For bases, the concentration of OH^- or $[OH^-]$ is greater than 1.0×10^{-7} M.

Aqueous HCl is an example of acidic solution. Hydrogen chloride (HCl) ionizes to produce H^+ and Cl^- ions upon dissolving in water. This **increases the concentration of H^+ ions** in the solution. According to **Le Chatelier's principle**, the equilibrium represented by

$$H_2O\,(l) \rightleftharpoons H^+\,(aq) + OH^-\,(aq)$$

$$HCl\,(g) \longrightarrow H^+\,(aq) + Cl^-\,(aq)$$

is **forced to the left**, towards the reactant. As a result, the **concentration of the hydroxide ion decreases**.

Now, consider KOH (aq), a **basic solution**. Solid potassium hydroxide (KOH) dissociates in water to yield potassium ions and hydroxide ions.

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$$

The **increase in concentration of the OH^- ions** will cause a decrease in the concentration of the H^+ ions.

No matter whether the aqueous solution is an acid, a base, or neutral:and the ion-product of $[H^+][OH^-]$ remains constant.

- For acidic solutions, $[H^+]$) is greater than $[OH^-]$.
- For basic solutions, [OH−] is greater than $[H^+]$.
- For neutral solutions, $[H_3O^+] = [OH-] = 1.0\times10^{-7}$M

This means that if you know $[H^+]$ for a solution, you can calculate what $[OH^-]$) has to be for the product to equal 1.0×10^{-14}, or if you know $[OH^-]$), you can calculate $[H^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{8.5.9}$$

Example 8.5.2

Hydrochloric acid (HCl) is a strong acid, meaning it is 100% ionized in solution. What is the $[H^+]$ and the $[OH^-]$ in a solution of 2.0×10^{-3} M HCl?

Solution

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

Known

- $[HCl] = 2.0\times10^{-3}$ M
- $Kw = 1.0\times10^{-14}$

Unknown

- $[H^+]$=?M
- $[OH^-]$=?M

Because HCl is 100% ionized, the concentration of H+ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one H+ ion and one Cl− ion. The concentration of OH− can then be determined from the [H+] and Kw.

Step 2: Solve.

$$[H^+]=2.0\times10^{-3}\,M$$

$$Kw = [H^+][OH^-] = 1.0\times10^{-14}$$

$$[OH^-] = Kw/[H^+] = 1.0\times10^{-14}/2.0\times10^{-3} = 5.0\times10^{-12}\,M$$

Step 3: Think about your result.

The $[H^+]$ is much higher than the $[OH^-]$ because the solution is acidic. As with other equilibrium constants, the unit for Kw is customarily omitted.

8.5.2 Exercise

Sodium hydroxide (NaOH) is a strong base. What is the $[H^+]$ and the $[OH^-]$ in a 0.001 M NaOH solution at 25 °C?

Answer

$[OH^-]$ = 0.001M or 1 x 10^{-3}M; $[H^+]=1\times10^{-11}$M.

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 8.5.3 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.

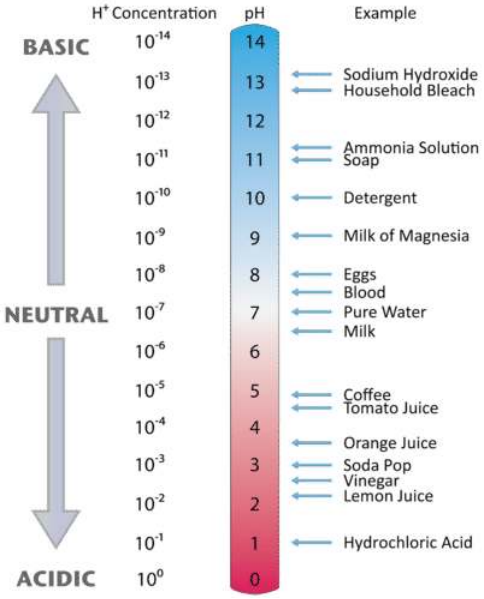


Figure 8.5.3: The relationship between $[H^+]$ and pH *values for several common materials.*

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 \(\PageIndex{4}.

Table **8.5.4**: Acidic, Basic and Neutral pH Values

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 8.5.3

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.

8.5.3 Exercise

Find the pH, given $[H^+]$ of the following:

a. 5.8×10^{-4} M
b. 1.0×10^{-7} M

Answer

a. 3.24

b. 7.00

Table **8.5.5** lists the pH of several common solutions. The most acidic among the listed solutions is battery acid with the lowest pH value (0.3). The most basic is 1M NaOH solution with the highest pH value of 14.0. Notice that some biological fluids (stomach acid and urine) are nowhere near neutral. You may also notice that many food products are slightly acidic. They are acidic because they contain solutions of weak acids. If the acid components of these foods were strong acids, the food would likely be inedible.

Table **8.5.5**: The pH Values of Some Common Solutions

Solution	pH
battery acid	0.3
stomach acid	1–2
lemon or lime juice	2.1
vinegar	2.8–3.0
Coca-Cola	3
wine	2.8–3.8
beer	4–5
coffee	5
milk	6
urine	6
pure H_2O	7
(human) blood	7.3–7.5
sea water	8
antacid (milk of magnesia)	10.5
NH_3 (1 M)	11.6
bleach	12.6
NaOH (1 M)	14.0

Example 8.5.4

Label each solution as acidic, basic, or neutral based only on the stated ***pH***.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Solution

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

8.5.4 Exercise

Identify each substance as acidic, basic, or neutral based only on the stated ***pH***.

a. human blood with ***pH*** = 7.4
b. household ammonia with ***pH*** = 11.0
c. cherries with ***pH*** = 3.6

Answer

a. slightly basic

b. basic

c. acidic

Acid Rain

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:

$$\mathrm{H_2O\,(l) + CO_2\,(g) \longrightarrow H_2CO_3\,(aq)} \tag{8.5.10}$$

$$\mathrm{H_2CO_3\,(aq) \rightleftharpoons H^+\,(aq) + HCO_3^-\,(aq)} \tag{8.5.11}$$

Acid rain is rainwater that has a pH of less than 5, due to a variety of nonmetal oxides, including CO_2, SO_2, SO_3, NO, and NO_2 being dissolved in the water and reacting with it to form not only carbonic acid, but sulfuric acid and nitric acid. The formation and subsequent ionization of sulfuric acid are shown here:

$$\mathrm{H_2O\,(l) + SO_3\,(g) \longrightarrow H_2SO_4\,(aq)} \tag{8.5.12}$$

$$\mathrm{H_2SO_4\,(aq) \longrightarrow H^+\,(aq) + HSO_4^-\,(aq)} \tag{8.5.13}$$

Carbon dioxide is naturally present in the atmosphere because we and most other organisms produce it as a waste product of metabolism. Carbon dioxide is also formed when fires release carbon stored in vegetation or when we burn wood or fossil fuels. Sulfur trioxide in the atmosphere is naturally produced by volcanic activity, but it also stems from burning fossil fuels, which have traces of sulfur, and from the process of "roasting" ores of metal sulfides in metal-refining processes. Oxides of nitrogen are formed in internal combustion engines where the high temperatures make it possible for the nitrogen and oxygen in air to chemically combine.

Acid rain is a particular problem in industrial areas where the products of combustion and smelting are released into the air without being stripped of sulfur and nitrogen oxides. In North America and Europe until the 1980s, it was responsible for the destruction of forests and freshwater lakes, when the acidity of the rain actually killed trees, damaged soil, and made lakes uninhabitable for all but the most acid-tolerant species. Acid rain also corrodes statuary and building facades that are made of marble and limestone (Figure 8.5.4). Regulations limiting the amount of sulfur and nitrogen oxides that can be released into the atmosphere by industry and automobiles have reduced the severity of acid damage to both natural and manmade environments in North America and Europe. It is now a growing problem in industrial areas of China and India.



(a) (b)

Figure 8.5.4: (a) Acid rain makes trees more susceptible to drought and insect infestation, and depletes nutrients in the soil. (b) It also is corrodes statues that are carved from marble or limestone. (credit a: modification of work by Chris M Morris; credit b: modification of work by "Eden, Janine and Jim"/Flickr)

Two photos are shown. Photograph a on the left shows the upper portion of trees against a bright blue sky. The tops of several trees at the center of the photograph have bare branches and appear to be dead. Image b shows a statue of a man that appears to from the revolutionary war era in either marble or limestone.

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.
- The ion-product of $[H^+][OH^-]$ in an aqueous solution remains constant.
- A ***p*H** value is simply the *negative* of the logarithm of the H^+ ion concentration ($-\log[H^+]$).
- The pH scale is used to succinctly communicate the acidity or basicity of a solution.
- A solution is **acidic** if pH < 7.
- A solution is **basic** if pH > 7.
- A solution is **neutral** if pH = 7.

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8.6: Buffers

Learning Objectives

- To define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have a marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

The mechanism involves a **buffer**, a solution that resists dramatic changes in pH. A *buffer* (or *buffered*) solution is one that resists a change in its pH when H^+ or OH^- ions are added or removed owing to some other reaction taking place in the same solution. Buffers do so by being composed of certain pairs of solutes: either a **weak acid plus its conjugate base** or a **weak base plus its conjugate acid**.

For example, a buffer can be composed of dissolved **acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate** ($NaC_2H_3O_2$). 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.

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. 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:

$$HC_2H_3O_2\,(aq) + OH^-\,(aq) \longrightarrow H_2O(\ell) + C_2H_3O_2^-\,(aq) \tag{8.6.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

> Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$H^+\,(aq) + C_2H_3O_2^-\,(aq) \longrightarrow HC_2H_3O_2\,(aq) \tag{8.6.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure 8.6.1 illustrates both actions of a buffer.

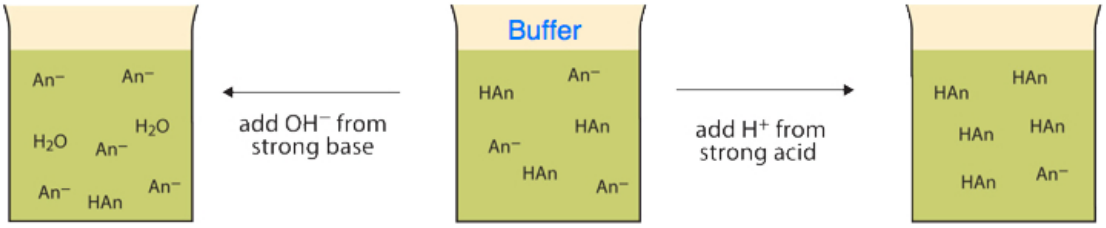


Figure **8.6.1**: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

A simple buffer system might be a 0.2 *M* solution of sodium acetate; the conjugate pair here is acetic acid HAc and its conjugate base, the acetate ion Ac^-. The idea is that this conjugate pair "pool" will be available to gobble up any small ($\leq 10^{-3}$ *M*) addition of H+ or OH^- that may result from other processes going on in the solution.

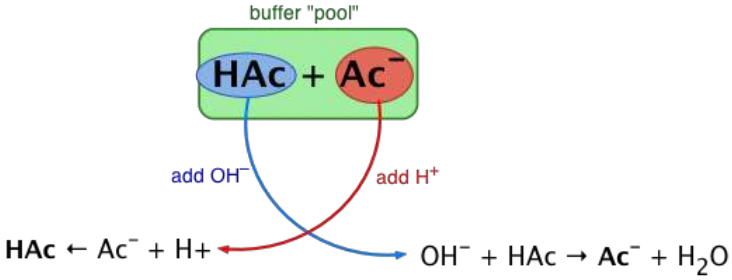


Figure **8.6.2**: How HAc/Ac buffer works

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain capacity. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

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:

$$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)} \tag{8.6.3}$$

while the ammonium ion [NH_4^+(aq)] can react with any hydroxide ions introduced by strong bases:

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(\ell)} \tag{8.6.4}$$

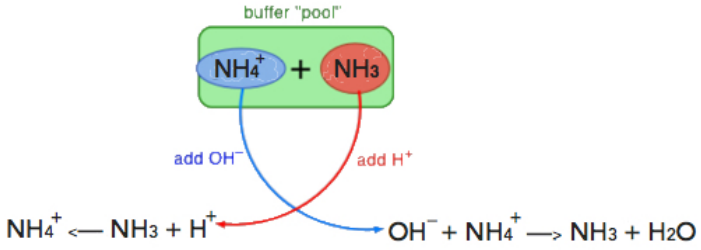


Figure **8.6.3**: How NH_4^+/NH_3 buffer works

Example 8.6.1

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.

? Exercise 8.6.1

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. H_3PO_4 and NaH_2PO_4
c. NH_3 and $(NH_4)_3PO_4$
d. NaOH and NaCl

Answer

b. H_3PO_4 (weak acid) and $H_2PO_4^-$ (conjugate base of H_3PO_4) make a buffer.

c. NH_3 (weak base) and NH_4^+ (conjugate acid of NH_3) make a buffer

✓ Food and Drink App: The Acid That Eases Pain

Although medicines are not exactly "food and drink," we do ingest them, so 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. 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. Because it is not given in Table 10.5.1, 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.

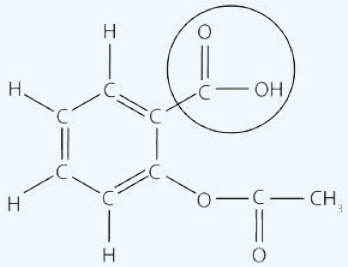


Figure 8.6.4 The Molecular Structure of Aspirin. The circled atoms are the acid part of the molecule.

Because the acid properties of aspirin may be problematic, many aspirin brands offer a "buffered aspirin" form of the medicine. In these cases, the aspirin also contains a buffering agent-usually MgO-that 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.

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.
- 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 (Figure \(\PageIndex{1}\)) is maintained by buffer systems.

Table 7.3.2: pH in Living Systems

Compartment	pH
Gastric Acid	1
Lysosomes	4.5
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

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write CO_2(aq)]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Medicine: The Buffer System in Blood

The normal pH of human blood is about 7.4. The carbonate buffer system in the blood uses the following equilibrium reaction:

$$\mathbf{CO_2}(g) + \mathbf{2\,H_2O}(l) \rightleftharpoons \mathbf{H_2CO_3}(aq) \rightleftharpoons \mathbf{HCO_3}^-(aq) + \mathbf{H_3O}^+(aq)$$

The concentration of carbonic acid, H_2CO_3 is approximately 0.0012 *M*, and the concentration of the hydrogen carbonate ion, $\mathbf{HCO_3}^-$, is around 0.024 *M*. Using the Henderson-Hasselbalch equation and the pK_a of carbonic acid at body temperature, we can calculate the pH of blood:

$$\mathbf{pH} = \mathbf{p}K_\mathbf{a} + \log\frac{[\mathbf{base}]}{[\mathbf{acid}]} = 6.1 + \log\frac{0.024}{0.0012} = 7.4$$

The fact that the H_2CO_3 concentration is significantly lower than that of the $\mathbf{HCO_3}^-$ ion may seem unusual, but this imbalance is due to the fact that most of the by-products of our metabolism that enter our bloodstream are acidic. Therefore, there must be a larger proportion of base than acid, so that the capacity of the buffer will not be exceeded.

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 fact, 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.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/− type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

Concept Review Exercise

1. Explain how a buffer prevents large changes in pH.

Answer

1. A buffer has components that react with both strong acids and strong bases to resist sudden changes in pH.

Exercises

1. Describe a buffer. What two related chemical components are required to make a buffer?
2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?
3. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. HCl and NaCl
    2. HNO_2 and $NaNO_2$
    3. NH_4NO_3 and HNO_3
    4. NH_4NO_3 and NH_3
4. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. H_3PO_4 and Na_3PO_4
    2. $NaHCO_3$ and Na_2CO_3
    3. $NaNO_3$ and $Ca(NO_3)_2$
    4. HN_3 and NH_3
5. For each combination in Exercise 3 that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.
6. For each combination in Exercise 4 that is a buffer, write the chemical equations for the reaction of the buffer components when a strong acid and a strong base is added.
7. The complete phosphate buffer system is based on four substances: H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}. What different buffer solutions can be made from these substances?

8. Explain why NaBr cannot be a component in either an acidic or a basic buffer.
9. Explain why $Mg(NO_3)_2$ cannot be a component in either an acidic or a basic buffer.

Answers

1. A buffer resists sudden changes in pH. It has a weak acid or base and a salt of that weak acid or base.
2. No. Combining a strong acid and a strong base will produce salt and water. Excess strong acid or strong base will not act as a buffer.
3. 1. not a buffer
   2. buffer
   3. not a buffer
   4. buffer

4. 1. not a buffer

2. buffer

3. not a buffer

4. not buffer

5. 3b: strong acid: $H^+ + NO_2^- \rightarrow HNO_2$; strong base: $OH^- + HNO_2 \rightarrow H_2O + NO_2^-$; 3d: strong acid: $H^+ + NH_3 \rightarrow NH_4^+$; strong base: $OH^- + NH_4^+ \rightarrow H_2O + NH_3$
6. 4b: strong acid: $H^+ + CO_3^{2-} \rightarrow HCO3^-$; strong base: $OH^- + HCO_3^- \rightarrow H_2O + CO_3^{2-}$;
7. Buffers can be made by combining H_3PO_4 and $H_2PO_4^-$, $H_2PO_4^-$ and HPO_4^{2-}, and HPO_4^{2-} and PO_4^{3-}.
8. NaBr splits up into two ions in solution, Na^+ and Br^-. Na^+ will not react with any added base knowing that NaOH is a strong base. Br- will not react with any added acid knowing that HBr is a strong acid. Because NaBr will not react with any added base or acid, it does not resist change in pH and is not a buffer.
9. $Mg(NO_3)_2$ includes two types of ions, Mg^{2+} and NO_3^-. $Mg(OH)_2$ is strong base and completely dissociates (100% falls apart), so Mg^{2+} will not react with any added base (0% combines with OH^-). HNO_3 is strong acid and completely dissociates (100% falls apart), so NO_3^- will not react with any added acid (0% combines with H^+). Because $Mg(NO_3)_2$ will not react with any added base or acid, it does not resist change in pH and is not a buffer.

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8.7.1: Acids and Bases (Exercises)

10.1: Arrhenius Definition of Acids and Bases

Concept Review Exercises

1. Give the Arrhenius definitions of an acid and a base.
2. What is neutralization?

Answers

1. Arrhenius acid: a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution; Arrhenius base: a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.
2. the reaction of an acid and a base

Exercises

1. Give two examples of Arrhenius acids.
2. Give two examples of Arrhenius bases.
3. List the general properties of acids.
4. List the general properties of bases.
5. Name each compound. (For acids, look up the name in Table 10.1.1. For bases, use the rules for naming ionic compounds from Chapter 3.)

    a. HBr(aq)

    b. $Ca(OH)_2$(aq)

    c. HNO_3(aq)

    d. $Fe(OH)_3$(aq)

6. Name each compound.

    a. HI(aq)

    b. $Cu(OH)_2$(aq)

    c. H_3PO_4(aq)

    d. CsOH(aq)

7. Write a balanced chemical equation for the neutralization of $Ba(OH)_2$(aq) with HNO_3(aq).

8. Write a balanced chemical equation for the neutralization of H_2SO_4(aq) with $Cr(OH)_3$(aq).

9. Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction.

10. Identify the salt produced in each acid-base reaction below. Then, balance the equation.

    a. $2HCl + Sr(OH)_2 \rightarrow 2H_2O$ + ??

    b. KNO_3; $HNO_3 + KOH \rightarrow$?? + H_2O

    c. $HF + Ca(OH)_2$ ---> ?? + H_2O

11. How many moles of sodium hydroxide (NaOH) are needed to neutralize 0.844 mol of acetic acid ($HC_2H_3O_2$)? (Hint: begin by writing a balanced chemical equation for the process.)

12. How many moles of perchloric acid ($HClO_4$) are needed to neutralize 0.052 mol of calcium hydroxide [$Ca(OH)_2$]? (Hint: begin by writing a balanced chemical equation for the process

13. Hydrazoic acid (HN_3) can be neutralized by a base.

a. Write the balanced chemical equation for the reaction between hydrazoic acid and calcium hydroxide.

b. How many milliliters of 0.0245 M $Ca(OH)_2$ are needed to neutralize 0.564 g of HN_3?

14. Citric acid ($H_3C_6H_5O_7$) has three hydrogen atoms that can form hydrogen ions in solution.

a. Write the balanced chemical equation for the reaction between citric acid and sodium hydroxide.

b. If an orange contains 0.0675 g of $H_3C_6H_5O_7$, how many milliliters of 0.00332 M NaOH solution are needed to neutralize the acid?

15. Magnesium hydroxide [$Mg(OH)_2$] is an ingredient in some antacids. How many grams of $Mg(OH)_2$ are needed to neutralize the acid in 158 mL of 0.106 M HCl(aq)? It might help to write the balanced chemical equation first.

16. Aluminum hydroxide [$Al(OH)_3$] is an ingredient in some antacids. How many grams of $Al(OH)_3$ are needed to neutralize the acid in 96.5 mL of 0.556 M H_2SO_4(aq)? It might help to write the balanced chemical equation first.

17. Write the balanced chemical equation for the reaction between HBr and $Ca(OH)_2$. What volume of 0.010 M HBr solution is be required to neutralize 25 mL of a 0.0100*M* $Ca(OH)_2$ solution?

18. Write the balanced chemical equation for the reaction between HNO_3 and KOH. What volume of 0.5M HNO_3 is required to neutralize 60 mL of 0.4M KOH solution?

Answers

1. HCl and HNO_3 (answers will vary)
2. NaOH and $Ca(OH)_2$ (answers will vary)
3. sour taste, react with metals, react with bases, and turn litmus red
4. bitter taste, feels slippery, react with acids and turn litmus blue
5. a. hydrobromic acid

b. calcium hydroxide

c. nitric acid

d. iron(III) hydroxide

6. a. hydroiodic acid

b. cupric hydroxide

c. phosphoric acid

d. cesium hydroxide

7. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O$

8. $3H_2SO_4(aq) + 2Cr(OH)_3(aq) \rightarrow Cr_2(SO_4)_3(aq) + 6H_2O$

9. $Mg(OH)_2 + 2HCl$ --> $MgCl_2 + 2H_2O$

10. a. $SrCl_2$; $2HCl + Sr(OH)_2 \rightarrow 2H_2O + SrCl_2$

b. KNO_3; $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

c. CaF_2; $2HF + Ca(OH)_2 \rightarrow CaF_2 + 2H_2O$

11. 0.844 mol

12. 0.104 mol

13. Part 1: $2HN_3(aq) + Ca(OH)_2 \rightarrow Ca(N_3)_2 + 2H_2O$

Part 2: 268 mL

14. Part 1: $H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow Na_3C_6H_5O_7(aq) + 3H_2O$

Part 2: 317.5 mL

15. 0.488 g

16. 2.79 g

17. $2HBr + Ca(OH)_2 \rightarrow CaBr_2 + 2H_2O$; 50 mL HBr

18. $HNO_3 + KOH \rightarrow KNO_3 + H_2O$; 48 mL HNO_3

10.2: Brønsted-Lowry Definition of Acids and Bases

Concept Review Exercise

1. Give the definitions of a Brønsted-Lowry acid and a Brønsted-Lowry base.

Answer

1. A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

Exercises

1. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

   $HCl(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

2. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

   $H_2O(\ell) + N_2H_4(aq) \rightarrow N_2H_5^+(aq) + OH^-(aq)$

3. Explain why a Brønsted-Lowry acid can be called a proton donor.
4. Explain why a Brønsted-Lowry base can be called a proton acceptor.
5. Write the chemical equation of the reaction of ammonia in water and label the Brønsted-Lowry acid and base.
6. Write the chemical equation of the reaction of methylamine (CH_3NH_2) in water and label the Brønsted-Lowry acid and base.
7. Demonstrate that the dissolution of HNO_3 in water is actually a Brønsted-Lowry acid-base reaction by describing it with a chemical equation and labeling the Brønsted-Lowry acid and base.
8. Identify the Brønsted-Lowry acid and base in the following chemical equation:

   $C_3H_7NH_2(aq) + H_3O^+(aq) \rightarrow C_3H_7NH_3^+(aq) + H_2O(\ell)$

9. Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in each of the following equations

   1. $\mathbf{NO_2^- + H_2O \longrightarrow HNO_2 + OH^-}$
   2. $\mathbf{HBr + H_2O \longrightarrow H_3O^+ + Br^-}$
   3. $\mathbf{HS^- + H_2O \longrightarrow H_2S + OH^-}$
   4. $\mathbf{H_2PO_4^- + OH^- \longrightarrow HPO_4^{2-} + H_2O}$
   5. $\mathbf{H_2PO_4^- + HCl \longrightarrow H_3PO_4 + Cl^-}$

10. Write the chemical equation for the reaction that occurs when cocaine hydrochloride ($C_{17}H_{22}ClNO_4$) dissolves in water and donates a proton to a water molecule. (When hydrochlorides dissolve in water, they separate into chloride ions and the appropriate cation.)

11. If codeine hydrobromide has the formula $C_{18}H_{22}BrNO_3$, what is the formula of the parent compound codeine?

Answers

1. HCl: Brønsted-Lowry acid; NH_3: Brønsted-Lowry base
2. H_2O: Brønsted-Lowry acid; N_2H_4: Brønsted-Lowry base
3. A Brønsted-Lowry acid gives away an H^+ ion—nominally, a proton—in an acid-base reaction.
4. A Brønsted-Lowry base accepts an H^+ ion (a proton) in an acid-base reaction.
5. $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ (here NH_3 = Brønsted-Lowry base; H_2O = Brønsted-Lowry acid)
6. $CH_3NH_2 + H_2O \rightarrow CH_3NH_3^+ + OH^-$ (here CH_3NH_2 = Brønsted-Lowry base; H_2O = Brønsted-Lowry acid)
7. $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$ (here HNO_3 = Brønsted-Lowry acid; H_2O = Brønsted-Lowry base)

8. $C_3H_7NH_2(aq) + H_3O^+(aq) \rightarrow C_3H_7NH_3^+(aq) + H_2O(\ell)$ (here H_3O^+ = Brønsted-Lowry acid; $C_3H_7NH_2$ = Brønsted-Lowry base)

10.3: Water - Both an Acid and a Base

Concept Review Exercises

1. Explain how water can act as an acid.
2. Explain how water can act as a base.

Answers

1. Under the right conditions, H_2O can donate a proton, making it a Brønsted-Lowry acid.
2. Under the right conditions, H_2O can accept a proton, making it a Brønsted-Lowry base.

Exercises

1. Is $H_2O(\ell)$ acting as an acid or a base?

   $H_2O(\ell) + NH_4^+(aq) \rightarrow H_3O^+(aq) + NH_3(aq)$

2. Is $H_2O(\ell)$ acting as an acid or a base?

   $CH_3^-(aq) + H_2O(\ell) \rightarrow CH_4(aq) + OH^-(aq)$

3. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some $C_2H_3O_2^-$ solutions, the following reaction can occur:

   $C_2H_3O_2^-(aq) + H_2O(\ell) \rightarrow HC_2H_3O_2(aq) + OH^-(aq)$

   Is H_2O acting as an acid or a base in this reaction?

4. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some NH_4^+ solutions, the following reaction can occur:

   $NH_4^+(aq) + H_2O \rightarrow NH_3(aq) + H_3O^+(aq)$

   Is H_2O acting as an acid or a base in this reaction?

5. Why is pure water considered neutral?

Answers

1. base
2. acid
3. acid
4. base

5. When water ionizes, equal amounts of H^+ (acid) and OH^- (base) are formed, so the solution is neither acidic nor basic: $H_2O(\ell) \rightarrow H^+(aq) + OH^-(aq)$

10.4: The Strengths of Acids and Bases

Concept Review Exercises

1. Explain the difference between a strong acid or base and a weak acid or base.
2. Explain what is occurring when a chemical reaction reaches equilibrium.
3. Define *pH*.

AnswerS

1. A strong acid or base is 100% ionized in aqueous solution; a weak acid or base is less than 100% ionized.
2. The overall reaction progress stops because the reverse process balances out the forward process.
3. pH is a measure of the hydrogen ion concentration.

Exercises

1. Name a strong acid and a weak acid. (Hint: use Table 10.4.1.)
2. Name a strong base and a weak base. (Hint: use Table 10.4.1.)
3. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. HF
    2. $HC_2H_3O_2$
    3. HCl
    4. $HClO_4$
4. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. H_2SO_4
    2. HSO_4^-
    3. HPO_4^{2-}
    4. HNO_3
5. Is each compound a strong base or a weak base? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. NH_3
    2. NaOH
    3. $Mg(OH)_2$
    4. $Cu(OH)_2$
6. Is each compound a strong base or a weak base? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. KOH
    2. H_2O
    3. $Fe(OH)_2$
    4. $Fe(OH)_3$
7. Write the chemical equation for the equilibrium process for each weak acid in Exercise 3.
8. Write the chemical equation for the equilibrium process for each weak acid in Exercise 4.
9. Write the chemical equation for the equilibrium process for each weak base in Exercise 5.
10. Write the chemical equation for the equilibrium process for each weak base in Exercise 6.
11. Which is the stronger acid—HCl(aq) or HF(aq)?
12. Which is the stronger base—KOH(aq) or $Ni(OH)_2$(aq)?
13. Consider the two acids in Exercise 11. For solutions that have the same concentration, which one would you expect to have a lower pH?
14. Consider the two bases in Exercise 12. For solutions that have the same concentration, which one would you expect to have a higher pH?
15. Consider the list of substances in Table \PageIndex3\PageIndex3.2"The pH Values of Some Common Solutions". What is the most acidic substance on the list that you have encountered recently?
16. Consider the list of substances in Table \PageIndex3\PageIndex3.2"The pH Values of Some Common Solutions". What is the most basic substance on the list that you have encountered recently?
17. Indicate whether solutions with the following pH values are acidic, basic, or neutral:
    1. pH = 9.4
    2. pH = 7.0
    3. pH = 1.2
    4. pH = 6.5

Answers

1. strong acid: HCl; weak acid: $HC_2H_3O_2$ (answers will vary)
2. strong base: NaOH; weak base: NH_3 (answers will vary)
3. 1. weak
    2. weak

3. strong
4. strong

4. 1. strong
2. weak
3. weak
4. strong

5. 1. weak
2. strong
3. strong
4. weak

6. 1. strong
2. weak
3. weak
4. weak

7. 3a: $HF(aq) \leftrightarrows H^+(aq) + F^-(aq)$; 3b: $HC_2H_3O_2(aq) \leftrightarrows H^+(aq) + C_2H_3O_2^-(aq)$

8. 4b: $HSO_4^-(aq) \leftrightarrows H^+(aq) + SO_4^{2-}(aq)$; 4c: $HPO_4^{2-}(aq) \leftrightarrows H^+(aq) + PO_4^{3-}(aq)$

9. 5a: $NH_3(aq) + H_2O \leftrightarrows NH_4^+(aq) + OH^-(aq)$; 5d: $Cu(OH)_2(aq) \leftrightarrows Cu^{2+}(aq) + 2OH^-(aq)$

10. 6b: $H_2O + H_2O \leftrightarrows H_3O^+(aq) + OH^-(aq)$; 6c: $Fe(OH)_2(aq) \leftrightarrows Fe^{2+}(aq) + 2OH^-(aq)$; 6d: $Fe(OH)3(aq) \leftrightarrows Fe^{3+}(aq) + 3OH^-(aq)$

11. HCl(aq)

12. KOH(aq)

13. HCl(aq)

14. KOH(aq)

15. (answers will vary)

16. (answers will vary)

17. 1. basic

2. neutral

3. acidic (strongly)

4. acidic (mildly)

10.5: Buffers

Concept Review Exercise

1. Explain how a buffer prevents large changes in pH.

Answer

1. A buffer has components that react with both strong acids and strong bases to resist sudden changes in pH.

Exercises

1. Describe a buffer. What two related chemical components are required to make a buffer?
2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?
3. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. HCl and NaCl
    2. HNO_2 and $NaNO_2$
    3. NH_4NO_3 and HNO_3

4. NH_4NO_3 and NH_3

4. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. H_3PO_4 and Na_3PO_4
    2. $NaHCO_3$ and Na_2CO_3
    3. $NaNO_3$ and $Ca(NO_3)_2$
    4. HN_3 and NH_3
5. For each combination in Exercise 3 that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.
6. For each combination in Exercise 4 that is a buffer, write the chemical equations for the reaction of the buffer components when a strong acid and a strong base is added.
7. The complete phosphate buffer system is based on four substances: H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}. What different buffer solutions can be made from these substances?
8. Explain why NaBr cannot be a component in either an acidic or a basic buffer.
9. Explain why $Mg(NO_3)_2$ cannot be a component in either an acidic or a basic buffer.

Answers

1. A buffer resists sudden changes in pH. It has a weak acid or base and a salt of that weak acid or base.
2. No. Combining a strong acid and a strong base will produce salt and water. Excess strong acid or strong base will not act as a buffer.
3. 1. not a buffer
    2. buffer
    3. not a buffer
    4. buffer

4. 1. not a buffer

2. buffer

3. not a buffer

4. not buffer

5. 3b: strong acid: $H^+ + NO_2^- \rightarrow HNO_2$; strong base: $OH^- + HNO_2 \rightarrow H_2O + NO_2^-$; 3d: strong acid: $H^+ + NH_3 \rightarrow NH_4^+$; strong base: $OH^- + NH_4^+ \rightarrow H_2O + NH_3$
6. 4b: strong acid: $H^+ + CO_3^{2-} \rightarrow HCO3^-$; strong base: $OH^- + HCO_3^- \rightarrow H_2O + CO_3^{2-}$;
7. Buffers can be made by combining H_3PO_4 and $H_2PO_4^-$, $H_2PO_4^-$ and HPO_4^{2-}, and HPO_4^{2-} and PO_4^{3-}.
8. NaBr splits up into two ions in solution, Na^+ and Br^-. Na^+ will not react with any added base knowing that NaOH is a strong base. Br- will not react with any added acid knowing that HBr is a strong acid. Because NaBr will not react with any added base or acid, it does not resist change in pH and is not a buffer.
9. $Mg(NO_3)_2$ includes two types of ions, Mg^{2+} and NO_3^-. $Mg(OH)_2$ is strong base and completely dissociates (100% falls apart), so Mg^{2+} will not react with any added base (0% combines with OH^-). HNO_3 is strong acid and completely dissociates (100% falls apart), so NO_3^- will not react with any added acid (0% combines with H^+). Because $Mg(NO_3)_2$ will not react with any added base or acid, it does not resist change in pH and is not a buffer.

Additional Exercises

1. The properties of a 1.0 M HCl solution and a 1.0 M $HC_2H_3O_2$ solution are compared. Measurements show that the hydrochloric acid solution has a higher osmotic pressure than the acetic acid solution. Explain why.
2. Of a 0.50 M HNO_3 solution and a 0.50 M $HC_2H_3O_2$ solution, which should have the higher boiling point? Explain why.
3. The reaction of sulfuric acid [$H_2SO_4(aq)$] with sodium hydroxide [NaOH(aq)] can be represented by two separate steps, with only one hydrogen ion reacting in each step. Write the chemical equation for each step.
4. The reaction of aluminum hydroxide [$Al(OH)_3(aq)$] with hydrochloric acid [HCl(aq)] can be represented by three separate steps, with only one hydroxide ion reacting in each step. Write the chemical equation for each step.

5. A friend brings you a small sample of an unknown chemical. Assuming that the chemical is soluble in water, how would you determine if the chemical is an acid or a base?
6. A neutral solution has a hydrogen ion concentration of about 1×10^{-7} M. What is the concentration of the hydroxide ion in a neutral solution?
7. The Lewis definitions of an acid and a base are based on electron pairs, not protons. A Lewis acid is an electron pair acceptor, while a Lewis base is an electron pair donor. Use Lewis diagrams to show that

   $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

   is an acid-base reaction in the Lewis sense as well as in the Arrhenius and Brønsted-Lowry senses.
8. Given the chemical reaction

   $NH_3(g) + BF_3(g) \rightarrow NH_3—BF_3(s)$

   show that the reaction illustrated by this equation is an acid-base reaction if we use the Lewis definitions of an acid and a base (see Exercise 7). The product contains a bond between the N and B atoms.

Answers

1. HCl is a strong acid and yields more ions in solution. $HC_2H_3O_2$ is a weak acid and undergoes partial ionization in solution.
2. HNO_3 is a strong acid while $HC_2H_3O_2$ is a weak acid. HNO_3 dissociates 100% and its solution contains more ions. The more ions the solution contains the lower is its vapor pressure; the higher temperature is required for it to boil.
3. $H_2SO_4 + NaOH \rightarrow NaHSO_4 + H_2O$; $NaHSO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$
4. $Al(OH)_3 + HCl \rightarrow Al(OH)_2Cl + H_2O$; $Al(OH)_2Cl + HCl \rightarrow Al(OH)Cl_2 + H_2O$; $Al(OH)Cl_2 + HCl \rightarrow AlCl3 + H_2O$
5. One way is to add it to $NaHCO_3$; if it bubbles, it is an acid. Alternatively, add the sample to litmus and look for a characteristic color change (red for acid, blue for base).
6. In a neutral solution, $[OH^-] = [H^+] = 1.0 \times 10^{-7}$ M
7. The O atom is donating an electron pair to the H^+ ion, making the base an electron pair donor and the acid an electron pair acceptor.
8. The N atom is donating a lone pair to B in BF_3, Hence NH_3 is the Lewis base and BF_3 is the Lewis acid.

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8.7.2: Acids and Bases (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

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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Index

A

acid strength
- 8.5: The Strengths of Acids and Bases

atomic mass
- 1.9: Nuclei of Atoms

B

base strength
- 8.5: The Strengths of Acids and Bases

C

chemical formula
- 2.8: Formulas for Binary Ionic Compounds

combination reaction
- 6.9: Some Types of Chemical Reactions

combustion reaction
- 6.9: Some Types of Chemical Reactions

D

decomposition reaction
- 6.9: Some Types of Chemical Reactions

duet rule
- 2.2: Two Types of Bonding

dynamic equilibrium
- 8.5: The Strengths of Acids and Bases

E

exercises
- 1.13.1: Chemistry, Matter, and Measurement (Exercises)
- 1.13.2: Elements, Atoms, and the Periodic Table (Exercises)
- 2.14.1: Ions, Molecules, and Compounds (Exercises)
- 2.14.2: Answers Ions, Molecules, and Compounds (Exercises)
- 2.14.3: More Practice with Nomenclature (Exercises)
- 3.12.E: Chemistry, Matter, and Measurement (Exercises)
- 4.7.1: Mole Practice with Guidence (Exercises)
- 4.7.2: Chemical Quantities (Exercises)
- 6.14.4: More Practice with Reactions (Exercises)

I

irradiation
- 1.10: Uses of Radioactive Isotopes

isotopes
- 1.9: Nuclei of Atoms

L

law of conservation of mass
- 5.2: The Law of Conservation of Matter

length
- 3.4: SI Length and Volume Units

M

measurements
- 1.4: Measurements

P

polyatomic ions
- 2.9: Polyatomic Ions

R

radioactive dating
- 1.10: Uses of Radioactive Isotopes

radioactive isotopes
- 1.10: Uses of Radioactive Isotopes

radioactivity
- 1.10: Uses of Radioactive Isotopes

redox reaction
- 6.10: Oxidation-Reduction (Redox) Reactions

Redox Reactions (Organic Chemistry)
- 6.11: Redox Reactions in Organic Chemistry and Biochemistry

S

scienfitic notation
- 3.3: Scientific Notation in Chemistry

stock system
- 2.7: Stock System Naming

stoichiometry
- 5.4: Quantitative Relationships Based on Chemical Equations

T

tracer
- 1.10: Uses of Radioactive Isotopes

U

units
- 1.4: Measurements

V

volume
- 3.4: SI Length and Volume Units

Glossary

Sample Word 1 | Sample Definition 1

Detailed Licensing

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By Page

- CVTC Basic Chemistry - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Matter, Atoms, and Elements - *Undeclared*
    - 1.1: Prelude to Chemistry, Matter, and Measurement - *CC BY-NC-SA 4.0*
    - 1.2: What is Chemistry? - *CC BY-NC-SA 4.0*
    - 1.3: The Classification of Matter - *CC BY-NC-SA 4.0*
    - 1.4: Measurements - *CC BY-NC-SA 4.0*
    - 1.5: Temperature Units and Conversions - *Undeclared*
    - 1.6: The Elements - *CC BY-NC-SA 4.0*
    - 1.7: Atomic Theory - *CC BY-NC-SA 4.0*
    - 1.8: The Structure of Atoms - *CC BY-NC-SA 4.0*
    - 1.9: Nuclei of Atoms - *CC BY-NC-SA 4.0*
    - 1.10: Uses of Radioactive Isotopes - *CC BY-NC-SA 4.0*
    - 1.11: Atomic Masses - *CC BY-NC-SA 4.0*
    - 1.12: The Periodic Table - *CC BY-NC-SA 4.0*
    - 1.13: Chapter 1 Exercises - *Undeclared*
      - 1.13.1: Chemistry, Matter, and Measurement (Exercises) - *Undeclared*
      - 1.13.2: Elements, Atoms, and the Periodic Table (Exercises) - *Undeclared*
  - 2: Chemical Bonding and Nomenclature - *Undeclared*
    - 2.1: Prelude to Ionic Bonding and Simple Ionic Compounds - *CC BY-NC-SA 4.0*
    - 2.2: Two Types of Bonding - *CC BY-NC-SA 4.0*
    - 2.3: Cations - *CC BY-NC 4.0*
    - 2.4: Anions - *CC BY-NC 4.0*
    - 2.5: Transition Metal Ions - *CC BY-NC 4.0*
    - 2.6: Naming Binary Ionic Compounds - *CC BY-NC 4.0*
    - 2.7: Stock System Naming - *CC BY-NC 4.0*
    - 2.8: Formulas for Binary Ionic Compounds - *CC BY-NC 4.0*
    - 2.9: Polyatomic Ions - *CC BY-NC 4.0*
    - 2.10: Names and Formulas of Ternary Ionic Compounds - *CC BY-NC 4.0*
    - 2.11: Naming Binary Molecular Compounds - *CC BY-NC 4.0*
    - 2.12: Naming Acids - *CC BY-NC 4.0*
    - 2.13: Names and Formulas of Bases - *CC BY-NC 4.0*
    - 2.14: Chapter 2 Exercises - *Undeclared*
      - 2.14.1: Ions, Molecules, and Compounds (Exercises) - *Undeclared*
      - 2.14.2: Answers Ions, Molecules, and Compounds (Exercises) - *Undeclared*
      - 2.14.3: More Practice with Nomenclature (Exercises) - *Undeclared*
  - 3: Measurement and Calculations - *Undeclared*
    - 3.1: SI Base Units - *CC BY-NC 4.0*
    - 3.2: Metric Prefixes - *CC BY-NC 4.0*
    - 3.3: Scientific Notation in Chemistry - *CC BY-NC 4.0*
    - 3.4: SI Length and Volume Units - *CC BY-NC 4.0*
    - 3.5: SI Mass and Weight Units - *CC BY-NC 4.0*
    - 3.6: SI Kinetic Energy Units - *CC BY-NC 4.0*
    - 3.7: Scientific Dimensional Analysis - *CC BY-NC 4.0*
    - 3.8: Metric Unit Conversions - *CC BY-NC 4.0*
    - 3.9: Derived Units - *CC BY-NC 4.0*
    - 3.10: Dosage Calculations - *CC BY-NC-SA 4.0*
    - 3.11: Density - *CC BY-NC 4.0*
    - 3.12: Chapter 3 Excercises - *Undeclared*
      - 3.12.E: Chemistry, Matter, and Measurement (Exercises) - *Undeclared*

- 4: Chemical Quantities - *Undeclared*
  - 4.1: Avogadro's Number - *CC BY-NC 4.0*
  - 4.2: Conversions Between Moles and Atoms - *CC BY-NC 4.0*
  - 4.3: Molar Mass - *CC BY-NC 4.0*
  - 4.4: Conversions Between Moles and Mass - *CC BY-NC 4.0*
  - 4.5: Conversions Between Mass and Number of Particles - *CC BY-NC 4.0*
  - 4.6: Percent Composition - *CC BY-NC 4.0*
  - 4.7: Chapter 4 Exercises - *Undeclared*
    - 4.7.1: Mole Practice with Guidence (Exercises) - *Undeclared*
    - 4.7.2: Chemical Quantities (Exercises) - *Undeclared*
    - 4.7.3: Answers to Chemical Quantities (Exercises) - *Undeclared*
- 5: Chemical Reactions and Quantities - *Undeclared*
  - 5.1: Prelude to Introduction to Chemical Reactions - *CC BY-NC-SA 4.0*
  - 5.2: The Law of Conservation of Matter - *CC BY-NC-SA 4.0*
  - 5.3: Chemical Equations - *CC BY-NC-SA 4.0*
  - 5.4: Quantitative Relationships Based on Chemical Equations - *CC BY-NC-SA 4.0*
  - 5.5: Mole-Mole Relationships in Chemical Reactions - *CC BY-NC-SA 4.0*
  - 5.6: Mole-Mass and Mass-Mass Problems - *CC BY-NC-SA 4.0*
- 6: Chemical Energy and Reaction Types - *Undeclared*
  - 6.1: Chemical Potential Energy - *CC BY-NC 4.0*
  - 6.2: Heat - *CC BY-NC 4.0*
  - 6.3: Exothermic and Endothermic Processes - *CC BY-NC 4.0*
  - 6.4: Chemical Reaction Rate - *CC BY-NC 4.0*
  - 6.5: Activation Energy - *CC BY-NC 4.0*
  - 6.6: Potential Energy Diagrams - *CC BY-NC 4.0*
  - 6.7: Factors Affecting Reaction Rate - *CC BY-NC 4.0*
  - 6.8: Catalysts - *CC BY-NC 4.0*
  - 6.9: Some Types of Chemical Reactions - *CC BY-NC-SA 4.0*
  - 6.10: Oxidation-Reduction (Redox) Reactions - *CC BY-NC-SA 4.0*
  - 6.11: Redox Reactions in Organic Chemistry and Biochemistry - *CC BY-NC-SA 4.0*
  - 6.12: Single Replacement Reactions - *CC BY-NC 4.0*
  - 6.13: Double Replacement Reactions - *CC BY-NC 4.0*
  - 6.14: Chapter 6 Exercises - *Undeclared*
    - 6.14.1: Chemical Energy (Exercises) - *Undeclared*
    - 6.14.2: Types of Reactions (Exercises) - *Undeclared*
    - 6.14.3: Answers Types of Reactions (Exercises) - *Undeclared*
    - 6.14.4: More Practice with Reactions (Exercises) - *Undeclared*
- 7: Solutions - *Undeclared*
  - 7.1: Structure of Water - *CC BY-NC 4.0*
  - 7.2: Structure of Ice - *CC BY-NC 4.0*
  - 7.3: Physical Properties of Water - *CC BY-NC 4.0*
  - 7.4: Solute and Solvent - *CC BY-NC 4.0*
  - 7.5: Dissolving Process - *CC BY-NC 4.0*
  - 7.6: Liquid-Liquid Solutions - *CC BY-NC 4.0*
  - 7.7: Electrolytes and Nonelectrolytes - *CC BY-NC 4.0*
  - 7.8: Dissociation - *CC BY-NC 4.0*
  - 7.9: Strong and Weak Electrolytes - *CC BY-NC 4.0*
  - 7.10: Solutions - *CC BY-NC-SA 4.0*
  - 7.11: Rate of Dissolving - *CC BY-NC 4.0*
  - 7.12: Saturated and Unsaturated Solutions - *CC BY-NC 4.0*
  - 7.13: How Temperature Influences Solubility - *CC BY-NC 4.0*
  - 7.14: Henry's Law - *CC BY-NC 4.0*
  - 7.15: Concentration - *CC BY-NC-SA 4.0*
  - 7.16: Molecular and Ionic Equations - *CC BY-NC 4.0*
  - 7.17: Net Ionic Equations - *CC BY-NC 4.0*
  - 7.18: Predicting Precipitates Using Solubility Rules - *CC BY-NC 4.0*
  - 7.19: Chapter 7 Exercises - *Undeclared*
    - 7.19.1: Solutions (Exercises) - *Undeclared*
    - 7.19.2: Answers to Solutions (Exercises) - *Undeclared*
- 8: Acids and Bases - *Undeclared*
  - 8.1: Prelude to Acids and Bases - *CC BY-NC-SA 4.0*
  - 8.2: Arrhenius Definition of Acids and Bases - *CC BY-NC-SA 4.0*
  - 8.3: Brønsted-Lowry Definition of Acids and Bases - *CC BY-NC-SA 4.0*
  - 8.4: Water - Both an Acid and a Base - *CC BY-NC-SA 4.0*
  - 8.5: The Strengths of Acids and Bases - *CC BY-NC-SA 4.0*
  - 8.6: Buffers - *CC BY-NC-SA 4.0*
  - 8.7: Chapter 8 Exercises - *Undeclared*
    - 8.7.1: Acids and Bases (Exercises) - *CC BY-NC-SA 4.0*
    - 8.7.2: Acids and Bases (Summary) - *CC BY-NC-SA 4.0*

- Back Matter - *Undeclared*
  - Index - *Undeclared*
  - Glossary - *Undeclared*
  - Detailed Licensing - *Undeclared*