CHEM100: FUNDAMENTALS OF CHEMISTRY



Brent Hamaoka
Palomar College

LibreTexts™

Palomar College

CHEM100: Fundamentals of Chemistry

Brent Hamaoka

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

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

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



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

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

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

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

This text was compiled on 01/12/2026

TABLE OF CONTENTS

Licensing

1: Introduction to Chemistry

- 1.1: The Scope of Chemistry
- 1.2: Chemicals Compose Ordinary Things
- 1.3: Hypothesis, Theories, and Laws
- 1.4: The Scientific Method - How Chemists Think
- 1.5: A Beginning Chemist - How to Succeed
- 1.6: Solid, Liquid, and Gas
- 1.7: Physical and Chemical Changes
- 1.8: Classification of Matter
- Index

4: Measurements, Conversion Factors, and Calculations

- 4.1: Taking Measurements
- 4.2: Scientific Notation: Writing Large and Small Numbers
- 4.3: Significant Figures: Writing Numbers to Reflect Precision
- 4.4: Significant Figures in Calculations
- 4.5: The Basic Units of Measurement
- 4.6: Problem Solving and Unit Conversions
- 4.7: Solving Multistep Conversion Problems
- 4.8: Units Raised to a Power
- 4.10: Energy and Heat Capacity Calculations
- 4.9: Density

2: Matter

- 2.4: Density and its Applications
- 2.1: In Your Room
- 2.2: What is Matter?
- 2.3: Classifying Matter According to Its State: Solid, Liquid, and Gas
- 2.5: Classifying Matter According to Its Composition
- 2.6: Differences in Matter: Physical and Chemical Properties
- 2.7: Changes in Matter: Physical and Chemical Changes

5: Atoms and Elements

- 5.1: Cutting Alumimun until you get Atoms
- 5.2: Indivisible: The Atomic Theory
- 5.3: The Nuclear Atom
- 5.4: The Properties of Protons, Neutrons, and Electrons
- 5.5: Elements: Defined by Their Number of Protons
- 5.6: Looking for Patterns: The Periodic Table
- 5.7: Ions - Losing and Gaining Electrons
- 5.8: Isotopes - When the Number of Neutrons Varies
- 5.9: Atomic Mass: The Average Mass of an Element's Atoms

6: Molecules and Compounds

- 6.1: Sugar and Salt
- 6.2: Compounds Display Constant Composition
- 6.3: Chemical Formulas: How to Represent Compounds
- 6.4: A Molecular View of Elements and Compounds
- 6.5: Writing Formulas for Ionic Compounds
- 6.6: Nomenclature: Naming Compounds
- 6.7: Naming Ionic Compounds
- 6.8: Naming Molecular Compounds
- 6.9: Naming Acids
- 6.10: Nomenclature Summary
- 6.11: Formula Mass: The Mass of a Molecule or Formula Unit

7: Quantitative Composition of Compounds

- 7.1: How Much Sodium?
- 7.2: Counting Nails by the Pound
- 7.3: Counting Atoms by the Gram
- 7.4: Counting Molecules by the Gram
- 7.5: Chemical Formulas as Conversion Factors
- 7.6: Mass Percent Composition of Compounds
- 7.7: Mass Percent Composition from a Chemical Formula
- 7.8: Calculating Empirical Formulas for Compounds
- 7.9: Calculating Molecular Formulas for Compounds

8: Chemical Reactions

- 8.1: Evidence of a Chemical Reaction
- 8.2: Chemical Equations
- 8.3: How to Write Balanced Chemical Equations
- 8.4: Classifying Chemical Reactions

12: Quantities in Chemical Reactions

- 12.1: Climate Change - Too Much Carbon Dioxide
- 12.2: Stoichiometry
- 12.3: Mole-to-Mole Conversions
- 12.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions
- 12.5: Limiting Reactant and Theoretical Yield
- 12.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants
- 12.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

14: Electrons in Atoms and the Periodic Table

- 14.1: The Perioidic Table
- 14.2: Electron Configurations
- 14.3: Electron Configurations and the Periodic Table (The Block Method)
- 14.4: Electron Configurations of Ions
- 14.5: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character

1.7: Classification of Matter

10: Acid-Base and Gas Evolution Reactions

- 10.1: Acids: Properties and Examples
- 10.2: Bases: Properties and Examples
- 10.3: Molecular Definitions of Acids and Bases
- 10.4: Reactions of Acids and Bases
- 10.5: Acid–Base and Gas Evolution Reactions

11: Redox Reactions

- 11.1: Oxidation and Reduction - Some Definitions
- 11.2: Electronegativity
- 11.3: Oxidation States: Electron Bookkeeping
- 11.4: Oxidation States & Redox Reactions
- 11.5: Half-Reactions
- 11.6: The Activity Series: Predicting Spontaneous Redox Reactions
- 11.7: Combustion Reactions

13: Solutions

- 13.1: Prelude - Tragedy in Cameroon
- 13.2: Solutions: Homogeneous Mixtures
- 13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy
- 13.4: Solutions of Gases in Water
- 13.5: Solution Concentration: Mass Percent
- 13.6: Solution Concentration: Molarity
- 13.7: Solution Dilution
- 13.8: Solution Stoichiometry
- 13.10: Water - Acid and Base in One
- 13.11: The pH and pOH Scales - Ways to Express Acidity and Basicity
- 13.9: Acid–Base Titration

15: Chemical Bonding

- 15.1: Representing Valence Electrons with Dots
- 15.2: Electrons are Transferred in Ionic Compounds
- 15.3: Lewis Structures: Electrons Shared
- 15.4: Lewis Structures: Counting Valence Electrons
- 15.5: Predicting the Shapes of Molecules
- 15.6: Electronegativity and Polarity - Why Oil and Water Don’t Mix

16: Gases

- 16.1: Extra-Long Straws
- 11.2: Kinetic Molecular Theory: Real Gases are Not Ideal
- 16.3: Pressure: The Result of Constant Molecular Collisions
- 16.4: Boyle’s Law: Pressure and Volume
- 16.5: Charles’s Law: Volume and Temperature
- 16.6: Gay-Lussac's Law: Temperature and Pressure
- 16.7: The Combined Gas Law: Pressure, Volume, and Temperature

- 16.8: Avogadro’s Law: Volume and Moles
- 16.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles

2.4: Solid, Liquid, and Gas

3: Introduction to Energy

- 3.1: Conservation of Mass - There is No New Matter
- 3.2: Energy
- 3.3: Energy and Chemical and Physical Change
- 3.4: Temperature: Random Motion of Molecules and Atoms
- 3.5: Temperature Changes - Heat Capacity

9: Precipitation Reactions

- 9.1: Aqueous Solutions and Solubility: Compounds Dissolved in Water
- 9.2: Precipitation Reactions
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

Chapter 3: Introduction to Energy

Index

Glossary

Detailed Licensing

Licensing

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

CHAPTER OVERVIEW

1: Introduction to Chemistry

Template:HideTOC

1.1: The Scope of Chemistry
1.2: Chemicals Compose Ordinary Things
1.3: Hypothesis, Theories, and Laws
1.4: The Scientific Method - How Chemists Think
1.5: A Beginning Chemist - How to Succeed
1.6: Solid, Liquid, and Gas
1.7: Physical and Chemical Changes
1.8: Classification of Matter
Index

1: Introduction to Chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

Front Matter

TitlePage
InfoPage

1: Introduction to Chemistry

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

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

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



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

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

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

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

This text was compiled on 01/12/2026

1.1: The Scope of Chemistry

Learning Objectives

- To recognize the breadth, depth, and scope of chemistry.
- Define chemistry in relation to other sciences.
- Identify the main disciplines of chemistry.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—the screen on which you are reading this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter. Because of this, chemistry is a science that has its fingers in just about everything. Being able to describe the ingredients in a cake and how they change when the cake is baked, for example, is chemistry!

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 **1.1.1**). 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.

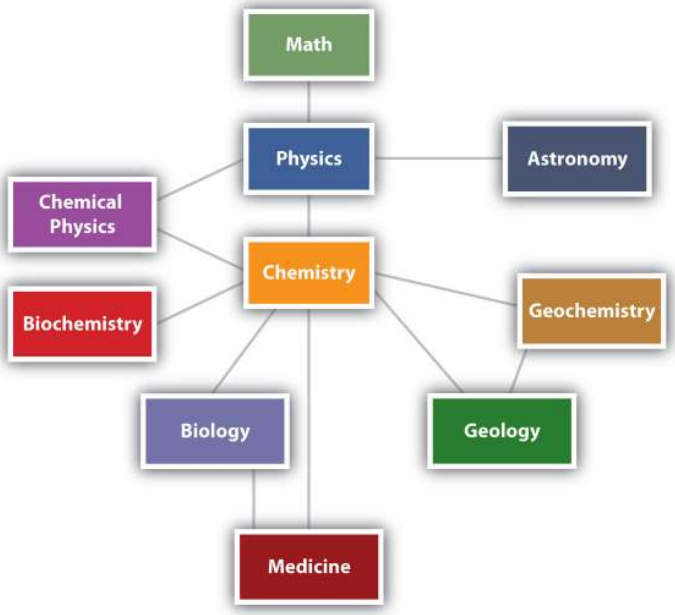


Figure **1.1.1**: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

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 **1.1.1** shows how many of the individual fields of science are related. At some level, all of these fields depend on matter because they all involve "stuff"; because of this, chemistry has been called the "central science", linking them all together.

> *There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed here.*

Example 1.1.1: Science Fields

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Solution:

a. Sculpture is not considered a science because it is not a study of some aspect of the natural universe.
b. Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

Exercise 1.1.1

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:

yes

Answer b:

yes

Answer c:

yes

Answer d:

no

Areas of Chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study:

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- **Analytical chemistry:** Analytical chemistry is the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so better treatments can be developed.



Figure **1.1.2**: *(left) Measurement of trace metals using atomic spectroscopy.(right) Measuring hormone concentrations.*

In practice, chemical research is often not limited to just one of the five major disciplines. A particular chemist may use biochemistry to isolate a particular chemical found in the human body such as hemoglobin, the oxygen carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to the areas of physical or analytical chemistry. Many chemists specialize in areas that are combinations of the main disciplines, such as bioinorganic chemistry or physical organic chemistry.

History of Chemistry

The **history** of chemistry is an interesting and challenging one. Very early chemists were often motivated mainly by the achievement of a specific goal or product. Making perfume or soaps did not need a lot of **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that we could all agree on). It is often difficult to figure out exactly what a particular person was using. However, the science developed over the centuries by trial and error.

Major progress was made in putting chemistry on a solid foundation when Robert Boyle (1637 - 1691) began his **research** in chemistry Figure **1.1.3**. He developed the basic ideas about the behavior of gases. He could then describe gases mathematically. Boyle also helped form the idea that small particles could combine to form molecules. Many years later, John Dalton used these ideas to develop the atomic theory.



Figure **1.1.3**: *Robert Boyle.*

The field of chemistry began to develop rapidly in the 1700's. Joseph Priestley (1733 - 1804) isolated and characterized several gases: oxygen, carbon monoxide, and nitrous oxide. It was later discovered that nitrous oxide ("laughing gas") worked as an anesthetic. This gas was used for that purpose for the first time in 1844 during a tooth extraction. Other gases discovered during that time were chlorine, by C.W. Scheele (1742 - 1786) and nitrogen, by Antoine Lavoisier (1743 - 1794). Lavoisier has been considered by many scholars to be the "father of chemistry". Among other accomplishments, he discovered the role of oxygen in combustion and definitively formulated the law of conservation of matter.

Chemists continued to discover new compounds in the 1800's. The science also began to develop a more theoretical foundation. John Dalton (1766 - 1844) put forth his atomic theory in 1807. This idea allowed scientists to think about chemistry in a much

more systematic way. Amadeo Avogadro (1776 - 1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas. A lot of effort was put forth in studying chemical reactions. These efforts led to new materials being produced. Following the invention of the battery by Alessandro Volta (1745 - 1827), the field of electrochemistry (both theoretical and applications) developed through major contributions by Humphry Davy (1778 - 1829) and Michael Faraday (1791 - 1867). Other areas of the discipline also progressed rapidly.

It would take a large book to cover developments in chemistry during the twentieth century and up to today. One major area of expansion was in the area of the chemistry of living processes. Research in photosynthesis in plants, the discovery and characterization of enzymes as biochemical catalysts, elucidation of the structures of biomolecules such as insulin and DNA - these efforts gave rise to an explosion of information in the field of biochemistry.

The practical aspects of chemistry were not ignored. The work of Volta, Davy, and Faraday eventually led to the development of batteries that provided a source of electricity to power a number of devices (Figure **1.1.4**).

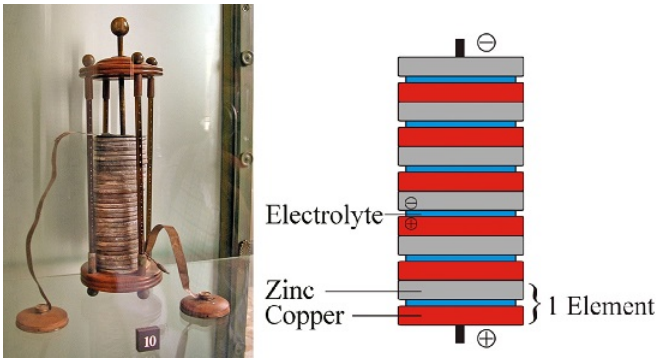


Figure **1.1.4**: *Battery developed by Volta. (CC BY-SA 3.0; (lef)* GuidoB and*(right)* Kkkdc)

Charles Goodyear (1800 - 1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles we have today. Louis Pasteur (1822 - 1895) pioneered the use of heat sterilization to eliminate unwanted microorganisms in wine and milk. Alfred Nobel (1833 - 1896) invented dynamite (Figure **1.1.5**). After his death, the fortune he made from this product was used to fund the Nobel Prizes in science and the humanities. J.W. Hyatt (1837 - 1920) developed the first plastic. Leo Baekeland (1863 - 1944) developed the first synthetic resin, widely used for inexpensive and sturdy dinnerware.



Figure **1.1.5**: *Dynamite explosion in Panama, Central America (1908).*

Today, chemistry continues to be essential to the development of new materials and technologies, from semiconductors for electronics to powerful new medicines, and beyond.

Summary

- Chemistry is the study of matter and the changes it undergoes and considers both *macroscopic* and *microscopic* information.
- Matter is anything that has mass and occupies space.
- The five main disciplines of chemistry are physical chemistry, organic chemistry, Inorganic chemistry, analytical chemistry and biochemistry.

- Many civilizations contributed to the growth of chemistry. A lot of early chemical research focused on practical uses. Basic chemistry theories were developed during the nineteenth century. New materials and batteries are a few of the products of modern chemistry.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

1.1: The Scope of Chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.2: Chemicals Compose Ordinary Things

Chemistry is the branch of science dealing with the structure, composition, properties, and the reactive characteristics of matter. Matter is anything that has mass and occupies space. Thus, chemistry is the study of literally everything around us – the liquids that we drink, the gasses we breathe, the composition of everything from the plastic case on your phone to the earth beneath your feet. Moreover, chemistry is the study of the transformation of matter. Crude oil is transformed into more useful petroleum products such as gasoline and kerosene by the process of refining. Some of these products are further transformed into plastics. Crude metal ores are transformed into metals, that can then be fashioned into everything from foil to automobiles. Potential drugs are identified from natural sources, isolated and then prepared in the laboratory. Their structures are systematically modified to produce the pharmaceuticals that have led to vast advances in modern medicine. Chemistry is at the center of all of these processes and chemists are the people that study the nature of matter and learn to design, predict and control these chemical transformations. Within the branches of chemistry you will find several apparent subdivisions. Inorganic chemistry, historically, focused on minerals and metals found in the earth, while organic chemistry dealt with carbon-containing compounds that were first identified in living things. Biochemistry is an outgrowth of the application of organic chemistry to biology and relates to the chemical basis for living things. In the later chapters of this text we will explore organic and biochemistry in a bit more detail and you will notice examples of organic compounds scattered throughout the text. Today, the lines between the various fields have blurred significantly and a contemporary chemist is expected to have a broad background in all of these areas.

In this chapter we will discuss some of the properties of matter, how chemists measure those properties and we will introduce some of the vocabulary that is used throughout chemistry and the other physical sciences.

Let's begin with matter. **Matter** is defined as any substance that has mass. It's important to distinguish here between weight and mass. Weight is the result of the pull of gravity on an object. On the Moon, an object will weigh less than the same object on Earth because the pull of gravity is less on the Moon. The mass of an object, however, is an inherent property of that object and does not change, regardless of location, gravitational pull, or whatever. It is a property that is solely dependent on the quantity of matter within the object.

Contemporary theories suggests that matter is composed of **atoms**. Atoms themselves are constructed from neutrons, protons and electrons, along with an ever-increasing array of other subatomic particles. We will focus on the neutron, a particle having no charge, the proton, which carries a positive charge, and the electron, which has a negative charge. Atoms are incredibly small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that's 28 sextillion). Because atoms and subatomic particles are so small, their mass is not readily measured using pounds, ounces, grams or any other scale that we would use on larger objects. Instead, the mass of atoms and subatomic particles is measured using atomic mass units (abbreviated amu). The atomic mass unit is based on a scale that relates the mass of different types of atoms to each other (using the most common form of the element carbon as a standard). The amu scale gives us a convenient means to describe the masses of individual atoms and to do quantitative measurements concerning atoms and their reactions. Within an atom, the neutron and proton both have a mass of one amu; the electron has a much smaller mass (about 0.0005 amu).



Figure **1.2.1**: Atoms are incredible small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that's 28 sextillion).

Atomic theory places the neutron and the proton in the center of the atom in the nucleus. In an atom, the nucleus is very small, very dense, carries a positive charge (from the protons) and contains virtually all of the mass of the atom. Electrons are placed in a diffuse cloud surrounding the nucleus. The electron cloud carries a net negative charge (from the charge on the electrons) and in a neutral atom there are always as many electrons in this cloud as there are protons in the nucleus (the positive charges in the nucleus are balanced by the negative charges of the electrons, making the atom neutral).

An atom is characterized by the number of neutrons, protons and electrons that it possesses. Today, we recognize at least 116 different types of atoms, each type having a different number of protons in its nucleus. These different types of atoms are called elements. The neutral element hydrogen (the lightest element) will always have one proton in its nucleus and one electron in the cloud surrounding the nucleus. The element helium will always have two protons in its nucleus. It is the number of protons in the nucleus of an atom that defines the identity of an element. Elements can, however, have differing numbers of neutrons in their nucleus. For example, stable helium nuclei exists that contain one, or two neutrons (but they all have two protons). These different types of helium atoms have different masses (3 or 4 amu) and they are called isotopes. For any given isotope, the sum of the numbers of protons and neutrons in the nucleus is called the mass number. All elements exist as a collection of isotopes, and the mass of an element that we use in chemistry, the atomic mass, is the average of the masses of these isotopes. For helium, there is approximately one isotope of Helium-3 for every million isotopes of Helium-4, hence the average atomic mass is very close to 4 (4.002602).

As different elements were discovered and named, abbreviations of their names were developed to allow for a convenient chemical shorthand. The abbreviation for an element is called its chemical symbol. A chemical symbol consists of one or two letters, and the relationship between the symbol and the name of the element is generally apparent. Thus helium has the chemical symbol He, nitrogen is N, and lithium is Li. Sometimes the symbol is less apparent but is decipherable; magnesium is Mg, strontium is Sr, and manganese is Mn. Symbols for elements that have been known since ancient times, however, are often based on Latin or Greek names and appear somewhat obscure from their modern English names. For example, copper is Cu (from cuprum), silver is Ag (from argentum), gold is Au (from aurum), and iron is (Fe from ferrum). Throughout your study of chemistry, you will routinely use chemical symbols and it is important that you begin the process of learning the names and chemical symbols for the common elements. By the time you complete General Chemistry, you will find that you are adept at naming and identifying virtually all of the 116 known elements. Table **1.2.1** contains a starter list of common elements that you should begin learning now!

Table 1.2.1: Names and Chemical Symbols for Common Elements

Element	Chemical Symbol	Element	Chemical Symbol
Hydrogen	H	Phosphorus	P
Helium	He	Sulfur	S
Lithium	Li	Chlorine	Cl
Beryllium	Be	Argon	Ar
Boron	B	Potassium	K
Carbon	C	Calcium	Ca
Nitrogen	N	Iron	Fe
Oxygen	O	Copper	Cu
Fluorine	F	Zinc	Zn
Neon	Ne	Bromine	Br
Sodium	Na	Silver	Ag
Magnesium	Mg	Iodine	I
Aluminum	Al	Gold	Au
Silicon	Si	Lead	Pb

The chemical symbol for an element is often combined with information regarding the number of protons and neutrons in a particular isotope of that atom to give the atomic symbol. To write an atomic symbol, you begin with the chemical symbol, then write the atomic number for the element (the number of protons in the nucleus) as a subscript, preceding the chemical symbol. Directly above this, as a superscript, now write the mass number for the isotope, that is, the total number of protons and neutrons in

the nucleus. Thus, for helium, the atomic number is 2 and there are two neutrons in the nucleus for the most common isotope making the atomic symbol $^{4}_{2}\mathrm{He}$. In the definition of the atomic mass unit, the "most common isotope of carbon", $^{12}_{6}\mathrm{C}$, is defined as having a mass of exactly 12 amu and the atomic masses of the remaining elements are based on their masses relative to this isotope. Chlorine (chemical symbol Cl) consists of two major isotopes, one with 18 neutrons (the most common, comprising 75.77% of natural chlorine atoms) and one with 20 neutrons (the remaining 24.23%). The atomic number of chlorine is 17 (it has 17 protons in its nucleus), therefore the chemical symbols for the two isotopes are $^{35}_{17}\mathrm{Cl}$ and $^{37}_{17}\mathrm{Cl}$.

When data is available regarding the natural abundance of various isotopes of an element, it is simple to calculate the average atomic mass. In the example above, $^{35}_{17}\mathrm{Cl}$ was the most common isotope with an abundance of 75.77% and $^{37}_{17}\mathrm{Cl}$ had an abundance of the remaining 24.23%. To calculate the average mass, first convert the percentages into fractions; that is, simply divide them by 100. Now, chlorine-35 represents a fraction of natural chlorine of 0.7577 and has a mass of 35 (the mass number). Multiplying these, we get (0.7577 × 35) = 26.51. To this, we need to add the fraction representing chlorine-37, or (0.2423 × 37) = 8.965; adding, (26.51 + 8.965) = 35.48, which is the weighted average atomic mass for chlorine. Whenever we do mass calculations involving elements or compounds (combinations of elements), we always need to use average atomic masses.

Contributors and Attributions

- Paul R. Young, Professor of Chemistry, University of Illinois at Chicago, Wiki: AskTheNerd; PRY@askthenerd.com - pyoung@uic.edu; ChemistryOnline.com
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

1.2: Chemicals Compose Ordinary Things is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.3: Hypothesis, Theories, and Laws

Learning Objectives

- Describe the difference between hypothesis, theory as scientific terms.
- Describe the difference between a theory and scientific law.

Although all of us have taken science classes throughout the course of our study, many people have incorrect or misleading ideas about some of the most important and basic principles in science. We have all heard of hypotheses, theories, and laws, but what do they really mean? Before you read this section, think about what you have learned about these terms before. What do these terms mean to you? What do you read contradicts what you thought? What do you read supports what you thought?

What is a Fact?

A fact is a basic statement establish by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

One of the most common terms used in science classes is a "hypothesis". The word can have many different definitions, depending on the context in which it is being used:

- "An educated guess" - because it provides a suggested solution based on evidence to be a scientific hypothesis
- Prediction - if you have ever carried out a science experiment, you probably made this type of hypothesis, in which you predicted the outcome of your experiment.
- Tentative or Proposed explanation - hypotheses can be suggestions about why something is observed, but in order for it to be scientific, we must be able to test the explanation to see if it works, if it is able to correctly predict what will happen in a situation, such as: if my hypothesis is correct, we should see ___ result when we perform ___ test.

> *A hypothesis is very tentative; it can be easily changed.*

What is a Theory?

The *United States National Academy of Sciences* describes what a theory is as follows:

> "Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."
>
> "A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact."

Not some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomenon. They aren't predictions (although we may use theories to make predictions). They are explanations why we observe something.
- Theories aren't likely to change. They have so much support and are able to explain satisfactorily so many observations, that they are not likely to change. Theories can, indeed, be facts. Theories can change, but it is a long and difficult process. In order for a theory to change, there must be many observations or evidence that the theory cannot explain.
- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight; it is not just one person's idea about something

> *Theories aren't likely to change.*

What is a Law?

Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are more overarching explanations of how nature works and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas is increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

What is a Belief?

A statement that is not scientifically provable. Beliefs may or may not be incorrect; they just are outside the realm of science to explore.

Laws vs. Theories

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has been accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, **theories are explanations** and **laws are patterns** we see in large amounts of data, frequently written as an equation. A theory will always remain a theory; a law will always remain a law.



Video* 1.3.1*: *What's the difference between a scientific law and theory?*

Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

1.3: Hypothesis, Theories, and Laws is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.4: The Scientific Method - How Chemists Think

Learning Objectives

- To identify the components of the scientific method

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments, which in turn lead to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.4.1).

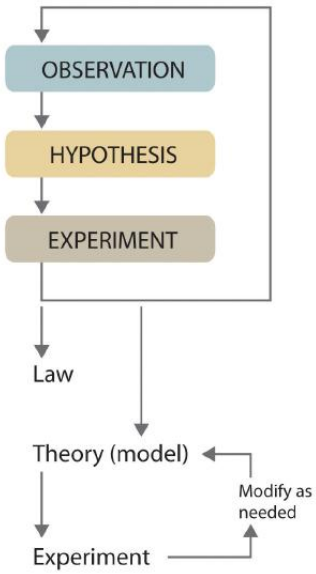


Figure 1.4.1: The Steps in the Scientific Method.

Step 1: Make observations

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

Step 2: Formulate a hypothesis

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

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

Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this

hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either support or refute it.

Step 3: Design and perform experiments

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

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether the original hypothesis is valid. In which case he can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified thus requiring further experimentation.

Step 5: Development into a law and/or theory

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

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

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

Because scientists can enter the cycle shown in Figure **1.4.1** at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

Example 1.4.1

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

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

Solution

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

Exercise 1.4.1

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

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."

b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter ever to play professional basketball.
e. Limestone is relatively insoluble in water but dissolves readily in dilute acid with the evolution of a gas.

Answer a:
experiment
Answer b:
law
Answer c:
theory
Answer d:
hypothesis
Answer e:
observation

Summary

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

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

Contributors and Attributions

-
- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

1.4: The Scientific Method - How Chemists Think is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.5: A Beginning Chemist - How to Succeed

Examples of the practical applications of chemistry are everywhere (Figure **1.5.1**). Engineers need to understand the chemical properties of the substances when designing biologically compatible implants for joint replacements or designing roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Archaeology and paleontology rely on chemical techniques to date bones and artifacts and identify their origins. Although law is not normally considered a field related to chemistry, forensic scientists use chemical methods to analyze blood, fibers, and other evidence as they investigate crimes. In particular, DNA matching—comparing biological samples of genetic material to see whether they could have come from the same person—has been used to solve many high-profile criminal cases as well as clear innocent people who have been wrongly accused or convicted. Forensics is a rapidly growing area of applied chemistry. In addition, the proliferation of chemical and biochemical innovations in industry is producing rapid growth in the area of patent law. Ultimately, the dispersal of information in all the fields in which chemistry plays a part requires experts who are able to explain complex chemical issues to the public through television, print journalism, the Internet, and popular books.

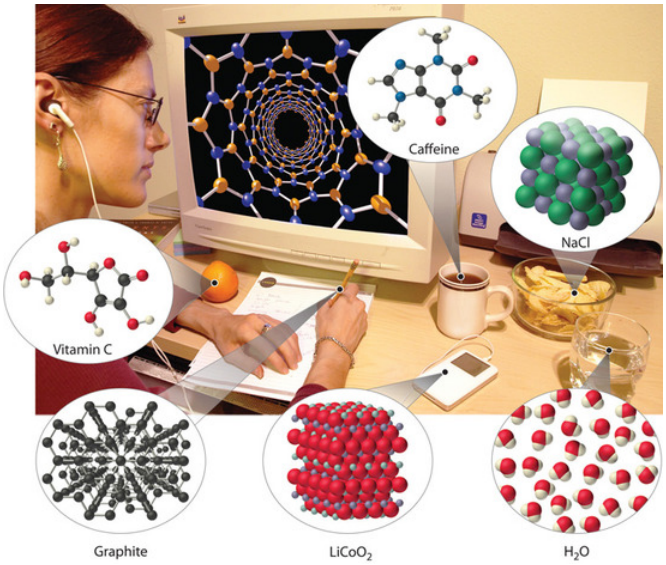


Figure **1.5.1**: Chemistry in Everyday Life. Although most people do not recognize it, chemistry and chemical compounds are crucial ingredients in almost everything we eat, wear, and use.

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

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

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

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

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

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

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



***Video* 1.5.1:** *How To Study Chemistry*

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

1.5: A Beginning Chemist - How to Succeed is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.6: Solid, Liquid, and Gas

Water can take many forms. At low temperatures (below $0°C$), it is a solid. When at "normal" temperatures (between $0°C$ and $100°C$), it is a liquid. While at temperatures above $100°C$, water is a gas (steam). The state the water is in depends upon the temperature. Each state (solid, liquid, and gas) has its own unique set of physical properties.

Matter and Its States

Matter typically exists in one of three states: **solid**, **liquid**, or **gas**. The state a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states.



Figure 1.6.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From top to bottom: quartz (solid), water (liquid), nitrogen dioxide (gas),

Note

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

Solid

Solids are defined by the following characteristics:

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

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

Liquid

Liquids have the following characteristics:

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

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



Mercury.

If we heat liquid mercury to its boiling point of **357°C**, and under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gas

Gases have the following characteristics:

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

Summary

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

Contributors and Attributions

-

1.6: Solid, Liquid, and Gas is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.7: Physical and Chemical Changes

Learning Outcomes

- Distinguish between physical and chemical changes.
- Give examples of physical and chemical changes.

Physical Changes

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.



Figure 10.4.1: *Melting ice in the Beaufort Sea.*

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust could not be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes. Firewood also represents an irreversible physical change since the pieces cannot be put back together to form the tree.



Figure 10.4.2: *Firewood being cut is a physical change because the composition doesn't change when being cut.*

Chemical Changes

When exposed to air, an object made of iron will eventually begin to rust (see figure below).



Figure 10.4.3: *Rust (iron oxide) forms on an unprotected iron surface.*

As the rust forms on the surface of the iron, it flakes off to expose more iron, which will continue to rust. Rust is clearly a substance that is different from iron. Rusting is an example of a chemical change.

A **chemical property** *describes the ability of a substance to undergo a specific chemical change*. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust. A more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are *burn*, *rot*, *explode*, and *ferment*. Chemical properties are very useful as a way of identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

A chemical change is also called a chemical reaction. A **chemical reaction** *is a process that occurs when one or more substances are changed into one or more new substances*. Zinc (Zn) is a silver-gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (S), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (ZnS). Pictured below are the substances involved in this reaction.

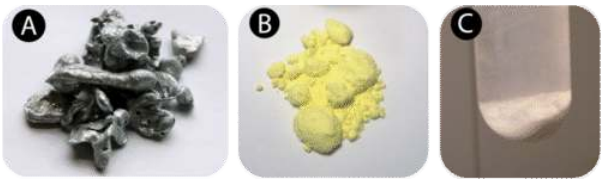


Figure 10.4.4: *Zinc (A) and sulfur (B) are two elements that undergo a chemical reaction when heated to form the the compound zinc sulfide (C).*

The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the reaction as:

$$\text{zinc} + \text{sulfur} \rightarrow \text{zinc sulfide} \tag{1.7.1}$$

A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:

$$\mathrm{Zn} + \mathrm{S} \rightarrow \mathrm{ZnS} \tag{1.7.2}$$

The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** *is a substance that is present at the start of a chemical reaction.* The substance(s) to the right of the arrow are called products. A **product** *is a substance that is present at the end of a chemical reaction*. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.

Recognizing Chemical Reactions

How can you tell if a chemical reaction is taking place? Certain visual clues indicate that a chemical reaction is likely (but not necessarily) occurring, including the following examples:

1. A change of color occurs during the reaction.
2. A gas is produced during the reaction.
3. A solid product, called a precipitate, is produced in the reaction.
4. A visible transfer of energy occurs in the form of light as a result of the reaction.

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced (see figure below). The production of a gas is also an indication that a chemical reaction may be occurring.



Figure 10.4.5: *Zinc reacts with hydrochloric acid to produce bubbles of hydrogen gas.*

When a colorless solution of lead (II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced (see figure below). A **precipitate** *is a solid product that forms from a reaction and settles out of a liquid mixture*. The formation of a precipitate may also indicate the occurrence of a chemical reaction.

$$\mathrm{Pb(NO_3)_2\,(aq) + 2KI\,(aq) \rightarrow PbI_2\,(s) + 2KNO_3\,(aq)} \tag{1.7.3}$$



Figure 10.4.6: *A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed.*

Contributors and Attributions

-
- Allison Soult, Ph.D. (Department of Chemistry, University of Kentucky)

This page titled 1.7: Physical and Chemical Changes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Allison Soult.

1.8: Classification of Matter

Learning Objectives

- To classify matter.

Chemists study the structures, physical properties, and chemical properties of material substances. These consist of **matter**, which is anything that occupies space and has mass. Gold and iridium are matter, as are peanuts, people, and postage stamps. Smoke, smog, and laughing gas are matter. Energy, light, and sound, however, are not matter; ideas and emotions are also not matter.

The **mass** of an object is the quantity of matter it contains. Do not confuse an object's mass with its **weight**, which is a force caused by the gravitational attraction that operates on the object. Mass is a fundamental property of an object that does not depend on its location.In physical terms, the mass of an object is directly proportional to the force required to change its speed or direction. A more detailed discussion of the differences between weight and mass and the units used to measure them is included in Essential Skills 1 (Section 1.9). Weight, on the other hand, depends on the location of an object. An astronaut whose mass is 95 kg weighs about 210 lb on Earth but only about 35 lb on the moon because the gravitational force he or she experiences on the moon is approximately one-sixth the force experienced on Earth. For practical purposes, weight and mass are often used interchangeably in laboratories. Because the force of gravity is considered to be the same everywhere on Earth's surface, 2.2 lb (a weight) equals 1.0 kg (a mass), regardless of the location of the laboratory on Earth.

Under normal conditions, there are three distinct states of matter: solids, liquids, and gases. **Solids** are relatively rigid and have fixed shapes and volumes. A rock, for example, is a solid. In contrast, **liquids** have fixed volumes but flow to assume the shape of their containers, such as a beverage in a can. **Gases**, such as air in an automobile tire, have neither fixed shapes nor fixed volumes and expand to completely fill their containers. Whereas the volume of gases strongly depends on their temperature and **pressure** (the amount of force exerted on a given area), the volumes of liquids and solids are virtually independent of temperature and pressure. Matter can often change from one physical state to another in a process called a **physical change**. For example, liquid water can be heated to form a gas called steam, or steam can be cooled to form liquid water. However, such changes of state do not affect the chemical composition of the substance.

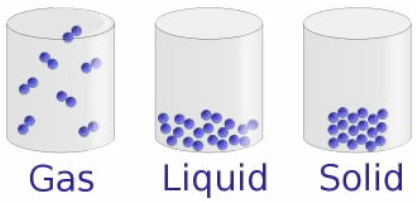


Figure used with permission from Wikipedia

Pure Substances and Mixtures

A pure chemical substance is any matter that has a fixed chemical composition and characteristic properties. Oxygen, for example, is a pure chemical substance that is a colorless, odorless gas at 25°C. Very few samples of matter consist of pure substances; instead, most are mixtures, which are combinations of two or more pure substances in variable proportions in which the individual substances retain their identity. Air, tap water, milk, blue cheese, bread, and dirt are all mixtures. If all portions of a material are in the same state, have no visible boundaries, and are uniform throughout, then the material is **homogeneous**. Examples of homogeneous mixtures are the air we breathe and the tap water we drink. Homogeneous mixtures are also called solutions. Thus air is a solution of nitrogen, oxygen, water vapor, carbon dioxide, and several other gases; tap water is a solution of small amounts of several substances in water. The specific compositions of both of these solutions are not fixed, however, but depend on both source and location; for example, the composition of tap water in Boise, Idaho, is not the same as the composition of tap water in Buffalo, New York. Although most solutions we encounter are liquid, solutions can also be solid. The gray substance still used by some dentists to fill tooth cavities is a complex solid solution that contains 50% mercury and 50% of a powder that contains mostly silver, tin, and copper, with small amounts of zinc and mercury. Solid solutions of two or more metals are commonly called alloys.

If the composition of a material is not completely uniform, then it is **heterogeneous** (e.g., chocolate chip cookie dough, blue cheese, and dirt). Mixtures that appear to be homogeneous are often found to be heterogeneous after microscopic examination. Milk, for example, appears to be homogeneous, but when examined under a microscope, it clearly consists of tiny globules of fat and protein dispersed in water. The components of heterogeneous mixtures can usually be separated by simple means. Solid-liquid mixtures such as sand in water or tea leaves in tea are readily separated by filtration, which consists of passing the mixture through a barrier, such as a strainer, with holes or pores that are smaller than the solid particles. In principle, mixtures of two or more solids, such as sugar and salt, can be separated by microscopic inspection and sorting. More complex operations are usually necessary, though, such as when separating gold nuggets from river gravel by panning. First solid material is filtered from river water; then the solids are separated by inspection. If gold is embedded in rock, it may have to be isolated using chemical methods.



Figure used with permission from Wikipedia

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

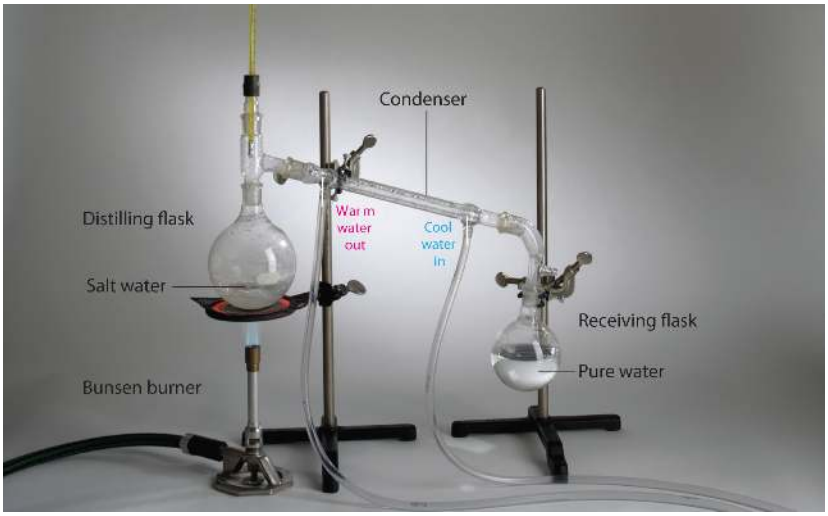


Figure 1.8.3: The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as

brandy or whiskey. (This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States!)

Crystallization separates mixtures based on differences in solubility, a measure of how much solid substance remains dissolved in a given amount of a specified liquid. Most substances are more soluble at higher temperatures, so a mixture of two or more substances can be dissolved at an elevated temperature and then allowed to cool slowly. Alternatively, the liquid, called the solvent, may be allowed to evaporate. In either case, the least soluble of the dissolved substances, the one that is least likely to remain in solution, usually forms crystals first, and these crystals can be removed from the remaining solution by filtration.



Figure 1.8.4: The Crystallization of Sodium Acetate from a Concentrated Solution of Sodium Acetate in Water. The addition of a small "seed" crystal (a) causes the compound to form white crystals, which grow and eventually occupy most of the flask. Video can be found here: https://www.youtube.com/watch?v=BLq5NibwV5g

Most mixtures can be separated into pure substances, which may be either elements or compounds. An **element**, such as gray, metallic sodium, is a substance that cannot be broken down into simpler ones by chemical changes; a **compound**, such as white, crystalline sodium chloride, contains two or more elements and has chemical and physical properties that are usually different from those of the elements of which it is composed. With only a few exceptions, a particular compound has the same elemental composition (the same elements in the same proportions) regardless of its source or history. The chemical composition of a substance is altered in a process called a **chemical change**. The conversion of two or more elements, such as sodium and chlorine, to a chemical compound, sodium chloride, is an example of a chemical change, often called a chemical reaction. Currently, about 118 elements are known, but millions of chemical compounds have been prepared from these 118 elements. The known elements are listed in the periodic table.

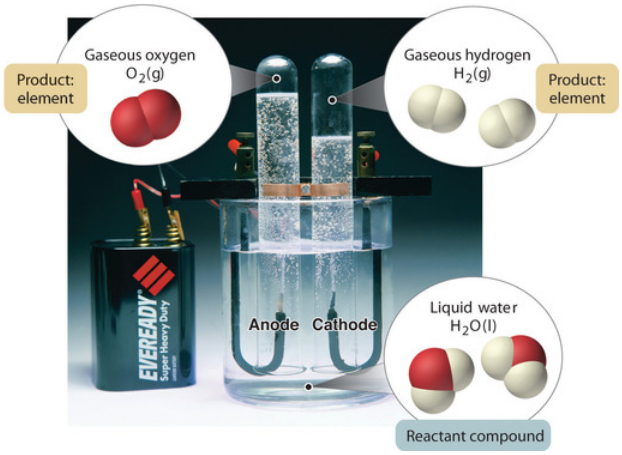


Figure 1.8.5: The Decomposition of Water to Hydrogen and Oxygen by Electrolysis. Water is a chemical compound; hydrogen and oxygen are elements.

In general, a reverse chemical process breaks down compounds into their elements. For example, water (a compound) can be decomposed into hydrogen and oxygen (both elements) by a process called electrolysis. In electrolysis, electricity provides the energy needed to separate a compound into its constituent elements (Figure 1.8.5). A similar technique is used on a vast scale to obtain pure aluminum, an element, from its ores, which are mixtures of compounds. Because a great deal of energy is required for

electrolysis, the cost of electricity is by far the greatest expense incurred in manufacturing pure aluminum. Thus recycling aluminum is both cost-effective and ecologically sound.

The overall organization of matter and the methods used to separate mixtures are summarized in Figure **1.8.6**.

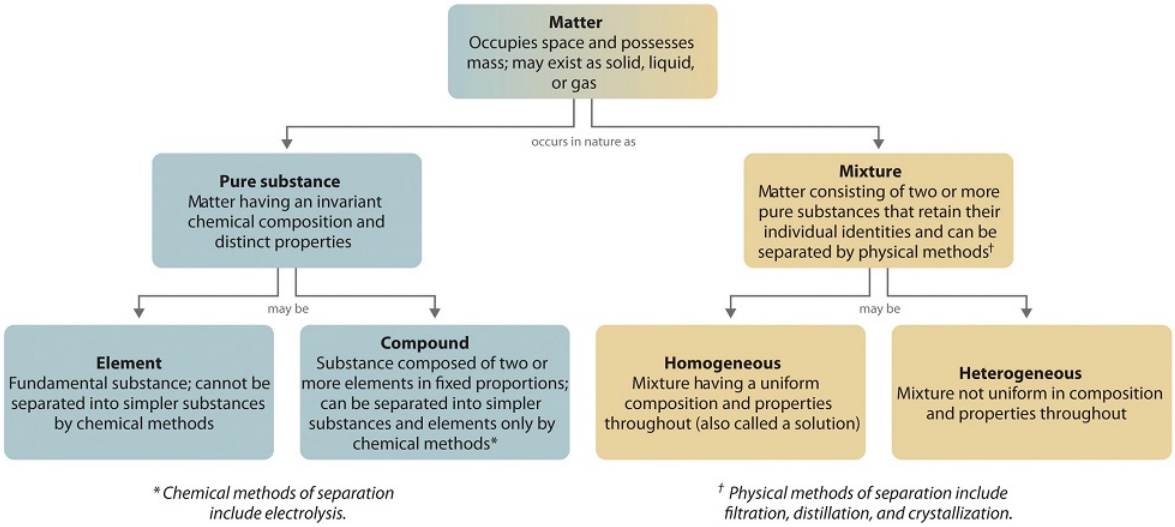


Figure **1.8.6**: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Example 1.8.1

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

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

Solution:

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

Exercise 1.8.1

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

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

Answer A

solution

Answer B

element

Answer C

heterogeneous mixture

Answer D

compound

Summary

Matter can be classified according to physical and chemical properties. Matter is anything that occupies space and has mass. The three states of matter are solid, liquid, and gas. A physical change involves the conversion of a substance from one state of matter to another, without changing its chemical composition. Most matter consists of mixtures of pure substances, which can be homogeneous (uniform in composition) or heterogeneous (different regions possess different compositions and properties). Pure substances can be either chemical compounds or elements. Compounds can be broken down into elements by chemical reactions, but elements cannot be separated into simpler substances by chemical means. The properties of substances can be classified as either physical or chemical. Scientists can observe physical properties without changing the composition of the substance, whereas chemical properties describe the tendency of a substance to undergo chemical changes (chemical reactions) that change its chemical composition. Physical properties can be intensive or extensive. Intensive properties are the same for all samples; do not depend on sample size; and include, for example, color, physical state, and melting and boiling points. Extensive properties depend on the amount of material and include mass and volume. The ratio of two extensive properties, mass and volume, is an important intensive property called density.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)

1.8: Classification of Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

Back Matter

Index

Index

A

acid–base titration
- 13.9: Acid–Base Titration

anion
- 5.7: Ions - Losing and Gaining Electrons

atomic mass unit
- 5.4: The Properties of Protons, Neutrons, and Electrons

Aufbau principle
- 14.2: Electron Configurations

B

Balancing a Chemical Equation
- 8.3: How to Write Balanced Chemical Equations

Bends
- 13.4: Solutions of Gases in Water

buoyancy
- 2.4: Density and its Applications

C

cation
- 5.7: Ions - Losing and Gaining Electrons

chemical change
- 1.7: Physical and Chemical Changes
- 1.8: Classification of Matter
- 2.7: Changes in Matter: Physical and Chemical Changes
- 1.7: Classification of Matter

chemical property
- 2.6: Differences in Matter: Physical and Chemical Properties

Coefficients and Subscripts
- 8.3: How to Write Balanced Chemical Equations

combination reaction
- 11.4: Oxidation States & Redox Reactions

complete ionic equation
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

compound
- 1.8: Classification of Matter
- 2.5: Classifying Matter According to Its Composition
- 1.7: Classification of Matter

conservation of energy
- 3.3: Energy and Chemical and Physical Change

conservation of mass
- 3.1: Conservation of Mass - There is No New Matter

crystallization
- 1.8: Classification of Matter
- 1.7: Classification of Matter

D

decomposition reaction
- 11.4: Oxidation States & Redox Reactions

density
- 2.4: Density and its Applications

disproportionation reaction
- 11.4: Oxidation States & Redox Reactions

distillation
- 1.8: Classification of Matter
- 1.7: Classification of Matter

E

electron configuration
- 14.2: Electron Configurations

element
- 1.8: Classification of Matter
- 2.5: Classifying Matter According to Its Composition
- 1.7: Classification of Matter

empirical formula
- 7.9: Calculating Molecular Formulas for Compounds

empirical formula mass
- 7.9: Calculating Molecular Formulas for Compounds

endothermic process
- 3.3: Energy and Chemical and Physical Change

exothermic process
- 3.3: Energy and Chemical and Physical Change

G

gas
- 1.8: Classification of Matter
- 1.7: Classification of Matter

H

heat capacity
- 3.5: Temperature Changes - Heat Capacity

Heat Capacity Calculations
- 4.10: Energy and Heat Capacity Calculations

heterogeneous mixture
- 1.8: Classification of Matter
- 1.7: Classification of Matter

homogeneous mixture
- 1.8: Classification of Matter
- 1.7: Classification of Matter

hund's rule
- 14.2: Electron Configurations

hydrometer
- 2.4: Density and its Applications

hypothesis
- 1.3: Hypothesis, Theories, and Laws

I

ions
- 5.7: Ions - Losing and Gaining Electrons

isotopes
- 5.8: Isotopes - When the Number of Neutrons Varies

L

law
- 1.3: Hypothesis, Theories, and Laws

liquid
- 1.8: Classification of Matter
- 1.7: Classification of Matter

M

mass
- 1.8: Classification of Matter
- 1.7: Classification of Matter

matter
- 1.8: Classification of Matter
- 2.2: What is Matter?
- 1.7: Classification of Matter

mixture
- 2.5: Classifying Matter According to Its Composition

MO
- 14.4: Electron Configurations of Ions

molarity
- 13.6: Solution Concentration: Molarity

molecular formula
- 7.9: Calculating Molecular Formulas for Compounds

N

net ionic equation
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

O

OpenChem
- 14.4: Electron Configurations of Ions

oxidation number
- 11.4: Oxidation States & Redox Reactions

oxidizing agent
- 11.4: Oxidation States & Redox Reactions

P

pauli exclusion principle
- 14.2: Electron Configurations

physical change
- 1.7: Physical and Chemical Changes
- 1.8: Classification of Matter
- 2.7: Changes in Matter: Physical and Chemical Changes
- 1.7: Classification of Matter

physical property
- 2.6: Differences in Matter: Physical and Chemical Properties

potential energy
- 3.2: Energy

pressure
- 1.8: Classification of Matter
- 1.7: Classification of Matter

R

redox reaction
- 11.4: Oxidation States & Redox Reactions
- 11.5: Half-Reactions

reducing agent
- 11.4: Oxidation States & Redox Reactions

relative abundances
- 5.9: Atomic Mass: The Average Mass of an Element's Atoms

S

Separation of Mixtures
- 2.7: Changes in Matter: Physical and Chemical Changes

significant figures
- 4.4: Significant Figures in Calculations

single replacement reaction
- 11.4: Oxidation States & Redox Reactions

solid
- 1.8: Classification of Matter
- 1.7: Classification of Matter

Solubility of gases
- 13.4: Solutions of Gases in Water

specific gravity
- 2.4: Density and its Applications

specific heat capacity
- 3.5: Temperature Changes - Heat Capacity

specific volume
- 2.4: Density and its Applications

spectator ions
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

Stability of Isotopes
- 5.8: Isotopes - When the Number of Neutrons Varies

Stock system
- 6.7: Naming Ionic Compounds

T

temperature
- 3.4: Temperature: Random Motion of Molecules and Atoms

theory
- 1.3: Hypothesis, Theories, and Laws

U

units of energy
- 3.2: Energy

V

valence electrons
- 14.2: Electron Configurations

W

weight
- 1.8: Classification of Matter
- 1.7: Classification of Matter

work
- 3.2: Energy

CHAPTER OVERVIEW

4: Measurements, Conversion Factors, and Calculations

Template:HideTOC

Chemistry, like all sciences, is quantitative. It deals with *quantities*, things that have amounts and units. Dealing with quantities is very important in chemistry, as is relating quantities to each other. In this chapter, we will discuss how we deal with numbers and units, including how they are combined and manipulated.

4.1: Taking Measurements
4.2: Scientific Notation: Writing Large and Small Numbers
4.3: Significant Figures: Writing Numbers to Reflect Precision
4.4: Significant Figures in Calculations
4.5: The Basic Units of Measurement
4.6: Problem Solving and Unit Conversions
4.7: Solving Multistep Conversion Problems
4.8: Units Raised to a Power
4.10: Energy and Heat Capacity Calculations
4.9: Density

4: Measurements, Conversion Factors, and Calculations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.1: Taking Measurements

Learning Objectives

- To 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 **4.1.1**); the nurse is also measuring.



Figure used with permission (GFDL; Pia von Lützau).

Chemists measure the properties of matter and express these measurements as quantities. A quantity is an amount of something and consists of a number and a unit. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5. If you ask a friend how far 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. *Both a number and a unit must be included to express a quantity properly.*

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

Example 4.1.1

Identify the number and the unit in each quantity.

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

Solution

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

Exercise 4.1.1

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 is 32, and the unit is fluid ounces.

Answer d:

The number is 98.6, and the unit is degrees Fahrenheit.

Key Takeaway

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

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

4.1: Taking Measurements is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.2: Scientific Notation: Writing Large and Small Numbers

Learning Objectives

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

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

$$N \times 10^n \tag{4.2.1}$$

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

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

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

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example 4.2.1.

Example 4.2.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

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

Solution

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, n = 2.	6.378×10^{2}
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, n = −2.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^{0} = 1$.)	7.86×10^{0}
d	Because the decimal point was moved four places to the left, n = 4.	1.2378×10^{4}
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^{1}
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^{3}
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Exercise 4.2.1

Convert each ordinary number to scientific notation or vice versa .

a. 67,000,000,000
b. 1,689
c. 12.6
d. 5.27×10^{4}
e. 1.0008×10^{6}

Answer a

6.7×10^{10}

Answer b

1.689×10^{3}

Answer c

1.26×10^{1}

Answer d

52,700.

Answer e

1,000,800.

Exercise 4.2.2

Convert each ordinary number to scientific notation or vice versa .

a. 0.000006567
b. −0.0004004
c. 0.000000000000123
d. 6.22×10^{-2}
e. 9.9×10^{-9}

Answer a

6.567×10^{-6}

Answer b

-4.004×10^{-4}

Answer c

1.23×10^{-13}

Answer d

0.0622

Answer 3

0.0000000099

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example **4.2.2** illustrates how to do this.

Example 4.2.2: Expressing Sums and Differences in Scientific Notation

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

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

Solution

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) = (1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) = (0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places: In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87×10^3
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) = (6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) = (69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.047×10^{-3}

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example **4.2.3**.

Example 4.2.3: Expressing Products and Quotients in Scientific Notation

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

a. $[(6.022 \times 10^{23})(6.42 \times 10^{-2})$

b. $\dfrac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$

c. $\dfrac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

	Explanation	Answer
a	In multiplication, we add the exponents: $$(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	3.87×10^{22}
b	In division, we subtract the exponents: $$\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}} = \frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^{4}$$ In converting 0.183×10^{4} to scientific notation, n has become more negative by 1 because the value of N has increased.	1.83×10^{3}
c	This problem has both multiplication and division: $$\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{(8.52 \times 10^{-2})} = \frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$$	4.7×10^{-31}

Exercise 4.2.3

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

a. $(6.78 \times 10^{4}) + (4.32 \times 10^{3})$

b. $\dfrac{8.67 \times 10^{-15}}{8.92 \times 10^{-20}}$

c. $(9.677 \times 10^{-5}) - (8.756 \times 10^{-6})$

d. $\dfrac{(3.46 \times 10^{-25})(9.1 \times 10)}{8.52 \times 10^{2}}$

Answer a

7.21 x 10^{4}

Answer b

9.72 x 10^{4}

Answer c

8.801 x 10^{-5}

Answer d

3.7 x 10^{-26}

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.2: Scientific Notation: Writing Large and Small Numbers is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.3: Significant Figures: Writing Numbers to Reflect Precision

Template:HideTOC

Learning Objectives

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

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

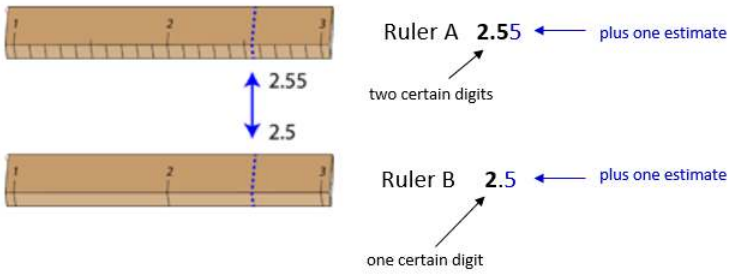


Figure 4.3.1: *Measurement with two different rulers.*

Measurement Uncertainty

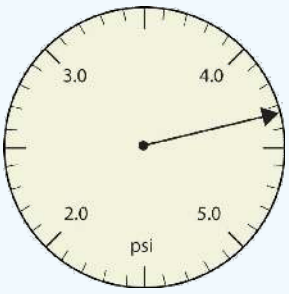
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest $0.1\ \text{g}$, other highly sensitive balances are capable of measuring to the nearest $0.001\ \text{g}$ or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 4.3.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

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

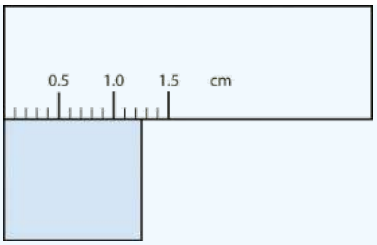
The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as $2.55\ \text{cm}$. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be $2.54\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

Example 4.3.1: Reporting Measurements to the Proper number of Significant Figures

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



a.

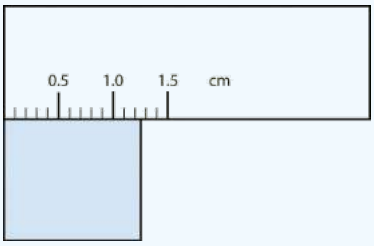
b. 

Solutions

	Explanation	Answer
a.	The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it's at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. The symbol psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire). The measurement is reported to three significant figures.	4.33 psi
b.	The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle's width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. The measurement is reported to three significant figures.	1.25 cm

Exercise 4.3.1

What would be the reported width of this rectangle?



Answer:

1.25 cm

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

Table 4.3.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.

4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous.
	  ◦ 1.4×10^3 has two significant figures.
	  ◦ 1.40×10^3 three significant figures.
	  ◦ 1.400×10^3 has four significant figures.

5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

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

Example 4.3.2

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

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

Solution

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

Exercise 4.3.2

Give the number of significant figures in each.

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

Answer a:

three significant figures.

Answer b:

four significant figures.

Answer c:

four significant figures.

Answer d:

Infinite (Exact number)

Answer e:

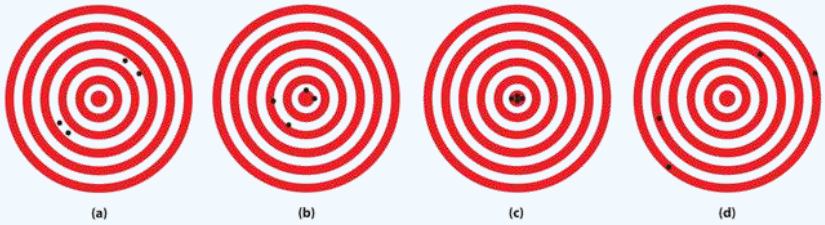
Ambiguous, could be two, three or four significant figure.

Accuracy and Precision

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

Example 4.3.3

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



Which target shows

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

Solution

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

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.3: Significant Figures: Writing Numbers to Reflect Precision is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.4: Significant Figures in Calculations

Template:HideTOC

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

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

Consider the measurement **207.518 m**. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table **4.4.1**.

Table 4.4.1: Rounding examples

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

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

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

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

Multiplication and Division

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

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

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

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

Example 4.4.1

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

a. 23.096×90.300
b. 125×9.000

SolutionS

	Explanation	Answer
a	The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	$\mathbf{2.0856 \times 10^3}$
b	The calculator gives 1,125 as the answer, but we limit it to three significant figures.	$\mathbf{1.13 \times 10^3}$

Addition and Subtraction

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

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

↑ limit final answer to the tenths column: 5.6

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

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

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

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

Example 4.4.2

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

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

SolutionS

	Explanation	Answer
a	The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$
b	The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2,001.06 = 2.001 \times 10^3$

Exercise 4.4.1

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

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

Answer a:

$0.240 = 2.40 \times 10^{-1}$

Answer b:

$1,437 = 1.437 \times 10^3$

Answer c:

$156 = 1.56 \times 10^2$

Answer d:

$2.14 = 2.14 \times 10^0$

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

Calculations Involving Multiplication/Division and Addition/Subtraction

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



***Video* 4.4.1:** *Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).*



***Video* 4.4.2:** https://www.youtube.com/watch?v=__csP0NtlGI

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

Example 4.4.3

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

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

SolutionS

	Explanation	Answer
a	2(1.008g) + 15.99 g = **Perform multiplication first.** 2 (1.008g *4 sig figs*) = **2.016 g** *4 sig figs* The number with the least number of significant figures is 1.008g; *the number 2 is an exact number and therefore has infinite number of significant figures*. **Then, perform the addition.** **2.016 g** *thousandths place* + 15.99 g *hundredths place (least precise)* = 18.006 g **Round the final answer.** Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)
b	137.3 s + 2 (35.45 s) = **Perform multiplication first.** 2 (35.45 s *4 sig figs*) = **70.90 s** *4 sig figs* The number with the least number of significant figures is 35.45; *the number 2 is an exact number and therefore has infinite number of significant figures.* **Then, perform the addition.** 137.3 s *tenths place (least precise)* + **70.90 s** *hundredths place* = 208.20 s **Round the final answer.** Round the final answer to the tenths place based on 137.3 s.	208.2 s
c	$\frac{118.7g}{2} - 35.5g$ = **Perform division first.** $\frac{118.7g}{2}$ *4 sig figs* = **59.35 g** *4 sig figs* The number with the least number of significant figures is 118.7g; the number 2 is an exact number and therefore has infinite number of significant figures. **Perform subtraction next.** **59.35 g** *hundredths place* − 35.5 g *tenths place (least precise)* = 23.85 g **Round the final answer.** Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

Exercise 4.4.2

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

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

Answer a

-5.62 s

Answer b

110.2 cm

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.4: Significant Figures in Calculations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.5: The Basic Units of Measurement

Template:HideTOC

Learning Objectives

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

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

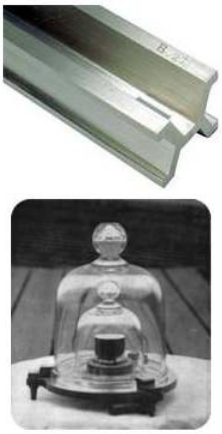


Figure 4.5.1: Meter standard (left) and Kilogram standard (right)

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 4.5.1.

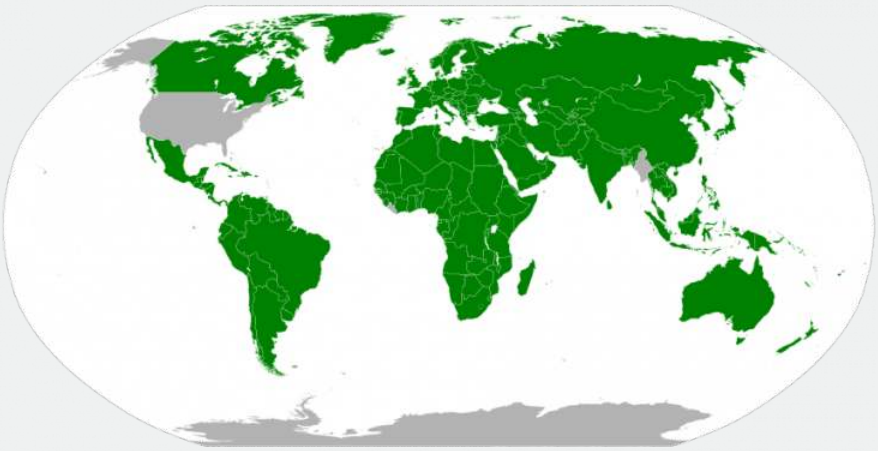
Table 4.5.1:: SI Base Units of Measurement

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

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

Unfortunately, The Metric System is Not Ubiquitous

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



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

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word μεγας, meaning "great". Table **4.5.2** lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 4.5.2: SI Prefixes

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

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

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

Example 4.5.1: Unit Abbreviations

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

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

SolutionS

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

Exercise 4.5.1

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

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

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.5: The Basic Units of Measurement is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.6: Problem Solving and Unit Conversions

Template:HideTOC

Learning Objectives

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

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

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can be used to convert meters into centimeters, or a more complex one can be used to convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. What always must be remembered is that a conversion factor has to represent a fact; this fact can either be simple or much more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you might be working with (Table 4.6.1).

Table 4.6.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	2.96 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume
*pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.		

Of course, there are other ratios which are not listed in Table 4.6.1. They may include:

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

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

$$1\ \text{cm} = \frac{1}{100}\ \text{m} = 10^{-2}\text{m} \tag{4.6.1}$$

or

$$100\ \text{cm} = 1\ \text{m} \tag{4.6.2}$$

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

$$\frac{100\ \text{cm}}{1\ \text{m}} = \frac{1\ \text{m}}{1\ \text{m}} \tag{4.6.3}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100\ \text{cm}}{1\ \text{m}} = \frac{1000\ \text{mm}}{1\ \text{m}} = \frac{1 \times 10^6 \mu\text{m}}{1\ \text{m}} = 1 \tag{4.6.4}$$

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

Performing Dimensional Analysis

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

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

$$3.55\ \text{m} \times \frac{100\ \text{cm}}{1\ \text{m}} \tag{4.6.5}$$

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

$$\frac{3.55\ \cancel{\text{m}}}{1} \times \frac{100\ \text{cm}}{1\ \cancel{\text{m}}} \tag{4.6.6}$$

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

$$\frac{3.55}{1} \times \frac{100\ \text{cm}}{1} = 355\ \text{cm} \tag{4.6.7}$$

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

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

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

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

$$3.55\ \text{m} \times \frac{1\ \text{m}}{100\ \text{cm}} = 0.0355 \frac{\text{m}^2}{\text{cm}} \tag{4.6.8}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure **4.6.1** shows a concept map for constructing a proper conversion.

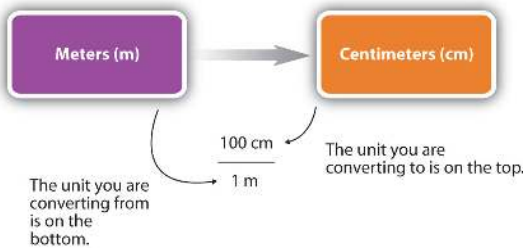


Figure **4.6.1**: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another.

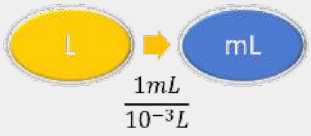
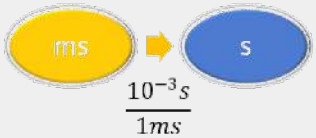
The general steps in performing dimensional analysis

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

Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

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

	Example 4.6.1	**Example 4.6.2**
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given"information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities	$1\,mL = 10^{-3}L$	$1\,ms = 10^{-3}s$
Prepare a concept map and use the proper conversion factor.	 $\frac{1mL}{10^{-3}L}$	 $\frac{10^{-3}s}{1ms}$
Cancel units and calculate.	$4.7\cancel{L} \times \frac{1\text{ mL}}{10^{-3}\ \cancel{L}} = 4,700\text{ mL}$ or $4.7\cancel{L} \times \frac{1,000\text{ mL}}{1\ \cancel{L}} = 4,700\text{ mL}$	$18\ \cancel{ms} \times \frac{10^{-3}\text{ s}}{1\ \cancel{ms}} = 0.018\text{ s}$ or $18\ \cancel{ms} \times \frac{1\text{ s}}{1,000\ \cancel{ms}} = 0.018\text{ s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

Exercise 4.6.1

Perform each conversion.

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

Answer a:

$1.01000x10^{-4}s$

Answer b:

$3.208x10^4g$

Answer c:

$1.53x10^2g$

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel units you don't want and produce units you do want.

Further Reading/Supplemental Links

- Tutorial: Vision Learning: Unit Conversions & Dimensional Analysis http://visionlearning.com/library/mo...?mid=144&1=&c3=

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.6: Problem Solving and Unit Conversions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.7: Solving Multistep Conversion Problems

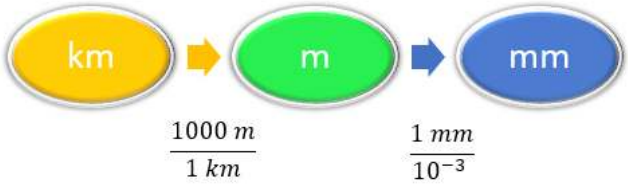
Template:HideTOC

Multiple Conversions

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

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

Concept Map



Calculation

$$54.7\ \cancel{\text{km}} \times \frac{1{,}000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1\ \text{mm}}{\cancel{10^{-3}\text{m}}} = 54{,}700{,}000\ \text{mm} \tag{4.7.1}$$

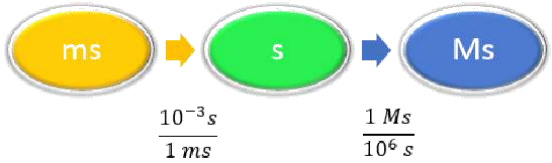
$$= 5.47 \times 10^{7}\ \text{mm} \tag{4.7.2}$$

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

Example 4.7.1: Unit Conversion

Convert 58.2 ms to megaseconds in one multistep calculation.

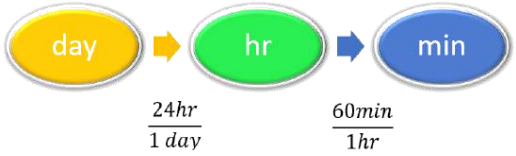
Solution

Steps for Problem Solving	Unit Conversion
Identify the "given"information and what the problem is asking you to "find."	Given: 58.2 ms Find: Ms
List other known quantities	$1ms = 10^{-3}s$ $1Ms = 10^{6}s$
Prepare a concept map	
Calculate	$58.2\ \cancel{ms} \times \frac{10^{-3}\ \cancel{s}}{1\ \cancel{ms}} \times \frac{1\ \text{Ms}}{1,000,000\ \cancel{s}} = 0.0000000582\ \text{Ms}$ $= 5.82 \times 10^{-8}\ \text{Ms}$ Neither conversion factor affects the number of significant figures in the final answer.

Example 4.7.2: Unit Conversion

How many seconds are in a day?

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given"information and what the problem is asking you to "find."	Given: 1 day Find: s
List other known quantities	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds
Prepare a concept map	
Calculate	$1\ \text{d} \times \frac{24\ \text{hr}}{1\ \text{d}} \times \frac{60\ \text{min}}{1\ \text{hr}} \times \frac{60\ \text{s}}{1\ \text{min}} = 86,400\ \text{s}$

Exercise 4.7.1

Perform each conversion in one multistep calculation.

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

https://chem.libretexts.org/@go/page/105263

Answer a:

$4.3007x10^{-14}kg$

Answer b:

83.75 ft

Answer c:

19 km

Career Focus: Pharmacist

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



A woman consulting with a pharmacist. This image was released by the National Cancer Institute, an agency part of the National Institutes of Health. (Public Domain; Rhoda Baer (Photographer) via NIH).

Summary

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

4.7: Solving Multistep Conversion Problems is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.8: Units Raised to a Power

Template:HideTOC

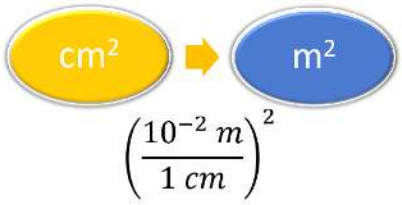
Learning Objectives

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

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

$$1\,cm^2 = (10^{-2}\,m)^2 = 10^{-4}\,m^2 \tag{4.8.1}$$

CONCEPT MAP



CALCULATION

$$1500\ \cancel{cm}^2 \times \left(\frac{10^{-2}\ \mathrm{m}}{1\ \cancel{cm}}\right)^2 = 0.15\ \mathrm{m}^2 \tag{4.8.2}$$

or

$$1500\ \cancel{cm}^2 \times \left(\frac{1\ \mathrm{m}}{100\ \cancel{cm}}\right)^2 = 0.15\ \mathrm{m}^2 \tag{4.8.3}$$

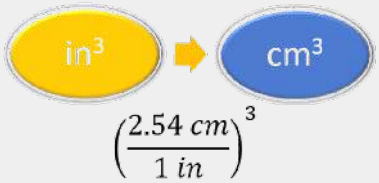
or

$$1500\ \cancel{cm}^2 \times \frac{1\ \mathrm{m}^2}{10,000\ \cancel{cm}^2} = 0.15\ \mathrm{m}^2 \tag{4.8.4}$$

Example 4.8.1: Volume of a sphere

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

Solution

Steps for Problem Solving	What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm^3)?
Identify the "given" information and what the problem is asking you to "find."	Given: radius = 4.30 in Find: cm^3 (volume)
Determine other known quantities	Volume of a sphere: V = $\frac{4}{3} \times \pi \times r^3$ = $\frac{4}{3} \times 3.1416 \times (4.30 in)^3$ = $333.04 in^3$
Prepare a concept map	
Calculate	$333.04\ in^3 \left(\frac{2.54 cm}{1 in}\right)^3 = 5.46 \times 10^3 cm^3$
Think about your result	A centimeter is a smaller unit than an inch, so the answer in cubic centimeter is larger than the given value in cubic inch.

Exercise 4.8.1

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

Answer:

495 km^2

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

4.8: Units Raised to a Power is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.10: Energy and Heat Capacity Calculations

Template:HideTOC

Learning Objectives

- To relate heat transfer to temperature change.

Heat is a familiar manifestation of transferring energy. When we touch a hot object, energy flows from the hot object into our fingers, and we perceive that incoming energy as the object being “hot.” Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as “cold.” In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T \tag{4.10.1}$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Every substance has a characteristic specific heat, which is reported in units of cal/g•°C or cal/g•K, depending on the units used to express ΔT. The specific heat of a substance is the amount of energy that must be transferred to or from 1 g of that substance to change its temperature by 1°. Table 4.10.1 lists the specific heats for various materials.

Table 4.10.1: *Specific Heats of Some Common Substances*

Substance	Specific Heat (J/g^oC)
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

The *direction* of heat flow is not shown in heat = $mc\Delta T$. If energy goes into an object, the total energy of the object increases, and the values of heat ΔT are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and ΔT are negative.

Example 4.10.1

A 15.0 g piece of cadmium metal absorbs 134 J of heat while rising from 24.0°C to 62.7°C. Calculate the specific heat of cadmium.

Solution

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

Known

- Heat $= q = 134\ \text{J}$
- Mass $= m = 15.0\ \text{g}$
- $\Delta T = 62.7^\circ\text{C} - 24.0^\circ\text{C} = 38.7^\circ\text{C}$

Unknown

- c_p of cadmium $=?\ \text{J/g}^\circ\text{C}$

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134\ \text{J}}{15.0\ \text{g} \times 38.7^\circ\text{C}} = 0.231\ \text{J/g}^\circ\text{C} \tag{4.10.2}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known (Table 4.10.1), they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g of water at 23.52°C was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation.

$$\Delta T = \frac{q}{c_p \times m} = \frac{813\ \text{J}}{4.18\ \text{J/g}^\circ\text{C} \times 60.0\ \text{g}} = 3.24^\circ\text{C} \tag{4.10.3}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^\circ\text{C} - 3.24^\circ\text{C} = 20.28^\circ\text{C} \tag{4.10.4}$$

Example 4.10.2

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Solution

We can use heat = $mc\Delta T$ to determine the amount of heat, but first we need to determine ΔT. Because the final temperature of the iron is 73.3°C and the initial temperature is 25.0°C, ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 73.3^\circ\text{C} - 25.0^\circ\text{C} = 48.3^\circ\text{C}$$

The mass is given as 150.0 g, and Table 7.3 gives the specific heat of iron as 0.108 cal/g•°C. Substitute the known values into heat = $mc\Delta T$ and solve for amount of heat:

$$\text{heat} = (150.0\ \text{g})\left(0.108\ \frac{\text{cal}}{\text{g} \cdot {}^\circ\text{C}}\right)(48.3^\circ\text{C}) = 782\ \text{cal} \tag{4.10.5}$$

Note how the gram and °C units cancel algebraically, leaving only the calorie unit, which is a unit of heat. Because the temperature of the iron increases, energy (as heat) must be flowing *into* the metal.

Exercise 4.10.1

What quantity of heat is transferred when a 295.5 g block of aluminum metal is cooled from 128.0°C to 22.5°C? What is the direction of heat flow?

Answer

Heat leaves the aluminim block.

Example 4.10.2

A 10.3 g sample of a reddish-brown metal gave off 71.7 cal of heat as its temperature decreased from 97.5°C to 22.0°C. What is the specific heat of the metal? Can you identify the metal from the data in Table **4.10.1**?

Solution

The question gives us the heat, the final and initial temperatures, and the mass of the sample. The value of ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 22.0°\text{C} - 97.5°\text{C} = -75.5°\text{C}$$

If the sample gives off 71.7 cal, it loses energy (as heat), so the value of heat is written as a negative number, −71.7 cal. Substitute the known values into heat = $mc\Delta T$ and solve for c:

$$-71.7 \text{ cal} = (10.3 \text{ g})(c)(-75.5°\text{C})$$

$$c = \frac{-71.7 \text{ cal}}{(10.3 \text{ g})(-75.5°\text{C})}$$

$$c = 0.0923 \text{ cal/g•°C}$$

This value for specific heat is very close to that given for copper in Table 7.3.

Exercise 4.10.2

A 10.7 g crystal of sodium chloride (NaCl) had an initial temperature of 37.0°C. What is the final temperature of the crystal if 147 cal of heat were supplied to it?

Answer a:

The

Summary

Specific heat calculations are illustrated.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.10: Energy and Heat Capacity Calculations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.9: Density

Template:HideTOC

Learning Objectives

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

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

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \tag{4.9.1}$$

or just

$$\rho = \frac{m}{V} \tag{4.9.2}$$

Based on this equation, it's clear that density can, and does, vary from element to element and substance to substance due to differences in the relationship of mass and volume. Pure water, for example, has a density of 0.998 g/cm^3 at 25° C. The average densities of some common substances are in Table **4.9.1**. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table 4.9.1: Densities of Common Substances

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

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

Example 4.9.1: Ethyl Alcohol

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

Solution

$$\rho = \frac{23.71002\,g}{30.2\,\cancel{mL}} \underbrace{\left(\frac{1000\,\cancel{mL}}{1\,L}\right)}_{\text{Conversion Factor}} = 0.785\,g/mL \tag{4.9.3}$$

Exercise 4.9.1

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

Answer a

$0.274\ kg/L$

Answer b

$1.049\ kg/L$

Density as a Conversion Factor

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

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

$$\frac{13.6\ g}{1\ mL} = 1 \tag{4.9.4}$$

and

$$\frac{1\ mL}{13.6\ g} = 1 \tag{4.9.5}$$

Which one do we use? It depends, as usual, on the units we need to cancel and introduce. For example, suppose we want to know the mass of 16 mL of mercury. We would use the conversion factor that has milliliters on the bottom (so that the milliliter unit cancels) and grams on top so that our final answer has a unit of mass:

$$2.0\ \cancel{mL} \times \frac{13.6\ g}{1\ \cancel{mL}} = 27.2\ g = 27\ g \tag{4.9.6}$$

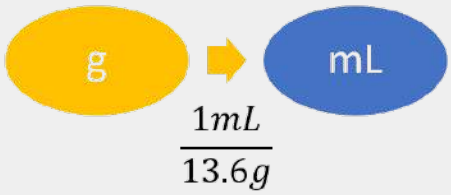
In the last step, we limit our final answer to two significant figures because the volume quantity has only two significant figures; the 1 in the volume unit is considered an exact number, so it does not affect the number of significant figures. The other conversion factor would be useful if we were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

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

Example 4.9.2: Mercury Thermometer Steps for Problem Solving

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

Solution

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

Exercise 4.9.2

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

Answer

77 L

Summary

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

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.9: Density is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

2: Matter

Template:HideTOC

- 2.4: Density and its Applications
- 2.1: In Your Room
- 2.2: What is Matter?
- 2.3: Classifying Matter According to Its State: Solid, Liquid, and Gas
- 2.5: Classifying Matter According to Its Composition
- 2.6: Differences in Matter: Physical and Chemical Properties
- 2.7: Changes in Matter: Physical and Chemical Changes

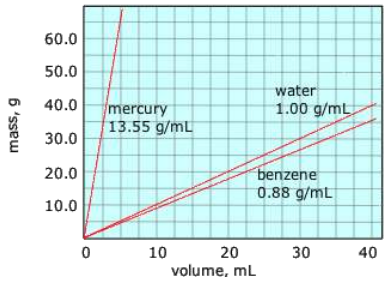
2: Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.4: Density and its Applications

The density of an object is one of its most important and easily-measured physical properties. Densities are widely used to identify pure substances and to characterize and estimate the composition of many kinds of mixtures. The purpose of this lesson is to show how densities are defined, measured, and utilized, and to make sure you understand the closely-related concepts of buoyancy and specific gravity, and the roles they play in our lives and the environment.

What is Density?

Most of us have long understood that "oil is lighter than water", or that iron is "heavier" than sugar. But in making such statements, we are implicitly comparing *equal volumes* of these substances: after all, we know that a cup of sugar will weigh more than a single ordinary steel nail. Mass and volume are measures of the *quantity* of a substance, and as such are defined as *extensive* properties of matter. The ratio of two extensive properties is always an *intensive* property — one that characterizes a particular kind of matter, independently of its size or mass. It is this ratio, (mass ÷ volume), that we are concerned with in this Module.



These plots show how the masses of three liquids vary with their volumes. Notice that

- the plots all have the same origin of (0,0): if the mass is zero, so is the volume;
- the plots are all straight lines, which signify direct proportionality.

The only difference between these plots is their slopes. Denoting mass and volume by m and V respectively, we can write the equation of each line as $m = \rho V$, where the slope ρ (Greek lower-case *rho*) is the proportionality constant that relates mass to volume. This quantity ρ is known as the *density*, which is usually defined as the mass per unit volume:

$$\rho = \frac{m}{V}. \tag{2.4.1}$$

The volume units *milliliter* (mL) and *cubic centimeter* (cm^3) are identical and are commonly used interchangeably.

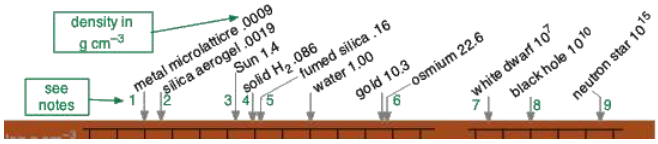
The general meaning of density is the amount of anything per unit volume. What we conventionally call the "density" is more precisely known as the "mass density".

Density can be expressed in any combination of mass and volume units; the most commonly seen units are grams per mL ($g\ mL^{-1}$, $g\ cm^{-3}$), or kilograms per liter.

$1\ kg\ m^{-3} = 10^{-3}\ g\ L^{-1} = 62.4\ lb\ ft^{-3}$

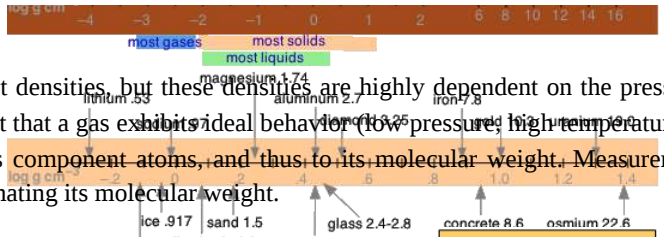
Densities of Common Substances

The range of densities encountered in the world spans a remarkably wide range, from essentially zero in outer space to the unimaginably huge values found in stellar bodies. These very high densities represent the ultimate limits of how much mass can be packed into a given volume. The following chart will give you some feeling for the values of density found in nature generally (top), in common solids (middle), and in gases and liquids (bottom). Please note that in order to depict reasonably wide ranges of values in limited space, **the density scales are *logarithmic***; thus zero on these scales corresponds to the density of water (10^0 = 1 g cm^{-3}). Densities listed for ordinary substances (including gases) are mostly those at around 20° C.



https://chem.libretexts.org/@go/page/3552

Solids, liquids and gases



In general, **gases** have the lowest densities, but these densities are highly dependent on the pressure and temperature which must always be specified. To the extent that a gas exhibits ideal behavior (low pressure, high temperature), the density of a gas is directly proportional to the masses of its component atoms, and thus to its molecular weight. Measurement of the density of a gas is a simple experimental way of estimating its molecular weight.

Densities of gases in g/L at 1 atm, 0°C	
chlorine	3.21
carbon dioxide	1.98
air (dry)	1.29
methane	0.72
hydrogen	0.09

Densities of liquids in g/mL at 1 atm, 25°C	
mercury	13.55
sulfuric acid	1.83
water	0.997
olive oil	0.92
ethanol	0.79

Densities of solids in g/mL at 1 atm, 25°C	
osmium	22.6
gold	19.3
iron	7.9
aluminum	2.70
sodium chloride	2.16
sugar	1.59
h.d. polyurethane	1.05
wood (pine)	.3 - .5
lithium	0.53

Liquids encompass an intermediate range of densities. Mercury, being a liquid metal, is something of an outlier. Liquid densities are largely independent of pressure, but they are somewhat temperature-sensitive.

The density range of **solids** is quite wide. Metals, whose atoms pack together quite compactly, have the highest densities, although that of lithium, the highest metallic element, is quite low. Composite materials such as wood and high-density polyurethane foam contain void spaces which reduce the average density.

How the temperature affects density

All substances tend to expand as they are heated, causing the same mass to occupy a greater volume, and thus lowering the density. For most solids, this expansion is relatively small, but it is far from negligible; for liquids, it is greater. The volumes of gases, as you may already know, are highly temperature-sensitive, and so, of course, are their densities.

What is the cause of thermal expansion? As molecules acquire thermal energy, they move about more vigorously. In condensed phases (liquids and solids), this motion has the character of an irregular kind of bumping or jostling that causes the average distances between the molecules to increase, thus leading to increased volume and smaller density.

Applications of density

Lava Lamps



Known more generically as "fluid motion lamps", these devices became popular in the 1970's and provide a nice, if somewhat mesmerizing illustration of density and buoyancy in action as the blobs of oozing goo move up and down in ever-changing shapes. These lamps consist of a container of water in which is placed a colored organic oily liquid that does not mix with water, thus constituting a second phase. The composition of the oil phase is such that its density is slightly greater than that of water at room temperature, so it normally resides at the bottom of the container. When the lamp is turned on, a heat source (usually an incandescent light bulb) concealed in the base of the container heats the oil phase. This reduces its density to a value below that of the water, causing blobs of oil to rise to the top of the container. Being now far removed from the heat source, the blobs cool down and sink back to the bottom, where they repeat the cycle.

Contributors and Attributions

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

This page titled 2.4: Density and its Applications is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Stephen Lower via source content that was edited to the style and standards of the LibreTexts platform.

2.1: In Your Room

Template:HideTOC

Matter is any substance that has mass and takes up space; this includes atoms and anything made up of these, but not other energy phenomena or waves such as light or sound. While this simple definition is easily applied, the way people view matter is often broken down into two characteristic length scales: the macroscopic and the microscopic.



Figure **2.1.1**: *A typical American university and college dormitory room in 2002. (CC BY-SA 3.0; Raul654).*

The macroscopic scale is the length scale on which objects or phenomena are large enough to be visible almost practically with the naked eye, without magnifying optical instruments. Everything that one can see, touch and handle in the dorm room of Figure **2.1.1** is in the macroscopic scale and to describe each of these objects only few macroscopic properties are required. However, each of these items can be decomposed into smaller microscopic scale properties.

The microscopic scale is the scale of objects and events smaller than those that can easily be seen by the naked eye, requiring a lens or microscope to see them clearly. All the everyday objects that we can bump into, touch or squeeze are ultimately composed of atoms. This ordinary atomic matter is in turn made up of interacting subatomic particles—usually a nucleus of protons and neutrons, and a cloud of orbiting electrons. Because of this, a large number of variables are needed to describe such a system which complicates the characterization.

Matter vs. MAss

Matter should not be confused with mass, as the two are not the same in modern physics. Matter is itself a physical substance of which systems may be composed, while mass is not a substance but rather a quantitative property of matter and other substances or systems.

Contributors and Attributions

- Wikipedia

2.1: In Your Room is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.2: What is Matter?

Template:HideTOC

Learning Objectives

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

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



Figure **2.2.1**: Everything from an ant, to a truck, to the earth and even the entire galaxy is composed of matter. Images used with permission from Wikipedia (CC_SA-BY-3.0; credit High Contrast)

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

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

Summary

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

Contributors and Attributions

- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

2.2: What is Matter? is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

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

Template:HideTOC

Learning Objectives

- To describe the solid, liquid and gas phases.

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



Figure **2.3.1**: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure **2.3.2** shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

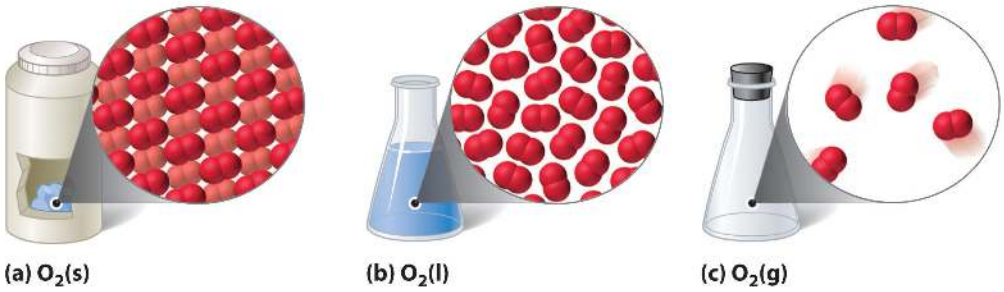
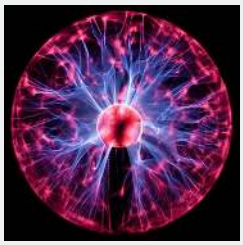


Figure **2.3.2**: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

Plasma: A Fourth State of Matter

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



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

Solids

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

Solids are defined by the following characteristics:

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

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

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure **2.3.3**. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, "without form") solids. Glass is one example of an amorphous solid.

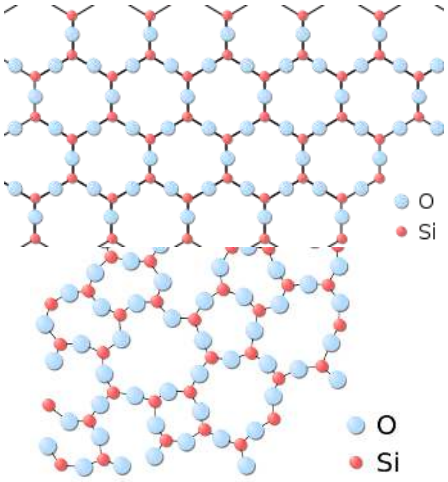


Figure **2.3.3**: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain)

Liquids

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

Liquids have the following characteristics:

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

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





Video* 2.3.1: Mercury boiling to become a gas.*

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

Gases

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

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

Table 2.3.1: Characteristics of the Three States of Matter

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

Example 2.3.1

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

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

Solution

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

Exercise 2.3.1

What state or states of matter does each statement describe?

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

Answer a:

solid

Answer b:

gas

Answer c:

solid

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

2.3: Classifying Matter According to Its State: Solid, Liquid, and Gas is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.5: Classifying Matter According to Its Composition

Template:HideTOC

Learning Objectives

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

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure **2.5.1**). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning it's the same everywhere) and properties that are constant throughout the sample (meaning there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances.A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

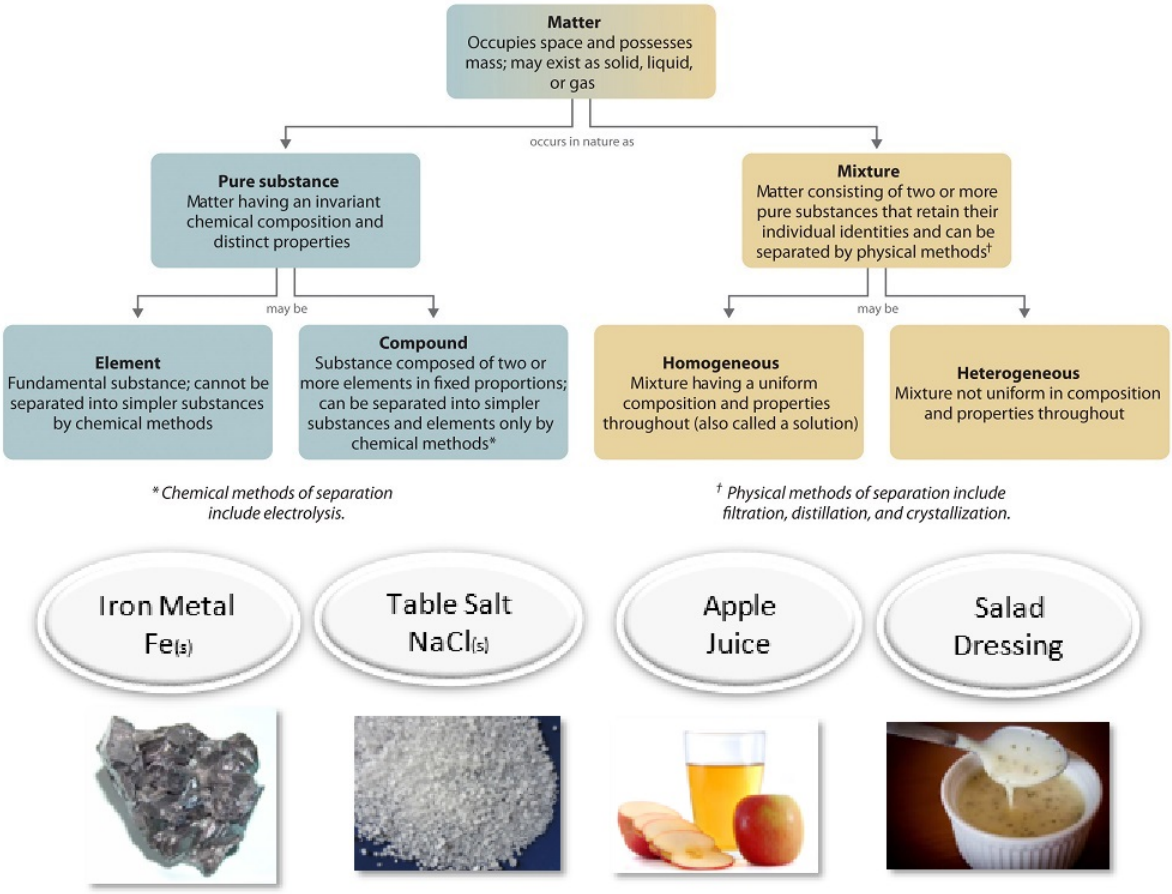


Figure **2.5.1**: *Relationships between the Types of Matter and the Methods Used to Separate Mixtures*

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount

into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

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

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

Phase

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

Example 2.5.1

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

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

Solution:

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

Exercise 2.5.1

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

a. white wine

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

Answer a:
homogeneous mixture (solution)

Answer b:
element

Answer c:
heterogeneous mixture

Answer d:
compound

Example 2.5.2

How would a chemist categorize each example of matter?

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

Solution

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

Exercise 2.5.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a:
a homogeneous mixture (solution), assume it'd filtered coffee

Answer b:
element

Answer c:
heterogeneous mixture.

Summary

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

Vocabulary

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

Contributors and Attributions

-
- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

2.5: Classifying Matter According to Its Composition is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.6: Differences in Matter: Physical and Chemical Properties

Learning Objectives

To separate physical from chemical properties.

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

Physical Property

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

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C. The average densities of some common substances are in Table **2.6.1**. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table **2.6.1**: Densities of Common Substances

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

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



Figure **2.6.1**: Pencil (left) and Diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

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

Chemical Properties

Chemical properties of matter describes its "potential" to undergo some chemical change or reaction by virtue of its composition. What elements, electrons, and bonding are present to give the potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually you should be able to look at the formula of a compound and state some chemical property. At this time this is very difficult to do and you are not expected to be able to do it. For example hydrogen has the potential to ignite and explode given the right conditions - this is a chemical property. Metals in general have they chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas - this is a chemical property.



Figure **2.6.2**: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust (Figure **2.6.2**). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 2.6.2: Contrasting Physical and Chemistry Properties

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

Example 2.6.1

Which of the following is a chemical property of iron?

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

Solution

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

Exercise 2.6.1*A*

Which of the following is a physical property of matter?

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

Answer

c

Exercise 2.6.1*B*

Which of the following is a chemical property?

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

Answer

a

Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

2.6: Differences in Matter: Physical and Chemical Properties is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.7: Changes in Matter: Physical and Chemical Changes

Template:HideTOC

Learning Objectives

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

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

Physical Change

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

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

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



Figure **2.7.1**: Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical as the the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass.

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

Chemical Change

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

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



Figure 2.7.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

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

Example 2.7.1

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

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

Solution

a. Physical: boiling and melting are physical changes. When water boils no bonds are broken or formed. The change could be written: $H_2O\,(l) \rightarrow H_2O\,(g)$

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

Exercise 2.7.1

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

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

Answer a:
physical change

Answer b:
chemical change

Answer c:
physical change

Answer d:
chemical change

Separating Mixtures Through Physical Changes

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

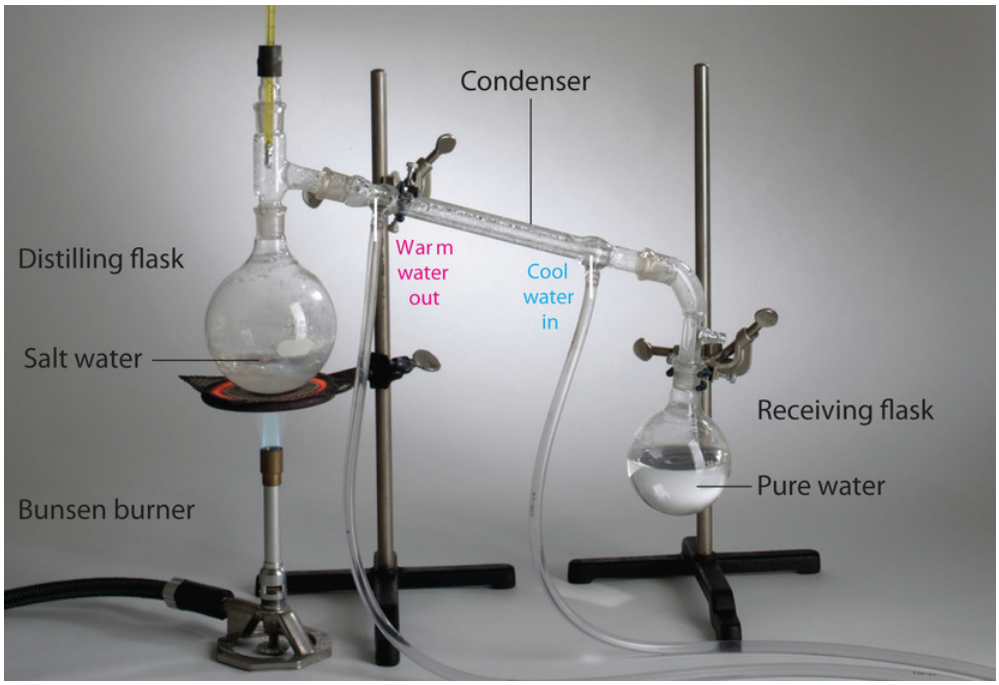


Figure **2.7.3***: The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.*

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

Another example for using physical properties to separate mixtures is filtration (Figure **2.7.4**). Filtration is any of various mechanical, physical or biological operations that separate solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.

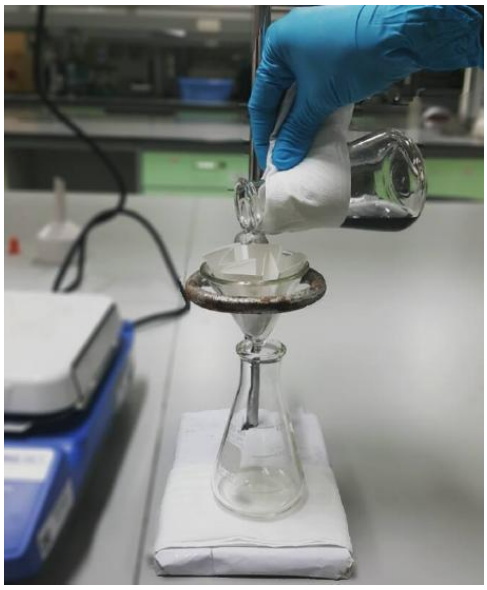


Figure **2.7.4***: Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).*

Summary

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

Contributors and Attributions

-
- Boundless (www.boundless.com)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

2.7: Changes in Matter: Physical and Chemical Changes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

5: Atoms and Elements

Template:HideTOC

- 5.1: Cutting Alumimun until you get Atoms
- 5.2: Indivisible: The Atomic Theory
- 5.3: The Nuclear Atom
- 5.4: The Properties of Protons, Neutrons, and Electrons
- 5.5: Elements: Defined by Their Number of Protons
- 5.6: Looking for Patterns: The Periodic Table
- 5.7: Ions - Losing and Gaining Electrons
- 5.8: Isotopes - When the Number of Neutrons Varies
- 5.9: Atomic Mass: The Average Mass of an Element’s Atoms

5: Atoms and Elements is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.1: Cutting Alumimun until you get Atoms

Template:HideTOC

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.)

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

5.1: Cutting Alumimun until you get Atoms is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.2: Indivisible: The Atomic Theory

Template:HideTOC

Learning Objectives

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

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

Democritus and the Greek Philosophers

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



Figure **5.2.1**: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right): British physicist and chemist John Dalton (1766-1844). Unlike the greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they're known today) was fully appreciated.

Greeks: "All thought and No Action"

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

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

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure **5.2.1**), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely

small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).



Figure 5.2.1 John Dalton was an English scientist who enunciated the modern atomic theory.

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

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

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

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

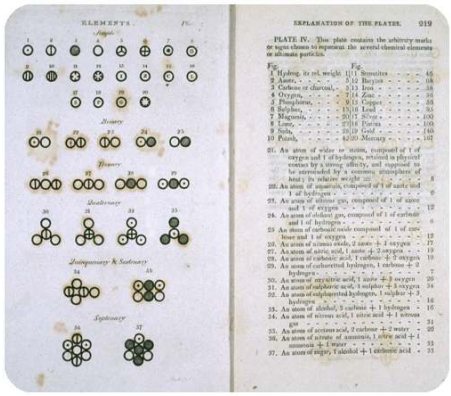
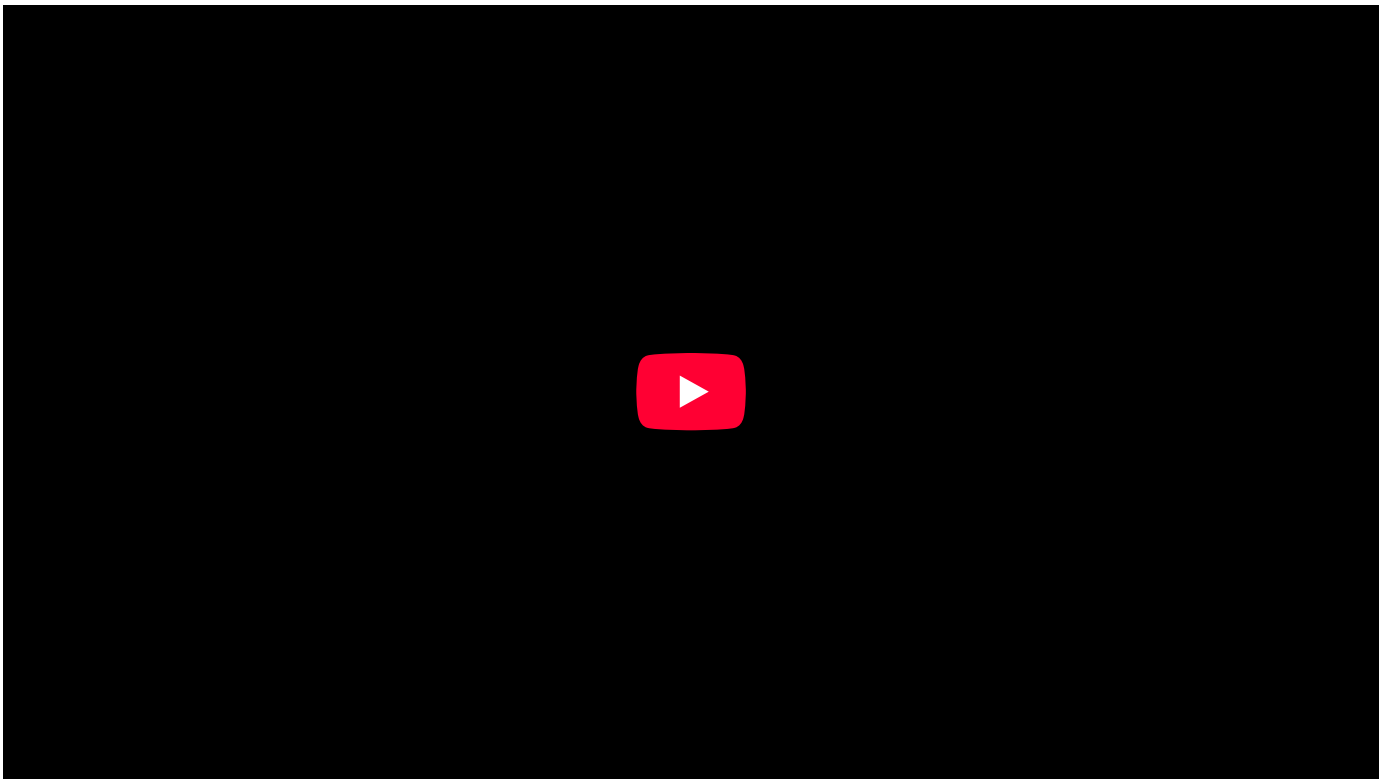


Figure 5.2.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images or as IBM research demonstrate in Video 5.2.1, control of individual atoms can be use used create animations.



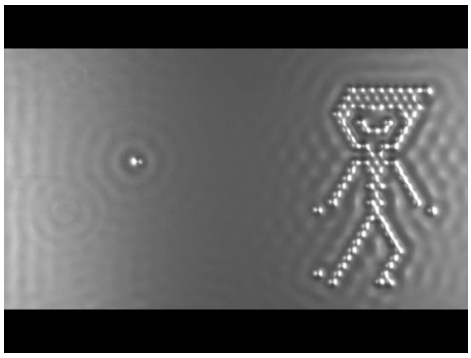


***Figure* 5.2.23**: A Boy And His Atom: The World's Smallest Movie

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

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.2: Indivisible: The Atomic Theory is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.3: The Nuclear Atom

Template:HideTOC

Learning Objectives

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

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



Figure **5.3.1**: *J.J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom*

Electrons and Plums

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

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

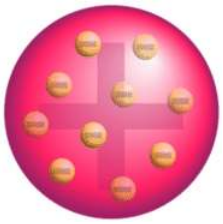


Figure **5.3.2**: *The "plum pudding" model.*

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

Atoms and Gold

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

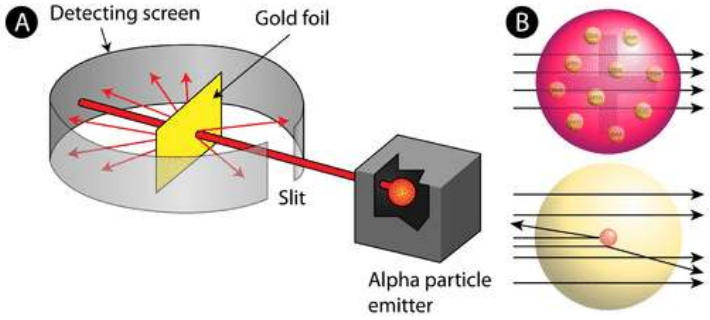


Figure **5.3.3**: *(A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top) all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).*

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed undeflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had

prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

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

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

Atomic Nucleus

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

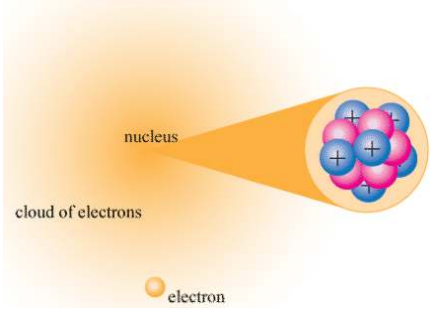


Figure **5.3.4**: The nuclear atom

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

Holding It All Together

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

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.

- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

Contributors and Attributions

-

5.3: The Nuclear Atom is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.4: The Properties of Protons, Neutrons, and Electrons

Template:HideTOC

Learning Objectives

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

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

Electrons

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

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



Figure 5.4.1: Electrons are much smaller than protons or neutrons. If an electrons were the mass of a penny, a proton or a neutrons would have the mass of a large bowling ball!

Protons

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

Neutrons

Atoms of all elements - except for most atoms of hydrogen - have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge - they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (**amu**). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{17} meters.

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

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

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

Table 5.4.1: Properties of Subatomic Particles

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

Table **5.4.1** gives the properties and locations of electrons, protons, and neutrons. The second column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Atomic mass units (**amu**) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly **1.0** in this unit system.

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

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they're neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.4: The Properties of Protons, Neutrons, and Electrons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.5: Elements: Defined by Their Number of Protons

Template:HideTOC

Learning Objectives

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

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



Figure 5.5.1: It is difficult to find qualities that are difference from each element and distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp)

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 5.5.1). If an atom has only one proton, we know it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

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

Table 5.5.1: Atoms of the First Six Elements

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

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of

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

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table 5.5.1). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table 5.5.1).

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

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

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons}) \tag{5.5.1}$$

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

Example 4.5.1

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

Solution

(number of protons) = 2 (Remember that an atom of helium always has 2 protons.)

(number of neutrons) = 2

mass number = (number of protons) + (number of neutrons)

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

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are **O** for oxygen, **Zn** for zinc, and **Fe** for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table 5.5.2.

Table 5.5.2: *Symbols and Latin Names for Elements*

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

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.5: Elements: Defined by Their Number of Protons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.6: Looking for Patterns: The Periodic Table

Template:HideTOC

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 **5.6.1**. 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.

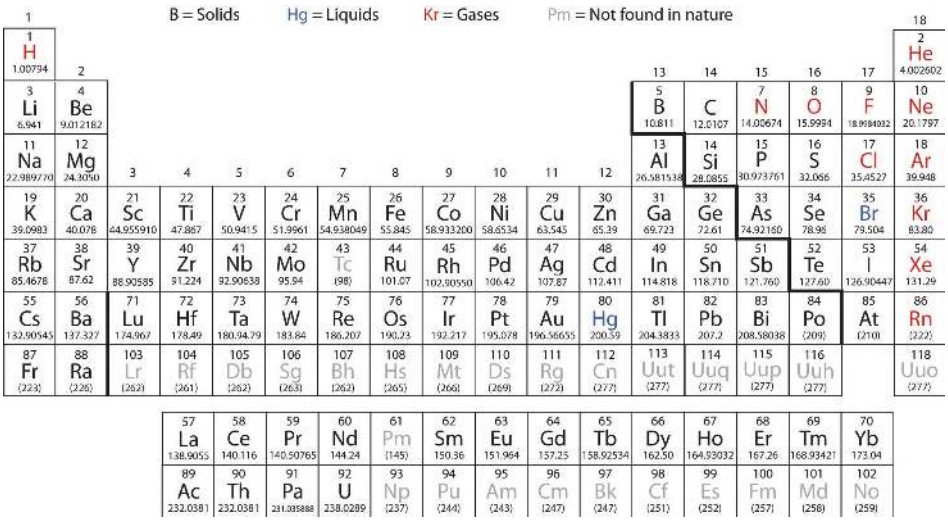


Figure **5.6.1**: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number.

Features of the Periodic Table

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

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

To Your Health: Radon

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

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building's inhabitants. Fortunately, specialized ventilation minimizes the

amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

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

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

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure **5.6.2**. 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 **5.6.2**, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

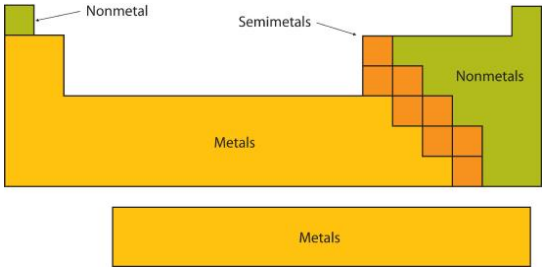


Figure **5.6.2**: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table.

Example 5.6.1

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure **5.6.1**, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

Exercise 5.6.1

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

Answer:

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure **5.6.3**. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

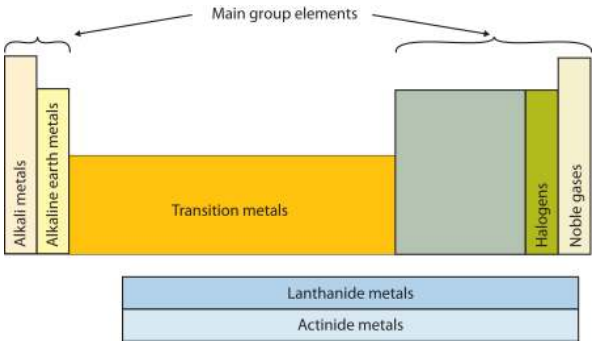


Figure **5.6.3**: Special Names for Sections of the Periodic Table.

Descriptive Names

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

Group 1: The Alkali Metals

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

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

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.



Example 5.6.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

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

Exercise 5.6.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a:
Fluorine is a halogen (Group 17)

Answer b:
Calcium is a alkaline earth metal (Group 2)

Answer c:
Krypton is a noble gas (Group 18)

Example 5.6.3: Classification of Elements

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

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

Solution

a. Lithium is a metal.

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

Exercise 5.6.3: Classification of Elements

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

a. F
b. U
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

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

Answer c:

Copper is a metal (and a transition metal too)

Summary

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

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

References

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

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.6: Looking for Patterns: The Periodic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.7: Ions - Losing and Gaining Electrons

Template:HideTOC

Learning Objectives

- Define the two types of ions.

Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge as a result because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called cations. Most metals become cations when they make ionic compounds.

Cations

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

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-} \tag{5.7.1}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure **5.7.1** is a graphical depiction of this process.

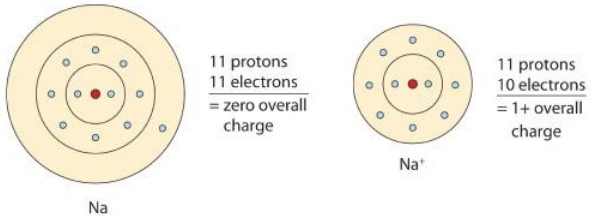


Figure **5.7.1**: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Anions

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

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

$$\mathrm{e^-} + \mathrm{Cl} \longrightarrow \mathrm{Cl^-} \tag{5.7.2}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure **5.7.2** is a graphical depiction of this process.

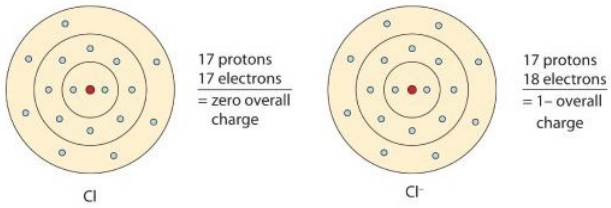


Figure **5.7.2**: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge.

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

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge. Figure **5.7.3** shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a multiply charged ion. The barium cation is written Ba^{2+}, not Ba^{+2}.

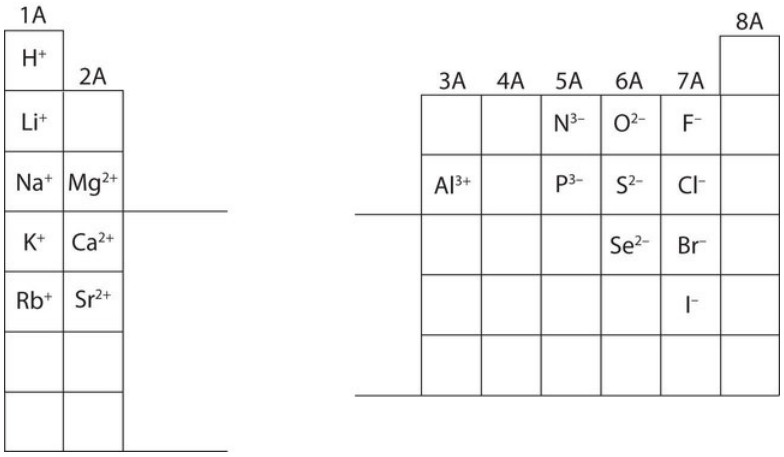


Figure **5.7.3**: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.7: Ions - Losing and Gaining Electrons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.8: Isotopes - When the Number of Neutrons Varies

Template:HideTOC

Learning Objectives

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

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

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 5.8.1. Most hydrogen atoms have just one proton and one electron and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

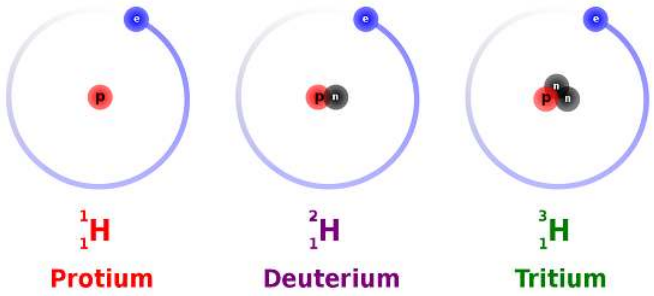


Figure 5.8.1: *The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh).*

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

Example 5.8.1: Lithium Isotopes

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

Solution

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

a.

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

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

b.

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

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

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

Stability of Isotopes

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

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

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

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

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



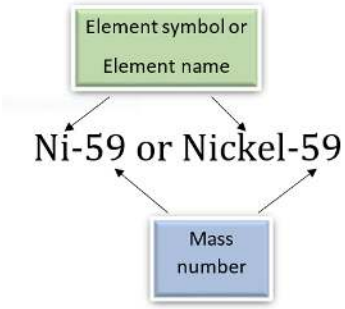
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\text{Ni} \tag{5.8.1}$$

$$^{238}_{92}\text{U} \tag{5.8.2}$$

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

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



Example 5.8.2: Krypton-40

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

Solution

$$\text{atomic number} = (\text{number of protons}) = 19 \tag{5.8.3}$$

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

$$\text{number of electrons} = 19 \tag{5.8.4}$$

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

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

$$\text{number of neutrons} = 40 - 19 = 21. \tag{5.8.5}$$

Example 5.8.3: Zinc-65

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

Solution

$$\text{number of protons} = 30 \tag{5.8.6}$$

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

$$\text{number of electrons} = 30 \tag{5.8.7}$$

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

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

$$\text{number of neutrons} = 65 - 30 = 35 \tag{5.8.8}$$

Exercise 5.8.3

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

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

Answer a:

27 protons, 27 electrons, 33 neutrons

Answer b:

11 protons, 11 electrons, 13 neutrons

Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.8: Isotopes - When the Number of Neutrons Varies is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

5.9: Atomic Mass: The Average Mass of an Element's Atoms

Template:HideTOC

Learning Objectives

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

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

Calculating Atomic Mass

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

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots \tag{5.9.1}$$

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

Example 5.9.1: Boron Isotopes

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

Solution

Boron has two isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots \tag{5.9.2}$$

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

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)\,(10) + (0.80)\,(11) \tag{5.9.3}$$

$$\text{Atomic mass} = 10.8\ \text{amu} \tag{5.9.4}$$

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

Example 5.9.2: Neon Isotopes

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

Solution

Neon has three isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots \tag{5.9.5}$$

- Isotope 1: $\%_1 = 0.9092$, $\text{mass}_1 = 19.99$
- Isotope 2: $\%_2 = 0.003$, $\text{mass}_2 = 20.99$
- Isotope 3: $\%_3 = 0.0885$, $\text{mass}_3 = 21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99) \tag{5.9.6}$$

$$\text{Atomic mass} = 20.17\ \text{amu} \tag{5.9.7}$$

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

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol **B**) is 10.8, which is what we calculated in Example **5.9.1**, and the atomic mass of neon (symbol **Ne**) is 20.8, which is what we calculated in Example **5.9.2**. Take time to notice that not all periodic tables have th atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

Exercise 5.9.1

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

Answer:

35.45 amu

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

5.9: Atomic Mass: The Average Mass of an Element's Atoms is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

6: Molecules and Compounds

Template:HideTOC

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multiatom combinations are called moleculesThe smallest part of a substance that has the physical and chemical properties of that substance.. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

6.1: Sugar and Salt
6.2: Compounds Display Constant Composition
6.3: Chemical Formulas: How to Represent Compounds
6.4: A Molecular View of Elements and Compounds
6.5: Writing Formulas for Ionic Compounds
6.6: Nomenclature: Naming Compounds
6.7: Naming Ionic Compounds
6.8: Naming Molecular Compounds
6.9: Naming Acids
6.10: Nomenclature Summary
6.11: Formula Mass: The Mass of a Molecule or Formula Unit

6: Molecules and Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.1: Sugar and Salt

Sodium chloride also known as table salt, is an ionic compound with the chemical formula $\mathbf{NaCl}$, representing a 1:1 ratio of sodium and chloride ions. It is commonly used as a condiment and food preservative. Salt can be created by adding two very reactive elements together: sodium ($\mathbf{Na\,(s)}$ metal and chlorine ($\mathbf{Cl_2\,(g)}$ gas.

$$\mathbf{2\,Na\,(s) + Cl_2\,(g) \rightarrow 2\,NaCl\,(s)} \tag{6.1.1}$$

The element sodium (Figure **6.1.1***a*) is a very reactive metal; given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine (Figure **6.1.1***b*) is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride (Figure **6.1.1***c*), known simply as salt.

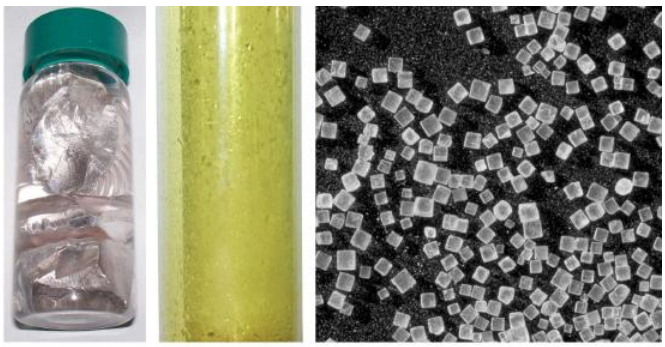


Figure **6.1.1**: Sodium + Chlorine = Sodium Chloride (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride—salt—which is necessary for our survival. Source: Photo on the left courtesy of reenhorn1 and photo in the center courtesy of Benjah-bmm27. Photo on the right © Thinkstock.

Salt is necessary for life. $\mathbf{Na^+}$ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. $\mathbf{Cl^-}$ ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavoring known, and one of the few rocks we eat. Clearly when the elemental sodium and chlorine combine (Equation 6.1.1), the resulting salt product has radically different properties (both physical and chemical). This reaction is spectacular to observed (Video **6.1.1**).



Video* 6.1.1*: *Making Table Salt using Sodium Metal and Chlorine gas*

Another compound is sugar, which is the generic name for sweet, soluble carbohydrates, many of which are used in food. Sugar has the chemical formulate $\mathbf{C_{12}H_{22}O_{11}}$ and is constructed from different elements than salt: carbon, hydrogen and oxygen. While sugar qualitatively resembles table salt (often confused in the kitchen), they two have distinctly different physical and chemical properties. There are various types of sugar derived from different sources. While sugar is made with carbon, hydrogen and oxygen, it is considerably hard to syntheses from the constituent elements than table salt (Equation 6.1.1). However, the thermal

decomposition is considerably easiler and can be represented as as a dehydration of sucrose to pure carbon and water vapor in Equation 6.1.2 and demonstrated in Video 6.1.2.

$$\mathrm{C_{12}H_{22}O_{11}\,(s) + heat \longrightarrow 12\,C\,(s) + 11\,H_2O\,(g)} \tag{6.1.2}$$



Video* 6.1.2*: *A science experiment in the kitchen shows what happens to sugar molecules when they are heated. The experiment did not disappoint!*

As with salt, sugar has radically different properties (both physical and chemical) than its constituent elements. That is a central feature of chemical reactions as this chapter will discuss.

Contributors and Attributions

- Wikipedia

Template:HideTOC

6.1: Sugar and Salt is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.2: Compounds Display Constant Composition

When building a hose, you start with a blueprint of what the house will look like. The plan states how many windows and what kind, how many doors and what style, how many rooms and what type (bedroom, kitchen, other). The blueprint shows how the different pieces will go together to make the house. As long as the blueprint is followed and exactly the same items are used, the houses will be identical.

Constant Composition stuff

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances - carbon and hydrogen.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** is a change that produces matter with a different composition. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes into carbon and water. Water is still a compound, but one which cannot be broken down into hydrogen and oxygen by heating. Instead, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a white solid which is essential for all living things (see below).

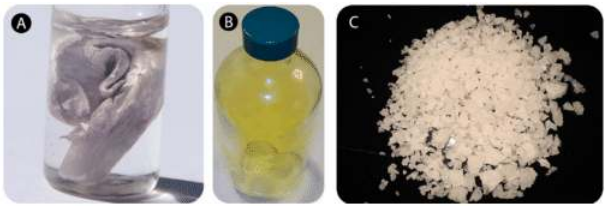


Figure **6.2.1**: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine.

Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- A chemical change is a change that produces matter with a different composition.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.2: Compounds Display Constant Composition is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.3: Chemical Formulas: How to Represent Compounds

Learning Objectives

- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a 2:1 ratio. The chemical formula for water is $\mathbf{H_2O}$. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is $\mathbf{H_2SO_4}$.

Certain groups of atoms are bonded together to form what is called a polyatomic ion that act as a single unit. Polyatomic ions are discussed in more detail in section 5.5. Polyatomic ions are enclosed in parenthesis followed by a subscript if more than one of the same ion exist in a chemical formula. The formula $Ca_3(PO_4)_2$ represents a compound with the following:

3 Ca atoms

+ 2 PO_4^{3-} ions

To count the total number of atoms for formulas with polyatomic ions enclosed in parenthesis, use the subscript as a multiplier for each atom or number of atoms.

$$Ca_3(PO_4)_{\color{red}2}$$

$$3\ Ca + {\color{red}2} \times 1\ P + {\color{red}2} \times 4\ O = 3\ Ca\ atoms + 2\ P\ atoms + 8\ O\ atoms$$

Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

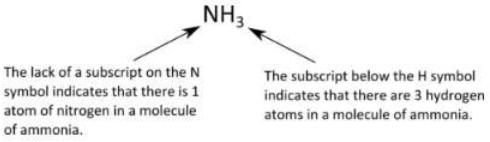


Figure **6.3.1**: *The molecular formula for ammonia.*

Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, so we will not concern ourselves with that point right now. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write $\mathbf{N_3H}$ for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $\mathbf{C_6H_{12}O_6}$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is $\mathbf{CH_2O}$. When chemists analyze an unknown compound, often the first step is to determine its empirical formula.

- molecular formula: $\mathbf{C_6H_{12}O_6}$
- empirical formula: $\mathbf{CH_2O}$

There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of $\mathbf{H_2O}$ or $\mathbf{P_2O_5}$, then the empirical formula is also the molecular formula.

Summary

- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Polyatomic ions in chemical formulas are enclosed in parentheses followed by a subscript if more than one of the same type of polyatomic ion exist.
- Molecular formulas do not indicate how the atoms are arranged in the molecule.
- The empirical formula tells the lowest whole-number ratio of elements in a compound. The empirical formula does not show the actual number of atoms.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.3: Chemical Formulas: How to Represent Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.4: A Molecular View of Elements and Compounds

Learning Objectives

- Classify substances as atomic elements, molecular elements, molecular compounds, or ionic compounds

Atomic Elements

Most elements exist with **individual atoms** as their basic unit. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

Molecular Elements

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called **molecules**. The smallest part of a substance that has the physical and chemical properties of that substance. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Table 6.4.1: Elements That Exist as Diatomic Molecules

Hydrogen, H	Oxygen	Nitrogen	Fluorine	Chlorine	Bromine	Iodine

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules. A molecule with only two atoms (Table **6.4.1**). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2, N_2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 (Figure **6.4.1**).

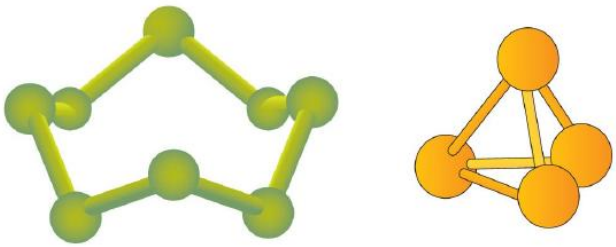


Figure **6.4.1**: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure **6.4.1** shows two examples of how we will be representing molecules in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

Ionic Compounds

The elements in the periodic table are divided into specific groupings; the metals, the non-metals, the semi-metals, and so on. These groupings are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). The basic unit of ionic compounds are called **formula units**.

Molecular Compounds

Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds. Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals) and silicon

dioxide (SiO_2) will also be a covalently bound molecule (a semi-metal and a non-metal). The basic unit of molecular compounds are called **molecules**.

Example 6.4.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. Fe
b. PCl_3
c. LiBr
d. P_4
e. oxygen gas

Solution

a. **Fe**, iron is an element that is represented with no subscript, so it is an **atomic element**.
b. **PCl_3** is made up of two nonmetals, so it is a **molecular compound**.
c. **LiBr** is made up of lithium, a metal and bromine a nonmetal, so it is an **ionic compound**.
d. **P_4** is a substance that is made up of four atoms of the same element, so it is a **molecular element**.
e. The formula for **oxygen gas** is **O_2** so it is a molecular element.

Exercise 6.4.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. I_2
b. He
c. H_2O
d. Al
e. CuCl

Answer a:
molecular element

Answer b:
atomic element

Answer c:
molecular compound

Answer d:
atomic element

Answer e:
ionic compound

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.4: A Molecular View of Elements and Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.5: Writing Formulas for Ionic Compounds

Learning Objectives

- Write the correct formula for an ionic compound.
- Recognize polyatomic ions in chemical formulas.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like $NaCl$, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

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

Example 6.5.1: Aluminum Nitride and Lithium Oxide

Write the formula for aluminum nitride and lithium oxide.

Solution

	Write the formula for aluminum nitride	**Write the formula for lithium oxide**
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Al^{3+} N^{3-}	Li^{+} O^{2-}
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

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

Example 6.5.2: The Crisscross Method for lead (IV) oxide

Write the formula for lead (IV) oxide.

Solution

Crisscross Method	Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Pb^{4+} O^{2-}
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all all subscripts that are 1.	PbO_2

Exercise 6.5.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

> *Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.*

Example 6.5.3: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Na^{+} S^{2-}
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all all subscripts that are 1.	Na_2S

Exercise 6.5.3

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a:
NaBr

Answer b:
LiCl

Answer c:
MgO

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Table 6.5.1 lists the most common polyatomic ions.

Table 6.5.1: Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 6.5.4: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Criss Cross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Ca^{2+} NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 6.5.5

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion

Solution

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 6.5.5

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

$MgCO_3$

Answer b:

$Al(CH_3COO)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 (or 2A) metal) and a nonmetallic element (bromine, a group 17 (or 7A) nonmetal). Therefore, it is most likely an ionic compound. (In fact, it *is* ionic.) In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 (or 5A), and oxygen, from group 16 (or 6A). It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

Example 6.5.6

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. NH_4Cl
d. OF_2

Solution

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic.	Na_2O, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.	PCl_3, not ionic
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic.	NH_4Cl, ionic
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.	OF_2, ionic

Exercise 6.5.6

Identify each compound as ionic or not ionic.

a. N_2O
b. $FeCl_3$
c. $(NH_4)_3PO_4$
d. $SOCl_2$

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Summary

Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.5: Writing Formulas for Ionic Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.6: Nomenclature: Naming Compounds

Nomenclature is the process of naming chemical compounds with different names so that they can be easily identified as separate chemicals. The primary function of chemical nomenclature is to ensure that a spoken or written chemical name leaves no ambiguity concerning which chemical compound the name refers to: each chemical name should refer to a single substance. A less important aim is to ensure that each substance has a single name, although a limited number of alternative names is acceptable in some cases. Preferably, the name also conveys some information about the structure or chemistry of a compound. A common name will often suffice to identify a chemical compound in a particular set of circumstances. To be more generally applicable, the name should indicate at least the chemical formula. To be more specific still, the three-dimensional arrangement of the atoms may need to be specified.

Contributors and Attributions

- Wikipedia (CC-BY-SA-3.0)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.6: Nomenclature: Naming Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.7: Naming Ionic Compounds

Learning Objectives

- To use the rules for naming ionic compounds.

After learning a few more details about the names of individual ions, you will be a step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system**, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 6.7.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table **6.7.1** lists the elements that use the common system, along with their respective cation names.

Table 6.7.2: Some Monatomic Anions

Ion	Name		Ion	Name
F^-	fluoride ion		O^{2-}	oxide ion
Cl^-	chloride ion		S^{2-}	sulfide ion
Br^-	bromide ion		P^{3-}	phosphide ion
I^-	iodide ion		N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table **6.7.2** lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

Example 6.7.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. SO_3^{2-}
d. NH_4^+
e. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

Exercise 6.7.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. SO_4^{2-}
d. Ba^{2+}
e. HCO_3^-

Answer a:
iron(II) ion

Answer b:
iron(III) ion

Answer c:
sulfate ion

Answer d:
barium ion

Answer e:
hydrogen carbonate ion or bicarbonate ion

Example 6.7.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Solution

a. Br^-
b. PO_4^{3-}
c. Cu^{2+}
d. Mg^{2+}

Exercise 6.7.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a:

F^-

Answer b:

CO_3^{2-}

Answer c:

Sn^{2+}

Answer d:

K^+

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

Now that we know how to name ions, we are ready to name ionic compounds. A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure **6.7.1** for the compound $BaCl_2$. The word *ion* is dropped from both parts.

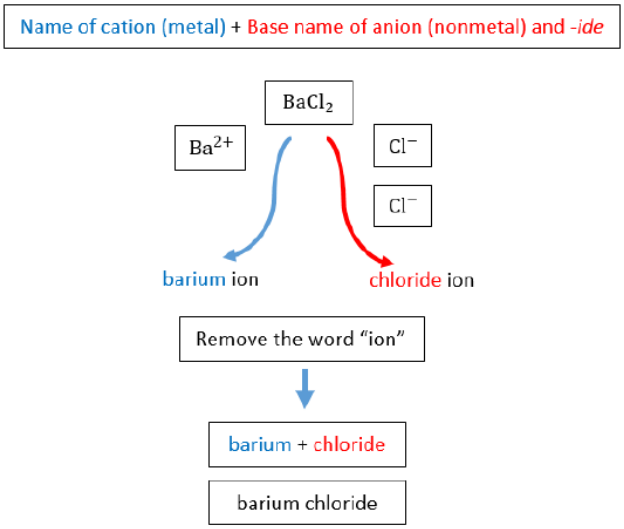


Figure **6.7.1**

Subscripts in the formula do not affect the name.

Example 6.7.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

Exercise 6.7.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

Answer c:

calcium phosphide

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure **6.7.2**).

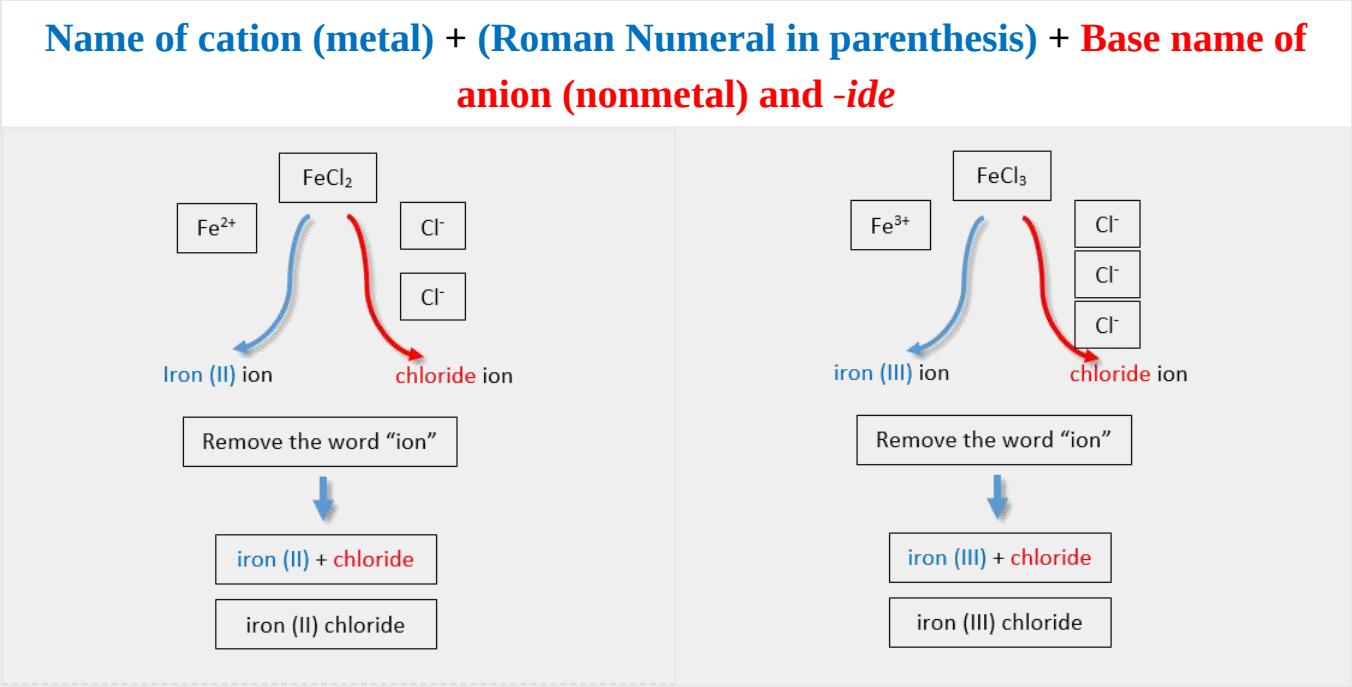


***Figure* 6.7.2:** *Naming the* $FeCl_2$ *and* $FeCl_3$ *Compounds in the Modern/Stock System.*

If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure **6.7.3**) .

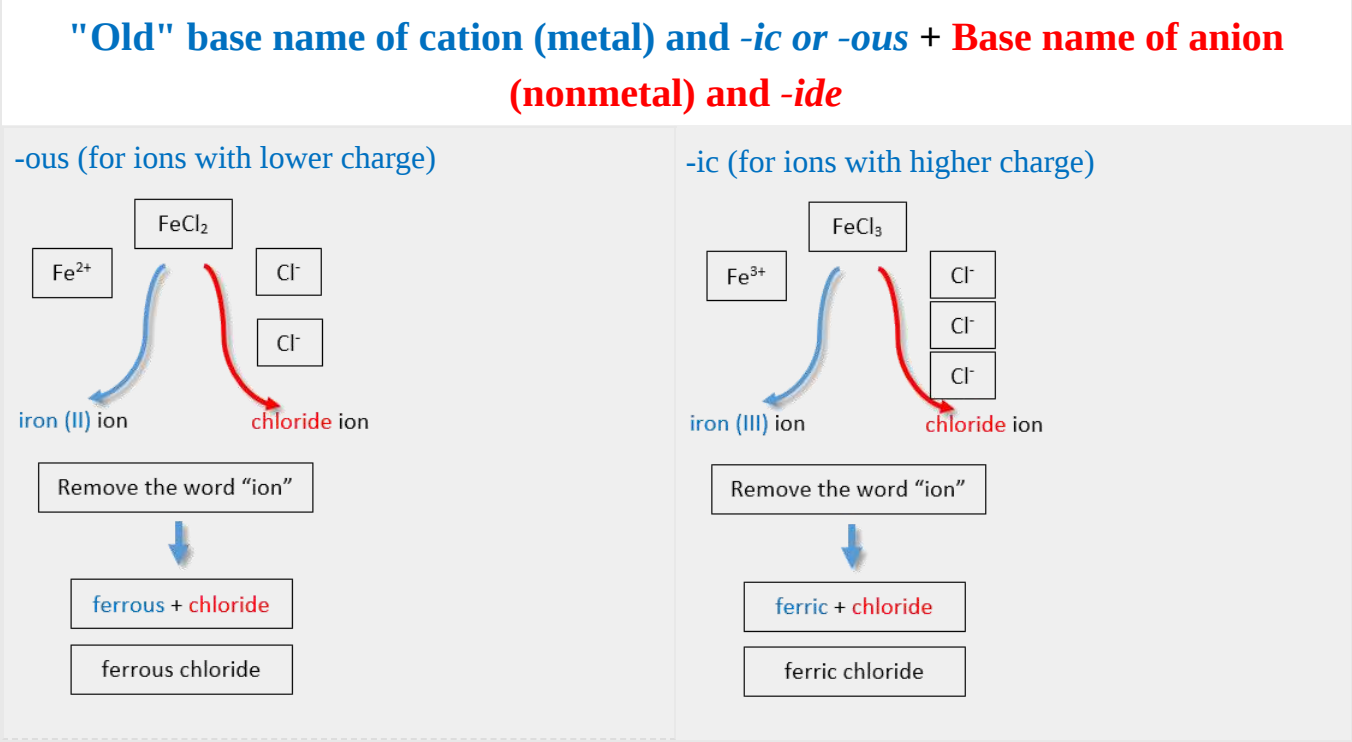


***Figure* 6.7.3:** *Naming the* $FeCl_2$ *and* $FeCl_3$ *Compounds in the Old/Common System.*

Example 6.7.4:

Name each ionic compound.

a. Co_2O_3
b. FeCl2

Solution

https://chem.libretexts.org/@go/page/105299

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloriide always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−. • This means that the one iron ion must have a 2+charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

Exercise 6.7.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium Sulfate compound in Figure **6.7.6**.

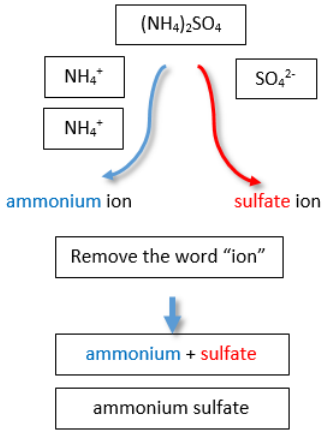


Figure **6.7.4**: *Naming Ionic Compounds with Polyatomic Ions*

Example 6.7.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound’s name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound’s name is iron(II) phosphate.	iron(II) phosphate

Exercise 6.7.5A

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Figure **6.7.1** is a synopsis of how to name simple ionic compounds.

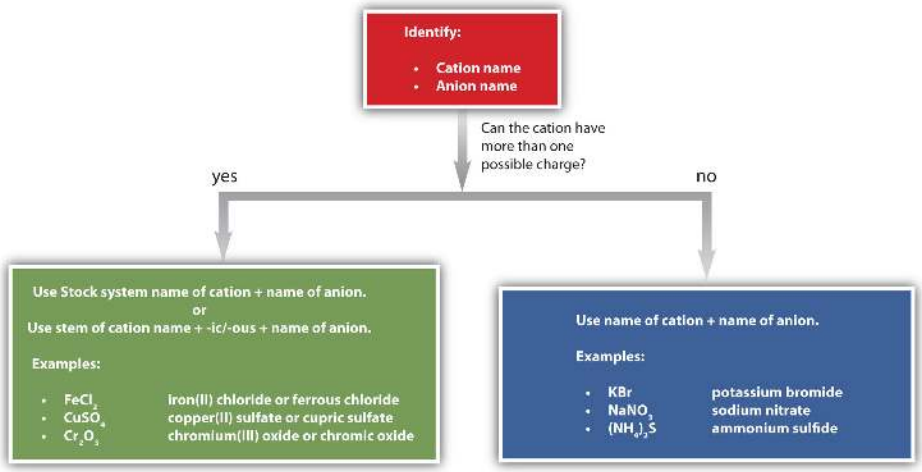


Figure **6.7.1**: A Guide to Naming Simple Ionic Compounds.

Exercise 6.7.5*B*

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

ammonium phosphate

Answer d:

gold(III) fluoride or auric fluoride

Answer e:

silver fluoride

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parenthes.
- Ternary compounds are composed of three or more elements.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.7: Naming Ionic Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.8: Naming Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

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

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **electrons** in such a way that a **bond** forms between 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.

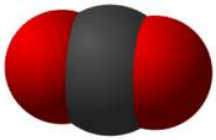


Figure 6.8.1: *Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.*

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

Naming Binary Molecular Compounds

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

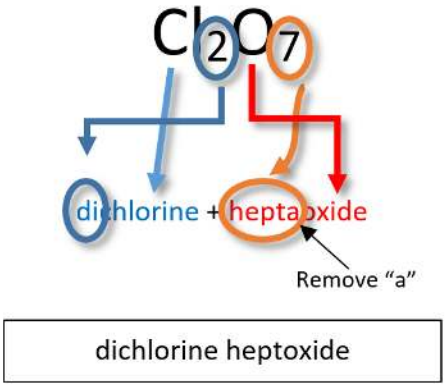


Figure 6.8.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 6.8.1 lists these numerical prefixes.

Table 6.8.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-
*This prefix is not used for the first element's name.	



Note

- Generally, the less-electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetr**o**xide instead of tetr**ao**xide.
- The prefix is "mono"is not added to the first element's name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table 6.8.2.

Table 6.8.2

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

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

Exercise 6.8.1

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system..

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.8: Naming Molecular Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.9: Naming Acids

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

Acids

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



Figure 6.9.1: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

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

$$HCl \rightarrow H^+ + Cl^- \tag{6.9.1}$$

Since acids produce H^+, cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic.

Naming Binary acids (in aqueous form)

A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. The acid name begins with the prefix *hydro-*. followed by the base name of the anion, followed by the suffix *-ic*.

Hydro- and Base name of Non metal and *-ic + acid*

Example: $HCl_{(aq)}$

Cl^-, chloride ion

hydrochloric acid

Naming Oxyacids

An **oxyacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

a. Oxyanions with ***-ite*** ending.

The name of the acid is the root of the anion followed by the suffix ***-ous***. There is no prefix.

Base name of oxyanion and *-ous+ acid*

Example: $H_2SO_{3(aq)}$

SO_3^{2-}, sulfite ion

sulfurous acid

b. Oxyanions with ***-ate*** ending.

The name of the acid is the root of the anion followed by the suffix ***-ic***. There is no prefix.

Base name of oxyanion and *-ic+ acid*

Example: $H_3PO_{4(aq)}$

PO_4^{3-} phosphate ion

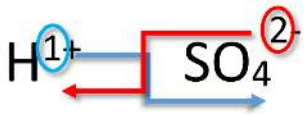
phosphoric acid

Note

The base name for sulfur containing oxyacid is ***sulfur-*** instead of just ***sulf-***. The same is true for a phosphorus containing oxyacid. The base name is ***phosphor-*** instead of simply ***phosph-***.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Another way to think about writing the correct formula is to utilize the crisscross method, shown below for sulfuric acid.



Formula: H_2SO_4

Figure 6.9.2: *Crisscross approach to writing formula for sulfuric acid.*

Summary

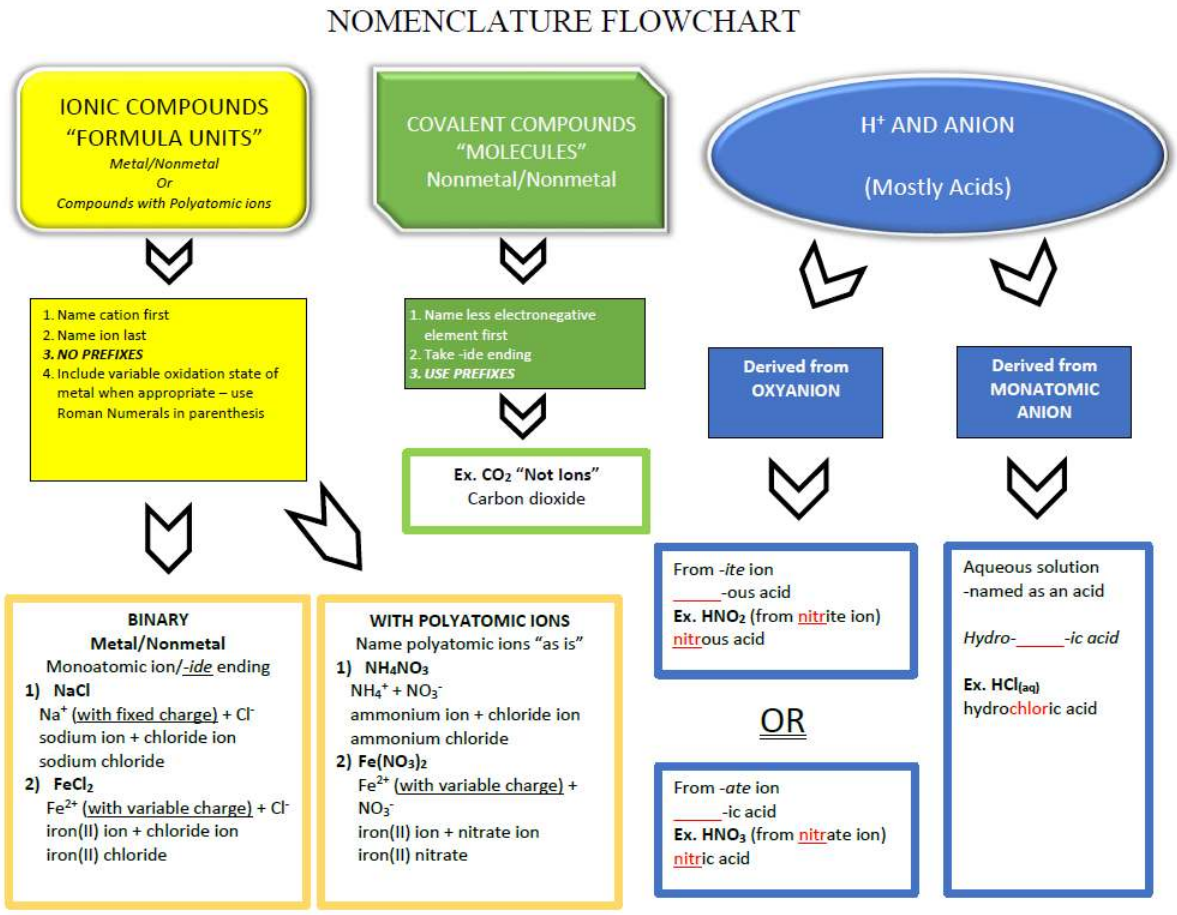
- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxyacids contain hydrogen, oxygen, and one other element.
- The name of the acid is based on the anion attached to the hydrogen.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.9: Naming Acids is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.10: Nomenclature Summary



Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

6.10: Nomenclature Summary is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.11: Formula Mass: The Mass of a Molecule or Formula Unit

Learning Objectives

- To determine the formula mass of an ionic compound.

One skill needed in future chapters is the ability to determine the mass of the formula of an ionic compound. This quantity is called the formula mass. The formula mass is obtained by adding the masses of each individual atom in the formula of the compound. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula mass of sodium chloride (NaCl). This formula mass is the sum of the atomic masses of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the masses to two decimal places:

Na: 22.99 amu

Cl: +35.34 amu

Total: 58.44 amu

To two decimal places, the formula mass of NaCl is 58.44 amu.

When an ionic compound has more than one anion or cation, you must remember to use the proper multiple of the atomic mass for the element in question. For the formula mass of calcium fluoride (CaF_2), we must multiply the mass of the fluorine atom by 2 to account for the two fluorine atoms in the chemical formula:

Ca: 1 x 40.08 = 40.08 amu

F: 2 x 19.00 = +38.00 amu

Total = 78.08 amu

The formula mass of CaF_2 is 78.08 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. For example, potassium nitrate (KNO_3) has one potassium atom, one nitrogen atom, and three oxygen atoms:

K: 1 x 39.10 = 39.10 amu

N: 1 x 14.00 = +14.00 amu

O: 3 x 16.00 = +48.00 amu

Total = 101.10 amu

The formula mass of KNO_3 is 101.10 amu.

> Potassium nitrate is a key ingredient in gunpowder and has been used clinically as a diuretic.

When a formula contains more than one polyatomic unit in the chemical formula, as in $Ca(NO_3)_2$, do not forget to multiply the atomic mass of every atom inside the parentheses by the subscript outside the parentheses. This is necessary because the subscript refers to the *entire polyatomic ion*. Thus, for $Ca(NO_3)_2$, the subscript 2 implies two complete nitrate ions, so we must sum the masses of two (1 × 2) nitrogen atoms and six (3 × 2) oxygen atoms, along with the mass of a single calcium atom:

Ca: 1 x 40.08 = 40.08 amu

N: 2 x 14.00 = +28.00 amu

O: 6 x 16.00 = +96.00 amu

Total = 164.08 amu

The key to calculating the formula mass of an ionic compound is to correctly count each atom in the formula and multiply the atomic masses of its atoms accordingly.

Example 6.11.1

Use the atomic masses (rounded to two decimal places) from the inside cover of this book to determine the formula mass for each ionic compound.

a. $FeCl_3$
b. $(NH_4)_3PO_4$

Solution

a.

Fe: 1 x 55.85 = 55.85 amu

Cl: 1 x 35.45 = +106.35 amu

Total = 162.20 amu

The formula mass of $FeCl_3$ is 162.2 amu.

b. When we distribute the subscript 3 through the parentheses containing the formula for the ammonium ion, we see that we have 3 nitrogen atoms and 12 hydrogen atoms. Thus, we set up the sum as follows:

N: 3 x 14.00 = 42.00 amu

H: 12 x 1.00 = +12.00 amu

P: 1 x 30.97 = +30.97 amu

O: 4 x 16.00 = +64.00 amu

Total = 148.97 amu

The formula mass for $(NH_4)_3PO_4$ is 149.0 amu.

Exercise 6.11.1

Use the atomic masses (rounded to two decimal places) from the inside cover of this book to determine the formula mass for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

Answer

a. 79.87 amu

b. 187.77 amu

c. 383.0 amu

To Your Health: Hydrates

Some ionic compounds have water (H_2O) incorporated within their formula unit. These compounds, called *hydrates*, have a characteristic number of water units associated with each formula unit of the compound. Hydrates are solids, not liquids or solutions, despite the water they contain.

To write the chemical formula of a hydrate, write the number of water units per formula unit of compound after its chemical formula. The two chemical formulas are separated by a vertically centered dot. The hydrate of copper(II) sulfate has five water units associated with each formula unit, so it is written as $CuSO_4 \cdot 5H_2O$. The name of this compound is copper(II) sulfate pentahydrate, with the *penta-* prefix indicating the presence of five water units per formula unit of copper(II) sulfate.



Figure used with permisison from Wikipedia

Hydrates have various uses in the health industry. Calcium sulfate hemihydrate ($CaSO_4 \cdot ½H_2O$), known as plaster of Paris, is used to make casts for broken bones. Epsom salt ($MgSO_4 \cdot 7H_2O$) is used as a bathing salt and a laxative. Aluminum chloride hexahydrate is an active ingredient in antiperspirants. The accompanying table lists some useful hydrates.

Table 6.11.1: Names and Formulas of Some Widely Used Hydrates

Formula	Name	Uses
$AlCl_3 \cdot 6H_2O$	aluminum chloride hexahydrate	antiperspirant
$CaSO_4 \cdot ½H_2O$	calcium sulfate hemihydrate (plaster of Paris)	casts (for broken bones and castings)
$CaSO_4 \cdot 2H_2O$	calcium sulfate dihydrate (gypsum)	drywall component
$CoCl_2 \cdot 6H_2O$	cobalt(II) chloride hexahydrate	drying agent, humidity indicator
$CuSO_4 \cdot 5H_2O$	copper(II) sulfate pentahydrate	fungicide, algicide, herbicide
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate (Epsom salts)	laxative, bathing salt
$Na_2CO_3 \cdot 10H_2O$	sodium carbonate decahydrate (washing soda)	laundry additive/cleane

KEY TAKEAWAY

- Formula masses of ionic compounds can be determined from the masses of the atoms in their formulas.

Contributors and Attributions

- Anonymous
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

6.11: Formula Mass: The Mass of a Molecule or Formula Unit is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

7: Quantitative Composition of Compounds

The following sections are concerned with the amounts of substances which participate in chemical reactions, the quantities of heat given off or absorbed when reactions occur, and the volumes of solutions which react exactly with one another. These seemingly unrelated subjects are discussed together because many of the calculations involving them are almost identical in form. The same is true of the density calculations and of the calculations involving molar mass and the Avogadro constant.

Template:HideTOC

7.1: How Much Sodium?
7.2: Counting Nails by the Pound
7.3: Counting Atoms by the Gram
7.4: Counting Molecules by the Gram
7.5: Chemical Formulas as Conversion Factors
7.6: Mass Percent Composition of Compounds
7.7: Mass Percent Composition from a Chemical Formula
7.8: Calculating Empirical Formulas for Compounds
7.9: Calculating Molecular Formulas for Compounds

7: Quantitative Composition of Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.1: How Much Sodium?

Why is Knowledge of Composition Important? everything in nature is either chemically or physically combined with other substances to know the amount of a material in a sample, you need to know what fraction of the sample it is Some Applications: the amount of sodium in sodium chloride for diet the amount of iron in iron ore for steel production the amount of hydrogen in water for hydrogen fuel the amount of chlorine in freon to estimate ozone depletion .



How much salt is in this salt packet? Image used wtih permission (CC-BY-SA; 3.0; Swilliams).

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

7.1: How Much Sodium? is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.2: Counting Nails by the Pound

Counting by Weighing and Avogadro's number

The size of molecule is so small that it is physically difficult if not impossible to directly count out molecules (Figure 7.2.1). However, we can count them indirectly by using a common trick of "counting by weighing".

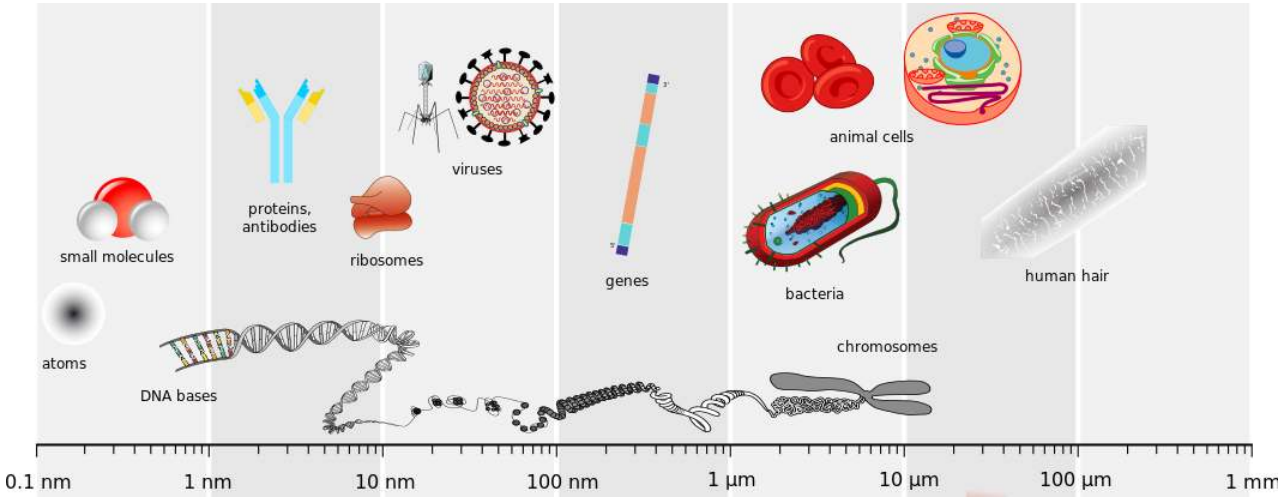


Figure 7.2.1: A comparison of the scales of various biological and technological objects. (CC BY-SA 3.0; Wikipedia)

Consider the example of counting nails in a big box at a hardware store. You need to estimate the number of nails in a box. The weight of an empty box is $213\ g$ and the weight of the box plus a bunch of big nails is $1340\ g$. Let's assume we know that the weight of one big nail is 0.450 g. I hope you aren't going to tear open the package and count the nails. We agree that

$$\text{mass of big nails} = 1340\ g - 213\ g = 1227\ g \tag{7.2.1}$$

Therefore

$$\text{Number of big nails in box} = \frac{1227\ g}{0.450\ g/\text{big nail}} = 2,726.6\ \text{big nails} = 2,730\ \text{big nails.} \tag{7.2.2}$$

You have just counted the number of big nails in the box by weighing them (rather than counting them individually).



Figure 7.2.2: Galvanized nails. Individually counting nails in a box would would require significant effort. Alternatively, we can cound them by weighing. (Public Domain; Wikipedia).

Now consider if the box of nails weighed the same, but box were filled with small nails with an individual mass of $0.23\ g/\text{small nail}$ instead? You would do the same math, but use a different denominator in Equation 7.2.2:

$$\text{Number of small nails in box} = \frac{1227\ g}{0.230\ g/\text{small nail}} = 5,$$
$$334.7\ \text{small nails} = 5,335\ \text{small nails.} \tag{7.2.3}$$

The individual mass is the conversion factor used the calculation and changes based on the nature of the nail (big or small). Let's ask ask a different question? How many *dozens* of nails are their in that same box of small nails described above.

If we knew the information from in Equation 7.2.3, we can just use the conversion of how many nails is in a dozen

$$\frac{5,335\,\text{small nails}}{12\text{small nails/dozen}} = 444.6\,\text{dozen small nails} \tag{7.2.4}$$

If we wanted to get this value from weighing we use the "dozen mass" instead of individual mass of

$$12 \times 0.23g = 2.76\,g/\text{dozen small nails.} \tag{7.2.5}$$

So following Equation 7.2.3 we get

$$\begin{aligned}\text{Number of dozens of small nails} &= \frac{1227\,g}{2.76\,g/\text{dozen small nails}} \\ &= 444.6\,\text{dozen small nails}\end{aligned} \tag{7.2.6}$$

and this is the same result as from Equation 7.2.4. These calculations demonstrate the difference between individual mass (i.e., per individual) and collective mass (e.g., per dozen or per gross). The collective mass of most importance to chemistry is *molar mass* (i.e., mass per mole or mass per 6.022×10^{23}).

Avogadro's Number

Avogadro's number is an accident of nature. It is the number of particles that delivers a mole of a substance. Avogadro's number = 6.022×10^{23}. The reason why the value is an accident of nature is that the mole is tied to the gram mass unit. The gram is a convenient mass unit because it matches human sizes. If we were a thousand times greater in size (like Paul Bunyan) we would find it handy to use kilogram amounts. This means the kilogram mole would be convenient. The number of particles handled in a kilogram mole is 1000 times greater. The kilo Avogadro number for the count of particles in a kilomole is 6.022×10^{26}.

If humans were tiny creatures (like Lilliputians) only 1/1000 our present size, milligrams would be more convenient. This means the milligram mole would be more useful. The number of particles handled in a milligram mole (millimole) would be 1/1000 times smaller. The milli Avogadro number for the count of particles in a millimole is 6.022×10^{20}.

What do you think would happen to Avogadro's number if the American system was used and amounts were measured in pound moles? Remember 1 pound = 454 grams. Avogadro's number would be larger by a factor of 454. A pound mole of hydrogen would weigh 1 pound which would be 454 grams. A gram mole of hydrogen weighs 1 gram and contains 6.022×10^{23} H atoms.

Molar Mass for Elements

You are able to read the periodic table and determine the average atomic mass for an element like carbon. The average mass is 12.01 amu. This mass is a ridiculously tiny number of grams. It is too small to handle normally. The molar mass of carbon is defined as the mass in grams that is numerically equal to the average atomic weight. This means

$$1g/molecarbon = 12.01\,g\,carbon \tag{7.2.7}$$

this is commonly written

$$1\,mol\,carbon = 12.01\,grams\,carbon. \tag{7.2.8}$$

This is the mass of carbon that contains 6.022×10^{23} carbon atoms.

Avogadro's number is 6.022×10^{23} particles.

This same process gives us the molar mass of any element. For example

- $1\,mol\,neon = 20.18\,g\,neon\,Ne$
- $1\,mol\,sodium = 22.99\,g\,sodium\,Na$

Molar Mass for Compounds

Example 7.2.1: molar mass of Water

The formulas for compounds are familiar to you. You know the formula for water is H_2O. It should be reasonable that the weight of a formula unit can be calculated by adding up the weights for the atoms in the formula.

Solution

The formula weight for water

weight from hydrogen + weight from oxygen

The formula weight for water

2 H atoms x 1.008 amu + 1 O atom x 16.00 amu = 18.016 amu

The molar mass for water

18.016 grams water or 18 grams to the nearest gram

Example 7.2.2: molar mass of Methane

The formula for methane the major component in natural gas is CH_4.

Solution

The formula weight for methane

weight from hydrogen + weight from carbon

The formula weight for methane

4 H atoms x 1.008 amu + 1 C atom x 12.01 amu = 16.04 amu

The molar mass for methane

16.04 grams per mole of methane

Example 7.2.3: Molar Mass of Ethyl Chloride

What is its molar mass for ethyl chloride CH_3CH_2Cl?

Solution

The *formula weight*

weight from hydrogen + weight from carbon + weight from chlorine

The formula weight

5 H atoms x 1.008 amu + 2 C atom x 12.01 amu + 35.5 amu = 64.5 amu

The molar mass for ethyl chloride

64.5 grams per mole of ethyl chloride

Example 7.2.4: Sulfur Dioxide

What is the molar mass for sulfur dioxide, $SO_2\,(g)$, a gas using in bleaching and disinfection processes?

Solution

Look up the atomic weight for each of the elements in the formula.

- 1 sulfur atom = 32. 07 amu
- 1 oxygen atom = 16.00 amu

Count the atoms of each element in the formula unit. .

- one sulfur atom
- two oxygen atoms

The formula weight

weight from sulfur + weight from oxygen

The formula weight

1 sulfur atom x (32. 07 amu) + 2 oxygen atoms x (16.00 amu)

The formula weight

SO_2 = 32. 07 amu + 32.00 amu = 64.07 amu = 64 amu SO_2

The molar mass for SO_2 is

64.07 grams of SO_2; 1 mol SO_2 = 64 grams per mole of SO_2

Exercise 7.2.1

What is the formula weight and molar mass for alum, $KAl(SO_4)_2 \cdot 12H_2O$?

Answer

1. Check the periodic table for the atomic masses for each atom in the formula
2. Count the number of each type of atom in formula.
3. Multiply the number of atoms by the atomic mass for each element.
4. Add up the masses for all of the elements.

element	average atomic mass	number of atoms in formula	rounded to nearest one unit for simplicity
potassium k	39.1 amu	1 =	39. amu
aluminum	26.98 amu	1	27. amu
sulfur	32.07 amu	2	64. amu
oxygen	16.00 amu	8 + 12 = 20	320. amu
hydrogen	1.008 amu	2 x 12 = 24	24. amu
			474. amu This is the formula weight.

Molar mass is 474 grams. This is a mass in grams that is numerically (474) the same as the formula weight.

1 mole alum $KAl(SO_4)_2 \cdot 12H_2O$ = 474 grams alum $KAl(SO_4)_2 \cdot 12H_2O$

7.2: Counting Nails by the Pound is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.3: Counting Atoms by the Gram

Learning Objectives

- Use Avogadro's number to convert to moles and vice versa given the number of particles of an element.
- Use the molar mass to convert to grams and vice versa given the number of moles of an element.

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number

In chemistry, it is impossible to deal with a single atom or molecule because we can't see them or count them or weigh them. Chemists have selected a number of particles with which to work that is convenient. Since molecules are extremely small, you may suspect this number is going to be very large and you are right. The number of particles in this group is $\mathbf{6.02 \times 10^{23}}$ particles and the name of this group is the **mole** (the abbreviation for ***mole*** is $\mathbf{mol}$). One mole of any object is $\mathbf{6.02 \times 10^{23}}$ of those objects. There is a very particular reason that this number was chosen and we hope to make that reason clear to you.

When chemists are carrying out chemical reactions, it is important that the relationship between the numbers of particles of each reactant is known. Any readily measurable mass of an element or compound contains an extraordinarily large number of atoms, molecules, or ions, so an extremely large numerical unit is needed to count them. The mole is used for this purpose.

A mole is defined as the amount of a substance that contains the number of carbon atoms in exactly 12 g of isotopically pure carbon-12. According to the most recent experimental measurements, this mass of carbon-12 contains 6.022142×10^{23} atoms, but for most purposes 6.022×10^{23} provides an adequate number of significant figures. Just as 1 mole of atoms contains 6.022×10^{23} atoms, 1 mole of eggs contains 6.022×10^{23} eggs. This number is called Avogadro's number, after the 19th-century Italian scientist who first proposed a relationship between the volumes of gases and the numbers of particles they contain.

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.022×10^{23}.

Converting Between Number of Atoms to Moles and Vice Versa

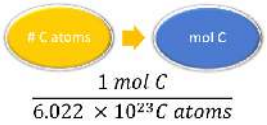
We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles by using the relationship

$$\mathbf{1\ mol\ X = 6.022 \times 10^{23}\ X\ atoms.} \tag{7.3.1}$$

Example 7.3.1: Moles of Carbon

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is $\mathbf{4.72 \times 10^{24}}$ atoms of carbon?

Solution

Steps for Problem Solving	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?
Identify the "given"information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C
List other known quantities	$1\ mol = 6.022 \times 10^{23}$ C atoms
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$4.72 \times 10^{24}\ \cancel{\text{C atoms}} \times \frac{1\ \text{mol C}}{6.02 \times 10^{23}\ \cancel{\text{C atoms}}} = 7.84\ \text{mol C}$ (7.3.2)
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

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 the element lithium is 6.94g, the molar mass of zinc is 65.38g, and the molar mass of gold is 196.97g. Each of these quantities contains 6.022×10^{23} atoms of that particular element. The units for molar mass are grams per mole or g/molg/mol. $1.00\ \text{mol}$ of carbon-12 atoms has a mass of $12.0\ \text{g}$ and contains 6.022×10^{23} atoms. 1.00 mole of any element has a mass numerically equal to its atomic mass in grams and contains 6.022×10^{23} particles. The mass, in grams, of 1 mole of particles of a substance is now called the **molar mass** (mass of 1.00 mole).

Converting Grams to Moles of an Element and Vice Versa

We can also convert back and forth between grams of an element and moles. The conversion factor for this is the molar mass of the substance. The **molar mass** is the ratio giving the number of grams for each one mole of the substance. This ratio is easily found by referring to the atomic mass of the element using the periodic table. This ratio has units of grams per mole or g/mol.

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 6.4.1 illustrates what conversion factor is needed and two examples are given below.

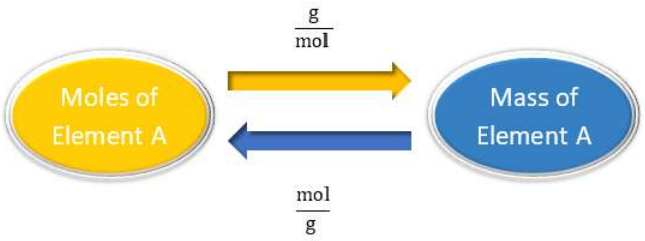


Figure 7.3.1: A Simple Flowchart for Converting Between Mass and Moles of an Element.

Example 7.3.2: Chromium

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

Steps for Problem Solving	**Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.**
Identify the "given"information and what the problem is asking you to "find."	Given: 0.560 mol Cr Find: g Cr
List other known quantities	1 mol Cr = 52.00g Cr
Prepare a concept map and use the proper conversion factor.	mol Cr → g Cr $\frac{52.00 g\ Cr}{1\ mol\ Cr}$
Cancel units and calculate.	$0.560\ \cancel{mol\ Cr} \times \frac{52.00\ g\ Cr}{1\ \cancel{mol\ Cr}} = 29.1\ g\ Cr \quad (7.3.3)$
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 mol

Example 7.3.3: Silicon

How many moles are in 107.6g of Si.

Solution

Steps for Problem Solving	How many moles are in 107.6g of Si.
Identify the "given"information and what the problem is asking you to "find."	Given: 107.6g Si Find: mol Si
List other known quantities	1 mol Si = 28.09g Si
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$107.6\ \cancel{g\ Si} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{g\ Si}} = 3.83\ \text{mol Si} \quad (7.3.4)$
Think about your result.	Since 1 mol of Si is 28.09g, 107.6 should be about 4 moles.

Exercise 7.3.1

a. How many moles are present in 100.0 g of Al?
b. What is the mass of 0.552 mol of Ag metal?

Answer a:

3.706 mol Al

Answer b:

59.5 g Ag

Summary

- A mole is defined as the amount of a substance that contains the number of carbon atoms in exactly 12 g of isotopically pure carbon-12.
- There are 6.02×10^{23} particles in 1.00 mole. This number is called Avogadro's number.
- The molar mass of an element can be found by referring to the atomic mass on a periodic table with units of g/mol.
- Using dimensional analysis, it is possible to convert between grams, moles, and the number of atoms or molecules.

Further Reading/Supplemental Links

- learner.org/resources/series61.html The ***learner.org*** website allows user to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called ***The Mole***.
- Using Avogadro's law, the mass of a substance can be related to the number of particles contained in that mass. The Mole: (www.learner.org/vod/vod_window.html?pid=803)
- Vision Learning tutorial: The Mole http://visionlearning.com/library/mo...p?mid-53&1=&c3=

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)

- Henry Agnew (UC Davis)

7.3: Counting Atoms by the Gram is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.4: Counting Molecules by the Gram

Learning Objectives

- Define molecular mass and formula mass.
- Perform conversions between mass and number of particles.

Molecular and Formula Masses

The molecular mass of a substance is the sum of the average masses of the atoms in one molecule of a substance. It is calculated by adding together the atomic masses of the elements in the substance, each multiplied by its subscript (written or implied) in the molecular formula. Because the units of atomic mass are atomic mass units, the units of molecular mass are also atomic mass units. The procedure for calculating molecular masses is illustrated in Example **7.4.1**.

Example 7.4.1: Ethanol

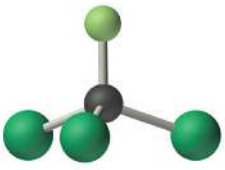
Calculate the molecular mass of ethanol, whose condensed structural formula is CH_3CH_2OH. Among its many uses, ethanol is a fuel for internal combustion engines

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: Ethanol molecule (CH_3CH_2OH) Find: molecular mass
Determine the number of atoms of each element in the molecule.	The molecular formula of ethanol may be written in three different ways: • CH_3CH_2OH (which illustrates the presence of an ethyl group • CH_3CH_{2-}, and an −OH group) • C_2H_5OH, and C_2H_6O; All show that ethanol has two carbon atoms, six hydrogen atoms, and one oxygen atom.
Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element.	1 C atom = 12.011 amu 1 H atom = 1.0079 amu 1 O atom = 15.9994 amu
Add together the masses to give the molecular mass.	2C: (2atoms)(12.011amu/atom) = 24.022 amu 6H: (6atoms)(1.0079amu/atom) = 6.0474amu +1O: (1atoms)(15.9994amu/atom) =15.9994amu C_2H_6O : molecular mass of ethanol = 46.069amu

Exercise 7.4.1: Freon

Calculate the molecular mass of trichlorofluoromethane, also known as Freon-11, whose condensed structural formula is CCl_3F. Until recently, it was used as a refrigerant. The structure of a molecule of Freon-11 is as follows:



Answer:

137.37 amu

Unlike molecules, which form covalent bonds, ionic compounds do not have a readily identifiable molecular unit. Therefore, for ionic compounds, the **formula mass** (also called the empirical formula mass) of the compound is used instead of the molecular mass. The formula mass is the sum of the atomic masses of all the elements in the empirical formula, each multiplied by its subscript (written or implied). It is directly analogous to the molecular mass of a covalent compound. The units are atomic mass units.

> *Atomic mass, molecular mass, and formula mass all have the same units: atomic mass units.*

Example 7.4.2: Calcium Phosphate

Calculate the formula mass of $Ca_3(PO_4)_2$, commonly called calcium phosphate. This compound is the principal source of calcium found in bovine milk.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: Calcium phosphate [$Ca_3(PO_4)_2$] formula unit Find: formula mass
Determine the number of atoms of each element in the molecule.	• The empirical formula—$Ca_3(PO_4)_2$—indicates that the simplest electrically neutral unit of calcium phosphate contains three Ca^{2+} ions and two PO_4^{3-} ions. • The formula mass of this molecular unit is calculated by adding together the atomic masses of three calcium atoms, two phosphorus atoms, and eight oxygen atoms.
Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element.	1 Ca atom = 40.078 amu 1 P atom = 30.973761 amu 1 O atom = 15.9994 amu
Add together the masses to give the formula mass.	3Ca: (3 atoms) (40.078 amu/atom)=120.234amu 2P: (2 atoms) (30.973761amu/atom)=61.947522amu + 8O: (8 atoms)(15.9994amu/atom)=127.9952amu Formula mass of $Ca_3(PO_4)_2$=310.177amu

Exercise 7.4.2: Silicon Nitride

Calculate the formula mass of Si_3N_4, commonly called silicon nitride. It is an extremely hard and inert material that is used to make cutting tools for machining hard metal alloys.



Si_3N_4 bearing parts. (Public Domain; David W. Richerson and Douglas W. Freitag; Oak Ridge National Laboratory).

Answer:

140.29 amu

Molar Mass

The molar mass of a substance is defined as the mass in grams of 1 mole of that substance. One mole of isotopically pure carbon-12 has a mass of 12 g. For an element, the molar mass is the mass of 1 mol of atoms of that element; for a covalent molecular compound, it is the mass of 1 mol of molecules of that compound; for an ionic compound, it is the mass of 1 mol of formula units. That is, the molar mass of a substance is the mass (in grams per mole) of 6.022×10^{23} atoms, molecules, or formula units of that substance. In each case, the number of grams in 1 mol is the same as the number of atomic mass units that describe the atomic mass, the molecular mass, or the formula mass, respectively.

> *The molar mass of any substance is its atomic mass, molecular mass, or formula mass in grams per mole.*

The periodic table lists the atomic mass of carbon as 12.011 amu; the average molar mass of carbon—the mass of 6.022×10^{23} carbon atoms—is therefore 12.011 g/mol:

Table 7.4.1: Molar Mass of Select Substances

Substance (formula)	Basic Unit	Atomic, Molecular, or Formula Mass (amu)	Molar Mass (g/mol)
carbon (C)	atom	12.011 (atomic mass)	12.011
ethanol (C_2H_5OH)	molecule	46.069 (molecular mass)	46.069
calcium phosphate [$Ca_3(PO_4)_2$]	formula unit	310.177 (formula mass)	310.177

Converting Between Grams and Moles of a Compound

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 ($CaCl_2$). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of $CaCl_2$ that you should measure as show in Example 7.4.3.

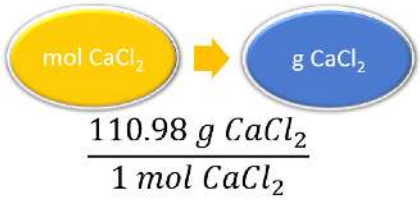
Example 7.4.3: Calcium Chloride

Calculate the mass of 3.00 moles of calcium chloride ($CaCl_2$).



Calcium chloride is used as a drying agent and as a road deicer.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: 3.00 moles of $CaCl_2$ Find: g $CaCl_2$
List other known quantities	1 ml $CaCl_2$ = 110.98 g $CaCl_2$
Prepare a concept map and use the proper conversion factor.	 $\frac{110.98\ g\ CaCl_2}{1\ mol\ CaCl_2}$
Cancel units and calculate.	$3.00\ \cancel{mol\ CaCl_2} \times \frac{110.98\ g\ CaCl_2}{1\ \cancel{mol\ CaCl_2}} = 333\ g\ CaCl_2$
Think about your result.	

Exercise 7.4.3: Calcium Oxide

What is the mass of $7.50\ mol$ of (calcium oxide) CaO?

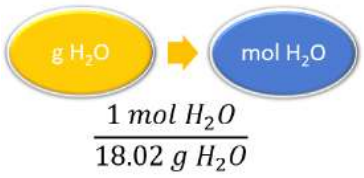
Answer:

420.60 g

Example 7.4.4: Water

How many moles are present in 108 grams of water?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: 108 g H_2O Find: mol H_2O
List other known quantities	$1\ mol\ H_2O = 18.02\ g\ H_2O$
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ H_2O}{18.02\ g\ H_2O}$
Cancel units and calculate.	$108\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 5.99\ mol\ H_2O$
Think about your result.	

https://chem.libretexts.org/@go/page/105308

Exercise 7.4.4: Nitrogen Gas

What is the mass of $7.50\ \text{mol}$ of Nitrogen gas N_2?

Answer:

105 g

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 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 moles. To convert from mass to number of particles or vice-versa, it will first require a conversion to moles as shown in Figure 7.4.1 and Example 7.4.5.

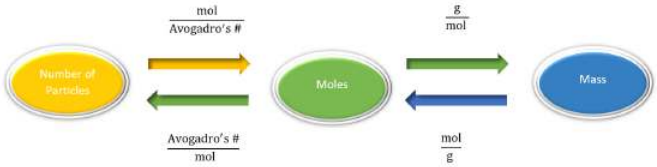
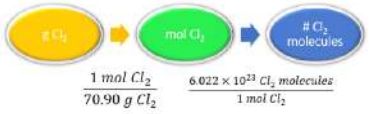


Figure 7.4.1: *Conversion from number of particles to mass or from mass to number of particles requires two steps.*

Example 7.4.5: Chlorine

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: 20.0 g Cl_2 Find: # Cl_2 molecules
List other known quantities	1 mol Cl_2 = 70.90 g Cl_2, 1mol Cl_2 = 6.022×10^{23} Cl_2 molecules
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$20.0\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{6.02 \times 10^{23}\ \text{molecules Cl}_2}{1\ \cancel{\text{mol Cl}_2}} = 1.70 \times 10^{23}\ \text{molecules Cl}_2$
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.

Exercise 7.4.5: Calcium Chloride

How many formula units are in 25.0 g of $CaCl_2$?

Answer:

1.36×10^{23} $CaCl_2$ formula units

Summary

- Calculations for formula mass and molecular mass are described.
- Calculations involving conversions between moles of a material and the mass of that material are described.
- Calculations are illustrated for conversions between mass and number of particles.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

7.4: Counting Molecules by the Gram is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.5: Chemical Formulas as Conversion Factors

Learning Objectives

- Use chemical formulas as conversion factors.

Figure **7.5.1** shows that we need two hydrogen atoms and one oxygen atom to make 1 water molecule. If we want to make two water molecules, we will need four hydrogen atoms and two oxygen atoms. If we want to make five molecules of water, we need 10 hydrogen atoms and 5 oxygen atoms. The ratio of atoms we will need to make any number of water molecules is the same: 2 hydrogen atoms to 1 oxygen atom.

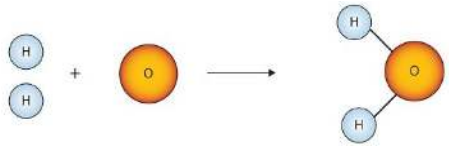


Figure **7.5.1** Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

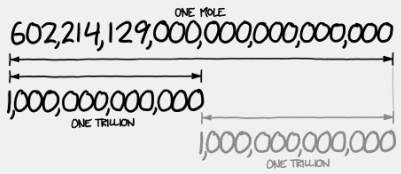
Using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of moles of molecules to the number of moles of atoms. For example, in 1 mol of water (H_2O) we can construct the relationships given in (Table **7.5.1**).

Table 7.5.1: Molecular Relationships for Water

1 Molecule of H_2O Has	1 Mol of H_2O Has	Molecular Relationships
2 H atoms	2 mol of H atoms	$\frac{2\text{ mol H atoms}}{1\text{ mol }H_2O\text{ molecules}}$ or $\frac{1\text{ mol }H_2O\text{ molecules}}{2\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }H_2O\text{ molecules}}$ or $\frac{1\text{ mol }H_2O\text{ molecules}}{1\text{ mol O atoms}}$

The MOle is big

A mole represents a very large number! The number 602,214,129,000,000,000,000,000 looks about twice as long as a trillion, which means it's about a trillion trillion.



(CC BY-SA NC; https://what-if.xkcd.com/4/).

A trillion trillion kilograms is how much a planet weighs. If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Table 7.5.2: Molecular and Mass Relationships for Ethanol

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular and Mass Relationships
2 C atoms	2 mol of C atoms	$\frac{2\text{ mol C atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{2\text{ mol C atoms}}$
6 H atoms	6 mol of H atoms	$\frac{6\text{ mol H atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{6\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{1\text{ mol O atoms}}$
2 (12.01 amu) C 24.02 amu C	2 (12.01 g) C 24.02 g C	$\frac{24.02\text{ g C}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{24.02\text{ g C}}$
6 (1.008 amu) H 6.048 amu H	6 (1.008 g) H 6.048 g H	$\frac{6.048\text{ g H}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{6.048\text{ g H}}$
1 (16.00 amu) O 16.00 amu O	1 (16.00 g) O 16.00 g O	$\frac{16.00\text{ g O}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{16.00\text{ g O}}$

The following example illustrates how we can use the relationships in Table **7.5.2** as conversion factors.

Example 7.5.1: Ethanol

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?

Solution

Steps for Problem Solving	If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?
Identify the "given"information and what the problem is asking you to "find."	Given: 2.5 mol C_2H_6O Find: mol C atoms
List other known quantities	1 mol C_2H_6O = 2 mol C
Prepare a concept map and use the proper conversion factor.	mol C_2H_6O → mol C atom $\frac{2\ mol\ C}{1\ mol\ C_2H_6O}$
Cancel units and calculate.	Note how the unit *mol* C_2H_6O *molecules* cancels algebraically. $2.5\ \cancel{\text{mol } C_2H_6O \text{ molecules}} \times \frac{2\ \text{mol C atoms}}{1\ \cancel{\text{mol } C_2H_6O \text{ molecules}}} = 5.0\ \text{mol C}$
Think about your result.	There are twice as many C atoms in one C_2H_6O molecule, so the final amount should be double.

Exercise 7.5.1

If a sample contains 6.75 mol of Na_2SO_4, how many moles of sodium atoms, sulfur atoms, and oxygen atoms does it have?

Answer:

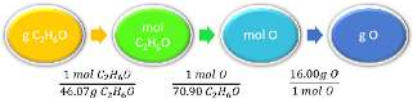
13.5 mol Na atoms, 6.75 mol S atoms, and 27.0 mol O atoms

The fact that 1 mol equals 6.022×10^{23} items can also be used as a conversion factor.

Example 7.5.2: Oxygen Mass

Determine the mass of Oxygen in 75.0g of C_2H_6O.

Solution

Steps for Problem Solving	Determine the mass of Oxygen in 75.0g of C_2H_6O
Identify the "given"information and what the problem is asking you to "find."	Given: 75.0g C_2H_6O Find: g O
List other known quantities	1 mol O = 16.0g O 1 mol C_2H_6O = 1 mol O 1 mol C_2H_6O = 46.07g C_2H_6O
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$75.0\ \cancel{g\ C_2H_6O} \times \frac{1\ \cancel{mol\ C_2H_6O}}{46.07\ \cancel{g\ C_2H_6O}} \times \frac{1\ \cancel{mol\ O}}{1\ \cancel{mol\ C_2H_6O}} \times \frac{16.00\ g\ O}{1\ \cancel{mol\ O}} =$
Think about your result.	

Exercise 7.5.2

a. How many molecules are present in 16.02 mol of C_4H_{10}? How many C atoms are in 16.02 mol?
b. How many moles of each type of atom are in 2.58 mol of Na_2SO_4?

Answer a:

9.647 x 10^{24} C_4H_{10} molecules and 3.859 x 10^{25} C atoms

Answer b:

5.16 mol Na atoms, 2.58 mol S atoms, and 10.3 mol O atoms

Summary

In any given formula the ratio of the number of moles of molecules (or formula units) to the number of moles of atoms can be used as a conversion factor.

7.5: Chemical Formulas as Conversion Factors is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.6: Mass Percent Composition of Compounds

Learning Objectives

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

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

Percent Composition

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

$$\%\ \text{by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\% \tag{7.6.1}$$

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

Example 7.6.1: Percent Composition from Mass Data

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

Solution

https://chem.libretexts.org/@go/page/105310

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given : Mass of compound = 20.00 g Mass of Zn = 16.07 g Find: % Composition (% Zn and %O)
List other known quantities	Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass. Mass of oxygen = 20.00 g - 16.07 g = 3.93 g O
Cancel units and calculate.	$$\%\,\text{Zn} = \frac{16.07\ \text{g Zn}}{20.00\ \text{g}} \times 100\% = 80.35\%\ \text{Zn} \quad (7.6.2)$$ $$\%\,\text{O} = \frac{3.93\ \text{g O}}{20.00\ \text{g}} \times 100\% = 19.65\%\ \text{O} \quad (7.6.3)$$ Calculate the percent by mass of each element by dividing the mass of that element by the mass of the compound and multiplying by 100%.
Think about your result.	The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Exercise 7.6.1

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

Answer

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

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

7.6: Mass Percent Composition of Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.7: Mass Percent Composition from a Chemical Formula

Learning Objectives

- Determine the percent composition of each element in a compound from the chemical formula.

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\% \quad (7.7.1)$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 7.7.1

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Solution

Steps for Problem Solving	Example 7.7.1
Identify the "given"information and what the problem is asking you to "find."	Given : Cl_2O_7 Find: % Composition (% Cl and %O)
List other known quantities	Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g Molar mass of Cl_2O_7 = 182.90 g/mol
Cancel units and calculate.	$$\%Cl = \frac{70.90\text{ g Cl}}{182.90\text{ g}} \times 100\% = 38.76\%\text{ Cl} \quad (7.7.2)$$ $$\%\text{ O} = \frac{112.00\text{ g O}}{182.90\text{ g}} \times 100\% = 61.24\%\text{ O} \quad (7.7.3)$$ 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%.
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 needed 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.

$$12.50\text{ g } Cl_2O_7 \times \frac{38.76\text{ g Cl}}{100\text{ g } Cl_2O_7} = 4.845\text{ g Cl} \quad (7.7.4)$$

$$12.50\ \text{g}\ Cl_2O_7 \times \frac{61.24\ \text{g O}}{100\ \text{g}\ Cl_2O_7} = 7.655\ \text{g O} \tag{7.7.5}$$

The sum of the two masses is **12.50 g**, the mass of the sample size.

Exercise 7.7.1

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a:

78.32% Ba and 21.67% F

Summary

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

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

7.7: Mass Percent Composition from a Chemical Formula is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.8: Calculating Empirical Formulas for Compounds

Learning Objectives

- Define empirical formula.
- Determine empirical formula from percent composition of a compound.

In the early days of chemistry, there were few tools for the detailed study of compounds. Much of the information regarding the composition of compounds came from the elemental analysis of inorganic materials. The "new" field of organic chemistry (the study of carbon compounds) faced the challenge of not being able to characterize a compound completely. The relative amounts of elements could be determined, but so many of these materials had carbon, hydrogen, oxygen, and possibly nitrogen in simple ratios. We did not know exactly how many of these atoms were actually in a specific molecule.

Determining Empirical Formulas

An empirical formula tells us the relative ratios of different atoms in a compound. The ratios hold true on the *molar* level as well. Thus, H_2O is composed of two atoms of hydrogen and 1 atom of oxygen. Likewise, **1.0 mole of H_2O** is composed **of 2.0 moles of hydrogen** and **1.0 mole of oxygen**. We can also work backwards from molar ratios since *if we know the molar amounts of each element in a compound we can determine the empirical formula*.

> *In a procedure called elemental analysis, an unknown compound can be analyzed in the laboratory in order to determine the percentages of each element contained within it. These percentages can be transformed into the mole ratio of the elements, which leads to the empirical formula.*

Note

Steps to determine empirical formula.

1. Assume a **100 g** sample of the compound so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.
3. In order to find a whole-number ratio, divide the moles of each element by whichever of the moles from step 2 is the smallest.
4. If all the moles at this point are whole numbers (or very close), the empirical formula can be written with the moles as the subscript of each element.
5. In some cases, one or more of the moles calculated in step 3 will not be whole numbers. Multiply each of the moles by the smallest whole number that will convert each into a whole number. Write the empirical formula.

Example 7.8.1

A compound of iron and oxygen is analyzed and found to contain **69.94%** iron and **30.06%** oxygen. Find the empirical formula of the compound.

Solution

Steps for Problem Solving	Example 7.8.1
Identify the "given"information and what the problem is asking you to "find."	Given: % of $\mathrm{Fe} = 69.94\%$ % of $\mathrm{O} = 30.06\%$ Find: Empirical formula $= \mathrm{Fe_?O_?}$
Calculate	
a. Assume a $100\ \mathrm{g}$ sample, convert the same % values to grams.	$69.94\ \mathrm{g\ Fe}$ (7.8.1) $30.06\ \mathrm{g\ O}$ (7.8.2)
b. Convert to moles.	$69.94\ \mathrm{g\ Fe} \times \frac{1\ \mathrm{mol\ Fe}}{55.85\ \mathrm{g\ Fe}} = 1.252\ \mathrm{mol\ Fe}$ (7.8.3) $30.06\ \mathrm{g\ O} \times \frac{1\ \mathrm{mol\ O}}{16.00\ \mathrm{g\ O}} = 1.879\ \mathrm{mol\ O}$ (7.8.4)
c. Divide both moles by the smallest of the results.	$\mathrm{Fe}: \frac{1.252\ \mathrm{mol}}{1.252}$ $\mathrm{O}: \frac{1.879\ \mathrm{mol}}{1.252}$ The "non- whole number" empirical formula of the compound is $\mathrm{Fe_1O_{1.5}}$
Multiply each of the moles by the smallest whole number that will convert each into a whole number.	Fe:O = 2 (1:1.5) = 2:3 Since the moles of O is still not a whole number, both moles can be multiplied by 2, while rounding to a whole number.
Write the empirical formula.	The empirical formula of the compound is $\mathrm{Fe_2O_3}$.
Think about your result.	The subscripts are whole numbers and represent the mole ratio of the elements in the compound. The compound is the ionic compound iron (III) oxide.

Exercise 7.8.1

Mercury forms a compound with chlorine that is 73.9% mercury and 26.1% chlorine by mass. What is the empirical formula?

Answer

$\mathrm{HgCl_2}$

Summary

- A process is described for the calculation of the empirical formula for a compound based on the percent composition of that compound.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

7.8: Calculating Empirical Formulas for Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.9: Calculating Molecular Formulas for Compounds

Learning Objectives

- To understand the difference between empirical formulas and molecular formulas
- To determine molecular formula from percent composition and molar mass of a compound.

Below, we see two carbohydrates: glucose and sucrose. Sucrose is almost exactly twice the size of glucose, although their empirical formulas are very similar. Some people could distinguish them on the basis of taste, but it's not a good idea to go around tasting chemicals. The best way is to determine the molar masses - this approach allows you to easily tell which compound is which.

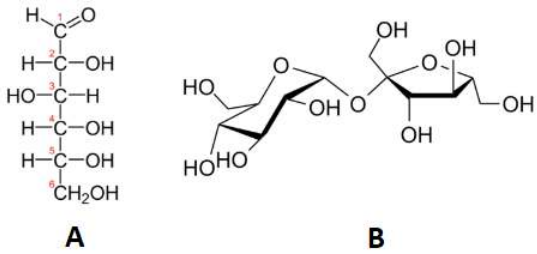


Figure **7.9.1**: (A) the molecular structure of glucose and (B) the molecular structure of sucrose.

Molecular Formulas

Molecular formulas give the kind and number of atoms of each element present in the molecular compound. In many cases, the molecular formula is the same as the empirical formula. The chemical formula will always be some *integer multiple* (n) of the empirical formula (i.e. integer multiples of the subscripts of the empirical formula).

$$\textbf{Molecular Formula} = n(\textbf{Empirical formula}) \tag{7.9.1}$$

therefore

$$n = \frac{\textbf{Molecular Formula}}{\textbf{Empirical Formula}} \tag{7.9.2}$$

The integer multiple, n can also be obtained by dividing the molar mass, MM, of the compound by the empirical formula mass, EFM (the molar mass represented by the empirical formula).

$$n = \frac{MM(molarmass)}{EFM(empiricalformulamolarmass)} \tag{7.9.3}$$

Table **7.9.1** shows the comparison between the empirical and molecular formula of methane, acetic acid, and glucose and the different values of n. The molecular formula of methane is $\mathbf{CH_4}$ and because it contains only one carbon atom, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole-number multiple of the empirical formula. Acetic acid is an organic acid that is the main component of vinegar. Its molecular formula is $\mathbf{C_2H_4O_2}$. Glucose is a simple sugar that cells use as a primary source of energy. Its molecular formula is $\mathbf{C_6H_{12}O_6}$. The structures of both molecules are shown in Figure **7.9.2**. They are very different compounds, yet both have the same empirical formula of $\mathbf{CH_2O}$.

Table 7.9.1: Molecular Formula and Empirical Formula of Various Compounds.

Name of Compound	Molecular Formula	Empirical Formula	n
Methane	$\mathbf{CH_4}$	$\mathbf{CH_4}$	1
Acetic acid	$\mathbf{C_2H_4O_2}$	$\mathbf{CH_2O}$	2
Glucose	$\mathbf{C_6H_{12}O_6}$	$\mathbf{CH_2O}$	6

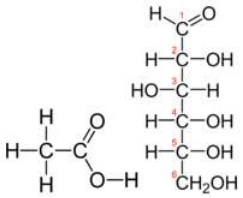


Figure 7.9.2: *Acetic acid (left) has a molecular formula of* $C_2H_4O_2$, *while glucose (right) has a molecular formula of* $C_6H_{12}O_6$. *Both have the empirical formula* CH_2O.

Empirical formulas can be determined from the percent composition of a compound as discussed in section 6.8. In order to determine its molecular formula, it is necessary to know the molar mass of the compound. Chemists use an instrument called a mass spectrometer to determine the molar mass of compounds. In order to go from the empirical formula to the molecular formula, follow these steps:

1. Calculate the **empirical formula mass (EFM)**.
2. Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.
3. Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.

Example 7.9.1

The empirical formula of a compound of boron and hydrogen is BH_3. Its molar mass is $27.7\ g/mol$. Determine the molecular formula of the compound.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: • Empirical formula $= BH_3$ • Molar mass $= 27.7\ g/mol$ Find: Molecular formula $=?$
Calculate the empirical formula mass (EFM).	$\text{Empirical formula mass (EFM)} = 13.84\ g/mol$ (7.9.4)
Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.	$69.94\ g\ Fe$ (7.9.5) $30.06\ g\ O$ (7.9.6) $\frac{\text{molar mass}}{\text{EFM}} = \frac{27.7}{13.84} = 2$ (7.9.7)
Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.	$BH_3 \times 2 = B_2H_6$ (7.9.8)
Write the molecular formula.	The molecular formula of the compound is B_2H_6.
Think about your result.	The molar mass of the molecular formula matches the molar mass of the compound.

Exercise **7.9.1**

Vitamin C (ascorbic acid) contains 40.92 % C, 4.58 % H, and 54.50 % O, by mass. The experimentally determined molecular mass is 176 amu. What is the empirical and chemical formula for ascorbic acid?

Answer: Empirical Formula

$C_3H_4O_3$

Answer: Molecular Formula

$C_6H_8O_6$

Summary

- A procedure is described that allows the calculation of the exact molecular formula for a compound.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

7.9: Calculating Molecular Formulas for Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

8: Chemical Reactions

Template:HideTOC

Topic hierarchy

8.1: Evidence of a Chemical Reaction

8.2: Chemical Equations

8.3: How to Write Balanced Chemical Equations

8.4: Classifying Chemical Reactions

8: Chemical Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.1: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

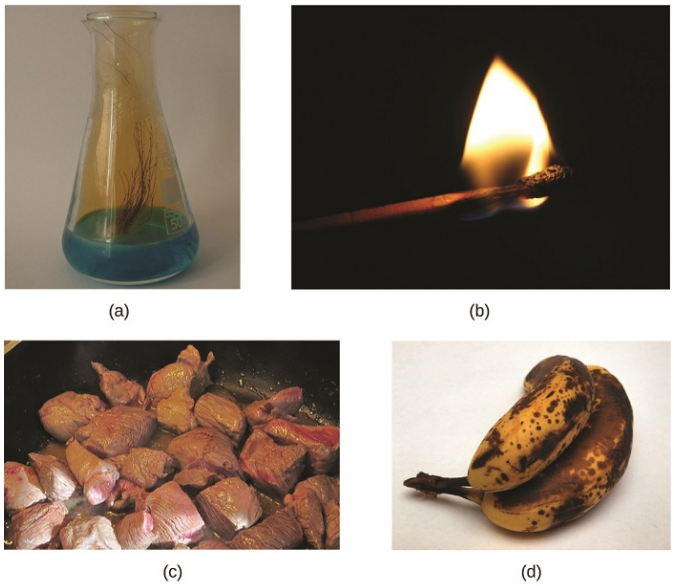


Figure **8.1.1**: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo)

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure **8.1.1**).



***Video*8.1.1*:** *Evidence of a Chemical Reaction*

Example 8.1.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

a. Freezing liquid Mercury
b. Adding yellow to blue to make green
c. Cutting a piece of paper into two pieces
d. Dropping a sliced orange into a vat of sodium dydroxide
e. Filling a balloon with natural air

Solution

a,b,c, & e involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is d

Exercise 8.1.1

Which of the following is a chemical reaction?

a. Painting a wall blue
b. A bicycle rusting
c. Ice cream melting
d. Scratching a key across a desk
e. Making a sand castle

Answer

b

Example 8.1.2: Evidence of a Chemical Reaction

Which of the following is not a chemical reaction?

a. Shattering Glass with a baseball
b. Corroding Metal
c. Fireworks Exploding
d. Lighting a match
e. Baking a cake

Solution

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is a.

Exercise 8.1.2

Which of the following is NOT a chemical reaction?

a. Frying an egg
b. Slicing carrots
c. A Macbook falling out of a window
d. Creating ATP in the human body
e. Dropping a fizzy tablet into a glass of water

Answer

b and c

Summary

Chemical reactions can be identified via a range of different observables including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties among others.

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

8.1: Evidence of a Chemical Reaction is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.2: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that, wheat is produced or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, some baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\text{Flour}+\text{Butter}+\text{Sugar}+\text{Vanilla}+\text{Baking Soda}+\text{Eggs}+\text{Chocolate Chips}}_{\text{Ingredients}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}} \tag{8.2.1}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $SO_2 + O_2$, are reactants and sulfur trioxide, SO_3, is the product.

$$\underbrace{2\,SO_2\,(g)+O_2\,(g)}_{\text{Reactants}} \rightarrow \underbrace{2\,SO_3\,(g)}_{\text{Products}} \tag{8.2.2}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants} \rightarrow \text{Products} \tag{8.2.3}$$

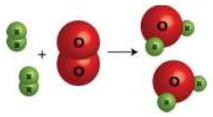
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. You will find there are others but these are the main ones that we need to know. Table **8.2.1** shows a listing of symbols used in chemical equations.

Table 8.2.1: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Desc
+	used to separate multiple reactants or products	(s)	reactant or produ
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or produ
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or prod
$\xrightarrow{Pt}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aque
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g) \tag{8.2.4}$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products and symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$Ca(NO_3)_2\,(aq) + 2NaOH\,(aq) \rightarrow Ca(OH)_2\,(s) + 2NaNO_3\,(aq) \tag{8.2.5}$$

How much easier is that to read? Let's try it in reverse? Look at the following reaction in shorthand and write the word equation for the reaction:

$$Cu\,(s) + AgNO_3\,(aq) \rightarrow Cu(NO_3)_2\,(aq) + Ag\,(s) \tag{8.2.6}$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know what symbols go on each side of the arrow and where the + signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).
3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning they are always found in pairs in nature. They include those elements listed in the table.

Table **8.2.1**: *Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

Example 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
b. Gaseous propane, $\mathrm{C_3H_8}$, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane ($\mathrm{C_3H_8}$) and oxygen ($\mathrm{O_2}$)

Product: carbon dioxide ($\mathrm{CO_2}$) and water ($\mathrm{H_2O}$)

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l) \tag{8.2.7}$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g) \tag{8.2.8}$$

Exercise 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. hydrogen gas reacts with nitrogen gas to produce gaseous ammonia
b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
c. copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a:

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

Answer b:

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

Answer c:

$Cu(s) + O_2(g) \rightarrow CuO(s)$

Summary

- A chemical reaction is the process in which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

8.2: Chemical Equations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.3: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products - they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

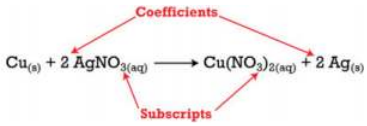
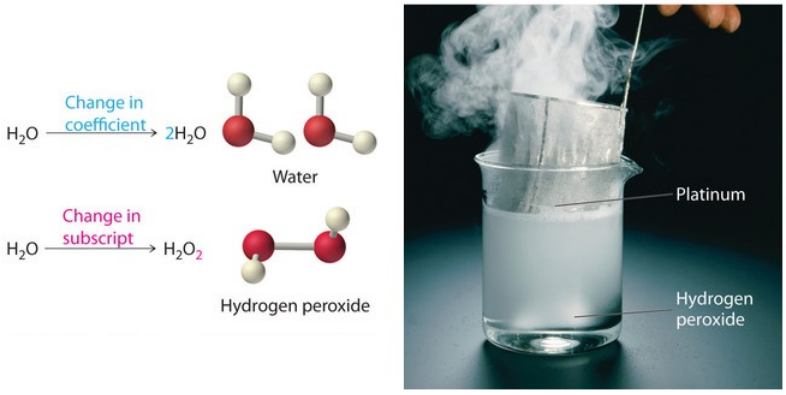


Figure **8.3.1**: Balancing Equations. *You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.*

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure **8.3.1**.



The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

Example 8.3.1: Combustion of Heptane

Balancing the chemical Equation for the combustion of Heptane (C_7H_{16})

$$C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Solution

1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients**.	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $C_7H_{16}\ (l) + O_2\ (g) \rightarrow \underline{7}\ CO_2\ (g) + H_2O\ (g)$ Reactants | Element/Polyatomic Ion | Products 7 | C | 7 b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $C_7H_{16}\ (l) + O_2\ (g) \rightarrow 7\ CO_2\ (g) + \underline{8}H_2O\ (g)$ Reactants | Element/Polyatomic Ion | Products 7 | C | 7 16 | H | 16
3. **Balance polyatomic ions as a unit**.	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms**.	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $C_7H_{16}\ (l) + \underline{11}\ O_2\ (g) \rightarrow 7\ CO_2\ (g) + 8H_2O\ (g)$ Reactants | Element/Polyatomic Ion | Products 7 | C | 7 16 | H | 16 22 | O | 22
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a

chemical equation is balanced.

Example 8.3.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Solution

https://chem.libretexts.org/@go/page/105318

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon

https://chem.libretexts.org/@go/page/105318

dioxide and water.

1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients**.	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $$C_8H_{18}\,(l) + O_2\,(g) \rightarrow \underline{\mathbf{8}}\,CO_2\,(g) + H_2O(g)$$

Reactants	Element/Polyatomic Ion	Products
8	C	8

b. Eighteen hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products:

$$C_8H_{18}\,(l) + O_2\,(g) \rightarrow 8CO_2\,(g) + \underline{\mathbf{9}}\,H_2O(g)$$

Reactants	Element/Polyatomic Ion	Products
8	C	8
18	H	18

3. **Balance polyatomic ions as a unit**.	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms**.	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $$C_8H_{18}\,(l) + \frac{25}{2}O_2\,(g) \rightarrow 8CO_2\,(g) + 9H_2O(g)$$

Reactants	Element/Polyatomic Ion	Products
8	C	8
18	H	18
25	O	25

The equation is now balanced, but we usually write equations with whole-number coefficients. We can eliminate the fractional

	coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $$\underline{\mathbf{2}}C_8H_{18}(l) + \underline{\mathbf{25}}O_2(g) \rightarrow \underline{\mathbf{16}}CO_2(g) + \underline{\mathbf{18}}H_2O(g)$$
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Reactants	Element/Polyatomic Ion	Products
16	C	16
36	H	36
50	O	50

Example 8.3.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$
2. **Adjust the coefficients**.	There are twice as many chloride ions in the reactants than in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $Pb(NO_3)_2(aq) + \underline{2}NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	1
2	Cl	2

3. **Balance polyatomic ions as a unit**.	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow \underline{2}NaNO_3(aq) + PbCl_2(s)$

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	2
2	Cl	2
2	NO_3^-	2

4. **Balance the remaining atoms**.	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow 2NaNO_3(aq) + PbCl_2(s)$

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	2
2	Cl	2
2	NO_3^-	2

Exercise 8.3.1

Is each chemical equation balanced?

a. $2Hg(\ell) + O_2(g) \rightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
c. $Mg(NO_3)_2(s) + 2Li(s) \rightarrow Mg(s) + 2LiNO_3(s)$

Answer a:

yes

Answer b:

no

Answer c:

yes

Exercise 8.3.2

Balance the following chemical equations.

a. $N_2\,(g) + O_2\,(g) \rightarrow NO_2\,(g)$
b. $Pb(NO_3)_2(aq) + FeCl_3(aq) \rightarrow Fe(NO_3)_3(aq) + PbCl_2(s)$
c. $C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Answer a:

$N_2\,(g) + 2O_2\,(g) \rightarrow 2NO_2\,(g)$

Answer b:

$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

Answer c:

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- Chemical reaction: The process in which one or more substances are changed into one or more new substances.
- Reactants: The starting materials in a reaction.
- Products: Materials present at the end of a reaction.
- Balanced chemical equation: A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- Subscripts: Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- Coefficient: A small whole number that appears in front of a formula in a balanced chemical equation.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

8.3: How to Write Balanced Chemical Equations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.4: Classifying Chemical Reactions

Learning Objectives

- Classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- Predict the products of simple reactions.

The chemical reactions we have described are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Most chemical reactions can be classified into one or more of five basic types: acid–base reactions, exchange reactions, condensation reactions (and the reverse, cleavage reactions), and oxidation–reduction reactions. The general forms of these five kinds of reactions are summarized in Table **8.4.1**, along with examples of each. It is important to note, however, that many reactions can be assigned to more than one classification, as you will see in our discussion.

Table 8.4.1*: Basic Types of Chemical Reactions*

Name of Reaction	General Form	Example(s)
Oxidation–Reduction (redox)	oxidant + reductant → reduced oxidant + oxidized reductant	$C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$
Acid–Base	acid + base → salt	$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
Exchange	AB + C → AC + B	$ZnCl_2(aq) + Mg(s) \rightarrow MgCl_2(aq) + Zn(s)$
	Double Replacement AB + CD → AD + CB	$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
Combination (Synthesis)	A + B → AB	$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
		$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$
Decomposition	AB → A + B	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

The classification scheme is only for convenience; the same reaction can be classified in different ways, depending on which of its characteristics is most important. Oxidation–reduction reactions, in which there is a net transfer of electrons from one atom to another, and condensation reactions are discussed in this section. Acid–base reactions and one kind of exchange reaction—the formation of an insoluble salt such as barium sulfate when solutions of two soluble salts are mixed together.

Combination Reactions

A **combination reaction** is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$$A + B \rightarrow AB \tag{8.4.1}$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

$$2Na\,(s) + Cl_2\,(g) \rightarrow 2NaCl\,(s) \tag{8.4.2}$$

Notice that in order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2).

One sort of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide.

$$2\text{Mg}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{MgO}\,(s) \tag{8.4.3}$$

Sulfur reacts with oxygen to form sulfur dioxide.

$$\text{S}\,(s) + \text{O}_2\,(g) \rightarrow \text{SO}_2\,(g) \tag{8.4.4}$$

When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide.

$$2\text{S}\,(s) + 3\text{O}_2\,(g) \rightarrow 2\text{SO}_3\,(g) \tag{8.4.5}$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide.

$$2\text{Fe}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{FeO}\,(s) \tag{8.4.6}$$

$$4\text{Fe}\,(s) + 3\text{O}_2\,(g) \rightarrow 2\text{Fe}_2\text{O}_3\,(s) \tag{8.4.7}$$

Example 8.4.1: Combustion of Solid Potassium

Potassium is a very reactive alkali metal that must be stored under oil in order to prevent it from reacting with air. Write the balanced chemical equation for the combination reaction of potassium with oxygen.

Solution

	Example 8.4.1
Plan the problem.	Make sure formulas of all reactants and products are correct before balancing the equation. Oxygen gas is a diatomic molecule. Potassium oxide is an ionic compound and so its formula is constructed by the crisscross method. Potassium as an ion becomes K^+, while the oxide ion is O^{2-}.
Solve.	The skeleton (unbalanced) equation: $\text{K}\,(s) + \text{O}_2\,(g) \rightarrow \text{K}_2\text{O}\,(s)$ The equation is then easily balanced with coefficients. $4\text{K}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{K}_2\text{O}\,(s)$
Think about your result.	Formulas are correct and the resulting combination reaction is balanced.

Combination reactions can also take place when an element reacts with a compound to form a new compound composed of a larger number of atoms. Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation:

$$2\text{CO}\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g) \tag{8.4.8}$$

Two compounds may also react to form a more complex compound. A very common example is the reactions of oxides with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide.

$$\text{CaO}\,(s) + \text{H}_2\text{O}\,(l) \rightarrow \text{Ca(OH)}_2\,(aq) \tag{8.4.9}$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$SO_3\,(g) + H_2O\,(l) \rightarrow H_2SO_4\,(aq) \tag{8.4.10}$$



Figure 8.4.1: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain.

Exercise 8.4.1

a. Write the chemical equation for the synthesis of silver bromide, $AgBr$.

b. Predict the products for the following reaction: $CO_2\,(g) + H_2O\,(l)$

Answer a:

$2Ag + Br_2 \rightarrow 2AgBr$

Answer b:

$CO_2\,(g) + H_2O\,(l) \rightarrow H_2CO_3$

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$AB \rightarrow A + B \tag{8.4.11}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas.

$$2HgO\,(s) \rightarrow 2Hg\,(l) + O_2\,(g) \tag{8.4.12}$$



Figure 8.4.2: *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.*

A reaction is also considered to be a decomposition reaction even when one or more of the produces are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide.

$$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g) \tag{8.4.13}$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water.

$$2NaOH\,(s) \rightarrow Na_2O\,(s) + H_2O\,(g) \tag{8.4.14}$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water.

$$H_2CO_3\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l) \tag{8.4.15}$$

Example 8.4.2: Electrolysis of Water

When an electric current is passed through pure water, it decomposes into its elements. Write a balanced equation for the decomposition of water.

Solution

	Example 8.4.2
Plan the problem.	Water is a binary compound composed of hydrogen and oxygen. The hydrogen and oxygen gases produced in the reaction are both diatomic molecules.
Solve.	The skeleton (unbalanced) equation: $H_2O\,(l) \xrightarrow{elec} H_2\,(g) + O_2\,(g)$ Note the abbreviation "**elec**" above the arrow to indicate the passage of an electric current to initiate the reaction. Balance the equation. $2H_2O\,(l) \xrightarrow{elec} 2H_2\,(g) + O_2\,(g)$
Think about your result.	The products are elements and the equation is balanced.

Exercise 8.4.2

Write the chemical equation for the decomposition of:

a. Al_2O_3
b. Ag_2S

Answer a:

$2Al_2O_3 \rightarrow 4Al + 3O_2$

Answer b:

$Ag_2S \rightarrow 2Ag + S$

Single Replacement Reactions

A third type of reaction is the single replacement reaction in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$A + BC \rightarrow AC + B \tag{8.4.16}$$

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:

$$Y + XZ \rightarrow XY + Z \tag{8.4.17}$$

where **Y** is a nonmetal and replaces the nonmetal **Z** in the compound with **X**.

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\mathrm{Mg}\,(s) + \mathrm{Cu(NO_3)_2}\,(aq) \rightarrow \mathrm{Mg(NO_3)_2}\,(aq) + \mathrm{Cu}\,(s) \tag{8.4.18}$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

Hydrogen Replacement

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$\mathrm{Zn}\,(s) + 2\mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\,(aq) + \mathrm{H_2}\,(g) \tag{8.4.19}$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal. Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All Group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2\mathrm{Na}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{NaOH}\,(aq) + \mathrm{H_2}\,(g) \tag{8.4.20}$$



Figure **8.4.2**: *(left) Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (right) Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite.*

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine.

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l) \tag{8.4.21}$$

The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

Example 8.4.3

What would be the products of the reaction between solid aluminum (**Al**) and iron (III) oxide ($\mathbf{Fe_2O_3}$)?

Solution

Explanation	Example 8.4.3
Plan the problem	To predict the products, we need to know that aluminum will replace iron and form aluminum oxide (the metal will replace the metal ion in the compound). Aluminum has a charge of $+3$ and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be $\mathbf{Al_2O_3}$. Since iron is replaced in the compound by aluminum, the iron will now be the single element in the products.
Solve	The unbalanced equation will be: $$\mathbf{Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe}$$ and the balanced equation will be: $$\mathbf{2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe}$$
Think about your result	This is a single replacement reaction, and when balanced the coefficients accurately reflect that the iron and aluminum have the same charge in this reaction.

Exercise 8.4.3

a. Write the chemical equation for the single replacement reaction between zinc solid and lead (II) nitrate solution to produce zinc nitrate solution and solid lead. (*Note: zinc forms ions with a $+2$ charge)
b. Predict the products for the following reaction: $\mathbf{Fe + CuSO_4}$ (in this reaction, assume iron forms ions with a $+2$ charge)

Answer a:

$$\mathbf{Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2}$$

Answer b:

$$\mathbf{Fe + CuSO_4 \rightarrow Cu + FeSO_4}$$

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:

$$\mathbf{AB + CD \rightarrow AD + BC} \tag{8.4.22}$$

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.

$$\mathbf{2KI\,(\mathit{aq}) + Pb(NO_3)_2\,(\mathit{aq}) \rightarrow 2KNO_3\,(\mathit{aq}) + PbI_2\,(\mathit{s})} \tag{8.4.23}$$

There are very strong attractive forces that occur between Pb^{2+} and I^- ions and the result is a brilliant yellow precipitate (Figure 8.4.3). The other product of the reaction, potassium nitrate, remains soluble.



Figure 8.4.3: Lead (II) iodide precipitates when potassium iodide is mixed with lead (II) nitrate (Equation 8.4.23). (CC BY-SA 3.0; PRHaney).

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas.

$$Na_2S\,(aq) + 2HCl\,(aq) \rightarrow 2NaCl\,(aq) + H_2S\,(g) \tag{8.4.24}$$

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.

$$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l) \tag{8.4.25}$$

Example 8.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $NaCN\,(aq) + HBr\,(aq) \rightarrow$ (hydrogen cyanide gas is formed)

Solution

	Example 8.4.4
Plan the problem	The production of a gas drives the reaction
Solve	The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partners, the balanced equation is: $$NaCN\,(aq) + HBr\,(aq) \rightarrow NaBr\,(aq) + HCN\,(g) \tag{8.4.26}$$
Think about your result	This is a double replacement reaction. All formulas are correct and the equation is balanced.

Exercise 8.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $(NH_4)_2SO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow$ (a precipitate of barium sulfate forms)

Answer a:

$$(NH_4)_2SO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow 2NH_4NO_3\,(aq) + BaSO_4\,(s) \tag{8.4.27}$$

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.

$$\mathrm{Na_2CO_3\,(aq) + 2HCl\,(aq) \rightarrow 2\ NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)} \tag{8.4.28}$$

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

8.4: Classifying Chemical Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

12: Quantities in Chemical Reactions

Template:HideTOC

So far, we have talked about chemical reactions in terms of individual atoms and molecules. Although this works, most of the reactions occurring around us involve much larger amounts of chemicals. Even a tiny sample of a substance will contain millions, billions, or a hundred billion billions of atoms and molecules. How do we compare amounts of substances to each other in chemical terms when it is so difficult to count to a hundred billion billion? Actually, there are ways to do this, which we will explore in this chapter. In doing so, we will increase our understanding of stoichiometry, which is the study of the numerical relationships between the reactants and the products in a balanced chemical reaction.

Topic hierarchy

12.1: Climate Change - Too Much Carbon Dioxide

12.2: Stoichiometry

12.3: Mole-to-Mole Conversions

12.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions

12.5: Limiting Reactant and Theoretical Yield

12.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants

12.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

12: Quantities in Chemical Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.1: Climate Change - Too Much Carbon Dioxide

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. Figure **12.1.1** shows CO_2 levels during the last three glacial cycles, as reconstructed from ice cores.

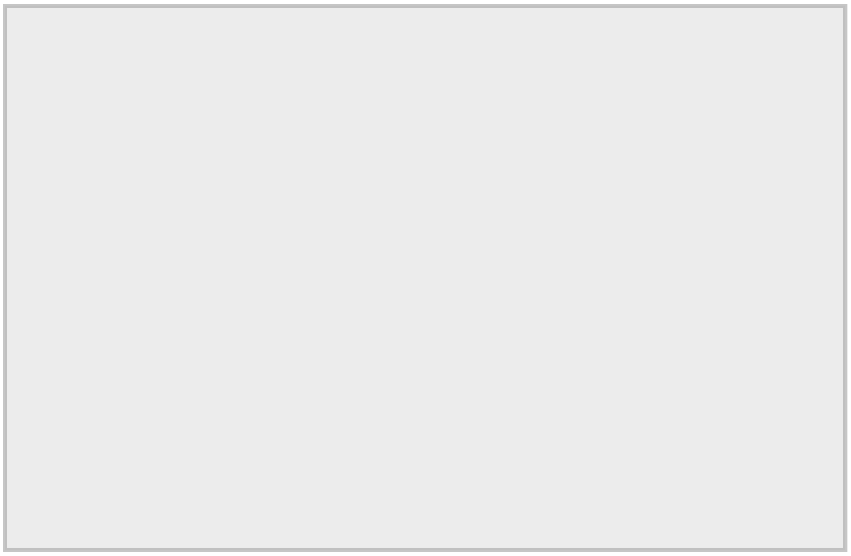


Figure **12.1.1**: PROXY (INDIRECT) MEASUREMENTS. Data source: Reconstruction from ice cores. Credit: NOAA

Carbon Dixode

Carbon dioxide ($\mathbf{CO_2}$) is the primary greenhouse gas emitted through human activities. In 2015, $\mathbf{CO_2}$ 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 $\mathbf{CO_2}$ to the atmosphere and by influencing the ability of natural sinks, like forests, to remove $\mathbf{CO_2}$ from the atmosphere. While $\mathbf{CO_2}$ 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 $\mathbf{CO_2}$ 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 $\mathbf{CO_2}$. An example of how $\mathbf{CO_2}$ can be generate consider the combustion of octane, a component of gasoline

$$\mathbf{2\,C_8H_{10}\,(l) + 25\,O_2\,(g) \rightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)} \qquad (12.1.1)$$

The balanced reaction in Equation 12.1.1 argues that for every two molecules of octane that are burned, 16 molecules of $\mathbf{CO_2}$ are generated.

Contributors and Attributions

- Earth Science Communications Team at NASA's Jet Propulsion Laboratory, California Institute of Technology
- EPA

12.1: Climate Change - Too Much Carbon Dioxide is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.2: Stoichiometry

Learning Objectives

- Explain the meaning of the term "stoichiometry".
- Determine the relative amounts of each substance in chemical equations.

You have learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this section you will explore the quantitative relationships that exist between the quantities of reactants and products in a balanced equation. This is known as stoichiometry.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry is actually Greek from two words: *στοικηιον*, which means "element", and \(\mu \epsilon \tau \rho \omicron \nu), which means "measure".

Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

$$\begin{array}{ccccccccc} 2\ CuSO_4 & + & 4\ KI & \rightarrow & 2\ CuI & + & 2\ K_2SO_4 & + & I_2 \\ \text{2 formula units } CuSO_4 & + & \text{4 formula units } KI & \rightarrow & \text{2 formula units } CuI & + & \text{2 formula units } K_2SO_4 & + & \text{1 molecule } I_2 \\ \text{2 moles } CuSO_4 & + & \text{4 moles } KI & \rightarrow & \text{2 moles } CuI & + & \text{2 moles } K_2SO_4 & + & \text{1 moles } I_2 \end{array}$$

The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 2 units of copper (II) sulfate ($CuSO_4$) we have, we need to have 4 units of potassium iodide (KI). For every two dozen copper (II) sulfates, we need 4 dozen potassium iodides. Because the unit "mole" is also a counting unit, we can interpret this equation in terms of moles, as well: For every two moles of copper (II) sulfate, we need 4 moles potassium iodide.

The production of ammonia (NH_3) from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g) \tag{12.2.1}$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

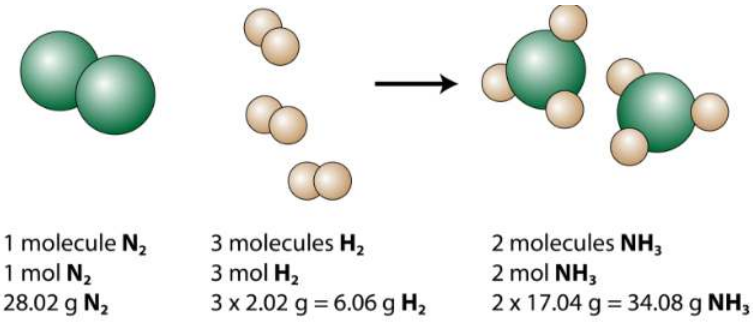


Figure 12.2.1: *This representation of the production of ammonia from nitrogen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.*

We see that 1 molecule of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia. This is the smallest possible relative amounts of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

The most useful quantity for counting particles is the mole. So if each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation.

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. **1 mol** of nitrogen has a mass of **28.02 g**, while **3 mol** of hydrogen has a mass of **6.06 g**, and **2 mol** of ammonia has a mass of **34.08 g**.

$$28.02\ \text{g}\ N_2 + 6.06\ \text{g}\ H_2 \rightarrow 34.08\ \text{g}\ NH_3 \tag{12.2.2}$$

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.

Example 12.2.1

For the combustion of ethane (\ce{C_2H_6}) equation below, i

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O \tag{12.2.3}$$

a. Indicate the number of formula units or molecules in the balanced equation.
b. Indicate the number of moles present in the balanced equation.

Solution

a. Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.
b. Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.

Exercise 12.2.1

For the following equation below, indicate the number of formula units or molecules, and the number of moles present in the balanced equation.

$$KBrO_3 + 6KI + 6HBr \rightarrow 7KBr + 3H_2O \tag{12.2.4}$$

Answer

One molecules of $KBrO_3$ plus six molecules of **KI** plus six molecules of **HBr** yields seven molecules of **KBr** plus three molecules of I_2 and three molecules of H_2O. One mole of $KBrO_3$ plus six moles of **KI** plus six moles of **HBr** yields seven moles of **KBr** plus three moles of I_2 plus three moles of H_2O.

Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- The coefficients of the balanced equation can be used to determine the ratio of moles of all substances in a reaction.

Vocabulary

- Stoichiometry: The calculation of quantitative relationships of the reactants and products in a balanced chemical equation.
- Formula unit: The empirical formula of an ionic compound.
- Mole ratio: The ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

12.2: Stoichiometry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.3: Mole-to-Mole Conversions

Learning Objectives

- Use a balanced chemical equation to determine molar relationships between 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

$$2H_2 + O_2 \rightarrow 2H_2O \tag{12.3.1}$$

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

$$4H_2 + 2O_2 \rightarrow 4H_2O \tag{12.3.2}$$

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

$$22H_2 + 11O_2 \rightarrow 22H_2O \tag{12.3.3}$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O \tag{12.3.4}$$

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

$$2\ mol\ H_2 + 1\ mol\ O_2 \rightarrow 2\ mol\ H_2O \tag{12.3.5}$$

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

$$2H_2 + O_2 \rightarrow 2H_2O \tag{12.3.6}$$

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.”

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:

$$\frac{2\ \text{mol H}_2}{1\ \text{mol O}_2} \text{ or } \frac{1\ \text{mol O}_2}{2\ \text{mol H}_2} \tag{12.3.7}$$

$$\frac{2\ \text{mol H}_2\text{O}}{1\ \text{mol O}_2} \text{ or } \frac{1\ \text{mol O}_2}{2\ \text{mol H}_2\text{O}} \tag{12.3.8}$$

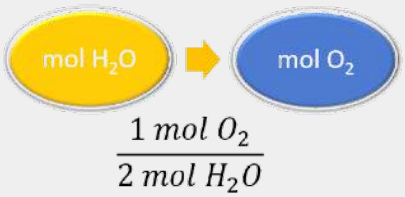
$$\frac{2\ \text{mol H}_2}{2\ \text{mol H}_2\text{O}} \text{ or } \frac{2\ \text{mol H}_2\text{O}}{2\ \text{mol H}_2} \tag{12.3.9}$$

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*.

Example 12.3.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?

Solution

Steps for Problem Solving	How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?
Find a balanced equation that describes the reaction	Unbalanced: $H_2 + O_2 \rightarrow H_2O$ Balanced: $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$
Identify the "given"information and what the problem is asking you to "find."	Given: moles H_2O Find: moles oxygen
List other known quantities	1 mol O_2 = 2 mol H_2O
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ O_2}{2\ mol\ H_2O}$
Cancel units and calculate.	$\cancel{27.6\ mol\ H_2O} \times \frac{1\ mol\ O_2}{\cancel{2\ mol\ H_2O}} = 13.8\ mol\ O_2$ To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

Example 12.3.2

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen
Find a balanced equation that describes the reaction	Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Balanced: $N_2 + \underline{3}H_2 \rightarrow \underline{2}NH_3$
Identify the "given"information and what the problem is asking you to "find."	Given: $\mathbf{H_2 = 4.20\ mol}$ Find: $\mathbf{mol}$ of $\mathbf{NH_3}$
List other known quantities	3 mol H_2 = 2 mol NH_3
Prepare a concept map and use the proper conversion factor.	 $\frac{2\ mol\ NH_3}{3\ mol\ H_2}$
Cancel units and calculate.	$\cancel{4.20\ mol\ H_2} \times \frac{2\ mol\ NH_3}{\cancel{3\ mol\ H_2}} = 2.80\ mol\ NH_3$ The reaction of $\mathbf{4.20\ mol}$ of hydrogen with excess nitrogen produces $\mathbf{2.80\ mol}$ of ammonia
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

Exercise 12.3.1

a. Given the following balanced chemical equation,

$$\mathbf{C_5H_{12} + 8\,O_2 \longrightarrow 5\,CO_2 + 6\,H_2O} \tag{12.3.10}$$

how many moles of H_2O can be formed if 0.0652 mol of C_5H_{12} were to react?

b. Balance the following unbalanced equation and determine how many moles of H_2O are produced when 1.65 mol of NH_3 react.

$$\mathbf{NH_3 + O_2 \longrightarrow N_2 + H_2O} \tag{12.3.11}$$

Answer a:

3.14 mol H_2O

Answer b:

$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol H_2O

Summary

- The balanced chemical reaction can be used to determine molar relationships between substances.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

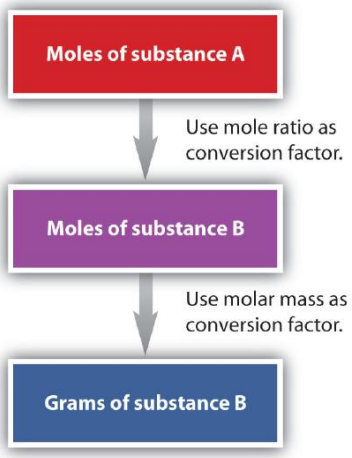
12.3: Mole-to-Mole Conversions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions

Learning Objectives

- 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

$$Fe_2O_3 + 3SO_3 \rightarrow Fe_2(SO_4)_3 \tag{12.4.1}$$

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:

$$3.59 \ \cancel{\text{mol Fe}_2\text{O}_3} \times \left(\frac{3 \text{ mol SO}_3}{1 \ \cancel{\text{mol Fe}_2\text{O}_3}} \right) = 10.77 \text{ mol SO}_3 \tag{12.4.2}$$

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:

$$10.77 \ \cancel{\text{mol SO}_3} \times \left(\frac{80.06 \text{ g SO}_3}{1 \ \cancel{\text{mol SO}_3}} \right) = 862 \text{ g SO}_3 \tag{12.4.3}$$

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:

$$3.59\ \cancel{mol\ Fe_2O_3} \times \underbrace{\left(\frac{3\ \cancel{mol\ SO_3}}{1\ \cancel{mol\ Fe_2O_3}}\right)}_{\text{converts to moles of } SO_3} \times \underbrace{\left(\frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}}\right)}_{\text{converts to grams of } SO_3} = 862\ g\ SO_3 \tag{12.4.4}$$

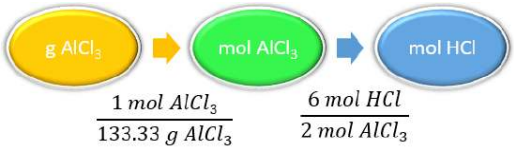
We get exactly the same answer when combining all the math steps together as we do when we calculate one step at a time.

Example 12.4.1: Generation of Aluminum Oxide

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

Steps for Problem Solving	Example 12.4.1
Identify the "given"information and what the problem is asking you to "find."	Given: 249 g $AlCl_3$ Find: moles HCl
List other known quantities	1 mol $AlCl_3$ = 133.33 g/mol 6 mol of HCl to 2 mol $AlCl_3$
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$249\ \cancel{g\ AlCl_3} \times \frac{1\ \cancel{mol\ AlCl_3}}{133.33\ \cancel{g\ AlCl_3}} \times \frac{6\ mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.60\ mol\ HCl$
Think about your result.	Since 249 g of $AlCl_3$ is less than 266.66 g, the mass for 2 moles of $AlCl_3$ and the relationship is 6 mol of HCl to 2 mol $AlCl_3$, the answer should be less than 6 moles of HCl.

Exercise 12.4.1: Generation of Aluminum Oxide

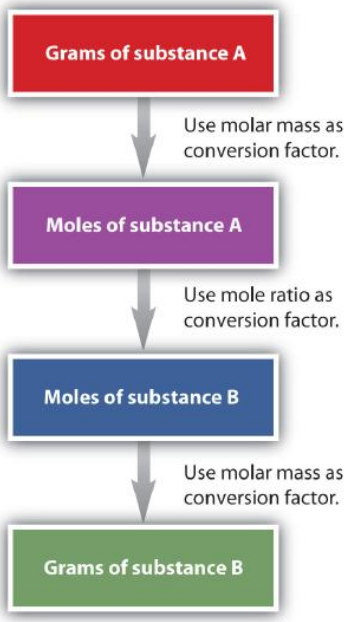
How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer

0.442 mol Al_2O_3

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:



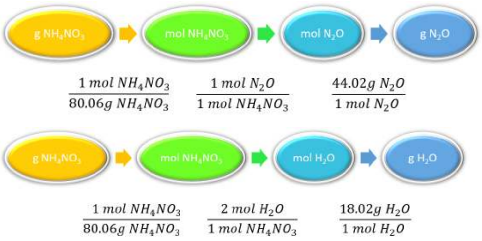
This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

Example 12.4.2: Decomposition of Ammonium Nitrate

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$\mathrm{NH_4NO_3}\,(s) \rightarrow \mathrm{N_2O}\,(g) + 2\mathrm{H_2O}\,(l)$$

In a certain experiment, $45.7\ \mathrm{g}$ of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Steps for Problem Solving	Example 12.4.2
Identify the "given"information and what the problem is asking you to "find."	Given: $45.7\ g\ NH_4NO_3$ Find: Mass N_2O =? g Mass H_2O =? g
List other known quantities	1 mol $NH_4NO_3 = 80.06\ g/mol$ 1 mol $N_2O = 44.02\ g/mol$ 1 mol $H_2O = 18.02\ g/mol$ 1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O
Prepare two concept maps and use the proper conversion factor.	
Cancel units and calculate.	$45.7\ g\ NH_4NO_3 \times \frac{1\ mol\ NH_4NO_3}{80.06\ g\ NH_4NO_3} \times \frac{1\ mol\ N_2O}{1\ mol\ NH_4NO_3} \times \frac{44.02\ g\ N_2O}{1\ mol\ N_2O} = 25.1\ g\ N_2O$ $45.7\ g\ NH_4NO_3 \times \frac{1\ mol\ NH_4NO_3}{80.06\ g\ NH_4NO_3} \times \frac{2\ H_2O}{1\ mol\ NH_4NO_3} \times \frac{18.02\ g\ H_2O}{1\ mol\ H_2O} = 20.6\ g\ H_2O$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

Exercise 12.4.2: Carbon Tetrachloride

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows.

$$CH_4(g) + 4\,Cl_2(g) \longrightarrow CCl_2(l) + 4\,HCl(l) \qquad (12.4.5)$$

How many grams of HCl are produced by the reaction of 100.0g of CH_4?

Answer

908.7g HCl

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

12.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.5: Limiting Reactant and Theoretical Yield

Learning Objectives

- Identify the limiting reactant (limiting reagent) in a given chemical reaction.
- Calculate how much product will be produced from the limiting reagent.
- Calculate how much reactant(s) remains when the reaction is complete.

In all the examples discussed thus far, the reactants were assumed to be present in stoichiometric quantities. Consequently, none of the reactants was left over at the end of the reaction. This is often desirable, as in the case of a space shuttle, where excess oxygen or hydrogen was not only extra freight to be hauled into orbit but also an explosion hazard. More often, however, reactants are present in mole ratios that are not the same as the ratio of the coefficients in the balanced chemical equation. As a result, one or more of them will not be used up completely but will be left over when the reaction is completed. In this situation, the amount of product that can be obtained is limited by the amount of only one of the reactants. The reactant that restricts the amount of product obtained is called the limiting reactant. The reactant that remains after a reaction has gone to completion is in excess.

Consider a nonchemical example. Assume you have invited some friends for dinner and want to bake brownies for dessert. You find two boxes of brownie mix in your pantry and see that each package requires two eggs. The balanced equation for brownie preparation is thus

$$\text{1 box mix} + \text{2 eggs} \rightarrow \text{1 batch brownies} \tag{12.5.1}$$

If you have a dozen eggs, which ingredient will determine the number of batches of brownies that you can prepare? Because each box of brownie mix requires two eggs and you have two boxes, you need four eggs. Twelve eggs is eight more eggs than you need. Although the ratio of eggs to boxes in is 2:1, the ratio in your possession is 6:1. Hence the eggs are the ingredient (reactant) present in excess, and the brownie mix is the limiting reactant. Even if you had a refrigerator full of eggs, you could make only two batches of brownies.



Figure **12.5.1**: The Concept of a Limiting Reactant in the Preparation of Brownies. For a chemist, the **balanced** chemical equation is the recipe that must be followed.

Phet: Reactants, products and leftovers

View this interactive simulation illustrating the concepts of limiting and excess reactants.

PhET: Reactants, Products and Leftovers

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$\mathrm{H_2} + \mathrm{Cl_2}(g) \rightarrow 2\,\mathrm{HCl}(g) \tag{12.5.2}$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen nonreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 3\ \text{mol}\ \mathrm{H_2} \times \frac{2\ \text{mol HCl}}{1\ \text{mol}\ \mathrm{H_2}} = 6\ \text{mol HCl} \tag{12.5.3}$$

Complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 2\ \text{mol}\ \mathrm{Cl_2} \times \frac{2\ \text{mol HCl}}{1\ \text{mol}\ \mathrm{Cl_2}} = 4\ \text{mol HCl} \tag{12.5.4}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure **12.5.2**).

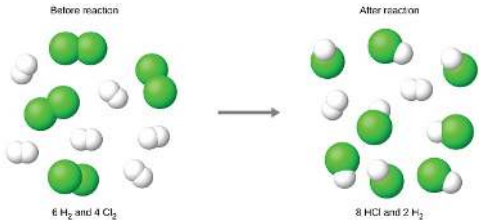


Figure **12.5.2**: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reactant; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reactant and which is in excess.

How to Identify the Limiting Reactant (Limiting Reagent)

There are two ways to determine the limiting reagent. One method is to find and compare the mole ratio of the reactants used in the reaction (Approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the **smallest** amount of product is the limiting reagent (Approach 2). This section will focus more on the second method.

Approach 1 (The "Reactant Mole Ratio Method"): Find the limiting reagent by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting reactant.

Approach 2 (The "The Product Method"): Find the limiting reagent by calculating and comparing the amount of product each reactant will produce.

1. Balance the chemical equation for the chemical reaction.
2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reactant.
5. The reactant that produces a larger amount of product is the excess reactant.
6. To find the amount of remaining excess reactant, subtract the mass of excess reactant consumed from the total mass of excess reagent given.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

Example 12.5.1: Identifying the Limiting Reactant

As an example consider the balanced equation

$$4C_2H_3Br_3 + 11O_2 \rightarrow 8CO_2 + 6H_2O + 6Br_2 \tag{12.5.5}$$

What is the limiting reactant if 76.4 grams of $C_2H_3Br_3$ reacted with 49.1 grams of O_2?

Solution

Using Approach 1:

Step 1: ***Balance the chemical equation***

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br

Step 2: Convert all given information into moles.

$$76.4\ \cancel{\text{g C}_2\text{H}_3\text{B}_3} \times \frac{1\ \text{mol C}_2\text{H}_3\text{B}_3}{266.72\ \cancel{\text{g C}_2\text{H}_3\text{B}_3}} = 0.286\ \text{mol C}_2\text{H}_3\text{Br}_3$$

$$49.1\ \cancel{\text{g O}_2} \times \frac{1\ \text{mol O}_2}{32.00\ \cancel{\text{g O}_2}} = 1.53\ \text{mol O}_2$$

Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.

Assuming that all of the oxygen is used up,

$$1.53\ \cancel{\text{mol O}_2} \times \frac{4\ \text{mol C}_2\text{H}_3\text{Br}_3}{11\ \cancel{\text{mol O}_2}} = 0.556\ \text{mol } C_2H_3Br_3 \text{ are required.}$$

Because 0.556 moles of $C_2H_3Br_3$ required > 0.286 moles of $C_2H_3Br_3$ available, $C_2H_3Br_3$ is the limiting reactant.

Using Approach 2:

Step 1: ***Balance the chemical equation***

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br

Step 2 and Step 3: Converting mass to moles and stoichiometry

$$76.4\ \cancel{\text{g C}_2\text{H}_3\text{Br}_3} \times \frac{1\ \cancel{\text{mol C}_2\text{H}_3\text{Br}_3}}{266.72\ \cancel{\text{g C}_2\text{H}_3\text{Br}_3}} \times \frac{8\ \cancel{\text{mol CO}_2}}{4\ \cancel{\text{mol C}_2\text{H}_3\text{Br}_3}} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}} = 25.2\ \text{g CO}_2$$

$$49.1\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{8\ \cancel{\text{mol CO}_2}}{11\ \cancel{\text{mol O}_2}} \times \frac{44.01\ \text{g CO}_2}{1\ \cancel{\text{mol CO}_2}} = 49.1\ \text{g CO}_2$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant

Therefore, by either method, $C_2H_3Br_3$ is the limiting reactant.

Example 12.5.2: Identifying the Limiting Reactant and the Mass of Excess Reactant

For example, in the reaction of magnesium metal and oxygen, calculate the mass of magnesium oxide that can be produced if 2.40 g Mg reacts with 10.0 g O_2. Also determine the amount of excess reactant. **MgO** is the only product in the reaction.

Solution

Following Approach 1:

Step 1: Balance the chemical equation

$$\underline{2}\ \text{Mg (s)} + \text{O}_2\ \text{(g)} \rightarrow \underline{2}\ \text{MgO (s)}$$

The balanced equation provides the relationship of 2 mol Mg to 1 mol O_2 to 2 mol MgO

Step 2 and Step 3: Converting mass to moles and stoichiometry

$$2.40\ \cancel{\text{g Mg}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \cancel{\text{mol MgO}}}{2\ \cancel{\text{mol Mg}}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = 3.98\ \text{g MgO}$$

$$10.0\ \cancel{g\ O_2} \times \frac{1\ \cancel{mol\ O_2}}{32.00\ \cancel{g\ O_2}} \times \frac{2\ \cancel{mol\ MgO}}{1\ \cancel{mol\ O_2}} \times \frac{40.31\ g\ MgO}{1\ \cancel{mol\ MgO}} = 25.2\ g\ MgO$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant

Mg produces less MgO than does O_2 (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reactant in this reaction.

Step 5: The reactant that produces a larger amount of product is the excess *reactant*

O_2 produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore O_2 is the excess reagent in this reaction.

Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reactant consumed from the total mass of excess reactant given.

Mass of excess reagent calculated using the limiting reactant:

$$2.40\ \cancel{g\ Mg} \times \frac{1\ \cancel{mol\ Mg}}{24.31\ \cancel{g\ Mg}} \times \frac{1\ \cancel{mol\ O_2}}{2\ \cancel{mol\ Mg}} \times \frac{32.00\ g\ O_2}{1\ \cancel{mol\ O_2}} = 1.58\ g\ O_2$$

OR

Mass of excess reactant calculated using the mass of the product:

$$3.98\ \cancel{g\ MgO} \times \frac{1\ \cancel{mol\ MgO}}{40.31\ \cancel{g\ MgO}} \times \frac{1\ \cancel{mol\ O_2}}{2\ \cancel{mol\ MgO}} \times \frac{32.0\ g\ O_2}{1\ \cancel{mol\ O_2}} = 1.58\ g\ O_2$$

Mass of total excess reactant given – mass of excess reactant consumed in the reaction

10.0g O_2 - (available) 1.58g O_2 (used) = 8.42g O_2 (excess)

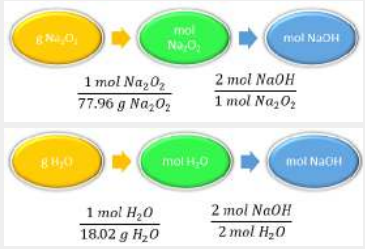
Therefore, O_2 is in excess.

Example 12.5.3: Limiting Reactant

What is the limiting reactant if 78.0 grams of Na_2O_2 were reacted with 29.4 grams of H_2O? The unbalanced chemical equation is

$$Na_2O_2\,(s) + H_2O\,(l) \longrightarrow NaOH\,(aq) + H_2O_2\,(l)$$

Solution

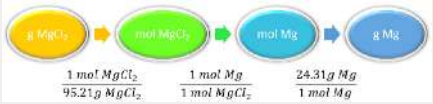
Steps for Problem Solving-The Product Method	**Example 12.5.1**
Identify the "given"information and what the problem is asking you to "find."	Given: 78.0 grams of Na_2O_2 29.4 g H_2O Find: limiting reactant
List other known quantities	1 mol Na_2O_2= 77.96 g/mol 1 mol H_2O = 18.02 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.
Balance the equation	Na_2O_2 (s) + 2H_2O (l) → 2NaOH (aq) + H_2O_2 (l) The balanced equation provides the relationship of 1 mol Na_2O_2 to 2 mol H_2O 2mol NaOH to 1 mol H_2O_2
Prepare a concept map and use the proper conversion factor	 Because the question only asks for the limiting reactant, we can perform two **mass-mole** calculations and determine which amou
Cancel units and calculate.	$$78.0\ g\ Na_2O_2 \times \frac{1\ mol\ Na_2O_2}{77.96\ g\ Na_2O_2} \times \frac{2\ mol\ NaOH}{1\ mol\ Na_2O_2} \times \frac{40\ g\ NaOH}{1\ mol\ NaOH} = 2.00\ mol\ NaOH$$ $$29.4\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mol\ NaOH}{2\ mol\ Na_2O_2} \times \frac{40\ g\ NaOH}{1\ mol\ NaOH} = 1.63\ mol\ NaOH$$ Therefore, H_2O is the *limiting reactant*.
Think about your result.	

Example 12.5.4: Limiting Reactant and Mass of Excess Reactant

A 5.00 g quantity of **Rb** are combined with 3.44 g of **$MgCl_2$** according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of **Mg** is formed, and what mass of what reactant is left over?

Steps for Problem Solving-The Product Method	Example 12.5.2
Steps for Problem Solving	A 5.00 g quantity of Rb is combined with 3.44 g of $MgCl_2$ according to this chemical reaction: $$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$ What mass of Mg is formed, and what mass of what reactant is left over?
Identify the "given"information and what the problem is asking you to "find."	Given: 5.00g Rb, 2.44g MgCl2 Find: mass of Mg formed, mass of remaining reactant
List other known quantities	• molar mass: Rb = 85.47 g/mol • molar mass: $MgCl_2$ = 95.21 g/mol • molar mass: Mg = 24.31 g/mol
Prepare concept maps and use the proper conversion factor.	Find mass Mg formed based on mass of Rb  Find mass of Mg formed based on mass of $MgCl_2$  Use limiting reactant to determine amount of excess reactant consumed 
Cancel units and calculate.	Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which $$5.00\ g\ Rb \times \frac{1\ mol\ Rb}{85.47\ g\ Rb} \times \frac{1\ mol\ Mg}{2\ mol\ Rb} \times \frac{24.31\ g\ Mg}{1\ mol\ Mg} = 0.711\ g\ Mg$$ $$3.44\ g\ MgCl_2 \times \frac{1\ mol\ MgCl_2}{95.21\ g\ MgCl_2} \times \frac{1\ mol\ Mg}{1\ mol\ MgCl_2} \times \frac{24.31\ g\ Mg}{1\ mol\ Mg} = 0.878\ g\ Mg$$ The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reactant. To determine how mu we have to do one more mass-mass calculation to determine what mass of $MgCl_2$ reacted with the 5.00 g of Rb and then subtra original amount. $$5.00\ g\ Rb \times \frac{1\ mol\ Rb}{85.47\ g\ Rb} \times \frac{1\ mol\ MgCl_2}{2\ mol\ Rb} \times \frac{95.21\ g\ MgCl_2}{1\ mol\ MgCl_2} = 2.78\ g\ MgCl_2\ \ reacted$$ Because we started with 3.44 g of $MgCl_2$, we have 3.44 g $MgCl_2$ − 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ left
Think about your result.	It usually is not possible to determine the limiting reagent using just the initial masses, as the reagents have different molar masse

Exercise 12.5.1

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?

$$\underbrace{22.7\ g}_{MgO(s)} + \underbrace{17.9\ g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer

H_2S is the limiting reagent; 1.5 g of MgO are left over.

Contributors and Attributions

- Sarick Shah (UCD)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

12.5: Limiting Reactant and Theoretical Yield is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants

Learning Objectives

- Calculate percentage or actual yields from known amounts of reactants

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

Percent Yield

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than would be predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, the maximum amount of product that could be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage.

$$\textbf{Percent Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \times 100\% \tag{12.6.1}$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

Typically, percent yields are understandably less than **100%** because of the reasons indicated earlier. However, percent yields greater than **100%** are possible if the measured product of the reaction contains impurities that cause its mass to be greater than it actually would be if the product was pure. When a chemist synthesizes a desired chemical, he or she is always careful to purify the products of the reaction. Example **12.6.1** illustrates the steps for determining percent yield.

Example 12.6.1: Decomposition of Potassium chlorate

Potassium chlorate decomposes upon slight heating in the presence of a catalyst according to the reaction below:

$$2KClO_3\,(s) \rightarrow 2KCl\,(s) + 3O_2\,(g)$$

In a certain experiment, **40.0 g $KClO_3$** is heated until it completely decomposes. The experiment is performed and the oxygen gas is collected and its mass is found to be **14.9 g**.

a. What is the theoretical yield of oxygen gas?
b. What is the percent yield for the reaction?

Solution

a. Calculation of theoretical yield

First, we will calculate the theoretical yield based on the stoichiometry.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Mass of **$KClO_3$ = 40.0 g**

Mass of O_2 collected = 14.9g

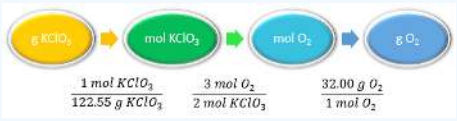
Find: Theoretical yield, g O_2

Step 2: List other known quantities and plan the problem.

1 mol $KClO_3$ = 122.55 g/mol

1 mol O_2 - 32.00 g/mol

Step 3: Apply stoichiometry to convert from the mass of a reactant to the mass of a product:



Step 4: Solve.

$$40.0\ \cancel{g\ KClO_3} \times \frac{1\ \cancel{mol\ KClO_3}}{122.55\ \cancel{g\ KClO_3}} \times \frac{3\ \cancel{mol\ O_2}}{2\ \cancel{mol\ KClO_3}} \times \frac{32.00\ g\ O_2}{1\ \cancel{mol\ O_2}} = 15.7\ g\ O_2$$

The theoretical yield of **O_2** is **15.7 g**, 15.6̲7 g unrounded.

Step 5: Think about your result.

The mass of oxygen gas must be less than the **40.0 g** of potassium chlorate that was decomposed.

b. Calculation of percent yield

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Theoretical yield =15.6̲7 g, use the unrounded number for the calculation.

Actual yield = 14.9g

Find: Percent yield, % Yield

Step 2: List other known quantities and plan the problem.

No other quantities needed

Step 3: Use the percent yield equation below

$$\textbf{Percent Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \times 100\%$$

Step 4: Solve.

$$\text{Percent Yield} = \frac{14.9\ \text{g}}{15.07\ \text{g}} \times 100\% = 94.9\%$$

Step 3: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under **100%**.

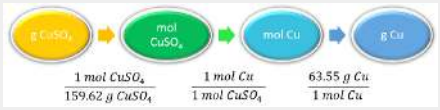
Example 12.6.2: Oxidation of Zinc

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$$

What is the percent yield?

Solution

Steps for Problem Solving-The Product Method	**Example 12.6.1**
Identify the "given"information and what the problem is asking you to "find."	Given: 1.274 g CuSO4 Actual yield = 0.392 g Cu Find: Percent yield
List other known quantities	1 mol CuSO4= 159.62 g/mol 1 mol Cu = 63.55 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.
Balance the equation	The chemical equation is already balanced. The balanced equation provides the relationship of 1 mol CuSO4 to 1 mol Zn to 1 mol Cu to 1 mol ZnSO4
Prepare a concept map and use the proper conversion factor.	 The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield (g Cu) is fou calculation based on the initial amount of $CuSO_4$.
Cancel units and calculate.	Double subscripts: use braces to clarify Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be $\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$ $\text{percent yield} = \left(\frac{0.392\ \text{g Cu}}{0.5072\ \text{g Cu}}\right) \times 100 = 77.3\%$
Think about your result.	Since the actual yield is slightly less than the theoretical yield, the percent yield is just under **100%**.

Exercise 12.6.1

What is the percent yield of a reaction that produces 12.5 g of the Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$CCl_4 + 2\,HF \rightarrow CF_2Cl_2 + 2\,HCl$$

Answer

48.3%

Summary

Theoretical yield is calculated based on the stoichiometry of the chemical equation. The actual yield is experimentally determined. The percent yield is determined by calculating the ratio of actual yield/theoretical yield.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

12.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

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, called 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 is 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.

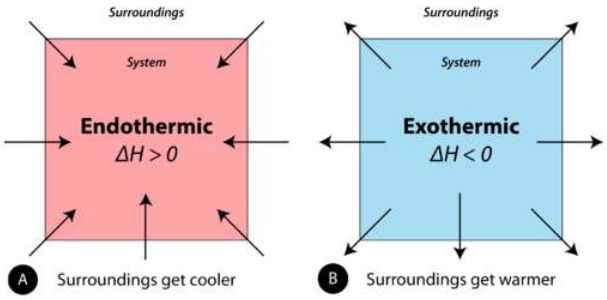


Figure 12.7.1: (A) Endothermic reaction. (B) Exothermic reaction.

Enthalpy

Heat changes in chemical reactions are often measured in the laboratory under conditions in which the reacting system is open to the atmosphere. In that case, the system is at a constant pressure. **Enthalpy** (H) is the heat content of a system at constant pressure. Chemists routinely measure changes in enthalpy of chemical systems as reactants are converted into products. The heat that is absorbed or released by a reaction at constant pressure is the same as the enthalpy change, and is given the symbol ΔH. Unless otherwise specified, all reactions in this material are assumed to take place at constant pressure.

The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products. The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds. Energy needs to be put into the system in order to break chemical bonds - they do not come apart spontaneously in most cases. Bond formation to produce products will involve release of energy. The change in enthalpy shows the trade-offs made in these two processes. Does it take more energy to break bonds than that needed to form bonds? If so, the reaction is endothermic and the enthalpy change is positive. If more energy is produced in bond formation than that needed for bond breaking, the reaction is exothermic and the enthalpy is negative.

Several factors influence the enthalpy of a system. Enthalpy is an extensive property, determined in part by the amount of material we work with. The state of reactants and products (solid, liquid, or gas) influences the enthalpy value for a system. The direction of the reaction affects the enthalpy value. A reaction that takes place in the opposite direction has the same numerical enthalpy value, but the opposite sign.

Thermochemical Equation

When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of $1\ \text{mol}$ of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of the balanced equation.

$$\mathrm{CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l) + 890.4\,kJ} \tag{12.7.1}$$

The equation tells us that $\mathrm{1\ mol}$ of methane combines with $\mathrm{2\ mol}$ of oxygen to produce $\mathrm{1\ mol}$ of carbon dioxide and $\mathrm{2\ mol}$ of water. In the process, $\mathrm{890.4\ kJ}$ is released and so it is written as a product of the reaction. A **thermochemical equation** is a chemical equation that includes the enthalpy change of the reaction. The process in the above thermochemical equation can be shown visually in the figure below.

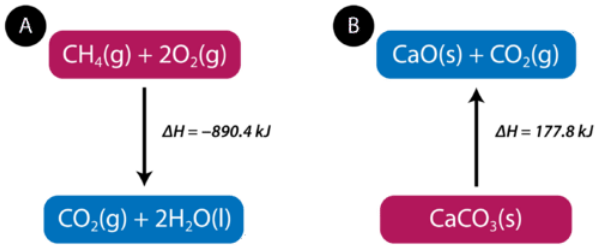


Figure 12.7.2: *(A) As reactants are converted to products in an exothermic reaction, enthalpy is released into the surroundings. The enthalpy change of the reaction is negative. (B) As reactants are converted to products in an endothermic reaction, enthalpy is absorbed from the surroundings. The enthalpy change of the reaction is positive.*

In the combustion of methane example, the enthalpy change is negative because heat is being released by the system. Therefore, the overall enthalpy of the system decreases. The **heat of reaction** is the enthalpy change for a chemical reaction. In the case above, the heat of reaction is $\mathrm{-890.4\ kJ}$. The thermochemical reaction can also be written in this way:

$$\mathrm{CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l)} \quad \Delta H = \mathrm{-890.4\,kJ} \tag{12.7.2}$$

Heats of reaction are typically measured in kilojoules. It is important to include the physical states of the reactants and products in a thermochemical equation as the value of the ΔH depends on those states.

Endothermic reactions absorb energy from the surroundings as the reaction occurs. When $\mathrm{1\ mol}$ of calcium carbonate decomposes into $\mathrm{1\ mol}$ of calcium oxide and $\mathrm{1\ mol}$ of carbon dioxide, $\mathrm{177.8\ kJ}$ of heat is absorbed. The process is shown visually in the figure above (B). The thermochemical reaction is shown below.

$$\mathrm{CaCO_3\,(s) + 177.8\,kJ \rightarrow CaO\,(s) + CO_2\,(g)} \tag{12.7.3}$$

Because the heat is absorbed by the system, the $\mathrm{177.8\ kJ}$ is written as a reactant. The heat of reaction is positive for an endothermic reaction.

$$\mathrm{CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)} \quad \Delta H = \mathrm{177.8\,kJ} \tag{12.7.4}$$

The way in which a reaction is written influences the value of the enthalpy change for the reaction. Many reactions are reversible, meaning that the product(s) of the reaction are capable of combining and reforming the reactant(s). If a reaction is written in the reverse direction, the sign of the ΔH changes. For example, we can write an equation for the reaction of calcium oxide with carbon dioxide to form calcium carbonate.

$$\mathrm{CaO\,(s) + CO_2\,(g) \rightarrow CaCO_3\,(s) + 177.8\,kJ} \tag{12.7.5}$$

The reaction is exothermic and thus the sign of the enthalpy change is negative.

$$\mathrm{CaO\,(s) + CO_2\,(g) \rightarrow CaCO_3\,(s)} \quad \Delta H = \mathrm{-177.8\,kJ} \tag{12.7.6}$$

Stoichiometric Calculations and Enthalpy Changes

Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems. Refer again to the combustion reaction of methane. Since the reaction of $\mathrm{1\ mol}$ of methane released $\mathrm{890.4\ kJ}$, the reaction of $\mathrm{2\ mol}$ of methane would release $\mathrm{2 \times 890.4\ kJ = 1781\ kJ}$. The reaction of $\mathrm{0.5\ mol}$ of methane would release $\frac{\mathrm{890.4\ kJ}}{2} = \mathrm{445.2\ kJ}$. As with other stoichiometry problems, the moles of a reactant or product can be linked to mass or volume.

Example 12.7.1

Sulfur dioxide gas reacts with oxygen to form sulfur trioxide in an exothermic reaction according to the following thermochemical equation.

$$2SO_2\,(g) + O_2\,(g) \rightarrow 2SO_3\,(g) + 198\ \text{kJ} \tag{12.7.7}$$

Calculate the enthalpy change that occurs when $58.0\ \text{g}$ of sulfur dioxide is reacted with excess oxygen.

Solution:

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

- Mass $SO_2 = 58.0\ \text{g}$
- Molar mass $SO_2 = 64.07\ \text{g/mol}$
- $\Delta H = -198\ \text{kJ}$ for the reaction of $2\ \text{mol}\ SO_2$

Unknown

- $\Delta H = ?\ \text{kJ}$

The calculation requires two steps. The mass of SO_2 is converted to moles. Then the moles of SO_2 is multiplied by the conversion factor of $\left(\frac{-198\ \text{kJ}}{2\ \text{mol}\ SO_2}\right)$.

Step 2: Solve.

$$\Delta H = -58.0\ \text{g}\ SO_2 \times \frac{1\ \text{mol}\ SO_2}{64.07\ \text{g}\ SO_2} \times \frac{-198\ \text{kJ}}{2\ \text{mol}\ SO_2} = -89.6\ \text{kJ} \tag{12.7.8}$$

Step 3: Think about your result.

The mass of sulfur dioxide is slightly less than $1\ \text{mol}$. Since $198\ \text{kJ}$ is released for every $2\ \text{mol}$ of SO_2 that reacts, the heat released when about $1\ \text{mol}$ reacts is one half of 198. The $89.6\ \text{kJ}$ is slightly less than half of 198. The sign of ΔH is negative because the reaction is exothermic.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

12.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

14: Electrons in Atoms and the Periodic Table

Template:HideTOC

14.1: The Perioidic Table

14.2: Electron Configurations

14.3: Electron Configurations and the Periodic Table (The Block Method)

14.4: Electron Configurations of Ions

14.5: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character

---

14: Electrons in Atoms and the Periodic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.1: The Perioidic Table

Learning Objectives

- Give the name and location of specific groups on the periodic table, including alkali metals, alkaline earth metals, noble gases, halogens, and transition metals.
- Explain the relationship between the chemical behavior of families in the periodic table and their electron configurations.
- Identify elements that will have the most similar properties to a given element.

The chemical behavior of atoms is controlled by their electron configuration. Since the families of elements were organized by their chemical behavior, it is predictable that the individual members of each chemical family will have similar electron configurations.

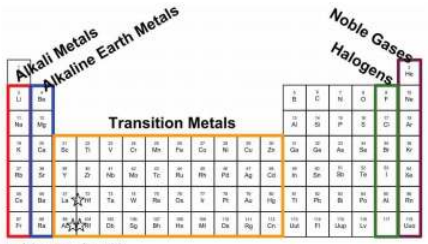
Families of the Periodic Table

Remember that Mendeleev arranged the periodic table so that elements with the most similar properties were placed in the same group. A **group** is a vertical column of the periodic table. All of the 1A elements have one valence electron. This is what causes these elements to react in the same ways as the other members of the family. The elements in 1A are all very reactive and form compounds in the same ratios with similar properties with other elements. Because of their similarities in their chemical properties, Mendeleev put these elements into the same group. Group 1A is also known as the **alkali metals**. Although most metals tend to be very hard, these metals are actually soft and can be easily cut.

Group 2A is also called the **alkaline earth metals**. Once again, because of their similarities in electron configurations, these elements have similar properties to each other. The same pattern is true of other groups on the periodic table. Remember, Mendeleev arranged the table so that elements with the most similar properties were in the same group on the periodic table.

It is important to recognize a couple of other important groups on the periodic table by their group name. Group 7A (or 17) elements are also called **halogens**. This group contains very reactive nonmetal elements.

The **noble gases** are in group 8A. These elements also have similar properties to each other, the most significant property being that they are extremely unreactive, rarely forming compounds. We will learn the reason for this later, when we discuss how compounds form. The elements in this group are also gases at room temperature.

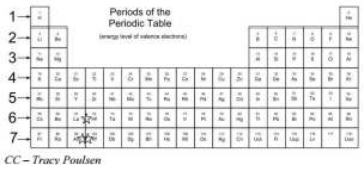


Families of the periodic table.
CC – Tracy Poulsen

An alternate numbering system numbers all of the s, p, and d block elements from 1-18. In this numbering system, group 1A is group 1; group 2A is group 2; the halogens (7A) are group 17; and the noble gases (8A) are group 18. You will come across periodic tables with both numbering systems. It is important to recognize which numbering system is being used and to be able to find the number of valence electrons in the main block elements regardless of which numbering system is being used.

Periods of the Periodic Table

If you can locate an element on the Periodic Table, you can use the element's position to figure out the energy level of the element's valence electrons. A **period** is a horizontal row of elements on the periodic table. For example, the elements sodium (Na) and magnesium (Mg) are both in period 3. The elements astatine (At) and radon (Rn) are both in period 6.



CC – Tracy Poulsen

Summary

- The vertical columns on the periodic table are called groups or families because of their similar chemical behavior.
- All the members of a family of elements have the same number of valence electrons and similar chemical properties.
- The horizontal rows on the periodic table are called periods.

Vocabulary

- Group (family): A vertical column in the periodic table.
- Alkali metals: Group 1A of the periodic table.
- Alkaline earth metals: Group 2A of the periodic table.
- Halogens: Group 7A of the periodic table.
- Noble gases: Group 8A of the periodic table.
- Transition elements: Groups 3 to 12 of the periodic table.

Contributors and Attributions

-

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

14.1: The Perioidic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.2: Electron Configurations

Learning Objectives

- To understand the basics of adding electrons to atomic orbitals
- To understand the basics of the Aufbau principle

The electron configuration of an element is the arrangement of its electrons in its atomic orbitals. By knowing the electron configuration of an element, we can predict and explain a great deal of its chemistry.

The Aufbau Principle

We construct the periodic table by following the aufbau principle (from German, meaning "building up"). First we determine the number of electrons in the atom; then we add electrons one at a time to the lowest-energy orbital available *without violating the Pauli principle*. We use the orbital energy diagram of Figure **14.2.1**, recognizing that each orbital can hold two electrons, one with spin up ↑, corresponding to $m_s = +½$, which is arbitrarily written first, and one with spin down ↓, corresponding to $m_s = −½$. A filled orbital is indicated by ↑↓, in which the electron spins are said to be *paired*. Here is a schematic orbital diagram for a hydrogen atom in its ground state:

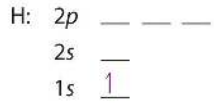
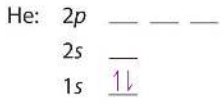


Figure **14.2.1**: One electron in.

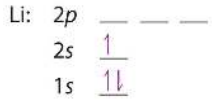
From the orbital diagram, we can write the electron configuration in an abbreviated form in which the occupied orbitals are identified by their principal quantum number *n* and their value of *l* (*s*, *p*, *d*, or *f*), with the number of electrons in the subshell indicated by a superscript. For hydrogen, therefore, the single electron is placed in the 1*s* orbital, which is the orbital lowest in energy (Figure **14.2.1**), and the electron configuration is written as $1s^1$ and read as "one-s-one."

A neutral helium atom, with an atomic number of 2 ($Z = 2$), has two electrons. We place one electron in the orbital that is lowest in energy, the 1*s* orbital. From the Pauli exclusion principle, we know that an orbital can contain two electrons with opposite spin, so we place the second electron in the same orbital as the first but pointing down, so that the electrons are paired. The orbital diagram for the helium atom is therefore



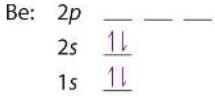
written as $1s^2$, where the superscript 2 implies the pairing of spins. Otherwise, our configuration would violate the Pauli principle.

The next element is lithium, with $Z = 3$ and three electrons in the neutral atom. We know that the 1*s* orbital can hold two of the electrons with their spins paired. Figure 6.29 tells us that the next lowest energy orbital is 2*s*, so the orbital diagram for lithium is

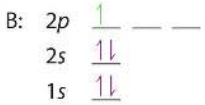


This electron configuration is written as $1s^22s^1$.

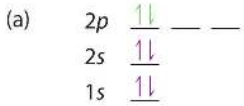
The next element is beryllium, with $Z = 4$ and four electrons. We fill both the 1*s* and 2*s* orbitals to achieve a $1s^22s^2$ electron configuration:

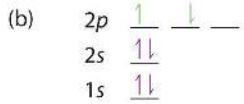


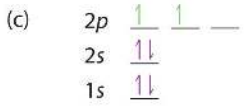
When we reach boron, with $Z = 5$ and five electrons, we must place the fifth electron in one of the 2*p* orbitals. Because all three 2*p* orbitals are degenerate, it doesn't matter which one we select. The electron configuration of boron is $1s^2 2s^2 2p^1$:



At carbon, with $Z = 6$ and six electrons, we are faced with a choice. Should the sixth electron be placed in the same 2*p* orbital that already has an electron, or should it go in one of the empty 2*p* orbitals? If it goes in an empty 2*p* orbital, will the sixth electron have its spin aligned with or be opposite to the spin of the fifth? In short, which of the following three orbital diagrams is correct for carbon, remembering that the 2*p* orbitals are degenerate?

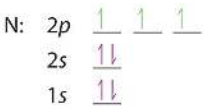






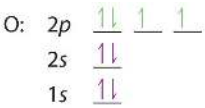
Because of electron-electron interactions, it is more favorable energetically for an electron to be in an unoccupied orbital than in one that is already occupied; hence we can eliminate choice a. Similarly, experiments have shown that choice b is slightly higher in energy (less stable) than choice c because electrons in degenerate orbitals prefer to line up with their spins parallel; thus, we can eliminate choice b. Choice c illustrates Hund's rule (named after the German physicist Friedrich H. Hund, 1896–1997), which today says that the lowest-energy electron configuration for an atom is the one that has the maximum number of electrons with parallel spins in degenerate orbitals. By Hund's rule, the electron configuration of carbon, which is $1s^2 2s^2 2p^2$, is understood to correspond to the orbital diagram shown in c. Experimentally, it is found that the ground state of a neutral carbon atom does indeed contain two unpaired electrons.

When we get to nitrogen ($Z = 7$, with seven electrons), Hund's rule tells us that the lowest-energy arrangement is

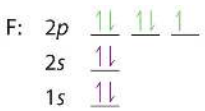


with three unpaired electrons. The electron configuration of nitrogen is thus $1s^2 2s^2 2p^3$.

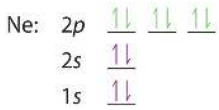
At oxygen, with $Z = 8$ and eight electrons, we have no choice. One electron must be paired with another in one of the 2*p* orbitals, which gives us two unpaired electrons and a $1s^2 2s^2 2p^4$ electron configuration. Because all the 2*p* orbitals are degenerate, it doesn't matter which one has the pair of electrons.



Similarly, fluorine has the electron configuration $1s^2 2s^2 2p^5$:



When we reach neon, with $Z = 10$, we have filled the $2p$ subshell, giving a $1s^2 2s^2 2p^6$ electron configuration:



Notice that for neon, as for helium, all the orbitals through the $2p$ level are completely filled. This fact is very important in dictating both the chemical reactivity and the bonding of helium and neon, as you will see.

Valence Electrons

As we continue through the periodic table in this way, writing the electron configurations of larger and larger atoms, it becomes tedious to keep copying the configurations of the filled inner subshells. In practice, chemists simplify the notation by using a bracketed noble gas symbol to represent the configuration of the noble gas from the preceding row because all the orbitals in a noble gas are filled. For example, [Ne] represents the $1s^2 2s^2 2p^6$ electron configuration of neon ($Z = 10$), so the electron configuration of sodium, with $Z = 11$, which is $1s^2 2s^2 2p^6 3s^1$, is written as $[Ne]3s^1$:

Neon	Z = 10	$1s^2 2s^2 2p^6$
Sodium	Z = 11	$1s^2 2s^2 2p^6 \mathbf{3}s^1 = [Ne]\mathbf{3}s^1$

Because electrons in filled inner orbitals are closer to the nucleus and more tightly bound to it, they are rarely involved in chemical reactions. This means that the chemistry of an atom depends mostly on the electrons in its outermost shell, which are called the valence electrons. The simplified notation allows us to see the valence-electron configuration more easily. Using this notation to compare the electron configurations of sodium and lithium, we have:

Sodium	$1s^2 2s^2 2p^6 3s^1 = [Ne]3s^1$
Lithium	$1s^2 2s^1 = [He]2s^1$

It is readily apparent that both sodium and lithium have one s electron in their valence shell. We would therefore predict that sodium and lithium have very similar chemistry, which is indeed the case.

As we continue to build the eight elements of period 3, the $3s$ and $3p$ orbitals are filled, one electron at a time. This row concludes with the noble gas argon, which has the electron configuration $[Ne]3s^2 3p^6$, corresponding to a filled valence shell.

Example 14.2.1: Electronic Configuration of Phoshorus

Draw an orbital diagram and use it to derive the electron configuration of phosphorus, $Z = 15$. What is its valence electron configuration?

Given: atomic number

Asked for: orbital diagram and valence electron configuration for phosphorus

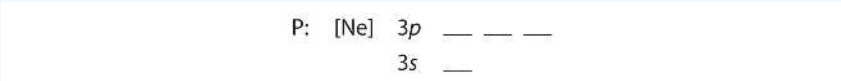
Strategy:

A. Locate the nearest noble gas preceding phosphorus in the periodic table. Then subtract its number of electrons from those in phosphorus to obtain the number of valence electrons in phosphorus.
B. Referring to Figure Figure **14.2.1**, draw an orbital diagram to represent those valence orbitals. Following Hund's rule, place the valence electrons in the available orbitals, beginning with the orbital that is lowest in energy. Write the electron configuration from your orbital diagram.
C. Ignore the inner orbitals (those that correspond to the electron configuration of the nearest noble gas) and write the valence electron configuration for phosphorus.

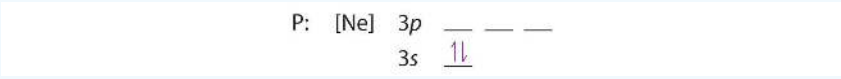
Solution:

A Because phosphorus is in the third row of the periodic table, we know that it has a [Ne] closed shell with 10 electrons. We begin by subtracting 10 electrons from the 15 in phosphorus.

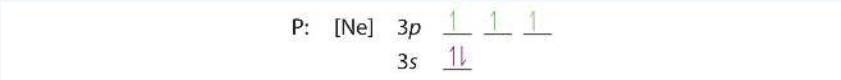
B The additional five electrons are placed in the next available orbitals, which Figure **14.2.1** tells us are the 3*s* and 3*p* orbitals:



Because the 3*s* orbital is lower in energy than the 3*p* orbitals, we fill it first:



Hund's rule tells us that the remaining three electrons will occupy the degenerate 3*p* orbitals separately but with their spins aligned:



The electron configuration is $[Ne]3s^2 3p^3$.

C We obtain the valence electron configuration by ignoring the inner orbitals, which for phosphorus means that we ignore the [Ne] closed shell. This gives a valence-electron configuration of $3s^2 3p^3$.

Exercise 14.2.1

Draw an orbital diagram and use it to derive the electron configuration of chlorine, $Z = 17$. What is its valence electron configuration?

Answer:

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

The general order in which orbitals are filled is depicted in Figure **14.2.2**. Subshells corresponding to each value of *n* are written from left to right on successive horizontal lines, where each row represents a row in the periodic table. The order in which the orbitals are filled is indicated by the diagonal lines running from the upper right to the lower left. Accordingly, the 4*s* orbital is filled prior to the 3*d* orbital because of shielding and penetration effects. Consequently, the electron configuration of potassium, which begins the fourth period, is $[Ar]4s^1$, and the configuration of calcium is $[Ar]4s^2$. Five 3*d* orbitals are filled by the next 10 elements, the transition metals, followed by three 4*p* orbitals. Notice that the last member of this row is the noble gas krypton ($Z = 36$), $[Ar]4s^2 3d^{10} 4p^6 = [Kr]$, which has filled 4*s*, 3*d*, and 4*p* orbitals. The fifth row of the periodic table is essentially the same as the fourth, except that the 5*s*, 4*d*, and 5*p* orbitals are filled sequentially.

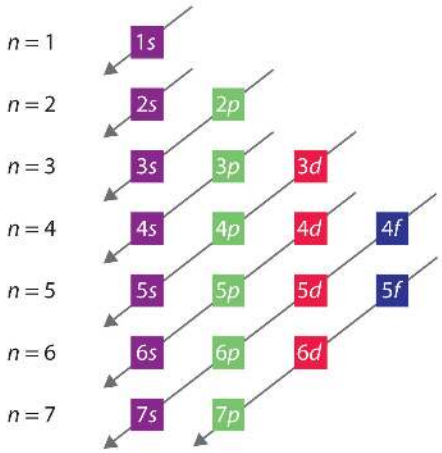


Figure **14.2.2**: Predicting the Order in Which Orbitals Are Filled in Multielectron Atoms. If you write the subshells for each value of the principal quantum number on successive lines, the observed order in which they are filled is indicated by a series of diagonal lines running from the upper right to the lower left.

The sixth row of the periodic table will be different from the preceding two because the 4*f* orbitals, which can hold 14 electrons, are filled between the 6*s* and the 5*d* orbitals. The elements that contain 4*f* orbitals in their valence shell are the lanthanides. When the 6*p* orbitals are finally filled, we have reached the next (and last known) noble gas, radon (Z = 86), $[Xe]6s^24f^{14}5d^{10}6p^6$ = [Rn]. In the last row, the 5*f* orbitals are filled between the 7*s* and the 6*d* orbitals, which gives the 14 actinide elements. Because the large number of protons makes their nuclei unstable, all the actinides are radioactive.

Example 14.2.2: Electron Configuration of Mercury

Write the electron configuration of mercury (Z = 80), showing all the inner orbitals.

Given: atomic number

Asked for: complete electron configuration

Strategy:

Using the orbital diagram in Figure **14.2.1** and the periodic table as a guide, fill the orbitals until all 80 electrons have been placed.

Solution:

By placing the electrons in orbitals following the order shown in Figure **14.2.2** and using the periodic table as a guide, we obtain

$1s^2$	row 1	2 electrons
$2s^22p^6$	row 2	8 electrons
$3s^23p^6$	row 3	8 electrons
$4s^23d^{10}4p^6$	row 4	18 electrons
$5s^24d^{10}5p^6$	row 5	18 electrons
	row 1–5	54 electrons

After filling the first five rows, we still have 80 − 54 = 26 more electrons to accommodate. According to Figure **14.2.2**, we need to fill the 6*s* (2 electrons), 4*f* (14 electrons), and 5*d* (10 electrons) orbitals. The result is mercury's electron configuration:

$$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10} = Hg = [Xe]6s^24f^{14}5d^{10}$$

with a filled 5*d* subshell, a $6s^24f^{14}5d^{10}$ valence shell configuration, and a total of 80 electrons. (You should always check to be sure that the total number of electrons equals the atomic number.)

Exercise 14.2.2: Electron Configuration of Flerovium

Although element 114 is not stable enough to occur in nature, two isotopes of element 114 were created for the first time in a nuclear reactor in 1999 by a team of Russian and American scientists. The element is named after the Flerov Laboratory of Nuclear Reactions of the Joint Institute for Nuclear Research in Dubna, Russia, where the element was discovered in 1998. The name of the laboratory, in turn, honors the Russian physicist Georgy Flyorov. Write the complete electron configuration for element 114.

Answer

$s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10} 7p^2$

The electron configurations of the elements are presented in Figure **14.2.2**, which lists the orbitals in the order in which they are filled. In several cases, the ground state electron configurations are different from those predicted by Figure **14.2.1**. Some of these anomalies occur as the 3*d* orbitals are filled. For example, the observed ground state electron configuration of chromium is [Ar]$4s^1 3d^5$ rather than the predicted [Ar]$4s^2 3d^4$. Similarly, the observed electron configuration of copper is [Ar]$4s^1 3d^{10}$ instead of [Ar]$s^2 3d^9$. The actual electron configuration may be rationalized in terms of an added stability associated with a half-filled (ns^1, np^3, nd^5, nf^7) or filled (ns^2, np^6, nd^{10}, nf^{14}) subshell. Given the small differences between higher energy levels, this added stability is enough to shift an electron from one orbital to another. In heavier elements, other more complex effects can also be important, leading to some of the additional anomalies indicated in Figure **14.2.2**. For example, cerium has an electron configuration of [Xe]$6s^2 4f^1 5d^1$, which is impossible to rationalize in simple terms. In most cases, however, these apparent anomalies do not have important chemical consequences.

> *Additional stability is associated with half-filled or filled subshells.*

Summary

Based on the Pauli principle and a knowledge of orbital energies obtained using hydrogen-like orbitals, it is possible to construct the periodic table by filling up the available orbitals beginning with the lowest-energy orbitals (the **aufbau principle**), which gives rise to a particular arrangement of electrons for each element (its **electron configuration**). **Hund's rule** says that the lowest-energy arrangement of electrons is the one that places them in degenerate orbitals with their spins parallel. For chemical purposes, the most important electrons are those in the outermost principal shell, the **valence electrons**.

14.2: Electron Configurations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.3: Electron Configurations and the Periodic Table (The Block Method)

SKILLS TO DEVELOP

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

Previously, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure **14.3.1**. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.

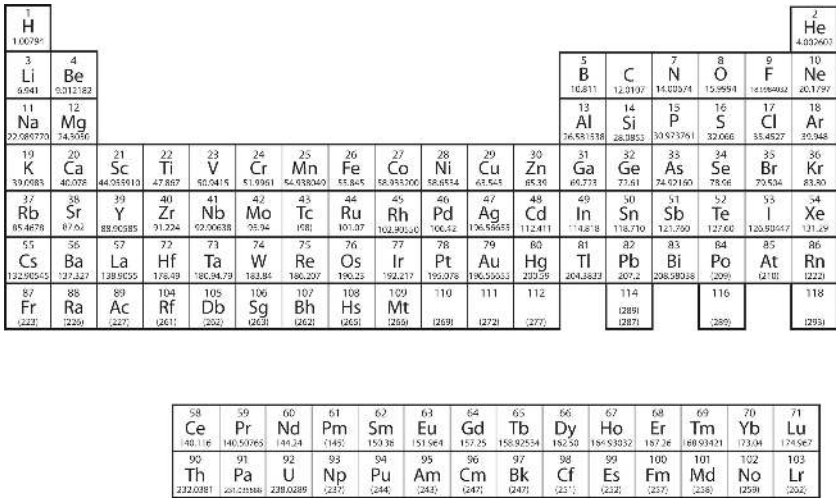


Figure **14.3.1**: The Periodic Table

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

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

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure **14.3.2**)

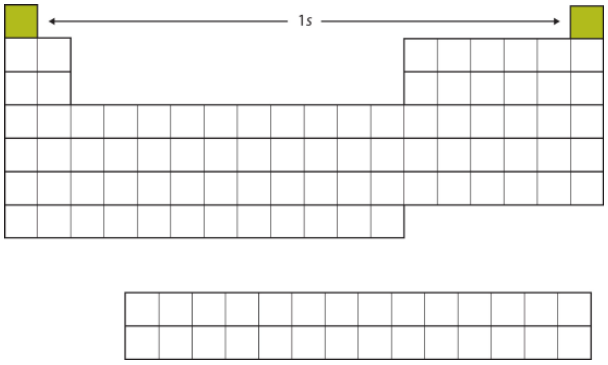


Figure **14.3.2**: The 1*s* Subshell. H and He represent the filling of the 1*s* subshell.

The next two electrons, for Li and Be, would go into the 2*s* subshell. Figure **14.3.3** shows that these two elements are adjacent on the periodic table.

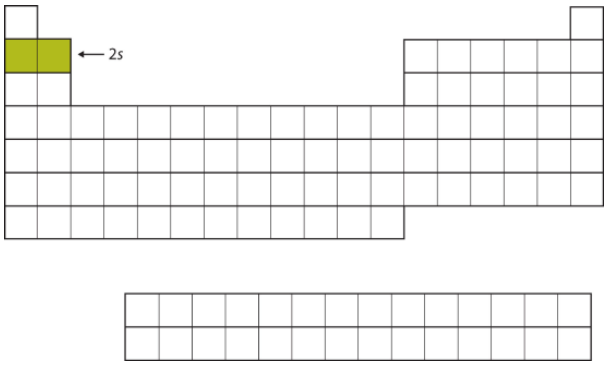


Figure 14.3.3: The 2*s* Subshell. In Li and Be, the 2*s* subshell is being filled.

For the next six elements, the 2*p* subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 14.3.4).

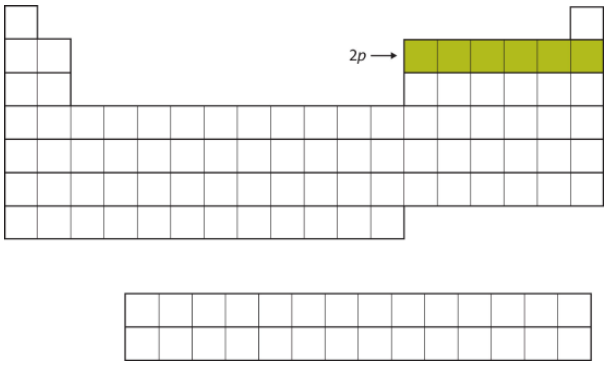


Figure 14.3.4: The 2*p* Subshell. For B through Ne, the 2*p* subshell is being occupied.

The next subshell to be filled is the 3*s* subshell. The elements when this subshell is being filled, Na and Mg, are back on the left side of the periodic table (Figure 14.3.5).

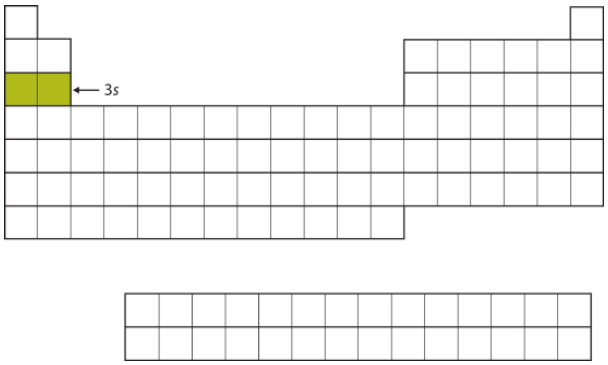


Figure 14.3.5: The 3*s* Subshell. Now the 3*s* subshell is being occupied.

Next, the 3*p* subshell is filled with the next six elements (Figure 14.3.6).

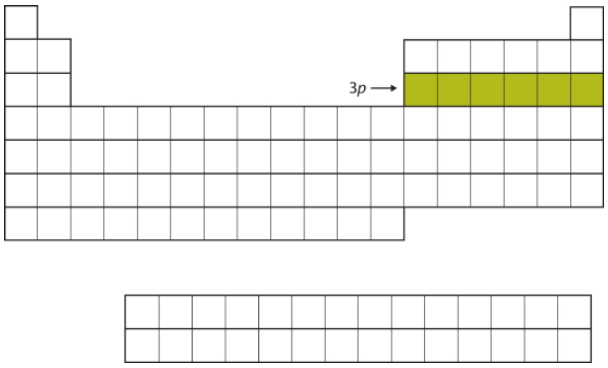


Figure 14.3.6: The 3*p* Subshell. Next, the 3*p* subshell is filled with electrons.

Instead of filling the 3*d* subshell next, electrons go into the 4*s* subshell (Figure 14.3.7).

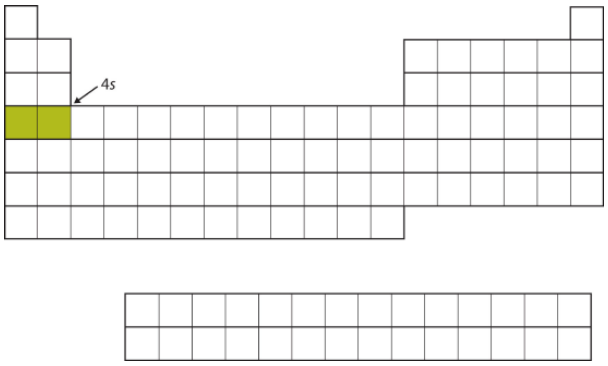


Figure 14.3.7: The 4*s* Subshell. The 4*s* subshell is filled before the 3*d* subshell. This is reflected in the structure of the periodic table.

After the 4*s* subshell is filled, the 3*d* subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure 14.3.8).

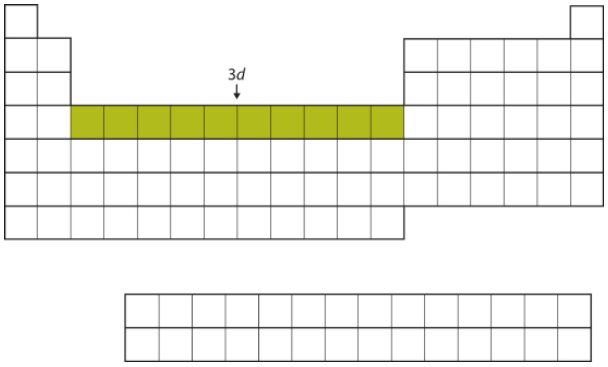


Figure 14.3.8: The 3*d* Subshell. The 3*d* subshell is filled in the middle section of the periodic table.

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

The first two columns on the left side of the periodic table are where the *s* subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** are the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 14.3.9 shows the blocks of the periodic table.

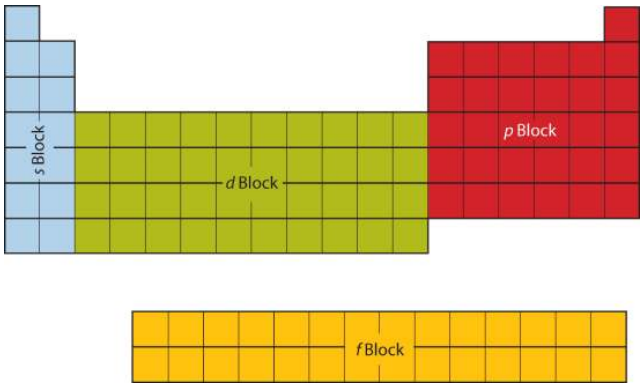


Figure **14.3.9**: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

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

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

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

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

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure **14.3.10**. It is in the fourth column of the *p* block. This means that its electron configuration should end in a p^4 electron configuration. Indeed, the electron configuration of Se is $[Ar]4s^23d^{10}4p^4$, as expected.

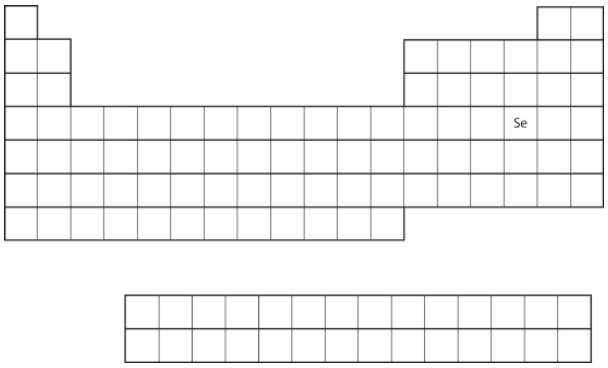


Figure **14.3.10**: Selenium on the Periodic Table

Example 14.3.1: Predicting Electron Configurations

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure **14.3.11**).

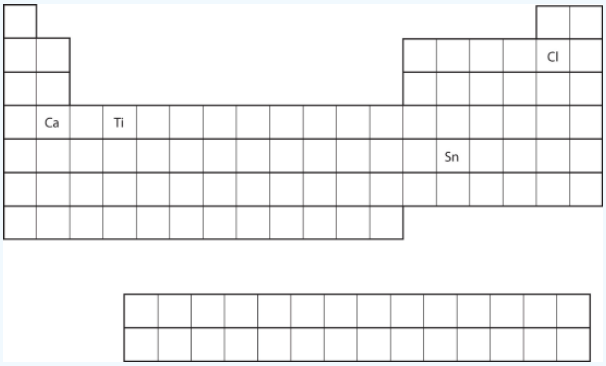


Figure **14.3.11**: Various Elements on the Periodic Table

a. Ca
b. Sn

Solution

a. Ca is located in the second column of the *s* block. We would expect that its electron configuration should end with s^2. Calcium's electron configuration is $[Ar]4s^2$.
b. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in p^2. Tin's electron configuration is $[Kr]5s^24d^{10}5p^2$.

Exercise 14.3.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom. Figure **14.3.11**.

a. Ti
b. Cl

Answer a

$[Ar]4s^23d^2$

Answer b

$[Ne]3s^23p^5$

Summary

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

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

14.3: Electron Configurations and the Periodic Table (The Block Method) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.4: Electron Configurations of Ions

14.4: Electron Configurations of Ions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.5: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character

Learning Objectives

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

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

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

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

$$as \downarrow PT, atomic\ radius \uparrow \tag{14.5.1}$$

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

$$as \rightarrow PT, atomic\ radius \downarrow \tag{14.5.2}$$

Figure 14.5.1 shows spheres representing the atoms of the *s* and *p* blocks from the periodic table to scale, showing the two trends for the atomic radius.

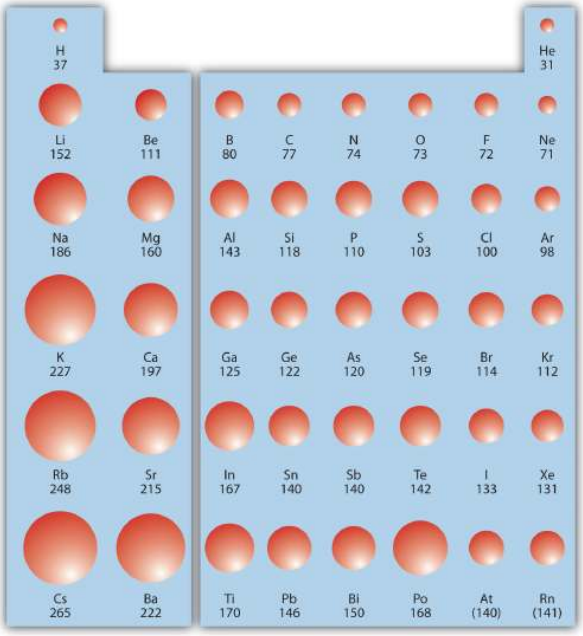


Figure 14.5.1: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

Example 14.5.1: Atomic Radii

Referring only to a periodic table and not to Figure **14.5.1**, which atom is larger in each pair?

a. Si or S
b. S or Te

Solution

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

Exercise 14.5.1: Atomic Radii

Referring only to a periodic table and not to Figure **14.5.1**, which atom is smaller, Ca or Br?

Answer

Br

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

$$A(g) \rightarrow A^{+}(g) + e^{-} \quad \Delta H \equiv IE \tag{14.5.3}$$

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

$$as \downarrow PT,\ IE \downarrow \tag{14.5.4}$$

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

$$as \rightarrow PT,\ IE \uparrow \tag{14.5.5}$$

Figure **14.5.2** shows values of IE versus position on the periodic table. Again, the trend isn't absolute, but the general trends going across and down the periodic table should be obvious.

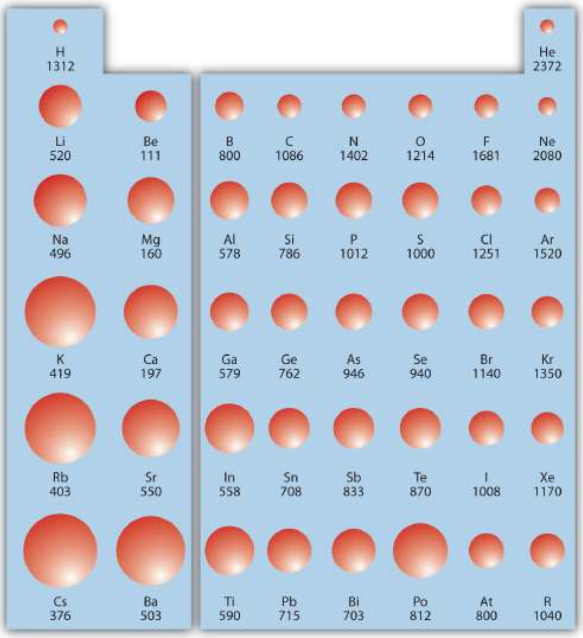


Figure **14.5.2**: Ionization Energy on the Periodic Table. Values are in kJ/mol.

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

- First Ionization Energy (IE_1):

$$A(g) \rightarrow A^+(g) + e^- \tag{14.5.6}$$

- Second Ionization Energy (IE_2):

$$A^+(g) \rightarrow A^{2+}(g) + e^- \tag{14.5.7}$$

- Third Ionization Energy (IE_3):

$$A^{2+}(g) \rightarrow A^{3+}(g) + e^- \tag{14.5.8}$$

and so forth.

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

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

$$Mg(g) \rightarrow Mg^+(g) + e^- \tag{14.5.9}$$

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

$$Mg^+(g) \rightarrow Mg^{2+}(g) + e^- \tag{14.5.10}$$

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

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^- \tag{14.5.11}$$

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

Example 14.5.2: Ionization Energies

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

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

Solution

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

Exercise 14.5.2: Ionization Energies

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

Answer

C

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

$$A(g) + e^- \rightarrow A^-(g) \quad \Delta H \equiv EA \tag{14.5.12}$$

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

$$as \rightarrow PT,\ EA \uparrow \tag{14.5.13}$$

There is not a definitive trend as you go down the periodic table; sometimes EA increases, sometimes it decreases. Figure **14.5.3** shows EA values versus position on the periodic table for the *s*- and *p*-block elements. The trend isn't absolute, especially considering the large positive EA values for the second column. However, the general trend going across the periodic table should be obvious.

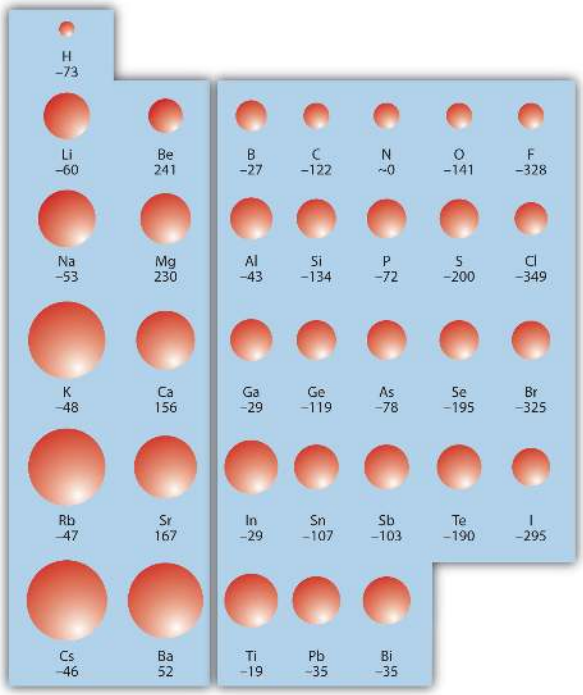


Figure **14.5.3**: Electron Affinity on the Periodic Table. Values are in kJ/mol.

Example 14.5.3: Electron Affinities

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

a. C or F
b. Na or S

Solution

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

Exercise 14.5.3: Electron Affinities

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

Answer

Br

Metallic Character

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

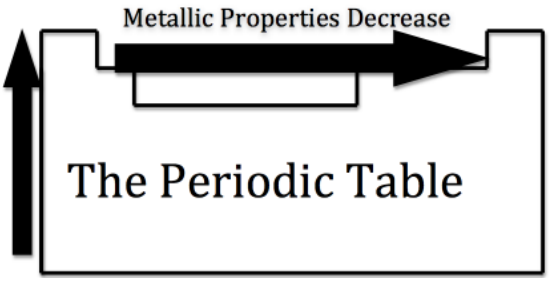


Figure 14.5.4: Courtesy of Jessica Thornton (UCD)

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

Uses in knowing the Periodic Properties of Elements

1. Predicting greater or smaller atomic size and radial distribution in neutral atoms and ions
2. Measuring and comparing ionization energies
3. Comparing electron affinities and electronegativities
    - Predicting redox potential
    - Comparing metallic character with other elements; its ability to form cations
    - Predicting what reaction may or may not occur due to the trends
    - Determining greater cell potential (sum of oxidation and reduction potential) between reactions
    - Completing chemical reactions according to trends

Summary

Certain properties—notably atomic radius, ionization energies, and electron affinities - can be qualitatively understood by the positions of the elements on the periodic table. Atomic size is the distance from the nucleus to the valence shell where the valence electrons are located. Atomic radius is a more definite and measurable way of defining atomic size. It is the distance from the center of one atom to the center of another atom in a homonuclear diatomic molecule. There are three factors that help in the prediction of the trends in the Periodic Table: number of protons in the nucleus, number of shells, and shielding effect. The atomic size increases from the top to the bottom in any group as a result of increases in all of the three factors. (As the number of energy levels increases, the size must increase.) Going across a period (from left to right), the number of protons increases and therefore the nuclear charge increases. (Going across a period, the number of electron energy levels remains the same but the number of electrons increases within these energy levels. Therefore the electrons are pulled in closer to the nucleus.)

Ionization energy is the energy required to remove the most loosely held electron from a gaseous atom or ion. Ionization energy generally increases across a period and decreases down a group. The effective nuclear charge is the charge of the nucleus felt by the valence electron. Electron affinity is the energy required (or released) when an electron is added to a gaseous atom or ion. Electron affinity generally increases going up a group and increases left to right across a period. Non-metals tend to have the highest electron affinities.

Contributors and Attributions

-

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

14.5: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1.7: Classification of Matter

Learning Objectives

- To classify matter.

Chemists study the structures, physical properties, and chemical properties of material substances. These consist of **matter**, which is anything that occupies space and has mass. Gold and iridium are matter, as are peanuts, people, and postage stamps. Smoke, smog, and laughing gas are matter. Energy, light, and sound, however, are not matter; ideas and emotions are also not matter.

The **mass** of an object is the quantity of matter it contains. Do not confuse an object's mass with its **weight**, which is a force caused by the gravitational attraction that operates on the object. Mass is a fundamental property of an object that does not depend on its location.In physical terms, the mass of an object is directly proportional to the force required to change its speed or direction. A more detailed discussion of the differences between weight and mass and the units used to measure them is included in Essential Skills 1 (Section 1.9). Weight, on the other hand, depends on the location of an object. An astronaut whose mass is 95 kg weighs about 210 lb on Earth but only about 35 lb on the moon because the gravitational force he or she experiences on the moon is approximately one-sixth the force experienced on Earth. For practical purposes, weight and mass are often used interchangeably in laboratories. Because the force of gravity is considered to be the same everywhere on Earth's surface, 2.2 lb (a weight) equals 1.0 kg (a mass), regardless of the location of the laboratory on Earth.

Under normal conditions, there are three distinct states of matter: solids, liquids, and gases. **Solids** are relatively rigid and have fixed shapes and volumes. A rock, for example, is a solid. In contrast, **liquids** have fixed volumes but flow to assume the shape of their containers, such as a beverage in a can. **Gases**, such as air in an automobile tire, have neither fixed shapes nor fixed volumes and expand to completely fill their containers. Whereas the volume of gases strongly depends on their temperature and **pressure** (the amount of force exerted on a given area), the volumes of liquids and solids are virtually independent of temperature and pressure. Matter can often change from one physical state to another in a process called a **physical change**. For example, liquid water can be heated to form a gas called steam, or steam can be cooled to form liquid water. However, such changes of state do not affect the chemical composition of the substance.

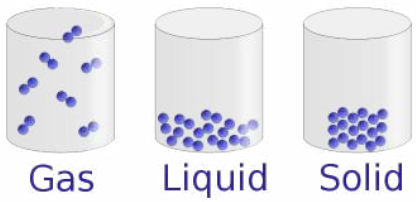


Figure used with permission from Wikipedia

Pure Substances and Mixtures

A pure chemical substance is any matter that has a fixed chemical composition and characteristic properties. Oxygen, for example, is a pure chemical substance that is a colorless, odorless gas at 25°C. Very few samples of matter consist of pure substances; instead, most are mixtures, which are combinations of two or more pure substances in variable proportions in which the individual substances retain their identity. Air, tap water, milk, blue cheese, bread, and dirt are all mixtures. If all portions of a material are in the same state, have no visible boundaries, and are uniform throughout, then the material is **homogeneous**. Examples of homogeneous mixtures are the air we breathe and the tap water we drink. Homogeneous mixtures are also called solutions. Thus air is a solution of nitrogen, oxygen, water vapor, carbon dioxide, and several other gases; tap water is a solution of small amounts of several substances in water. The specific compositions of both of these solutions are not fixed, however, but depend on both source and location; for example, the composition of tap water in Boise, Idaho, is not the same as the composition of tap water in Buffalo, New York. Although most solutions we encounter are liquid, solutions can also be solid. The gray substance still used by some dentists to fill tooth cavities is a complex solid solution that contains 50% mercury and 50% of a powder that contains mostly silver, tin, and copper, with small amounts of zinc and mercury. Solid solutions of two or more metals are commonly called alloys.

If the composition of a material is not completely uniform, then it is **heterogeneous** (e.g., chocolate chip cookie dough, blue cheese, and dirt). Mixtures that appear to be homogeneous are often found to be heterogeneous after microscopic examination. Milk, for example, appears to be homogeneous, but when examined under a microscope, it clearly consists of tiny globules of fat and protein dispersed in water. The components of heterogeneous mixtures can usually be separated by simple means. Solid-liquid mixtures such as sand in water or tea leaves in tea are readily separated by filtration, which consists of passing the mixture through a barrier, such as a strainer, with holes or pores that are smaller than the solid particles. In principle, mixtures of two or more solids, such as sugar and salt, can be separated by microscopic inspection and sorting. More complex operations are usually necessary, though, such as when separating gold nuggets from river gravel by panning. First solid material is filtered from river water; then the solids are separated by inspection. If gold is embedded in rock, it may have to be isolated using chemical methods.



Figure used with permission from Wikipedia

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

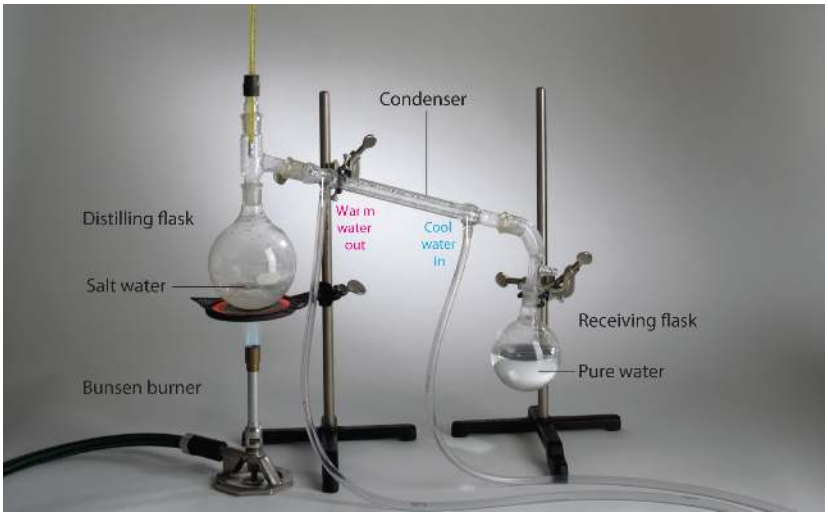


Figure 1.7.3: The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as

brandy or whiskey. (This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States!)

Crystallization separates mixtures based on differences in solubility, a measure of how much solid substance remains dissolved in a given amount of a specified liquid. Most substances are more soluble at higher temperatures, so a mixture of two or more substances can be dissolved at an elevated temperature and then allowed to cool slowly. Alternatively, the liquid, called the solvent, may be allowed to evaporate. In either case, the least soluble of the dissolved substances, the one that is least likely to remain in solution, usually forms crystals first, and these crystals can be removed from the remaining solution by filtration.



Figure **1.7.4**: The Crystallization of Sodium Acetate from a Concentrated Solution of Sodium Acetate in Water. The addition of a small "seed" crystal (a) causes the compound to form white crystals, which grow and eventually occupy most of the flask. Video can be found here: https://www.youtube.com/watch?v=BLq5NibwV5g

Most mixtures can be separated into pure substances, which may be either elements or compounds. An **element**, such as gray, metallic sodium, is a substance that cannot be broken down into simpler ones by chemical changes; a **compound**, such as white, crystalline sodium chloride, contains two or more elements and has chemical and physical properties that are usually different from those of the elements of which it is composed. With only a few exceptions, a particular compound has the same elemental composition (the same elements in the same proportions) regardless of its source or history. The chemical composition of a substance is altered in a process called a **chemical change**. The conversion of two or more elements, such as sodium and chlorine, to a chemical compound, sodium chloride, is an example of a chemical change, often called a chemical reaction. Currently, about 118 elements are known, but millions of chemical compounds have been prepared from these 118 elements. The known elements are listed in the periodic table.

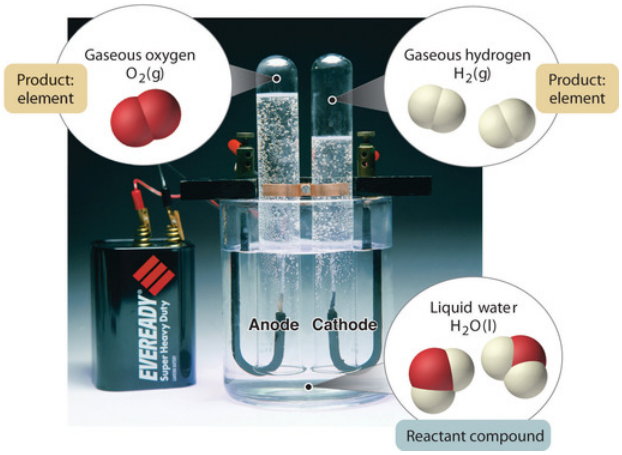


Figure **1.7.5**: The Decomposition of Water to Hydrogen and Oxygen by Electrolysis. Water is a chemical compound; hydrogen and oxygen are elements.

In general, a reverse chemical process breaks down compounds into their elements. For example, water (a compound) can be decomposed into hydrogen and oxygen (both elements) by a process called electrolysis. In electrolysis, electricity provides the energy needed to separate a compound into its constituent elements (Figure **1.7.5**). A similar technique is used on a vast scale to obtain pure aluminum, an element, from its ores, which are mixtures of compounds. Because a great deal of energy is required for

electrolysis, the cost of electricity is by far the greatest expense incurred in manufacturing pure aluminum. Thus recycling aluminum is both cost-effective and ecologically sound.

The overall organization of matter and the methods used to separate mixtures are summarized in Figure **1.7.6**.

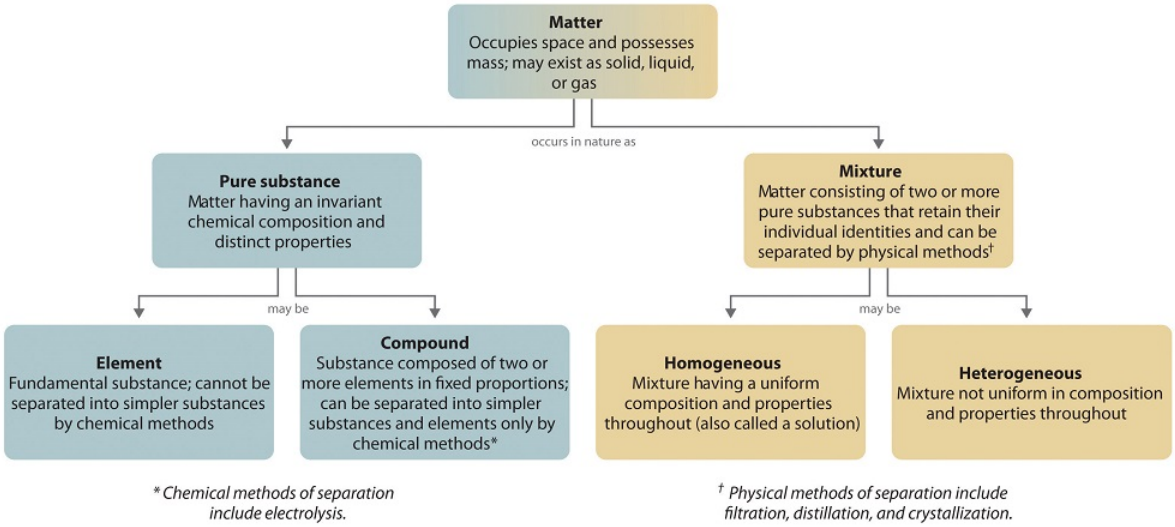


Figure **1.7.6**: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Example 1.7.1

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

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

Solution:

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

Exercise 1.7.1

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

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

Answer A

solution

Answer B

element

Answer C

heterogeneous mixture

Answer D

compound

Summary

Matter can be classified according to physical and chemical properties. Matter is anything that occupies space and has mass. The three states of matter are solid, liquid, and gas. A physical change involves the conversion of a substance from one state of matter to another, without changing its chemical composition. Most matter consists of mixtures of pure substances, which can be homogeneous (uniform in composition) or heterogeneous (different regions possess different compositions and properties). Pure substances can be either chemical compounds or elements. Compounds can be broken down into elements by chemical reactions, but elements cannot be separated into simpler substances by chemical means. The properties of substances can be classified as either physical or chemical. Scientists can observe physical properties without changing the composition of the substance, whereas chemical properties describe the tendency of a substance to undergo chemical changes (chemical reactions) that change its chemical composition. Physical properties can be intensive or extensive. Intensive properties are the same for all samples; do not depend on sample size; and include, for example, color, physical state, and melting and boiling points. Extensive properties depend on the amount of material and include mass and volume. The ratio of two extensive properties, mass and volume, is an important intensive property called density.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)

1.7: Classification of Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

10: Acid-Base and Gas Evolution Reactions

Template:HideTOC

10.1: Acids: Properties and Examples

10.2: Bases: Properties and Examples

10.3: Molecular Definitions of Acids and Bases

10.4: Reactions of Acids and Bases

10.5: Acid–Base and Gas Evolution Reactions

10: Acid-Base and Gas Evolution Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.1: Acids: Properties and Examples

Learning Objectives

- Examine properties of acids.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{10.1.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

It should not be hard for you to name several common acids, but you might find that listing bases is just a little more difficult. Here's a partial list of some common acids and bases, along with some chemical formulas:

Table 10.1.1: Common Acids and Their Uses

Chemist Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) and stomach acid is HCl)	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.
sulfuric acid, H_2SO_4		Used in car batteries, and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fiz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $C_6H_8O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$	aspirin	The active ingredient in aspirin.

In the next section we will begin to examine what, exactly makes an acid an acid, and what makes a base act as a base. Take a look at the formulas given in the first table and take a guess.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. A colorless, highly pungent solution of hydrogen chloride (HCl) in water. Hydrochloric acid is usually prepared by treating $\mathbf{HCl}$ with water.

$$\mathbf{HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{10.1.2}$$

Hydrochloric acid can therefore be used to prepare chloride salts. Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases.

Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula $\mathbf{H_2SO_4}$. Sulfuric acid is a diprotic acid and has a wide range of applications including in domestic acidic drain cleaners,[l]as an electrolyte in lead-acid batteries and in various cleaning agents. It is also a central substance in the chemical industry.



Figure **10.1.1**: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating

property, removing water (H_2O) from other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 14.2.1).

Nitric Acid

Nitric acid (HNO_3) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide ($NO_2\,(g)$) with water.

$$3\,NO_2\,(g) + H_2O\,(l) \longrightarrow 2\,HNO_3\,(aq)^+\,NO\,(g) \tag{10.1.3}$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals (e.g., nitric acid with magnesium, manganese or zinc will liberate H_2 gas):

$$Mg + 2\,HNO_3 \longrightarrow Mg(NO_3)_2 + H_2 \tag{10.1.4}$$

$$Mn + 2\,HNO_3 \longrightarrow Mn(NO_3)_2 + H_2 \tag{10.1.5}$$

$$Zn + 2\,HNO_3 \longrightarrow Zn(NO_3)_2 + H_2 \tag{10.1.6}$$

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard posed by it is chemical burns as it carries out acid hydrolysis with proteins (amide) and fats (ester) which consequently decomposes living tissue (Figure 10.1.2). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin



Figure* 10.1.2*: *Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).*

Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula H_2CO_3 and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of $H_2CO_3\,(aq)$. Carbonic acid, which is a weak acid, forms two kinds of salts, the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve producing calcium bicarbonate which leads to many limestone features such as stalactites and stalagmites. Carbonic acid is a polyprotic acid — specifically it is diprotic meaning it has two protons which may dissociate from the parent molecule.

When carbon dioxide dissolves in water it exists in chemical equilibrium (discussed in Chapter 15) producing carbonic acid:

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \tag{10.1.7}$$

The reaction can be pushed to favor the reactants to generate $CO_2\,(g)$ from solution, which is key to the bubbles observed in carbonated beverages (Figure 10.1.3).



Figure 10.1.3: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

Formic Acid

Formic acid (HCO_2H) is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally, most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

Citric Acid

Citric acid ($C_6H_8O_7$) is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the TCA cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is used as a flavoring and preservative in food and beverages, especially soft drinks.

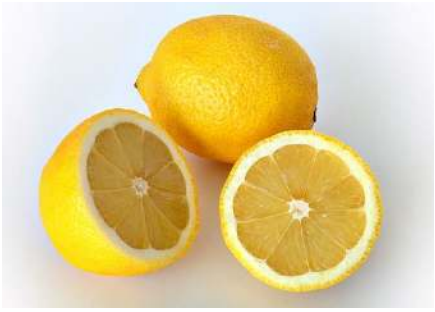


Figure 10.1.4: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years.

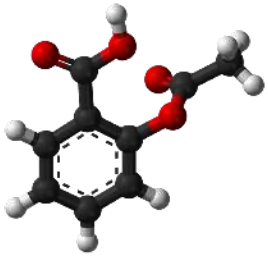


Figure 10.1.5: Ball-and-stick model of the aspirin molecule. (Public Domain; Ben Mills).

Aspirin is a white, crystalline, weakly acidic substance.

Summary

A brief summary of key aspects of several key acids commonly encountered by students was given. Acids are a distinct class of compounds because of the properties of their aqueous solutions.

Contributors and Attributions

-
- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

10.1: Acids: Properties and Examples is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.2: Bases: Properties and Examples

Learning Objectives

- Examine properties of bases.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the $\mathbf{HCl}$ in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.



Figure $\mathbf{10.2.1}$: Phenolphthalein indicator in presence of base.

Warning

Tasting chemicals and touching them are NOT good lab practices and should be avoided - in other words, don't do this at home.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula $\mathbf{NaOH}$. It is a white solid ionic compound consisting of sodium cations $\mathbf{Na^+}$ and hydroxide anions $\mathbf{OH^-}$.



Sample of sodium hydroxide as pellets in a watchglass. (Public Domain; Walkerma).

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction

$$\mathbf{NaOH\,(s)\ \rightarrow Na^+\,(aq) + OH^-\,(aq)} \tag{10.2.1}$$

The resulting solution is usually colorless and odorless and feels slippery when it comes in contact with skin.

Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula **KOH**, and is commonly called caustic potash. Along with sodium hydroxide (NaOH), this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its corrosive nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathrm{KOH\,(s)\ \rightarrow K^{+}\,(aq) + OH^{-}\,(aq)} \tag{10.2.2}$$

Concentrated aqueous solutions are sometimes called *potassium lyes*.

Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $\mathrm{Mg(OH)_2}$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.



Bottle of Antacid tablets. (CC BY 2.,5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $\mathrm{Mg(OH)_2}$. Although a weak concentration of dissociated ions can be found in solution:

$$\mathrm{Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}\,(aq) + 2\,OH^{-}\,(aq)} \tag{10.2.3}$$

Calcium Hydroxide

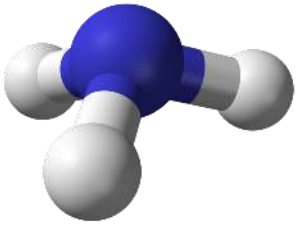
Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathrm{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$\mathrm{Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,OH^{-}\,(aq)} \tag{10.2.4}$$

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathrm{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form



Ball-and-stick model of the ammonia molecule. (Public Domain; Ben Mills).

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from $\mathbf{H_2O}$ to yield ammonium and hydroxide ions

$$\mathbf{NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)} \tag{10.2.5}$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate.

Table **10.2.1**: Common Bases and Its Uses

Some Common Bases	Uses
sodium hydroxide, NaOH (lye or caustic soda)	Used in the manufacture of soaps and detergents and as the main ingredient in oven and drain cleaners.
potassium hydroxide, KOH (lye or caustic potash)	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$ (slaked lime)	Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil.
aluminum hydroxide	Used in water purification and as an ingredient in antacids.
ammonia, NH_3	Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

Contributors and Attributions

-
- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

10.2: Bases: Properties and Examples is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.3: Molecular Definitions of Acids and Bases

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major classifications of substances known as acids or bases. The Arrhenius definition states that an acid produces H^+ in solution and a base produces OH^-. This theory was developed by Svante Arrhenius in 1883. Later, two more sophisticated and general theories were proposed. These are the Brønsted-Lowry and the Lewis definitions of acids and bases. The Lewis theory is discussed elsewhere.

The Arrhenius Theory of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. An **Arrhenius acid** is a compound that **increases the concentration of H^+ ions** that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$\mathrm{HCl\,(aq)\ \rightarrow H^+\,(aq) + Cl^-\,(aq)} \tag{10.3.1}$$

In this reaction, hydrochloric acid (***HCl***) dissociates completely into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$\mathrm{HCl\,(aq) + H_2O\,(l)\ \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{10.3.2}$$

An **Arrhenius base** is a compound that **increases the concentration of OH^- ions** that are present when added to water. The dissociation is represented by the following equation:

$$\mathrm{NaOH\ \ (aq)\ \rightarrow Na^+\ \ (aq) + OH^-\ \ (aq)} \tag{10.3.3}$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium (***Na^+***) and hydroxide (***OH^-***) ions when dissolved in water, thereby releasing OH^- ions into solution.

> *Arrhenius acids are substances which produce hydrogen ions in solution and Arrhenius bases are substances which produce hydroxide ions in solution.*

Limitations to the Arrhenius Theory

The Arrhenius theory has many more limitations than the other two theories. The theory does not explain the weak base ammonia (NH_3), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself. Also, the Arrhenius definition of acid and base is limited to aqueous (i.e., water) solutions.

The Brønsted-Lowry Theory of Acids and Bases

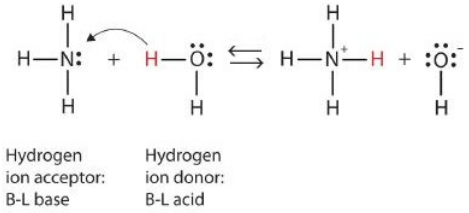
In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, **a Brønsted-Lowry acid** is a **proton donor (PD)**, while **a Brønsted-Lowry base** is a **proton acceptor (PA)**.

> *A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.*

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$NH_3\,(aq) + H_2O\,(\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq) \tag{10.3.4}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:

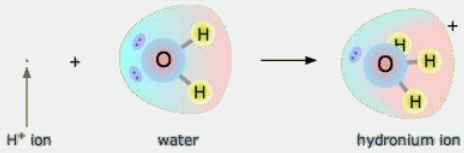


Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

The Hydronium IOn

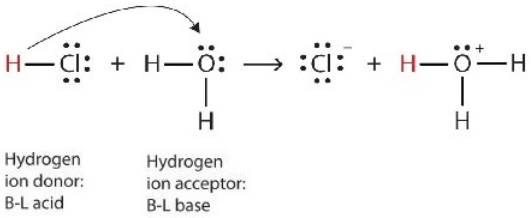
A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+ to represent the hydronium ion.



With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$HCl\,(g) + H_2O\,(\ell) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq) \tag{10.3.5}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.

- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.
- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Example 10.3.1

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:

$$\mathbf{C_6H_5NH_2\,(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3^+\,(aq) + OH^-\,(aq)}$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

Exercise 10.3.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$\mathbf{H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+} \tag{10.3.6}$$

Answer:

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: H_2O

Exercise 10.3.2

Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. HPO_4^{2-}
c. H_3PO_4
d. NH_4^+
e. $CH_3NH_3^+$

Answer

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H^+. This eliminates $\mathbf{HCl}$, $\mathbf{H_3PO_4}$, $\mathbf{NH_4^+}$ and $\mathbf{CH_3NH_3^+}$ because they are Bronsted-Lowry acids. They all give away protons. In the case of $\mathbf{HPO_4^{2-}}$, consider the following equation:

$$\mathbf{HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow PO_4^{3-}\,(aq) + H_3O^+\,(aq)}$$

Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-}. Now consider the following equation:

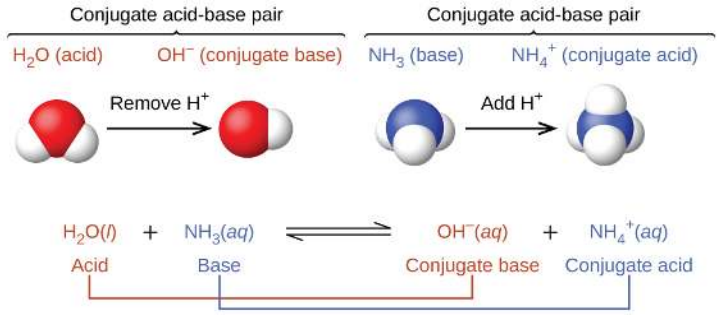
$$\mathbf{HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow H_2PO_4^- + OH^-\,(aq)}$$

In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $H_2PO_4^-$ and OH^-. Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) $\mathbf{HPO_4^{2-}}$.**

Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, $\mathbf{OH^-}$, and the conjugate acid of ammonia, $\mathbf{NH_4^+}$:



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 $\boldsymbol{H_2O/OH^-}$.

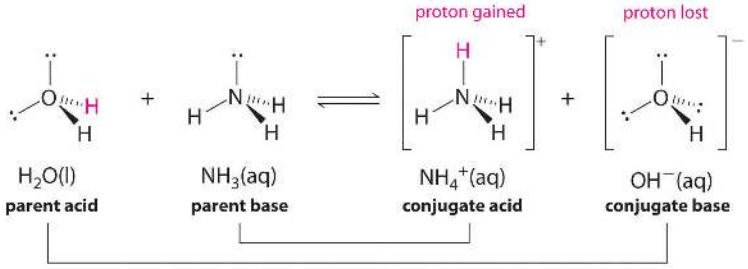


Figure **10.3.1**. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

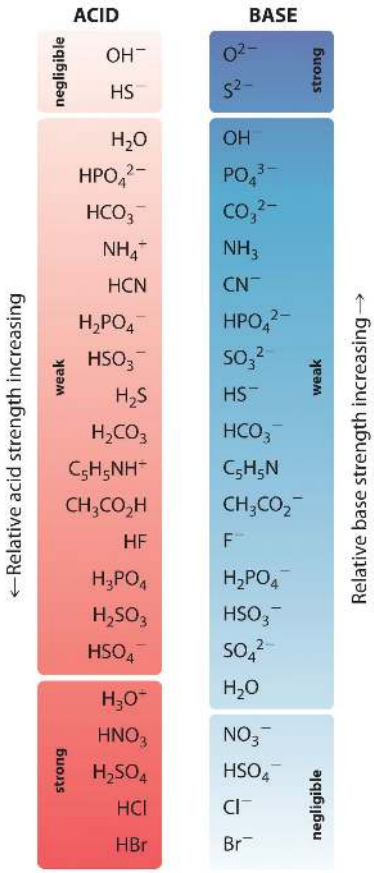


Figure 10.3.1: The Relative Strengths of Some Common Conjugate Acid–Base Pairs

The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

Example 10.3.2

Identify the conjugate acid-base pairs in this equilibrium.

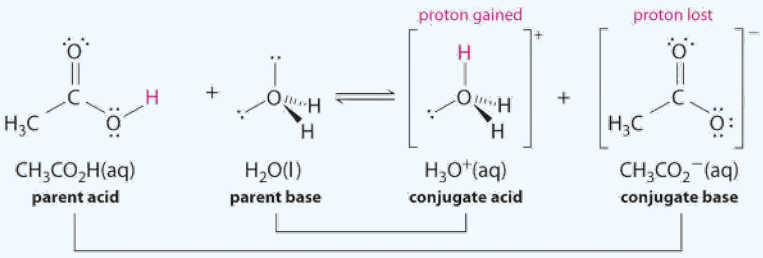
$$CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-$$

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).



https://chem.libretexts.org/@go/page/105390

Example 10.3.3

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 10.3.3

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)

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

10.3: Molecular Definitions of Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.4: Reactions of Acids and Bases

Learning Objectives

- Write acid-base neutralization reactions

Neutralization Reactions

What happens when an acid such as HCl is mixed with a base such as NaOH:

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)} \tag{10.4.1}$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than H^+ and a negative ion other than the hydroxide ion, OH^-. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$\mathrm{H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow Na^+\,(aq) + Cl^-\,(aq) + H_2O\,(l)} \tag{10.4.2}$$

Removing the spectator ions we get the net ionic equation:

$$\mathrm{H^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O\,(l)} \tag{10.4.3}$$

When a strong acid and a strong base are combined in the proper amounts - when $[H^+]$ equals $[OH^-]$\) - a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. We will not go into details here, but generally if a strong acid is mixed with a weak base there the resulting solution will be slightly acidic; if a strong base is mixed with a weak acid the solution will be slightly basic.



Example 10.4.1: propionic acid + calcium hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$].

Solution

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$
Balance the equation. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$ to provide H^+ ions.	$\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$

Exercise 10.4.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer

$$\mathrm{Ba(OH)_2\,(s) + 2\,CH_3CO_2H\ (aq) \longrightarrow Ba(CH_3CO_2)_2\,(aq) + 2\,H_2O\,(l)} \tag{10.4.4}$$

Acids and Bases React with Metals

Acids react with most metals to form a salt and hydrogen gas. As discussed in Chapter 7, metals that are more active than acids can undergo a single diplacement reaction.

For example, zinc metal reacts with hydrochloric acid producing zinc chloride and hydrogen gas.

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)} \tag{10.4.5}$$

Bases react with certain metals like zinc or aluminum for example to also produce hydrogen gas. Sodium hydroxide reacts with zinc and water to form sodium zincate and hydrogen gas.

$$\mathrm{Zn\,(s) + 2\,NaOH\,(aq) + 2\,H_2O\,(l) \longrightarrow Na_2Zn(OH)_4\,(aq) + H_2\,(g).} \tag{10.4.6}$$

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

10.4: Reactions of Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

10.5: Acid–Base and Gas Evolution Reactions

Learning Objectives

- Identify when a reaction will evolve a gas.

Neutralization Reactions

Acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen ion containing compound and the base is a hydroxide ion containing compound, water is also a product. The general reaction is as follows:

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{salt} \tag{10.5.1}$$

The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:

$$\mathrm{NaOH\,(aq) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O\ (\ell)} \tag{10.5.2}$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$\mathrm{2\,NaOH\,(aq) + H_2SO_4\,(aq) \rightarrow Na_2SO_4\,(aq) + 2\,H_2O\ (\ell)} \tag{10.5.3}$$

Example 10.5.1: Neutralizing Nitric Acid

Nitric acid (HNO_3(aq)) can be neutralized by calcium hydroxide ($Ca(OH)_2$(aq)). Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Solution

		Solution
Write the unbalanced equation.	This is a double displacement reaction, so the cations and anions swap to create new products.	$Ca(OH)_2(aq) + HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(\ell)$
Balance the equation.	Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of HNO_3 to provide H^+ ions	$Ca(OH)_2(aq) + \underline{2}HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + \underline{2}H_2O(\ell)$
Additional step: Identify the salt		The salt formed is calcium nitrate.

Exercise 10.5.1

Hydrocyanic acid ($\mathrm{HCN\,(aq)}$) can be neutralized by potassium hydroxide ($\mathrm{KOH\,(aq)}$). Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Answer

$$\mathrm{KOH\,(aq) + HCN\,(aq) \longrightarrow KCN\,(aq) + H_2O(\ell)}$$

Gas Evolving Reactions

A gas evolution reaction is a chemical process that produces a gas, such as oxygen or carbon dioxide. In the following examples, an acid reacts with a carbonate, producing salt, carbon dioxide, and water, respectively. For example, nitric acid reacts with sodium carbonate to form sodium nitrate, carbon dioxide, and water (Table 10.5.1):

$$\mathrm{2\,HNO_3(aq) + Na_2CO_3(aq) \longrightarrow 2\,NaNO_3(aq) + CO_2(g) + H_2O(l)} \tag{10.5.4}$$

Sulfuric acid reacts with calcium carbonate to form calcium sulfate, carbon dioxide, and water:

$$\mathrm{H_2SO_4\,(aq) + CaCO_3\,(aq) \longrightarrow CaSO_4\,(aq) + CO_2(g) + H_2O\,(l)} \tag{10.5.5}$$

Hydrochloric acid reacts with calcium carbonate to form calcium chloride, carbon dioxide, and water:

$$\mathrm{2\,HCl\,(aq) + CaCO_3\,(aq) \longrightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)} \tag{10.5.6}$$

Figure 10.5.1 demonstrates this type of reaction:

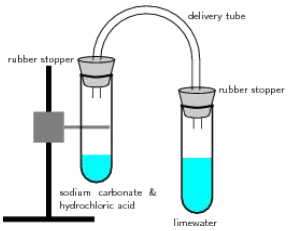


Figure 10.5.1: Reaction of acids with carbonates. In this reaction setup, lime water is poured into one of the test tubes and sealed with a stopper. A small amount of hydrochloric acid is carefully poured into the remaining test tube. A small amount of sodium carbonate is added to the acid, and the tube is sealed with a rubber stopper. The two tubes are connected. As a result of the acid-carbonate reaction, carbon dioxide is produced and the lime water turns milky.

In this reaction setup, lime water is poured into one of the test tubes and sealed with a stopper. A small amount of hydrochloric acid is carefully poured into the remaining test tube. A small amount of sodium carbonate is added to the acid, and the tube is sealed with a rubber stopper. The two tubes are connected. As a result of the acid-carbonate reaction, carbon dioxide is produced and the lime water turns milky.

Table 10.5.1: Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
sulfide	none	H_2S	$2\,HCl\,(aq) + K_2S \rightarrow H_2S\,(g) + 2\,KCl\,(aq)$
carbonates and bicarbonates	H_2CO_3	CO_2	$2\,HCl\,(aq) + K_2CO_2 \rightarrow H_2O\,(l) + CO_2\,(g)$
sulfites and bisulfites	H_2SO_3	SO_2	$2\,HCl\,(aq) + K_2SO_2 \rightarrow H_2O\,(l) + SO_2\,(g)$
ammonia	NH_4OH	NH_3	$NH_4Cl\,(aq) + KOH \rightarrow H_2O\,(l) + NH_3\,(g)$

The test tube on the right contains limewater (a solution of calcium hydroxide, $Ca(OH)_2$). On the left, a solution of hydrochloric acid has been added to a solution of sodium carbonate to generate $CO_2(g)$. The test tubes are sealed with rubber stoppers and connected with a delivery tube. As the reaction proceeds, the limewater on the right turns from clear to milky; this is due to the $CO_2(g)$ reacting with the aqueous calcium hydroxide to form calcium carbonate, which is only slightly soluble in water. The entire experiment is illustrated in the following video:



***Video* 10.5.1:** *Carbon Dioxide (CO_2) & Limewater (Chemical Reaction). Carbon dioxide (formed from bicarbonate) and an acid react with limewater, turning it from clear to milky.*

When this experiment is repeated with nitric or sulfuric acid instead of HCl, it yields the same results: the clear limewater turns milky, indicating the production of carbon dioxide. Another method to chemically generate gas is the oxidation of metals in acidic solutions. This reaction will yield a metal salt and hydrogen gas.

$$2\,HCl\,(aq) + Zn\,(s) \rightarrow ZnCl_2\,(aq) + H_2\,(g) \tag{10.5.7}$$

Here, hydrochloric acid oxidizes zinc to produce an aqueous metal salt and hydrogen gas bubbles. Recall that oxidation refers to a loss of electrons, and reduction refers to the gain of electrons. In the above redox reaction, neutral zinc is oxidized to Zn^{2+}, and the acid, H^+, is reduced to $H_2(g)$. The oxidation of metals by strong acids is another common example of a gas evolution reaction.

Contributors and Attributions

- Boundless (www.boundless.com)
- Wikipedia (CC-BY-SA-3.0)
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

10.5: Acid–Base and Gas Evolution Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

11: Redox Reactions

Template:HideTOC

The rusting of an old car. A burning campfire. A toy battery-operated car. The chemical processes in your body that break down carbohydrates to produce water, carbon dioxide and energy. The ripening of fruit. It's not easy to see what all of these types of reactions have in common, but they all belong to a very important category of chemical reactions known as oxidation-reduction (or redox) reactions.

11.1: Oxidation and Reduction - Some Definitions
11.2: Electronegativity
11.3: Oxidation States: Electron Bookkeeping
11.4: Oxidation States & Redox Reactions
11.5: Half-Reactions
11.6: The Activity Series: Predicting Spontaneous Redox Reactions
11.7: Combustion Reactions

11: Redox Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.1: Oxidation and Reduction - Some Definitions

"Redox" is short for "oxidation and reduction", two complimentary types of chemical reactions. The term **oxidation** originally referred to substances combining with oxygen, as happens when an iron bar rusts or a campfire log burns. We often refer to these two examples as corrosion and combustion. **Reduction** originally referred to the process of converting metal ores to pure metals, a process that is accompanied by a reduction in the mass of the ore.

These two terms have broader meanings now. In all oxidation-reduction reactions an exchange of electrons occurs - one substance loses electrons while something else gains them. That is the key to understanding redox reactions. We'll define these terms below.



Figure **11.1.1***: Reaction of copper wire in a silver nitrate solution.*

A simple demonstration of a redox reaction involves placing a solid piece of copper wire in a silver nitrate solution. Within minutes the wire begins to look fuzzy or furry, as small silver crystals begin to form on the wire. Meanwhile, the originally clear silver nitrate solution begins to take on a pale bluish tint. Furthermore, if the crystals are shaken off of the wire we see that the wire partially disintegrated.

The overall equation for our demonstration describes the events:

$$\mathrm{Cu\,(s) + 2\,AgNO_3\,(aq) \longrightarrow Cu(NO_3)_2\,(aq) + 2\,Ag\,(s)} \tag{11.1.1}$$

Remember that when we have aqueous solutions of ionic compounds, the ions are really present as separate ions, not as bonded particles. So we can write an expanded equation:

$$\mathrm{Cu\,(s) + 2\,Ag^+\,(aq) + 2\,NO_3^-\,(aq) \longrightarrow Cu^{2+}\,(aq) + 2\,NO_3^-\,(aq) + 2\,Ag\,(s)} \tag{11.1.2}$$

Finally we remove spectator ions from the equation. Remember that spectator ions are ions that undergo no change during a reaction. Notice the nitrate ions, NO_3^- – they start as aqueous ions and end up exactly the same.

Removing the spectator ions gives us our **net ionic equation:**

$$\mathrm{Cu\,(s) + 2\,Ag^+\,(aq) \longrightarrow Cu^{2+}\,(aq) + 2\,Ag\,(s)} \tag{11.1.3}$$

We can now see a bit more clearly what changes are occurring with this reaction.

Oxidation of Copper Metal to Make Copper Ions

Copper began as a neutral atom with no charge, but changed into an ion with a charge of +2. An atom becomes a positive ion by losing electrons:

$$\mathrm{Cu\,(s) \longrightarrow Cu^{2+}\,(aq) + 2\,e^-} \tag{11.1.4}$$

Notice that copper began as a solid but is converted into aqueous ions - this is why the copper wire disintegrates. We say that copper was **oxidized** because is has lost electrons (i.e., electrons appear on the product side of the Equation 11.1.4).

Reduction of Silver Ions to Make Silver Metal

Silver was converted from an ion with a charge of +1, Ag+, to a neutral atom, Ag. The only way an ion can undergo this change is to gain an electron:

$$Ag^+(aq) + e^- \longrightarrow Ag(s) \quad (11.1.5)$$

Notice that solid silver is formed - this is what causes the fuzzy appearance to begin appearing on the wire - solid silver crystals. Silver has gained electrons - it has been reduced (i.e., electrons appear on the reactant side of Equation 11.1.5).

The electrons that silver gained had to come from somewhere - they came from copper. Conversely, a substance such as copper can only lose electrons if there is something else that will take them up, the silver ions. **One cannot occur without the other**. This exchange of electrons is what defines an **oxidation - reduction reaction**.

Definition: Oxidation

oxidation is the **loss** of electrons.

Definition: Reduction

Reduction the **gain** of electrons

You will be learning several new terms in this unit and it is important that you learn them very quickly. You may find it useful to have some tricks to help you remember these terms including the LEO/GER.

LEO the Lion Says GER

- LEO: Loss of Electrons is Oxidation
- GER: Gain of Electrons is Reduction



11.1: Oxidation and Reduction - Some Definitions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.2: Electronegativity

Linus Pauling described electronegativity as “the power of an atom in a molecule to attract electrons to itself.”[1] Basically, the electronegativity of an atom is a relative value of that atom's ability to attract election density toward itself when it bonds to another atom. The higher the electronegative of an element, the more that atom will attempt to pull electrons towards itself and away from any atom it bonds to. The main properties of an atom dictate it's electronegativity are it's atomic number as well as its atomic radius. The trend for electronegativity is to increase as you move from left to right and bottom to top across the periodic table. This means that the most electronegative atom is Fluorine and the least electronegative is Francium.

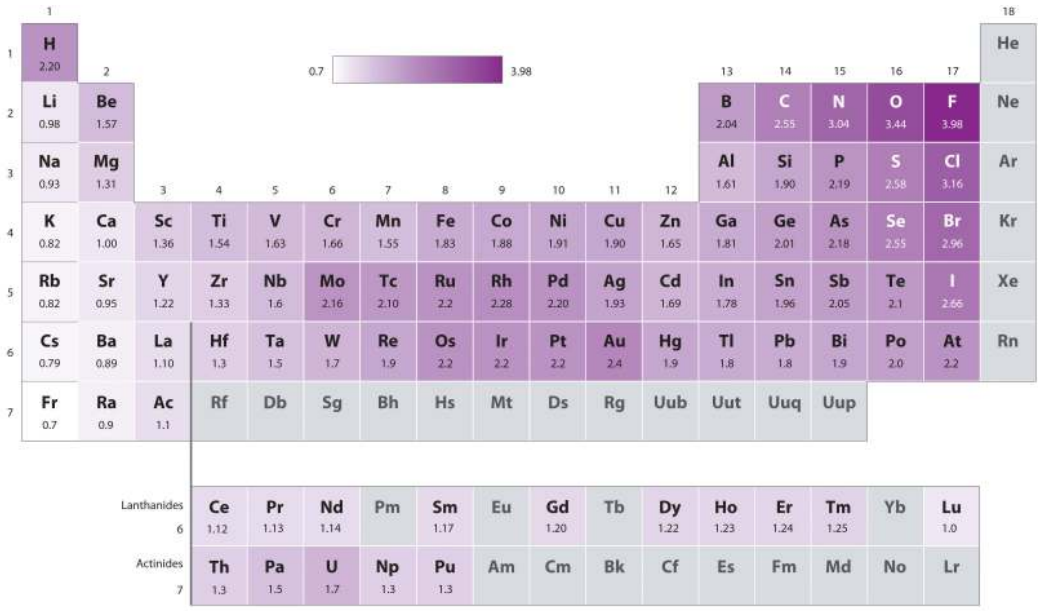
There are a few different 'types' of electronegativity which differ only in their definitions and the system by which they assign values for electronegativity. For example, there is Mulliken electronegativity which is defined as "the average of the ionization energy and electron affinity of an atom"[3], which as we will see, differs slightly from Pauling's definition of electronegativity.

Pauling Electronegativity

Linus Pauling was the original scientist to describe the phenomena of electronegativity. By assigning a value of 4.0 to Fluorine (the most electronegative element), Pauling was able to set up relative values for all of the elements. This was when he first noticed the trend that the electronegativity of an atom was determined by it's position on the periodic table and that the electronegativity tended to increase as you moved left to right and bottom to top along the table. The range of values for Pauling's scale of electronegativity ranges from Fluorine (most electronegative = 4.0) to Francium (least electronegative = 0.7). Furthermore, if the electronegativity difference between two atoms is very large, then the bond type tends to be more ionic, however if the difference in electronegativity is small then it is a nonpolar covalent bond.

Linus Pauling (1901-1994)

When he was nine, Pauling’s father died, and his mother tried to convince him to quit school to support the family. He did not quit school, but was later denied a high school degree, and had to work several jobs to put himself through college. Pauling would go on to become one of the most influential chemists of the century if not all time. He won two Nobel Prizes, one for chemistry in 1954 and one for peace in 1962.

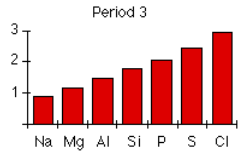


The attraction that a bonding pair of electrons feels for a particular nucleus depends on:

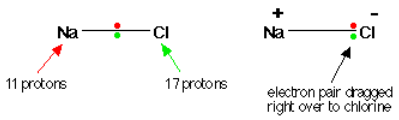
- the number of protons in the nucleus;
- the distance from the nucleus;
- the amount of screening by inner electrons.

Trends in electronegativity across a period

The positively charged protons in the nucleus attract the negatively charged electrons. As the number of protons in the nucleus increases, the electronegativity or attraction will increase. Therefore electronegativity **increases** from **left to right** in a row in the periodic table. This effect only holds true for a row in the periodic table because the attraction between charges falls off rapidly with distance. The chart shows electronegativities from sodium to chlorine (ignoring argon since it does not does not form bonds).



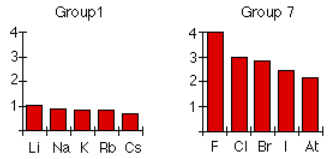
Why does electronegativity increase across a period? Consider sodium at the beginning of period 3 and chlorine at the end (ignoring the noble gas, argon). Think of sodium chloride as if it were covalently bonded.



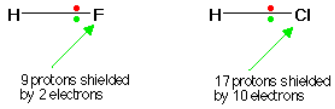
Both sodium and chlorine have their bonding electrons in the 3-level. The electron pair is screened from both nuclei by the 1s, 2s and 2p electrons, but the chlorine nucleus has 6 more protons in it. It is no wonder the electron pair gets dragged so far towards the chlorine that ions are formed. Electronegativity increases across a period because the number of charges on the nucleus increases. That attracts the bonding pair of electrons more strongly.

Trends in electronegativity down a group

As you go down a group, electronegativity decreases. (If it increases up to fluorine, it must decrease as you go down.) The chart shows the patterns of electronegativity in Groups 1 and 7.



Why does electronegativity fall as you go down a group? As you go **down** a group, electronegativity **decreases** because the bonding pair of electrons is increasingly distant from the attraction of the nucleus. Consider the hydrogen fluoride and hydrogen chloride molecules:



The bonding pair is shielded from the fluorine's nucleus only by the $1s^2$ electrons. In the chlorine case it is shielded by all the $1s^2 2s^2 2p^6$ electrons. In each case there is a net pull from the center of the fluorine or chlorine of +7. But fluorine has the bonding pair in the 2-level rather than the 3-level as it is in chlorine. If it is closer to the nucleus, the attraction is greater.

? Exercise 11.2.1

Predict the order or increasing electronegativity from the following elements

a. F, Li, C, O
b. Te, Cl, S, Se
c. Cs, At, Tl, I

References

1. Zumdahl, Steven S. "Chemical Principles" 5th Edition. Houghton Mifflin Company 2005. Chapter 13.2 "Electronegativity" pgs. 587-590.
2. Housecroft, Catherine E. et. al. "Inorganic Chemistry" 3rd Edition. Pearson Education Limited 2008. CHampter 2.5 "Electronegativity Values" pgs. 42-44
3. International Union of Pure and Applied Chemistry. "Electronegativity". http://goldbook.iupac.org/E01990.html.

Contributors and Attributions

- Jim Clark (Chemguide.co.uk)

11.2: Electronegativity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- **Pauling Electronegativity** by Matthew Salem is licensed CC BY 4.0.
- **Current page** is licensed CC BY-NC-SA 4.0.

11.3: Oxidation States: Electron Bookkeeping

Redox reactions are all about electrons being transferred from one substance to another, so it would be useful if we had a system for keeping track of what gains and what loses electrons, and how many electrons are involved. We do - our record-keeping system is called **Oxidation Numbers.** You may remember electronegativity from the previous section. When determining the oxidation number on an individual atom, the shared electrons are assigned to the most electronegative atom.

Electronegativity

The ability of a bonded atom to attract shared electrons towards itself

When two atoms of different elements are bonded together by a covalent bond (sharing electrons), the electrons are generally not shared equally between the two atoms due to differences in their electronegativity. Think of this as a tug-of-war for electrons. Sometimes both atoms pull with equal strength on shared electrons; other times there is a clearly stronger player who will pull the electrons closer to itself.

Consider the bond between a hydrogen atom (with one valence electron) and an oxygen atom (with its six valence electrons):



Because oxygen has a higher electronegativity than hydrogen, the shared electrons are closer to the oxygen atom than to the hydrogen atom. This is not the total transfer of electrons that would create an ion, but partial charges do form - the hydrogen end of the bond is partially positive (+1) because it has partially lost one electron, and the oxygen end of the H–O is partially negative (-1) because it has partially gained the one electron from hydrogen:



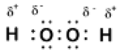
Our molecule is incomplete however, because there is a lone electron around oxygen. Let's add one more hydrogen to complete our water molecule:



We see that each of the two hydrogen has "lost" one electron to oxygen. Oxygen has "gained" two electrons - one from each hydrogen. Again, these are not true ions but it is useful to think of them in the same way.

Charges given to atoms in a molecule in this way are called **oxidation numbers**. We can use oxidation numbers to keep track of where electrons are in a molecule, and how they move during a reaction. In our water example, hydrogen is assigned an oxidation number of +1 because each individual hydrogen has "lost" one electron. Oxygen has an oxidation number of +2 because the single oxygen atom has "gained" a total of two electrons, one from each hydrogen.

Here is another molecule involving hydrogen and oxygen - hydrogen peroxide, H_2O_2:



In hydrogen peroxide, each hydrogen still has an oxidation number of +1 because each hydrogen "gives up" a single electron to oxygen. Oxygen, however, now has an oxidation number of -1 because each oxygen gains just one electron from its neighboring hydrogen. The electrons between the two identical oxygen atoms are shared equally, so there is no partial charge resulting from that bond.

Oxidation Numbers

A positive or negative number assigned to an atom in a molecule or ion that reflects a partial gain or loss of electrons

Knowing the oxidation number of each individual element in a molecule will be a key step in our understanding of redox reactions. Fortunately it will not usually involve drawing electron dot

diagrams. Instead there are a series of rules that we can use to determine oxidation numbers. Here are the main rules:

1. The oxidation number of a pure element (by itself, and not an ion) is zero.

Element	Oxidation Number
Na	0
H_2	0
O_2	0
	0

2. The oxidation number of a monatomic ion (by itself or as part of an ionic compound) is equal to its charge.
Alkali metals - elements in the first column of the periodic table - will always have an oxidation number of +1; Alkali metals (column 2) are almost always +2

Ionic Compound	Ions	Charge
NaCl	Na^+	+1
	Cl^-	-1
Mg_3N_2	Mg^{+2}	+2
	N^{-3}	-3

3. The oxidation number of hydrogen is almost always +1 when it is in a compound.

Compound	element	O: N
HCl	H	
	Cl	
H_2S	H	
	S	

4. The oxidation number of oxygen is almost always -2 when it is in a compound
The exceptions:
- peroxides, such as hydrogen peroxide. In peroxides oxygen has an oxidation number of -1.
- when oxygen is combined with fluorine it's oxidation number is +2

Compound	element	O: N
MgO magnesium oxide	Mg	
	O	
Na_2O sodium oxide	Na	
	O	
Na_2O_2 sodium peroxide	Na	
	O	

5. The sum of the oxidation numbers in a compound is zero.
To determine the oxidation number of Mn in Mn_2O_7 we must work backwards:
- We know each oxygen is -2 (Rule 4)
- 7 oxygen gives a total of:

-2 × 7 atoms = -14 total

Since the sum of oxidation numbers must be zero, the total oxidation number of Mn must be +14 to cancel out oxygen's -14, but since there are 2 Mn atoms, each **individual atom** will have an oxidation number of +7:

$$\frac{+14 \text{ total}}{2 \text{ Mn atoms}} = +7$$

Compound	element	Oxidation Number	Number atoms
Mg_3N_2	Mg	+2	3
	N	-3	2
		SUM	
Mn_2O_7	Mn	+7	2
	O	-2	7
		SUM	
Cl_2O_3	Cl	+3	2
	O	-2	3
		SUM	

6. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on that ion.
Again, work backwards to determine the oxidation number of any non-oxygen or non-hydrogen atom.
To determine the oxidation number of Cr in $Cr_2O_7^{2-}$:
- Oxygen will be -2 (Rule 4), for a total of:

Compound	element	Oxidation Number	Number atoms
NO_3^-	N	+5	1

-2 × 7 = -14

- Since the sum of the oxidation numbers will be -2 (the charge on the entire ion), the total for all Cr must be +12 because:

+12 + (-14) = -2

- Since there is are two Cr, each Cr will have an oxidation number of +6

= +7

$$\frac{+12}{2} = +6$$

	O	-2	3	
		SUM		
$Cr_2O_7^{2-}$	Cr	+6	2	
	O	-2	7	
		SUM		
SO_4^{2-}	S	+6	1	
	O	-2	4	
		SUM		

It is important to note that oxidation number always refers to ***each individual atom*** in the compound, not to the total for that element. For example in H_2O - the total positive "charge" for both hydrogen atoms will be +2 (which balances with the -2 from oxygen), but EACH HYDROGEN has an oxidation number of +1.

Example 11.3.1: Assigning Oxidation Numbers

Assign oxidation numbers to the atoms in each substance.

a. Br_2
b. SiO_2
c. $Ba(NO_3)_2$

Solution

a. Br_2 is the elemental form of bromine. Therefore, by rule 1, each atom has an oxidation number of 0.
b. By rule 3, oxygen is normally assigned an oxidation number of −2. For the sum of the oxidation numbers to equal the charge on the species (which is zero), the silicon atom is assigned an oxidation number of +4.
c. The compound barium nitrate can be separated into two parts: the Ba^{2+} ion and the nitrate ion. Considering these separately, the Ba^{2+} ion has an oxidation number of +2 by rule 2. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of −2, and there are three oxygens. According to rule 4, the sum of the oxidation number on all atoms must equal the charge on the species, so we have the simple algebraic equation

$$x + 3(-2) = -1$$

where x is the oxidation number of the nitrogen atom and −1 represents the charge on the species. Evaluating,

$$x + (-6) = -1 x = +5$$

Thus, the oxidation number on the N atom in the nitrate ion is +5.

Exercise 11.3.1: Oxidation States in Phosphoric Acids

Assign oxidation numbers to the atoms in H_3PO_4.

Answer

H = +1, O = −2, P = +5

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction are thus also defined in terms of increasing or decreasing oxidation numbers, respectively.

Example 11.3.2: Oxidation States in Sodium Chloride

Identify what is being oxidized and reduced in this redox equation:

$$2\,Na + Cl_2 \longrightarrow 2\,NaCl$$

Solution

Consider the reactants. Because both reactants are the elemental forms of their atoms, the Na and Cl atoms as reactants have oxidation numbers of 0. In the ionic product, the sodium ions have an oxidation number of +1, while the chloride ions have an oxidation number of −1:

We note that the sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; chlorine is decreasing its oxidation number from 0 to −1, so it is being reduced:

Because oxidation numbers are changing, this is a redox reaction. Note that the total number of electrons lost by the sodium (two, one lost from each atom) is gained by the chlorine atoms (two, one gained for each atom).

Exercise 11.3.2

Identify what is being oxidized and reduced in this redox equation.

$$C + O_2 \longrightarrow CO_2$$

Answer

C is being oxidized from 0 to +4; O is being reduced from 0 to −2.

In this introduction to oxidation-reduction reactions, we are using straightforward examples. However, oxidation reactions can become quite complex; the following equation represents a redox reaction:



To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced? This is also an example of a net ionic reaction; spectator ions that do not change oxidation numbers are not displayed in the equation.

Contributors and Attributions

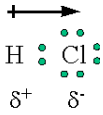
•

11.3: Oxidation States: Electron Bookkeeping is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.4: Oxidation States & Redox Reactions

Electron transfer is one of the most basic processes that can happen in chemistry. It simply involves the movement of an electron from one atom to another. Many important biological processes rely on electron transfer, as do key industrial transformations used to make valuable products. In biology, for example, electron transfer plays a central role in respiration and the harvesting of energy from glucose, as well as the storage of energy during photosynthesis. In society, electron transfer has been used to obtain metals from ores since the dawn of civilization.

Oxidation state is a useful tool for keeping track of electron transfers. It is most commonly used in dealing with metals and especially with transition metals. Unlike metals from the first two columns of the periodic table, such as sodium or magnesium, transition metals can often transfer different numbers of electrons, leading to different metal ions (e.g., sodium is generally found as Na^+ and magnesium is almost always Mg^{2+}, but manganese could be Mn^{2+}, Mn^{3+}, and so on, as far as Mn^{7+}). Oxidation state is a number assigned to an element in a compound according to some rules. This number enable us to describe oxidation-reduction reactions, and balancing redox chemical reactions. When a covalent bond forms between two atoms with different electronegativities the shared electrons in the bond lie closer to the more electronegative atom:



The **oxidation number** of an atom is the charge that results when the electrons in a covalent bond are assigned to the more electronegative atom and is the charge an atom would possess ***if*** the bonding were ionic. In HCl (above) the oxidation number for the hydrogen would be +1 and that of the Cl would be -1.

Example 11.4.1

Determine which element is oxidized and which element is reduced in the following reactions (be sure to include the Oxidation State of each):

a. $\mathbf{Zn + 2\,H^+ \longrightarrow Zn^{2+} + H_2}$
b. $\mathbf{2\,Al + 3\,Cu^{2+} \longrightarrow 2\,Al^{3+} + 3\,Cu}$
c. $\mathbf{CO_3^{2-} + 2\,H^+ \longrightarrow CO_2 + H_2O}$

Solution

a. Zn is oxidized (Oxidation number: 0 → +2); H^+ is reduced (Oxidation number: +1 → 0)
b. Al is oxidized (Oxidation number: 0 → +3); Cu^{2+} is reduced (+2 → 0)
c. This is not a redox reaction because each element has the same oxidation number in both reactants and products: O= -2, H= +1, C= +4.

An atom is oxidized if its oxidation number increases, and an atom is reduced if its oxidation number decreases. The atom that is oxidized is the reducing agent, and the atom that is reduced is the oxidizing agent. (Note: the oxidizing and reducing agents can be the same element or compound).

Oxidation Numbers and Nomenclature

Compounds of the alkali (oxidation number +1) and alkaline earth metals (oxidation number +2) are typically ionic in nature. Compounds of metals with higher oxidation numbers (e.g., tin +4) tend to form molecular compounds

- In ionic and covalent molecular compounds usually the ***less electronegative element is given first***.
- In ionic compounds the names are given which refer to ***the oxidation (ionic) state***
- In molecular compounds the names are given which refer ***to the number of molecules present in the compound***

Figure 11.4.1: Example of nomenclature based on oxidation states.

Ionic		Molecular	
MgH_2	magnesium hydride	H_2S	dihydrogen sulfide
FeF_2	iron(II) fluoride	OF_2	oxygen difluoride
Mn_2O_3	manganese(III) oxide	Cl_2O_3	dichlorine trioxide

An oxidation-reduction (redox) reaction is a type of chemical reaction that involves a transfer of electrons between two species. An oxidation-reduction reaction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron. Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting.

Oxidation-Reduction Reaction Examples

Redox reactions are comprised of two parts, a reduced half and an oxidized half, that *always* occur together. The reduced half gains electrons and the oxidation number decreases, while the oxidized half loses electrons and the oxidation number increases. Simple ways to remember this include the mnemonic devices **OIL RIG,** meaning "*oxidation is loss*" and "*reduction is gain*," and **LEO says GER,** meaning "*loss of* e^- = *oxidation*" and "*gain of* e^- = *reduced*." There is no net change in the number of electrons in a redox reaction. Those given off in the oxidation half reaction are taken up by another species in the reduction half reaction.

The two species that exchange electrons in a redox reaction are given special names. The ion or molecule that accepts electrons is called the **oxidizing agent**; by accepting electrons it causes the oxidation of another species. Conversely, the species that donates electrons is called the **reducing agent**; when the reaction occurs, it reduces the other species. In other words, what is oxidized is the reducing agent and what is reduced is the oxidizing agent. (Note: the oxidizing and reducing agents can be the same element or compound, as in disproportionation reactions).



Figure **11.4.1**: A thermite reaction taking place on a cast iron skillet. A thermite reaction, using about 110 g of the mixture, taking place. The cast-iron skillet was destroyed in the process. Image by Schuyler S and used with permission by Wikipedia.

A good example of a redox reaction is the thermite reaction, in which iron atoms in ferric oxide lose (or give up) O atoms to Al atoms, producing Al_2O_3 (Figure 20.1.1).

$$Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(l)} \tag{11.4.1}$$

Another example of the redox reaction is the reaction between zinc and copper sulfate.

Example 11.4.2: Identifying Oxidized Elements

Using the equations from the previous examples, determine what is oxidized in the following reaction.

$$Zn + 2H^+ \rightarrow Zn^{2+} + H_2 \tag{11.4.2}$$

Solution

The Oxidation State of H changes from +1 to 0, and the Oxidation State of Zn changes from 0 to +2. Hence, Zn is oxidized and acts as the reducing agent.

Example 11.4.3: Identifying Reduced Elements

What is reduced species in this reaction?

$$Zn + 2H^+ \rightarrow Zn^{2+} + H_2 \tag{11.4.3}$$

Solution

The Oxidation State of H changes from +1 to 0, and the Oxidation State of Zn changes from 0 to +2. Hence, H+ ion is reduced and acts as the oxidizing agent.

Combination Reactions

Combination reactions are among the simplest redox reactions and, as the name suggests, involves "combining" elements to form a chemical compound. As usual, oxidation and reduction occur together. The general equation for a combination reaction is given below:

$$A + B \rightarrow AB \tag{11.4.4}$$

Example 11.4.4: Combination Reaction

Equation: $H_2 + O_2 \rightarrow H_2O$

Calculation: $0 + 0 \rightarrow$ **(2)(+1) + (-2) = 0**

Explanation:

In this equation both H_2 and O_2 are free elements; following Rule #1, their Oxidation States are 0. The product is H_2O, which has a total Oxidation State of 0. According to Rule #6, the Oxidation State of oxygen is usually -2. Therefore, the Oxidation State of H in H_2O must be +1.

Decomposition Reactions

A decomposition reaction is the reverse of a combination reaction, the breakdown of a chemical compound into individual elements:

$$AB \rightarrow A + B \tag{11.4.5}$$

Example 11.4.5: Decomposition Reaction

Consider the decomposition of water:

$$H_2O \rightarrow H_2 + O_2 \tag{11.4.6}$$

Calculation:

$$(2)(+1) + (-2) = 0 \rightarrow 0 + 0 \tag{11.4.7}$$

In this reaction, water is "decomposed" into hydrogen and oxygen. As in the previous example the H_2O has a total Oxidation State of 0; thus, according to Rule #6 the Oxidation State of oxygen is usually -2, so the Oxidation State of hydrogen in H_2O must be +1.

Single Replacement Reactions

A single replacement reaction involves the "replacing" of an element in the reactants with another element in the products:

$$A + BC \rightarrow AB + C \tag{11.4.8}$$

Example 11.4.6: Single Replacement Reaction

Equation:

$$Cl_2 + NaBr \rightarrow NaCl + Br_2 \tag{11.4.9}$$

Calculation: (0) + ((+1) + (-1) = 0) -> ((+1) + (-1) = 0) + 0
Explanation: In this equation, Br is replaced with Cl, and the Cl atoms in Cl_2 are reduced, while the Br ion in NaBr is oxidized.

Double Replacement Reactions

A double replacement reaction is similar to a double replacement reaction, but involves "replacing" two elements in the reactants, with two in the products:

$$AB + CD \rightarrow AD + CB \tag{11.4.10}$$

Example 11.4.7: Double Replacement Reaction

The reaction of gasous hydrogen chloride and iron oxide is a double replacement reaction. Write the expected reaction for this chemistry equation.

Solution

$$\mathrm{Fe_2O_3 + 6\,HCl \longrightarrow 2\,FeCl_3 + 3\,H_2O} \tag{11.4.11}$$

In this equation, Fe and H trade places, and oxygen and chlorine trade places.

Combustion Reactions

Combustion reactions almost always involve oxygen in the form of O_2, and are almost always exothermic, meaning they produce heat. Chemical reactions that give off light and heat and light are colloquially referred to as "burning."

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O \tag{11.4.12}$$

Although combustion reactions typically involve redox reactions with a chemical being oxidized by oxygen, many chemicals "burn" in other environments. For example, both titanium and magnesium burn in nitrogen as well:

$$2Ti_{(s)} + N_{2(g)} \rightarrow 2TiN_{(s)} \tag{11.4.13}$$

$$3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)} \tag{11.4.14}$$

Moreover, chemicals can be oxidized by other chemicals than oxygen, such as Cl_2 or F_2; these processes are also considered combustion reactions

Disproportionation Reactions

Disproportionation Reactions: In some redox reactions a single substance can be both oxidized and reduced. These are known as disproportionation reactions, with the following general equation:

$$2A \rightarrow A^{+n} + A^{-n} \tag{11.4.15}$$

Where n is the number of electrons transferred. Disproportionation reactions do not need begin with neutral molecules, and can involve more than two species with differing oxidation states (but rarely).

Example 11.4.8: Disproportionation Reaction

Disproportionation reactions have some practical significance in everyday life, including the reaction of hydrogen peroxide, H_2O_2 poured over a cut. This a decomposition reaction of hydrogen peroxide, which produces oxygen and water. Oxygen is present in all parts of the chemical equation and as a result it is both oxidized and reduced. The reaction is as follows:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g) \tag{11.4.16}$$

Explanation: On the reactant side, H has an Oxidation State of +1 and O has an Oxidation State of -1, which changes to -2 for the product H_2O (oxygen is reduced), and 0 in the product O_2 (oxygen is oxidized).

Summary

Oxidation signifies a loss of electrons and reduction signifies a gain of electrons. Balancing redox reactions is an important step that changes in neutral, basic, and acidic solutions. Tthe types of redox reactions: Combination and decomposition, Displacement reactions (single and double), Combustion, Disproportionation. The oxidizing agent undergoes reduction and the reducing agent undergoes oxidation.

References

1. Petrucci, et al. *General Chemistry: Principles & Modern Applications*. 9th ed. Upper Saddle River, New Jersey: Pearson/Prentice Hall, 2007.
2. Sadava, et al. *Life: The Science of Biology*. 8th ed. New York, NY. W.H. Freeman and Company, 2007

Contributors and Attributions

- Christopher Spohrer (UCD), Christina Breitenbuecher (UCD), Luvleen Brar (UCD)
- Chris P Schaller, Ph.D., (College of Saint Benedict / Saint John's University)

11.4: Oxidation States & Redox Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.5: Half-Reactions

Learning Objectives

- To identify a chemical reaction as an oxidation-reduction reaction.

When zinc metal is submerged into a quantity of aqueous HCl, the following reaction occurs (Figure 11.5.1):

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow H_2\,(g) + ZnCl\ 2\,(aq)} \tag{11.5.1}$$

This is one example of what is sometimes called a *single replacement reaction* because Zn replaces H in combination with Cl.



Figure 11.5.1: Zinc Metal plus Hydrochloric Acid. It is fairly obvious that zinc metal reacts with aqueous hydrochloric acid! The bubbles are hydrogen gas.

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow H_2\,(g) + Zn^{2+}\,(aq) + 2\,Cl^-\,(aq)} \tag{11.5.2}$$

Viewed this way, the net reaction seems to be a charge transfer between zinc and hydrogen atoms. (There is no net change experienced by the chloride ion.) In fact, electrons are being transferred from the zinc atoms to the hydrogen atoms (which ultimately make a molecule of diatomic hydrogen), changing the charges on both elements.

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called oxidation. The gain of electrons is called reduction. Because any loss of electrons by one substance must be accompanied by a gain in electrons by something else, oxidation and reduction always occur together. As such, electron-transfer reactions are also called oxidation-reduction reactions, or simply **redox reactions**. The atom that loses electrons is **oxidized**, and the atom that gains electrons is **reduced**. Also, because we can think of the species being oxidized as causing the reduction, the species being oxidized is called the reducing agent, and the species being reduced is called the oxidizing agent.

> *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,

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)} \tag{11.5.3}$$

zinc atoms are oxidized to Zn^{2+}. The half reaction for the oxidation reaction, omitting phase labels, is as follows:

$$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-} \tag{11.5.4}$$

This half reaction is balanced in terms of the number of zinc atoms, and it also shows the two electrons that are needed as products to account for the zinc atom losing two negative charges to become a 2+ ion. With half reactions, there is one more item to balance:

the overall charge on each side of the reaction. 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:

$$2\,H^+ + 2\,e^- \longrightarrow H_2 \tag{11.5.5}$$

There are two hydrogen atoms on each side, and the two electrons written as reactants serve to neutralize the 2+ charge on the reactant hydrogen ions. Again, the overall charge on both sides is zero.

The overall reaction is simply the combination of the two half reactions and is shown by adding them together.

$$\begin{aligned} &Zn \rightarrow Zn^{2+} + \cancel{2e^-} \\ &\cancel{2H^+ + 2e^-} \rightarrow H_2 \\ \hline &Zn + 2H^+ \rightarrow Zn^{2+} + H_2 \end{aligned}$$

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 11.5.1: 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.

Solution

We start by using symbols of the elements and ions to represent the reaction:

$$Ag^+ + Al \longrightarrow Ag + Al^{3+} \tag{11.5.6}$$

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:

$$Ag^+ + e^- \longrightarrow Ag \tag{11.5.7}$$

Aluminum is oxidized, losing three electrons to change from Al to Al^{3+}:

$$Al \longrightarrow Al^{3+} + 3\,e^- \tag{11.5.8}$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{aligned} &3\left(Ag^+ + e^- \rightarrow Ag\right) \\ &\underline{\qquad\qquad Al \rightarrow Al^{3+} + 3e^-} \\ &3Ag^+ + Al \rightarrow 3Ag + Al^{3+} \end{aligned}$$

Now the equation is balanced, not only in terms of elements but also in terms of charge.

Exercise 11.5.1

Write and balance the redox reaction that has calcium ions and potassium metal as reactants and calcium metal and potassium ions as products.

> *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:

$$\mathrm{Cd\,(s) + 2\,NiOOH\,(s) + 2\,H_2O(\ell) \longrightarrow Cd(OH)_2\,(s) + 2\,Ni(OH)\ 2\,(s)} \tag{11.5.9}$$

The cadmium is oxidized, while the nickel atoms in NiOOH are reduced. Except for the water, all the substances in this reaction are solids, allowing NiCad batteries to be recharged hundreds of times before they stop operating. Unfortunately, NiCad batteries are fairly heavy batteries to be carrying around in a pacemaker. Today, the lighter lithium/iodine battery is used instead. The iodine is dissolved in a solid polymer support, and the overall redox reaction is as follows:

$$\mathrm{2\,Li\,(s) + I_2\,(s) \longrightarrow 2\,LiI\,(s)} \tag{11.5.10}$$

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 **11.5.1**: A small button battery like this is used to power a watch, pacemaker, or calculator. from Gerhard H Wrodnigg (via Wikipedia)

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 start with the same reactant.)

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

3. a. $Ca \rightarrow Ca^{2+} + 2e^-$

   $2H^+ + 2e^- \rightarrow H_2$

   Combined: $Ca + 2H^+ \rightarrow Ca^{2+} + H_2$

   b. $Br_2 + 2e^- \rightarrow 2Br^-$

   $2I^- \rightarrow I_2 + 2e^-$

   Combined: $Br_2 + 2I^- \rightarrow 2Br^- + I_2$

11.5: Half-Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.6: The Activity Series: Predicting Spontaneous Redox Reactions

Learning Objectives

- Use the activity series to predict if a reaction will occur.

We see below two metals that can be exposed to water. The picture on the left is of sodium, which gives a violent reaction when it comes in contact with water. The picture on the right is of silver, a metal so unreactive with water that it can be made into drinking vessels.



Figure **11.6.1**: On the left, sodium reacts with water. On the right, silver in the form of cups do not react with water.

The Activity Series

Single-replacement reactions only occur when the element that is doing the replacing is more reactive than the element that is being replaced. Therefore, it is useful to have a list of elements in order of their relative reactivities. The **activity series** is a list of elements in decreasing order of their reactivity. Since metals replace other metals, while nonmetals replace other nonmetals, they each have a separate activity series. The table below is an activity series of most common metals and of the halogens.

Table 11.6.1: Activity Series of Metals in Aqueous Solutions

most active (easily oxidized – readily lose electrons))		
lithium	Li	These metals displace hydrogen from water $Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_2 + H_{2\ (g)}$ These elements are very reactive and react readily to form compounds
potassium	K	
barium	Ba	
calcium	Ca	
sodium	Na	
magnesium	Mg	These metals displace hydrogen from acids $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_2 + H_{2(g)}$
aluminum	A	
zinc	Zn	
chromium	Cr	
iron	Fe	
cadmium	Cd	
nickel	Ni	
tin	Sn	
lead	Pb	
hydrogen	H	
copper	Cu	These metals do not displace hydrogen from acids or water These elements are more stable, and form compounds less readily than do those higher in the table.
silver	Ag	
mercury	Hg	
platinum	Pt	
gold	Au	
least active		

For a single-replacement reaction, a given element is capable of replacing an element that is below it in the activity series. This can be used to predict if a reaction will occur. Suppose that small pieces of the metal nickel were placed into two separate aqueous solutions: one of iron (III) nitrate and one of lead (II) nitrate. Looking at the activity series, we see that nickel is below iron, but above lead. Therefore, the nickel metal will be capable of replacing the lead in a reaction, but will not be capable of replacing iron.

$$\mathrm{Ni}\,(s) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow \mathrm{Ni(NO_3)_2}\,(aq) + \mathrm{Pb}\,(s) \tag{11.6.1}$$

$$\mathrm{Ni}\,(s) + \mathrm{Fe(NO_3)_3}\,(aq) \rightarrow \mathrm{NR\ (no\ reaction)} \tag{11.6.2}$$

In the descriptions that accompany the activity series of metals, a given metal is also capable of undergoing the reactions described below that section. For example, lithium will react with cold water, replacing hydrogen. It will also react with steam and with acids, since that requires a lower degree of reactivity.

Steps for Problem Solving	**Example 11.6.1** Use the activity series to predict if the reaction below will occur. If not, write **NR**. If the reaction does occur, write the products of the reaction and balance the equation. $\mathrm{Al}\,(s) + \mathrm{Zn(NO_3)_2}\,(aq) \rightarrow$	**Example 11.6.2** Use the activity series to predict if the reaction below will occur. If not, write **NR**. If the reaction does occur, write the products of the reaction and balance the equation. $\mathrm{Ag}\,(s) + \mathrm{HCl}\,(aq) \rightarrow$
Plan the problem.	Compare the placements of aluminum (the element doing the replacing) and zinc (the element being replaced) on the activity series.	Compare the placements of silver (the element doing the replacing) and hydrogen (the element being replaced) on the activity series.
Solve	Since aluminum is above zinc, it is capable of replacing it and a reaction will occur. The products of the reaction will be aqueous aluminum nitrate and solid zinc. Take care to write the correct formulas for the products before balancing the equation. Aluminum adopts a $+3$ charge in an ionic compound, so the formula for aluminum nitrate is $\mathrm{Al(NO_3)_3}$. The balanced equation is: $$\mathrm{2Al}\,(s) + \mathrm{3Zn(NO_3)_2}\,(aq) \rightarrow \mathrm{2Al(NO_3)_3}\,(aq) + \mathrm{3Zn}\,(s) \quad (11.6.3)$$	Since silver is below hydrogen, it is not capable of replacing hydrogen in a reaction with an acid. $$\mathrm{Ag}\,(s) + \mathrm{HCl}\,(aq) \rightarrow \mathrm{NR} \quad (11.6.4)$$

Exercise 11.6.1

Use the activity series to predict if the following reactions will occur. If not, write **NR**. If the reaction does occur, write the products of the reaction and balance the equation.

a. Would it be possible to store a silver spoon in a zinc nitrate solution? That is, will the following reaction occur:

$$Ag_{(s)} + Zn(NO_3)_{2(aq)} \rightarrow \quad (11.6.5)$$

b. Would it be possible to store a silver nitrate solution in a copper container? That is, will the following reaction occur:

$$Cu_{(s)} + AgNO_{3(aq)} \rightarrow \quad (11.6.6)$$

Summary

- Metals and halogens are ranked according to their ability to displace other metals or halogens below them in the series.

Contributors and Attributions

-

11.6: The Activity Series: Predicting Spontaneous Redox Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.7: Combustion Reactions

Roasting marshmallows over an open fire is a favorite past-time for campers, outdoor cook-outs, and just gathering around a fire in the backyard. The trick is to get the marshmallow a nice golden brown without catching it on fire. Too often we are not successful and we see the marshmallow burning on the stick - a combustion reaction taking place right in front of us.

Combustion Reactions

A **combustion reaction** is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. Combustion reactions must involve $\mathbf{O_2}$ as one reactant. The combustion of hydrogen gas produces water vapor.

$$\mathbf{2H_2}\,(g) + \mathbf{O_2}\,(g) \rightarrow \mathbf{2H_2O}\,(g) \tag{11.7.1}$$

Notice that this reaction also qualifies as a combination reaction.

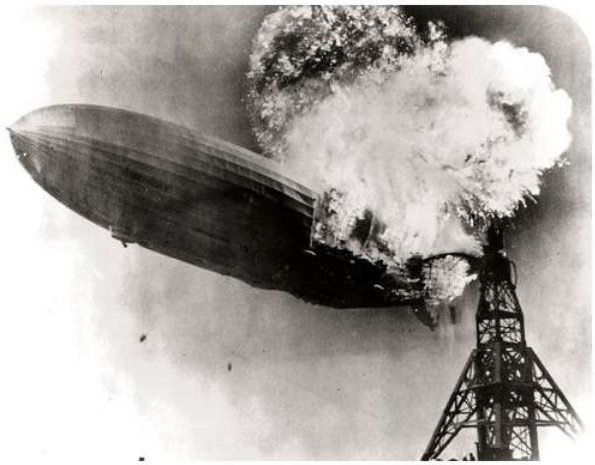


Figure 11.6.1: *Explosion of the* Hindenberg.

The *Hindenberg* was a hydrogen-filled airship that suffered an accident upon its attempted landing in New Jersey in 1937. The hydrogen immediately combusted in a huge fireball, destroying the airship and killing 36 people. The chemical reaction was a simple one: hydrogen combining with oxygen to produce water.

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. propane $(\mathbf{C_3H_8})$ is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$\mathbf{C_3H_8}\,(g) + \mathbf{5O_2}\,(g) \rightarrow \mathbf{3CO_2}\,(g) + \mathbf{4H_2O}\,(g) \tag{11.7.2}$$

Example 11.6.1

Ethanol can be used as a fuel source in an alcohol lamp. The formula for ethanol is $\mathbf{C_2H_5OH}$. Write the balanced equation for the combustion of ethanol.

Solution

Step 1: Plan the problem.

Ethanol and oxygen are the reactants. As with a hydrocarbon, the products of the combustion of an alcohol are carbon dioxide and water.

Step 2: Solve.

Write the skeleton equation:

$$\mathbf{C_2H_5OH}\,(l) + \mathbf{O_2}\,(g) \rightarrow \mathbf{CO_2}\,(g) + \mathbf{H_2O}\,(g) \tag{11.7.3}$$

Balance the equation.

$$C_2H_5OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 3H_2O\,(g) \tag{11.7.4}$$

Step 3: Think about your result.

Combustion reactions must have oxygen as a reactant. Note that the water is produced is in the gas rather than the liquid state because of the high temperatures that accompany a combustion reaction.

Summary

- Combustion reaction is defined and examples are given.

Contributors and Attributions

-

11.7: Combustion Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

13: Solutions

Template:HideTOC

Solutions play a very important role in many biological, laboratory, and industrial applications of chemistry. Of particular importance are solutions involving substances dissolved in water, or aqueous solutions. Solutions represent equilibrium systems, and the lessons learned in our last unit will be of particular importance again. Quantitative measurements of solutions are another key component of this unit. Solutions can involve all physical states - gases dissolved in gases (the air around us), solids dissolved in solids (metal alloys), liquids dissolved in solids (amalgams - liquid mercury dissolved in another metal such as silver, tin or copper). In this unit we will almost exclusively be concerned with aqueous solutions - substances dissolved in water.

13.1: Prelude - Tragedy in Cameroon
13.2: Solutions: Homogeneous Mixtures
13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy
13.4: Solutions of Gases in Water
13.5: Solution Concentration: Mass Percent
13.6: Solution Concentration: Molarity
13.7: Solution Dilution
13.8: Solution Stoichiometry
13.10: Water - Acid and Base in One
13.11: The pH and pOH Scales - Ways to Express Acidity and Basicity
13.9: Acid–Base Titration

13: Solutions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.1: Prelude - Tragedy in Cameroon

Lake Nyos is a crater lake in the Northwest Region of Cameroon, and is a deep lake high on the flank of an inactive volcano in the Oku volcanic plain along the Cameroon line of volcanic activity. A volcanic dam impounds the lake waters. A pocket of magma lies beneath the lake and leaks carbon dioxide (CO_2) into the water, changing it into carbonic acid. Nyos is one of only three known exploding lakes to be saturated with carbon dioxide in this way.

In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (Figure 13.1.1), a deep lake in a volcanic crater. It is believed that the lake underwent a turnover due to gradual heating from below the lake, and the warmer, less-dense water saturated with carbon dioxide reached the surface. Consequently, tremendous quantities of dissolved CO_2 were released, and the colorless gas, which is denser than air, flowed down the valley below the lake and suffocated humans and animals living in the valley.



(a) (b)

Figure 13.1.1: (a) It is believed that the 1986 disaster that killed more than 1700 people near Lake Nyos in Cameroon resulted when a large volume of carbon dioxide gas was released from the lake. (b) A CO_2 vent has since been installed to help outgas the lake in a slow, controlled fashion and prevent a similar catastrophe from happening in the future. (credit a: modification of work by Jack Lockwood; credit b: modification of work by Bill Evans)

Following the Lake Nyos tragedy, scientists investigated other African lakes to see if a similar phenomenon could happen elsewhere. Lake Kivu in Democratic Republic of Congo, 2,000 times larger than Lake Nyos, was also found to be supersaturated, and geologists found evidence for outgassing events around the lake about every thousand years.

Contributors and Attributions

- Wikipedia
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Template:HideTOC

13.1: Prelude - Tragedy in Cameroon is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.2: Solutions: Homogeneous Mixtures

skills to develop

- Learn some terminology involving solutions.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious.

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO_2 in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent. Table **13.2.1** lists some common types of solutions, with examples of each.

Table 13.2.1: 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



Figure **13.2.1**: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Example 13.2.1: Sugar Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 13.2.1

A solution is made by dissolving 3.33 g of HCl(g) in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer

solute: HCl(g); solvent: CH_3OH

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.2: Solutions: Homogeneous Mixtures is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy

Learning Objectives

- Define electrolytes and non electrolytes
- Explain why solutions form.
- Discuss the idea of water as the "universal solvent".
- Explain how water molecules attract ionic solids when they dissolve in water.

We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. We also know that solutions have constant composition and we can also vary this composition up to a point to maintain the homogeneous nature of the solution. But how exactly do solutions form? Why is it that oil and water will not form a solution and yet vinegar and water will? Why could we dissolve table salt in water but not in vegetable oil? The reasons why solutions will form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Solubility and Saturation

Table salt (NaCl) readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table **13.3.1** lists the solubilities of some simple ionic compounds. These solubilities vary widely: NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table 13.3.1: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated but rather concentrated. In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

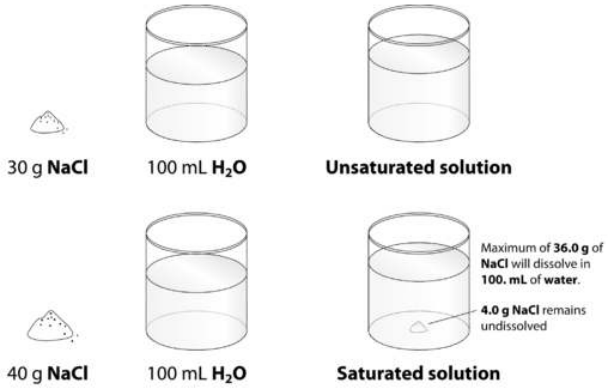


Figure 13.3.1: When 30.0 g of NaCl is added to 100 mL, it all dissolves, forming an unsaturated solution. When 40.0 g is added, 36.0 g dissolves and 4.0 g remains undissolved, forming a saturated solution.

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.



Electrolyte Solutions: Dissolved Ionic Solids

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte (good conductor). If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte (do not conduct electricity as well).

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 13.3.1).

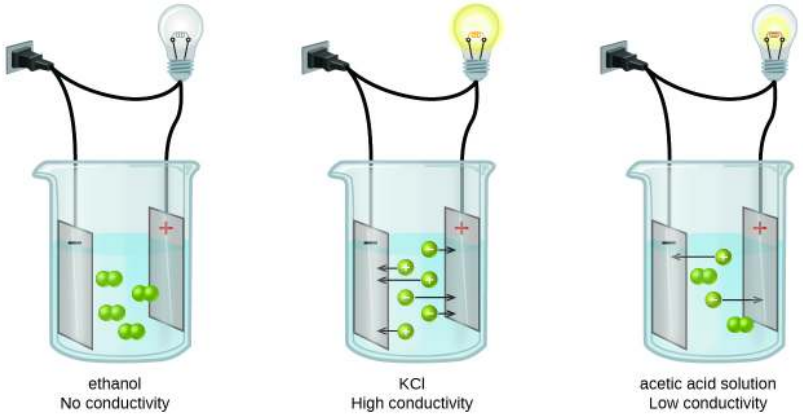


Figure 13.3.1: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Water and other polar molecules are attracted to ions, as shown in Figure 13.3.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

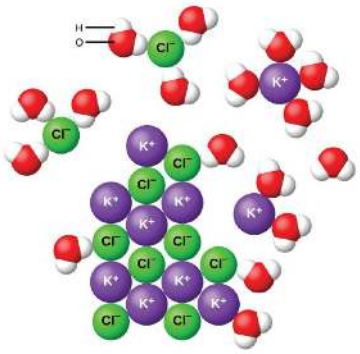


Figure 13.3.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Example 13.3.1: Identifying IOnic Compounds

Which compound(s) will dissolve in solution to separate into ions?

a. **LiF**
b. $\mathbf{P_2F_5}$
c. $\mathbf{C_2H_5OH}$

Solution

LiF will separate into ions when dissolved in solution, because it is an ionic compound. $\mathbf{P_2F_5}$ and $\mathbf{C_2H_5OH}$ are both covalent and will stay as molecules in a solution.

Exercise 13.3.1

Which compounds will dissolve in solution to separate into ions?

a. $C_6H_{12}O_{11}$, glucose
b. CCl_4

c. $CaCl_2$
d. $AgNO_3$

Answer:

c & d

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 $100\ g$ of solvent. The solubility of sodium chloride in water is $36.0\ g$ per $100\ g$ water at $20°C$. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (Figure $13.3.4$).

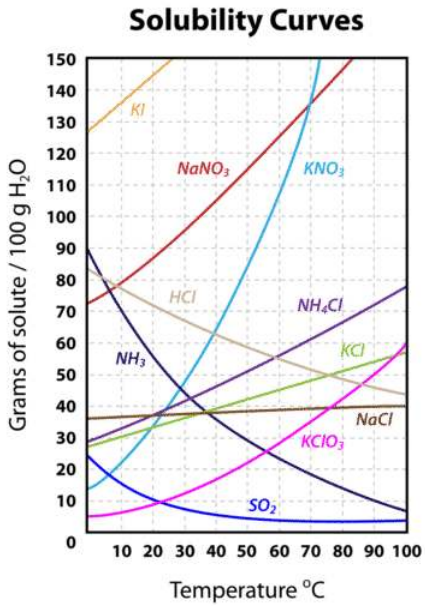


Figure $13.3.4$: *Solubility curves for several compounds.*

Notice how the temperature dependence of $NaCl$ is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of $NaCl$. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances - HCl, NH_3, and SO_2 - have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ g$ of KNO_3 is added to $100\ g$ of water at $30°C$. According to the solubility curve, approximately $48\ g$ of KNO_3 will dissolve at $30°C$. This means that the solution will be saturated since $48\ g$ is less than $80\ g$. We can also determine that there will be $80 - 48 = 32\ g$ of undissolved KNO_3 remaining at the bottom of the container. Now suppose that this saturated solution is heated to $60°C$. According to the curve, the solubility of KNO_3 at $60°C$ is about $107\ g$. Now the solution is unsaturated since it contains only the original $80\ g$ of dissolved solute. Now suppose the solution is cooled all the way down to $0°C$. The solubility at $0°C$ is about $14\ g$, meaning that $80 - 14 = 66\ g$ of the KNO_3 will recrystallize.

Summary

- Solubility is a specific amount of solute that can dissolve in a given amount of solvent.
- Saturated and unsaturated solutions are defined.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).
- Water is considered as the the universal solvent since it can dissolve both ionic and polar solutes, as well as some nonpolar solutes (in very limited amounts).
- The solubility of a solid in water increases with an increase in temperature.

Vocabulary

- Miscible: Liquids that have the ability to dissolve in each other.
- Immiscible: Liquids that do not have the ability to dissolve in each other.
- Electrostatic attraction: The attraction of oppositely charged particles.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.4: Solutions of Gases in Water

Learning Objectives

- Predict that the solubility of a gas decreases with an increase in temperature.

In an earlier module of this chapter, the effect of intermolecular attractive forces on solution formation was discussed. The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility. For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium, but 100 times less than the solubility of chloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the liquid hydrocarbon hexane, C_6H_{14}, is approximately 20 times greater than it is in water.

Other factors also affect the solubility of a given substance in a given solvent. Temperature is one such factor, with gas solubility typically decreasing as temperature increases (Figure **13.4.1**). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

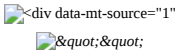


Figure **13.4.1**: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (Figure **13.4.2**).

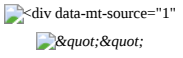


Figure **13.4.2**: (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service)

The solubility of a gaseous solute is also affected by the partial pressure of solute in the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at this pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution in the form of small bubbles (Figure **13.4.3**). At this point, the beverage is *supersaturated* with carbon dioxide and, with time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become "flat."

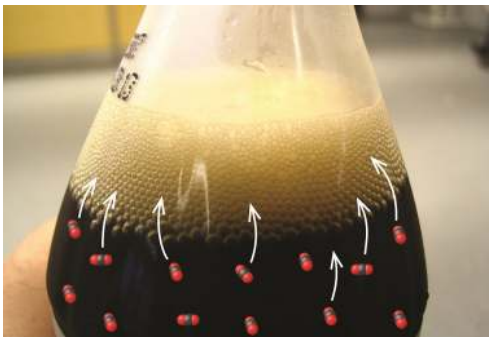


Figure 1**13.4.3**: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

"Fizz"

The dissolution in a liquid, also known as fizz usually involves carbon dioxide under high pressure. When the pressure is reduced, the carbon dioxide is released from the solution as small bubbles, which causes the solution to become effervescent, or fizzy. A common example is the dissolving of carbon dioxide in water, resulting in carbonated water.

Carbon dioxide is weakly soluble in water, therefore it separates into a gas when the pressure is released. This process is generally represented by the following reaction, where a pressurized dilute solution of carbonic acid in water releases gaseous carbon dioxide at decompression:

$$H_2CO_{3(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)} \tag{13.4.1}$$

In simple terms, it is the result of the chemical reaction occurring in the liquid which produces a gaseous product.

Contributors and Attributions

- Wikipedia
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.4: Solutions of Gases in Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.5: Solution Concentration: Mass Percent

Learning Objectives

- Express the amount of solute in a solution in various concentration units.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms whose meanings depend on various factors.

Introduction

Concentration is the measure of how much of a given substance is mixed with another substance. Solutions can be said to be dilute or concentrated. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution.

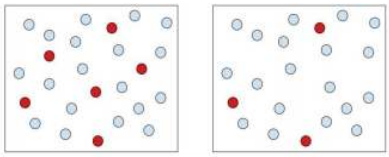


Figure **13.5.1**: *The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).*

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. What you have to do is take the frozen juice from inside these containers and usually empty it into 3 or 4 times the container size full of water to mix with the juice concentrate and make your container of juice. Therefore, you are diluting the concentrated juice. When we talk about solute and solvent, the concentrated solution has a lot of solute versus the dilute solution that would have a smaller amount of solute.

The terms "concentrated" and "dilute" provide qualitative methods of describing concentration. Although qualitative observations are necessary and have their place in every part of science, including chemistry, we have seen throughout our study of science that there is a definite need for quantitative measurements in science. This is particularly true in solution chemistry. In this section, we will explore some quantitative methods of expressing solution concentration.

Mass Percent

There are several ways of expressing the concentration of a solution by using a percentage. The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \text{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{13.5.1}$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving $25.0\ \text{g}$ of sugar into $100.0\ \text{g}$ of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\ \text{g sugar}}{125.0\ \text{g solution}} \times 100\% = 20.0\%\ \text{sugar} \tag{13.5.2}$$

Example 13.5.1

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\%\ \mathrm{m/m} = \frac{36.5\ \mathrm{g}}{355\ \mathrm{g}} \times 100\% = 10.3\%$$

Exercise 13.5.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer

7.90 %

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g\ \text{solute}}{100\ g\text{solution}}$ or $\frac{100\ g\text{solution}}{g\ \text{solute}}$

For example, if you need to make $3000.0\ \mathrm{g}$ of a 5.00% solution of sodium chloride, the mass of solute needs to be determined.

Solution

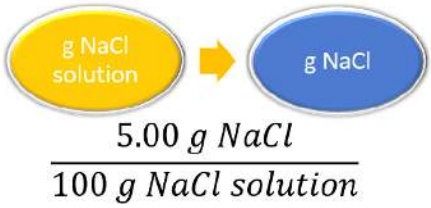
Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:



$$\frac{5.00\ g\ NaCl}{100\ g\ NaCl\ solution}$$

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0\ \cancel{g\ NaCl\ solution} \times \frac{5.00\ g\ NaCl}{100\ \cancel{g\ NaCl\ solution}} = 150.0 g\ NaCl \tag{13.5.3}$$

You would need to weigh out $150\ \mathrm{g}$ of NaCl and add it to $2850\ \mathrm{g}$ of water. Notice that it was necessary to subtract the mass of the NaCl $(150\ \mathrm{g})$ from the mass of solution $(3000\ \mathrm{g})$ to calculate the mass of the water that would need to be added.

Exercise 13.5.1

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg , 3.00 % (by mass) H_2O_2 solution?

Answer

180. g H_2O_2

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.5: Solution Concentration: Mass Percent is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.6: Solution Concentration: Molarity

Learning Objectives

- Use molarity to determine quantities in chemical reactions.
- Use Molarity as a conversion factor in calculations.

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{13.6.1}$$

The symbol for molarity is **M** or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[Ag^+]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to calculate with but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is therefore

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl} \tag{13.6.2}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” You would read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> *Be sure to note that molarity is calculated as the total volume of the **entire** solution, not just volume of solvent! The solute contributes to total volume.*

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4 \, \cancel{g HCl} \times \frac{1 \, mol \, HCl}{36.5 \, \cancel{g HCl}} = 0.614 \, mol \, HCl \tag{13.6.3}$$

Now we can use the definition of molarity to determine a concentration:

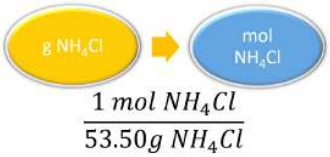
$$M = \frac{0.614 \, mol \, HCl}{1.56 L \, solution} = 0.394 \, M HCl \tag{13.6.4}$$

Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

Example 13.6.1

A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: Mass $= 42.23\ g\ NH_4Cl$ Volume solution $= 500.0\ mL = 0.5000\ L$ Find: Molarity = ? M
List other known quantities.	Molar mass $NH_4Cl = 53.50\ g/mol$
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  $\frac{1\ mol\ NH_4Cl}{53.50g\ NH_4Cl}$ 2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters. $M = \frac{mol\ NH_4Cl}{L\ solution}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \cancel{g\ NH_4Cl} \times \frac{1\ mol\ \cancel{NH_4Cl}}{53.50\ g\ NH_4Cl} = 0.7893\ mol\ NH_4Cl$ (13.6.5) $\frac{0.7893\ mol\ NH_4Cl}{0.5000\ L\ solution} = 1.579\ M$ (13.6.6)
Think about your result.	The molarity is $1.579\ M$, meaning that a liter of the solution would contain $1.579\ mol\ NH_4Cl$. Four significant figures are appropriate.

Exercise 13.6.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer

1.57 M $C_6H_{12}O_6$

Exercise 13.6.2

What is the concentration, in mol/L, where $137\ g$ of $NaCl$ has been dissolved in enough water to make $500\ mL$ of solution?

Answer

4.69 M NaCl

Using Molarity in Calculations

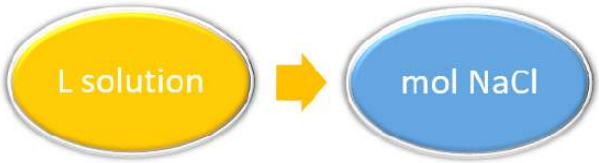
Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition.

Determining moles of solute given the concentration and volume of a solution

For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

Solution



$$\frac{.887\ mol\ NaCl}{1\ L\ solution}$$

$$0.108\ \cancel{L\ NaCl} \times \frac{0.887\ mol\ NaCl}{1\cancel{L\ solution}} = 0.0958\ mol\ NaCl \tag{13.6.7}$$

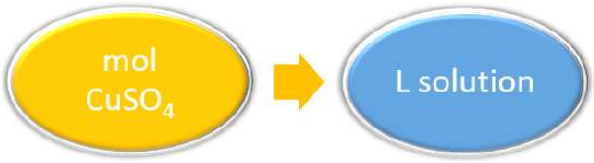
If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Determining volume of a solution given the concentration and moles of solute

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:



$$\frac{1\ L\ solution}{2.35\ mol\ CuSO_4}$$

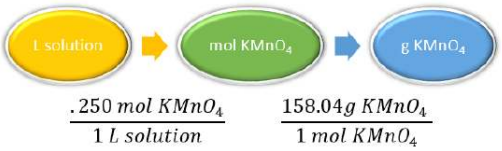
$$4.88\ \cancel{mol\ CuSO_4} \times \frac{1\ L\ solution}{2.35\ \cancel{mol\ CuSO_4}} = 2.08\ L\ of\ solution \tag{13.6.8}$$

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams. The following example illustrates this.

Example 13.6.2

A chemist needs to prepare $3.00\ \text{L}$ of a $0.250\ \text{M}$ solution of potassium permanganate (KMnO_4). What mass of KMnO_4 does she need to make the solution?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: Molarity $= 0.250\ \mathrm{M}$ Volume $= 3.00\ \mathrm{L}$ Find: Mass $\mathrm{KMnO_4} = ?\ \mathrm{g}$
List other known quantities.	Molar mass $\mathrm{KMnO_4} = 158.04\ \mathrm{g/mol}$ 0.250 mol $KMnO_4$ to 1 L of $KMnO_4$ solution
Plan the problem.	
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $\mathrm{mol\ KMnO_4} = 0.250\ \mathrm{M\ KMnO_4} \times 3.00\ \mathrm{L} = 0.750\ \mathrm{mol\ KMnO_4}$ (13.6.9) $3.00\ \cancel{L\ solution} \times \frac{0.250\ \cancel{\mathrm{mol\ KMnO_4}}}{1\ \cancel{L\ solution}} \times \frac{158.04\ \mathrm{g\ KMnO_4}}{1\ \cancel{\mathrm{mol\ KMnO_4}}} = 119\ \mathrm{g\ KMnO_4}$ (13.6.10)
Think about your result.	When $119\ \mathrm{g}$ of potassium permanganate is dissolved into water to make $3.00\ \mathrm{L}$ of solution, the molarity is $0.250\ \mathrm{M}$.

Exercise 13.6.3

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer

1.74 L

Exercise 13.6.4

Answer the problems below using concentration as a conversion factor.

a. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
b. What volume of 1.50 M HCl solution contains 10.0 g of hydrogen chloride?

Answer a

6.12 g

Answer b

183 mL or 0.183L

How to indicate contentrations

- square brackets are often used to represent concentration, e.g., [NaOH] = 0.50 M.
- use the capital letter M for molarity, not a lower case m (this is a different concentration unit called *molality*).



Watch as the Flinn Scientific Tech Staff demonstrates "How To Prepare Solutions."

An important note - there are many different ways you can set up and solve your chemistry equations. Some students prefer to answer multi-step calculations in one long step; others prefer to work out each step individually. Neither method is necessarily better or worse than the other method - whatever makes most sense to ***you*** is the one you should use. We will typically use unit analysis (also called dimension analysis or factor analysis).

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.6: Solution Concentration: Molarity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.7: Solution Dilution

Learning Objectives

- Explain how concentrations can be changed in the lab
- Understand how stock solutions are used in the laboratory

We are often concerned with how much solute is dissolved in a given amount of solution. We will begin our discussion of solution concentration with two related and relative terms - **dilute** and **concentrated**.

- A **dilute** solution is one in which there is a relatively small amount of solute dissolved in the solution.
- A **concentrated** solution contains a relatively large amount of solute.

These two terms do not provide any quantitative information (actual numbers) - but they are often useful in comparing solutions in a more general sense. These terms also do not tell us whether or not the solution is saturated or unsaturated, or whether the solution is "strong" or "weak". These last two terms will have special meanings when we discuss acids and bases, so be careful not to confuse these terms.

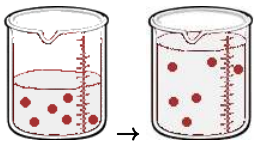
Stock Solutions

It is often necessary to have a solution whose concentration is very precisely known. Solutions containing a precise mass of solute in a precise volume of solution are called **stock (or standard) solutions.** To prepare a standard solution a piece of lab equipment called a volumetric flask should be used. These flasks range in size from 10 mL to 2000 mL are are carefully calibrated to a single volume. On the narrow stem is a **calibration mark**. The precise mass of solute is dissolved in a bit of the solvent and this is added to the flask. Then enough solvent is added to the flask until the level reaches the calibration mark.

Often it is convenient to prepare a series of solutions of known concentrations by first preparing a single **stock solution** as described in the previous section. **Aliquots** (carefully measured volumes) of the stock solution can then be diluted to any desired volume. In other cases it may be inconvenient to weigh accurately a small enough mass of sample to prepare a small volume of a dilute solution. Each of these situations requires that a solution be diluted to obtain the desired concentration.

Dilutions of Stock (or Standard) Solutions

Imagine we have a salt water solution with a certain concentration. That means we have a certain amount of salt (a certain mass or a certain number of moles) dissolved in a certain volume of solution. Next we willl dilute this solution - we do that by adding more water, not more salt:



Before Dilution After Dilution

The molarity of solution 1 is

$$M_1 = \frac{\text{moles}_1}{\text{liter}_1} \tag{13.7.1}$$

and the molarity of solution 2 is

$$M_1 = \frac{\text{moles}_1}{\text{liter}_1} \tag{13.7.2}$$

rearrange the equations to find moles:

$$\text{moles}_1 = M_1\text{liter}_1 \tag{13.7.3}$$

and

$$\text{moles}_2 = M_2\text{liter}_2 \tag{13.7.4}$$

What stayed the same and what changed between the two solutions? By adding more water, we changed the volume of the solution. Doing so also changed it's concentration. **However, the number of moles of solute did not change.** So,

$$moles_1 = moles_2 \tag{13.7.5}$$

Therefore

$$\boxed{M_1V_1 = M_2V_2} \tag{13.7.6}$$

where

- M_1 and M_2 are the concentrations of the original and diluted solutions and
- V_1 and V_2 are the volumes of the two solutions

Preparing dilutions is a common activity in the chemistry lab and elsewhere. Once you understand the above relationship, the calculations are easy to do.

Suppose that you have $100.\ \text{mL}$ of a $2.0\ \text{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \text{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2.

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \text{M} \times 100.\ \text{mL}}{500.\ \text{mL}} = 0.40\ \text{M HCl} \tag{13.7.7}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves deciding how much of a highly concentrated solution is required to make a desired quantity of solution of lesser concentration. The highly concentrated solution is typically referred to as the stock solution.

Example 13.7.1: Diluting NITRIC ACID

Nitric acid (HNO_3) is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \text{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \text{L}$ of a $0.50\ \text{M}$ solution?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: M_1, Stock $\text{HNO}_3 = 16\ \text{M}$ $V_2 = 8.00\ \text{L}$ $M_2 = 0.50\ \text{M}$ Find: Volume stock HNO_3 $(V_1) = ?\ \text{L}$
List other known quantities	none
Plan the problem	First, rearrange the equation algebraically to solve for V_1. $$V_1 = \frac{M_2 \times V_2}{M_1} \tag{13.7.8}$$
Calculate and cancel units	Now substitute the known quantities into the equation and solve. $$V_1 = \frac{0.50\ \text{M} \times 8.00\ \text{L}}{16\ \text{M}} = 0.25\ \text{L} = 250\ \text{mL} \tag{13.7.9}$$
Think about your result.	$250\ \text{mL}$ of the stock HNO_3 needs to be diluted with water to a final volume of $8.00\ \text{L}$. The dilution is by a factor of 32 to go from $16\ \text{M}$ to $0.5\ \text{M}$.

Exercise 13.7.1

A 0.885 M solution of KBr whose initial volume is 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

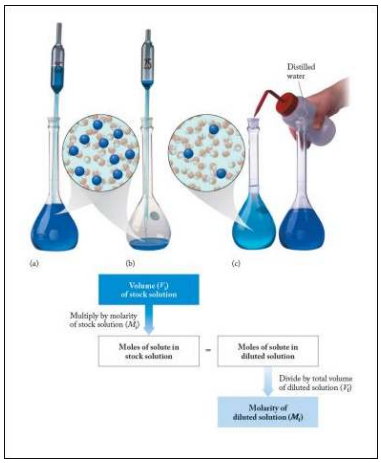
Answer

135.4 mL

Diluting and Mixing Solutions



How to Dilute a Solution by CarolinaBiological



Note

The calculated volume will have the same dimensions as the input volume, and dimensional analysis tells us that in this case we don't need to convert to liters, since L cancels when we divide M (mol/L) by M (mol/L).

Contributors and Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

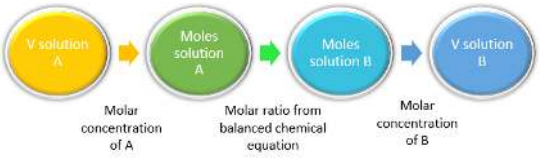
13.7: Solution Dilution is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.8: Solution Stoichiometry

Learning Objectives

- Determine amounts of reactants or products in aqueous solutions.

As we learned in Chapter 7, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are "switched" (they *replace* each other). Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and hence their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead(II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead(II) chloride.

$$Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow PbCl_2\,(s) + 2\,NaNO_3\,(aq) \tag{13.8.1}$$

In the reaction shown above, if we mixed .123 L of a 1.00 M solution of NaCl with 1.50 M solution of $Pb(NO_3)_2$, we could calculate the volume of $Pb(NO_3)_2$ that reacted to completely precipitate the Pb^{2+} ions.

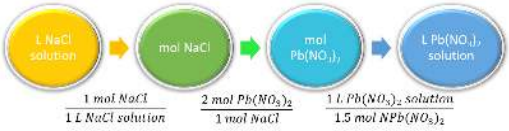
The molar concentration can also be expressed as the following:

1.00M NaCl = $\frac{1.00\ mol\ NaCl}{1\ L\ NaCl\ solution}$

and

1.50M $Pb(NO_3)_2$ = $\frac{1.00\ mol\ Pb(NO_3)_2}{1\ L\ Pb(NO_3)_2\ solution}$

First, we must examine the reaction stoichiometry. In this reaction, one mole of $Pb(NO_3)_2$ reacts with two moles of NaCl to give one mole of $PbCl_2$. Thus, the concept map utilizing the stoichiometric ratios is:



so the volume of lead (II) nitrate that reacted is calculated as:

$$.123\ L\ NaClsolution \times \frac{1\ L\ NaCl\ solution}{1.00\ mol\ NaCl} \times \frac{1\ mol\ Pb(NO_3)_2}{2\ mol\ NaCl} \times \frac{1\ L\ Pb(NO_3)_2\ solution}{1\ mol\ Pb(NO_3)_2} = 0.4\ Pb(NO_3)_2\ solution \tag{13.8.2}$$

Example 13.8.1:

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate the Ba^{2+} in the solution?

Solution

Steps for Problem Solving	Example 13.8.1
Identify the "given"information and what the problem is asking you to "find."	Given: 275 mL $BaCl_2$ 0.250 M $BaCl_2$ or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M Na_2SO_4 or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume Na_2SO_4
Set-up and balance the chemical equation	$Na_2SO_4(aq) + BaCl_2(aq)$ $BaSO_4(s) + \underline{2}$ NaCl (aq) An insoluble product is formed after the reaction.
List other known quantities	1 mol of Na_2SO_4 to 1 mol $BaCl_2$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$275\ mL\ BaCl_2\ solution \times \frac{1\ L}{1000\ mL} \times \frac{0.250\ mol\ BaCl_2}{1\ BaCl_2\ solution} \times \frac{1\ mol\ Na_2SO_4}{1\ mol\ BaCl_2} \times \frac{1\ L\ Na_2SO_4\ solution}{0.500\ mol Na_2SO_4}$
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than unit is correct.

Exercise 13.8.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

Answer:

0.250 L LiOH solution

Contributors and Attributions

- Paul R. Young, Professor of Chemistry, University of Illinois at Chicago, Wiki: AskTheNerd; PRY@askthenerd.com - pyoung@uic.edu; ChemistryOnline.com
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.8: Solution Stoichiometry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.10: Water - Acid and Base in One

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H_3O^+ and OH^- in aqueous solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4^+ + OH^- \tag{13.10.1}$$

where H_2O acts as an acid (in red).

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^- \tag{13.10.2}$$

where H_2O acts as an base (in blue).

It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{13.10.3}$$

This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**.

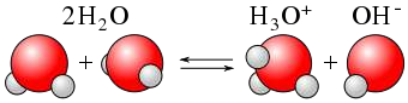


Figure 13.10.1: Autoionization of water, giving hydroxide and hydronium ions.

At this level, the concentration of both $H_3O^+(aq)$ and $OH^-(aq)$ in a sample of pure H_2O is about $1.0 \times 10^{-7}\ M$ (at room temperature). If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H_3O^+]=[OH^-] = 1.0 \times 10^{-7} \tag{13.10.4}$$

for *any* sample of pure water because H_2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14}:

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14} \tag{13.10.5}$$

- For acids, the concentration of $H_3O^+(aq)$ (i.e., $[H_3O^+]$) is greater than $1.0 \times 10^{-7}\ M$.
- For bases the concentration of $OH^-(aq)$ (i.e., $[OH^-]$) is greater than $1.0 \times 10^{-7}\ M$.

However, the *product* of the two concentrations—$[H_3O^+][OH^-]$—is *always* equal to 1.0×10^{-14}, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{13.10.6}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{13.10.7}$$

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$) has to be for the product to equal 1.0×10^{-14}, or if you know $[OH^-]$), you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

Warning: Temperature Matters

The degree of autoionization of water (Equation 13.10.3) and hence the value of K_w changes with temperature, so Equations 13.10.4 - 13.10.7 are accurate only at room temperature.

Example 13.10.1: Hydroxide ConCentration

What is $[OH^-]$) of an aqueous solution if $[H_3O^+]$ is $1.0 \times 10^{-4} M$?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[H_3O^+] = 1.0 \times 10^{-4}\ M$ Find: $[OH^-]$ = ? M
List other known quantities	none
Plan the problem	Using the expression for K_w, (Equation 13.10.7), rearrange the equation algebraically to solve for $[OH^-]$. $$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$
Calculate	Now substitute the known quantities into the equation and solve. $$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} M$$ It is assumed that the concentration unit is molarity, so $[OH^-]$ is 1.0×10^{-10} M.
Think about your result.	The concentration of the acid is high (> 1 x 10^{-7} M) so $[OH^-]$ should be low.

Exercise 13.10.1

What is $[OH^-]$ in a 0.00032 M solution of H_2SO_4?

Hint

Assume **both** protons ionize from the molecule.. although this is not the case.

Answer:

$3.1 \times 10^{-11}\ M$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H_3O^+ or OH^- ions in the formula unit because $[H_3O^+]$ or $[OH^-]$) may not be the same as the concentration of the acid or base itself.

Example 13.10.2: Hydronium Concentration

What is $[H_3O^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: $[Ca(OH)_2] = 0.0044\,M$ Find: $[H_3O^+]$ = ? M
List other known quantities	We begin by determining $[OH^-]$. The concentration of the solute is 0.0044 M, but because $Ca(OH)_2$ is a strong base, there are two OH^- ions in solution for every formula unit dissolved, so the actual $[OH^-]$ is two times this, $$[OH^-] = 2 \times 0.0044\ \text{M} = 0.0088\,\text{M}\cdot$$
Plan the problem	Use the expression for K_w (Equation 13.10.7) and rearrange the equation algebraically to solve for $[H_3O^+]$. $$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$$
Calculate	Now substitute the known quantities into the equation and solve. $$[H_3O^+] = \frac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$$ $[H_3O^+]$ has decreased significantly in this basic solution.
Think about your result.	The concentration of the base is high (> 1 x 10^{-7} M) so Extra close brace or missing open brace should be low.

Exercise 13.10.2

What is $[H_3O^+]$ of an aqueous solution if $[OH^-]$ is $1.0 \times 10^{-9}\,M$?

Answer:

1.0×10^{-5} M

In any aqueous solution, the product of $[H_3O^+]$ and $[OH^-]$ equals 1.0×10^{-14} (at room temperature).

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.10: Water - Acid and Base in One is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.11: The pH and pOH Scales - Ways to Express Acidity and Basicity

Learning Objectives

- Define pH and pOH
- Determine the pH of acidic and basic solutions.
- Determine the hydronium ion concentration and pOH from pH.

As we have seen, $[H_3O^+]$ and $[OH^-]$ values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of $[H_3O^+]$:

$$pH = -\log[H_3O^+] \tag{13.11.1}$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H_3O^+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the pH scale. The range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. Figure **13.11.1** illustrates this relationship, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H_3O^+]$, which will give a positive value for pH.

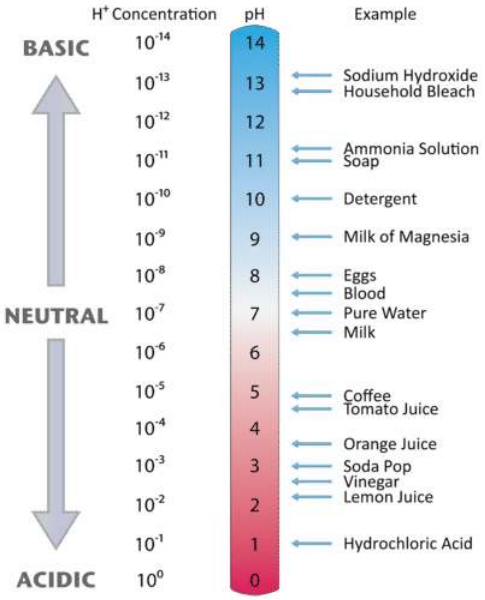


Figure **13.11.1**: The pH *values for several common materials.*

Warning: The Ph scale has no limits

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H_3O^+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the pH scale. The range of values from 0 to 14 that describes the acidity or basicity of a solution.. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral.

Example 13.11.1

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Answer

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

Exercise 13.11.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

a. human blood with pH = 7.4
b. household ammonia with pH = 11.0
c. cherries with pH = 3.6

Answer a:

basic

Answer b:

basic

Answer c:

acidic

Calculating pH from Hydronium Concentration

The pH of solutions can be determined by using logarithms as illustrated in the next example for stomach acid. Stomach acid is a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}\ M$, what is the pH of the solution?

$$\begin{aligned} \text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) = 2.92 \end{aligned}$$

Logarithms

To get the **log value** on your calculator, enter the number (in this case, the hydronium ion concentration) first, then press the LOG key.

If the number is 1.0 x 10^{-5} (for $[H_3O^+]$ = 1.0 x 10^{-5} M) you should get an answer of "-5"

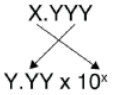
If you get a different answer, or an error, try pressing the LOG key before you enter the number.

Example 13.11.2: Converting Ph to Hydronium Concentration

Find the pH, given the $[H_3O^+]$ of the following:

a. 1 ×10^{-3} M
b. 2.5 ×10^{-11} M
c. 4.7 ×10^{-9} M

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: a. $[H_3O^+]$ =1 × 10^{-3} M b. $[H_3O^+]$ =2.5 ×10^{-11} M c. $[H_3O^+]$ = 4.7 ×10^{-9} M Find: ? pH
Plan the problem	Need to use the expression for pH (Equation 13.11.1) pH = - log $[H_3O^+]$
Calculate	Now substitute the known quantity into the equation and solve. a. pH = - log $[1 \times 10^{-3}]$ = 3.0 (1 decimal places since 1 has 1 significant figure) b. pH = - log [2.5 ×10^{-11}] = 10.60 (2 decimal places since 2.5 has 2 significant figures) c. pH = - log [4.7 ×10^{-9}] = 8.30 (2 decimal places since 4.7 has 2 significant figures) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 13.11.2

Find the pH, given $[H_3O^+]$ of the following:

a. 5.8 ×10^{-4} M
b. 1.0×10^{-7}

Answer a:

3.22

Answer b:

7.00

Calculating Hydronium Concentration from pH

Sometimes you need to work "backwards" - you know the pH of a solution and need to find $[H_3O^+]$, or even the concentration of the acid solution. How do you do that? To convert pH into $[H_3O^+]$ we solve Equation 13.11.1 for $[H_3O^+]$. This involves taking the antilog (or inverse log) of the negative value of pH .

$$[H_3O^+] = \text{antilog}(-pH) \tag{13.11.2}$$

or

$$[H_3O^+] = 10^{-pH} \tag{13.11.3}$$

As mentioned above, different calculators work slightly differently - make sure you can do the following calculations using ***your*** calculator.

Calculator Skills

We have a solution with a pH = 8.3. What is $[H_3O^+]$?

With some calculators you will do things in the following order:

1. Enter 8.3 as a negative number (use the key with both the +/- signs, not the subtraction key)
2. Use your calculator's 2nd or Shift or INV (inverse) key to type in the symbol found above the LOG key. The shifted function should be 10^x.
3. You should get the answer **5.0×10^{-9}**

Other calculators require you to enter keys in the order they appear in the equation.

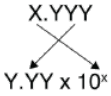
1. Use the Shift or second function to key in the 10^x function.
2. Use the +/- key to type in a negative number, then type in 8.3
3. You should get the answer **5.0×10^{-9}**

If neither of these methods work, try rearranging the order in which you type in the keys. Don't give up - you must master your calculator!

Example 13.11.3: Calculating Hydronium Concentration from pH

Find the hydronium ion concentration in a solution with a pH of 12.6. Is this solution an acid or a base? How do you know?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: pH = 12.6 Find: $[H_3O^+]$ = ? M
Plan the problem	Need to use the expression for $[H_3O^+]$ (Equation 13.11.3) $[H_3O^+] = \text{antilog}(-pH)$ or $[H_3O^+] = 10^{-pH}$
Calculate	Now substitute the known quantity into the equation and solve. $[H_3O^+] = \text{antilog}(12.\underline{60}) = \underline{2.5} \times 10^{-13}$ M (2 significant figures since 4.7 has 12.60 2 decimal places) or $[H_3O^+] = 10^{-12.\underline{60}} = \underline{2.5} \times 10^{-13}$ M (2 significant figures since 4.7 has 12.60 2 decimal places) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 13.11.3

If moist soil has a pH of 7.84, what is $[H_3O^+]$ of the soil solution?

Answer

1.5 x 10^{-8} M

The pOH scale

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration.

$$\mathrm{pOH} = -\log\left[\mathrm{OH}^-\right] \tag{13.11.4}$$

The relation between the hydronium and hydroxide ion concentrations expressed as p-functions is easily derived from the K_w expression:

$$K_w = [\mathrm{H_3O^+}][\mathrm{OH}^-] \tag{13.11.5}$$

$$-\log K_w = -\log([\mathrm{H_3O^+}][\mathrm{OH}^-]) = -\log[\mathrm{H_3O^+}] + -\log[\mathrm{OH}^-] \tag{13.11.6}$$

$$\mathrm{p}K_w = \mathrm{pH} + \mathrm{pOH} \tag{13.11.7}$$

At 25 °C, the value of K_w is 1.0×10^{-14}, and so:

$$14.00 = \mathrm{pH} + \mathrm{pOH} \tag{13.11.8}$$

The hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7}\ M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{13.11.9}$$

$$\mathrm{pOH} = -\log[\mathrm{OH}^-] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{13.11.10}$$

And so, at this temperature, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities less than $1.0 \times 10^{-7}\ M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7}\ M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

Example 13.11.4:

Find the pOH of a solution with a pH of 4.42.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: pH =4.42 Find: ? pOH
Plan the problem	Need to use the expression pOH = 14 - pH
Calculate	Now substitute the known quantity into the equation and solve. pOH=14−4.42=9.58
Think about your result	The pH is that of an acidic solution, and the resulting pOH is the difference after subtracting from 14. The answer has two significant figures because the given pH has two decimal places.

Exercise 13.11.4

The pH of a solution is 8.22. What is the pOH?

Answer

5.78

The diagram below shows all of the interrelationships between [H3O+][H3O+], [OH−][OH−], pH, and pOH.

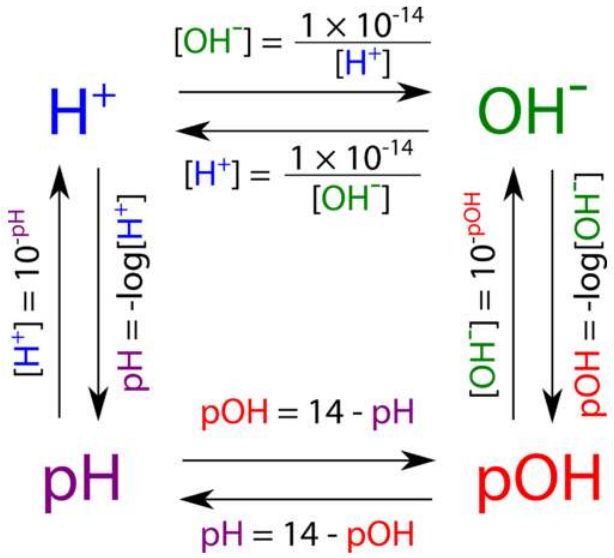


Figure **13.11.1**: *Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH.*

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
-
- Template:OpenStax
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.11: The pH and pOH Scales - Ways to Express Acidity and Basicity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.9: Acid–Base Titration

Learning Objectives

- Understand the basics of acid-base titrations
- Understand the use of indicators
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base. The indicator will signal, by color change, when the base has been neutralized (when $[H^+] = [OH^-]$). At that point - called the **equivalence point** or **end point** - the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

It is important to accurately measure volumes when doing titrations. The instrument you would use is called a burette (or buret).



Figure 13.9.1: Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\text{\# mol HCl} = (0.02566\ \text{L})(0.1078\ \text{M}) = 0.002766\ \text{mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:

$$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$$

So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766\ \cancel{mol\ HCl} \times \frac{1\ mol\ NaOH}{1\ \cancel{mol\ HCl}} = 0.002766\ mol\ NaOH \tag{13.9.1}$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766\ \cancel{mol\ HCl} \times \frac{40.00\ g\ NaOH}{1\ \cancel{mol\ HCl}} = 0.1106\ g\ NaOH \tag{13.9.2}$$

This is type of calculation is performed as part of a titration.

Example 13.9.1: Equivalence Point

What mass of $Ca(OH)_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3? The balanced chemical equation is as follows:

$$2\,HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2\,H_2O \tag{13.9.3}$$

Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

\# moles HNO_3 = (0.04402 L)(0.0885 M) = 0.00390 mol HNO_3

Using the balanced chemical equation, we can determine the number of moles of $Ca(OH)_2$ present in the analyte:

$$0.00390\ \cancel{mol\ HNO_3} \times \frac{1\ mol\ Ca(OH)_2}{2\ \cancel{mol\ HNO_3}} = 0.00195\ mol\ Ca(OH)_2 \tag{13.9.4}$$

Then we convert this to a mass using the molar mass of $Ca(OH)_2$:

$$0.00195 \cancel{mol\ Ca(OH)_2} \times \frac{74.1\ g\ Ca(OH)_2}{\cancel{mol\ Ca(OH)_2}} = 0.144\ g\ Ca(OH)_2 \tag{13.9.5}$$

Exercise 13.9.1

What mass of $H_2C_2O_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:

$$H_2C_2O_4 + 2\,NaOH \longrightarrow Na_2C_2O_4 + 2\,H_2O \tag{13.9.6}$$

Answer:

0.182 g

Exercise 13.9.2

If 25.00 mL of HCl solution with a concentration of 0.1234 M is neutralized by 23.45 mL of NaOH, what is the concentration of the base?

Answer:

0.1316 M NaOH

Exercise 13.9.3

A 20.0 mL solution of strontium hydroxide, $Sr(OH)_2$, is placed in a flask and a drop of indicator is added. The solution turns color after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the $Sr(OH)_2$ solution?

Answer:

$3.12\times 10^{-2}\ M$ $Sr(OH)_2$



Video 14.6.1: *https://www.youtube.com/watch?v=IKP9UayAmH0*

Indicator Selection for Titrations

What indicator is used depends on what type of titration you are performing. An indicator should be chosen that will change color when enough of one substance (acid or base) has been added to exactly use up the other substance. Only when a strong acid and a strong base are produced will the resulting solution be neutral. The three main types of acid-base titrations, and suggested indicators, are:

Titration between . . .	Indicator	Explanation
strong acid and strong base	any	
strong acid and weak base	methyl orange	changes color in the acidic range (3.2 - 4.4)
weak acid and strong base	phenolphthalein	changes color in the basic range (8.2 - 10.6)

Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all the analyte has reacted with the titrant.

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

13.9: Acid–Base Titration is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

15: Chemical Bonding

Template:HideTOC

How do atoms make compounds? Typically they join together in such a way that they lose their identities as elements and adopt a new identity as a compound. These joins are called *chemical bonds*. But how do atoms join together? Ultimately, it all comes down to electrons. Before we discuss how electrons interact, we need to introduce a tool to simply illustrate electrons in an atom.

15.1: Representing Valence Electrons with Dots

15.2: Electrons are Transferred in Ionic Compounds

15.3: Lewis Structures: Electrons Shared

15.4: Lewis Structures: Counting Valence Electrons

15.5: Predicting the Shapes of Molecules

15.6: Electronegativity and Polarity - Why Oil and Water Don’t Mix

---

15: Chemical Bonding is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.1: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram or a Lewis diagram or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (It does not matter what order the positions are used.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H\cdot} \tag{15.1.1}$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\mathbf{\dot{H}} \ \textit{or} \ \mathbf{\cdot H} \ \textit{or} \ \mathbf{\underset{\cdot}{H}} \tag{15.1.2}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He\!:} \tag{15.1.3}$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1*s* subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li\cdot} \tag{15.1.4}$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be\!:} \tag{15.1.5}$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\mathbf{\dot{B}e\!:} \tag{15.1.6}$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\mathbf{\cdot\dot{C}\!:} \tag{15.1.7}$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\mathbf{\cdot\dot{\underset{\cdot}{N}}\!:} \tag{15.1.8}$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that a side has no more than two electrons.

$$\mathbf{\cdot\ddot{\underset{\cdot}{O}}\!:} \tag{15.1.9}$$

Fluorine and neon have seven and eight dots, respectively:

$$\mathbf{:\ddot{\underset{\cdot}{F}}\!:} \tag{15.1.10}$$



(15.1.11)

With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the $n = 3$ shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

Example 15.1.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3*s* electrons:



(15.1.12)

2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the $n = 4$ shell, there are six electrons. Its electron dot diagram is as follows:



(15.1.13)

Exercise 15.1.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a



Answer b



Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis electron dot diagrams for ions have less (for cations) or more (for anions) dots than the corresponding atom.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.1: Representing Valence Electrons with Dots is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.2: Electrons are Transferred in Ionic Compounds

Learning Objectives

- State the octet rule.
- Define *ionic bond*.
- Illustrate the transfer of electrons in an ionic compound.

In Section 4.7 we saw how ions are formed by losing electrons to make cations or by gaining electrons to form anions. The astute reader may have noticed something: Many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The trend that atoms like to have eight electrons in their valence shell is called the **octet rule**. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the $n = 2$ shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:} \tag{15.2.1}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5 \tag{15.2.2}$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \rightsquigarrow \cdot\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:} \tag{15.2.3}$$

resulting in two ions—the Na^+ ion and the Cl^- ion:

$$\mathbf{Na}\cdot^{+} \qquad \mathbf{:}\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^{-} \tag{15.2.4}$$

$$[Ne] \qquad [Ne]\,3s^2 3p^6 \tag{15.2.5}$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Na}\cdot^{+} + \mathbf{:}\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^{-} \rightarrow Na^+Cl^- \;\; or \;\; NaCl \tag{15.2.6}$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

$$\mathbf{Mg^{2+} + \left[:\ddot{O}:\right]^{2-}} \quad Mg^{2+}O^{2-} \text{ or } MgO \tag{15.2.7}$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na\cdot \sim \cdot\ddot{O}:} \tag{15.2.8}$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



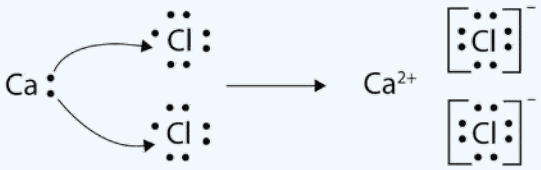
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

Example 15.2.1: Synthesis of calcium chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from *Ca* atoms and *Cl* atoms.

Solution

A *Ca* atom has two valence electrons, while a *Cl* atom has seven electrons. A *Cl* atom needs only one more to complete its octet, while *Ca* atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one *Ca* atom. The transfer process looks as follows:

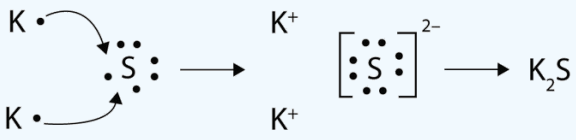


The oppositely charged ions attract each other to make $CaCl_2$.

Exercise 15.2.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from *K* atoms and *S* atoms.

Answer:



https://chem.libretexts.org/@go/page/105348

Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.2: Electrons are Transferred in Ionic Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.3: Lewis Structures: Electrons Shared

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

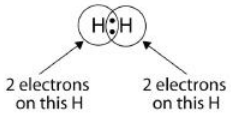
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H} \tag{15.3.1}$$

The two H atoms can share their electrons:

$$\mathbf{H:H} \tag{15.3.2}$$

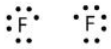
We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:

H–H

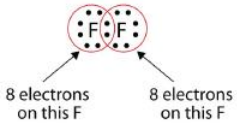
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.



Note that each F atom has a complete octet around it now:

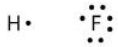


We can also write this using a dash to represent the shared electron pair:

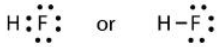


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

Example 15.3.1:

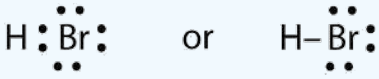
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:



The two atoms can share their unpaired electron:



Exercise 15.3.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer:



When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far and there are still some electrons hanging out there unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of $\mathbf{C_2H_4}$. The carbon atoms are bonded together, with each carbon also being bonded to two hydrogen atoms.

$$\text{two } \mathbf{C} \text{ atoms } = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four } \mathbf{H} \text{ atoms } = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

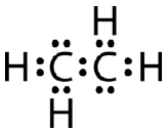


Figure **15.3.1***: Incorrect dot structure of ethene.*

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

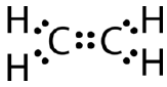


Figure **15.3.2**: *Correct dot structure for ethene.*

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

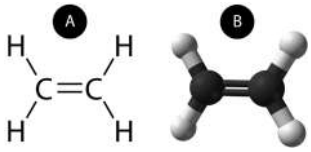


Figure **15.3.3**: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

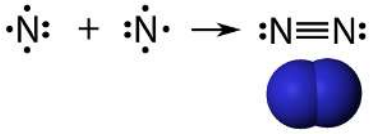


Figure **15.3.4**: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributors and Attributions

-
- Anonymous
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.3: Lewis Structures: Electrons Shared is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.4: Lewis Structures: Counting Valence Electrons

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions:

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

  For CO_3^{2-}, for example, we add two electrons to the total because of the -2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. **Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).**

- These electrons will usually be lone pairs.

5. **If any electrons are left over, place them on the central atom.**

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and each atom has an octet of electrons except for hydrogen (with two electrons).

The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 15.4.1: Water

Write the Lewis Structure for H_2O.

Solution

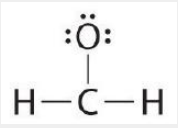
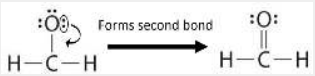
https://chem.libretexts.org/@go/page/105350

Steps for Writing Lewis Structures	Example **15.4.1**
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary
7. Final check	The Lewis structure gives oxygen an octet and each hydrogen two electrons,

Example 15.4.2

Write the Lewis structure for the CH_2O molecule

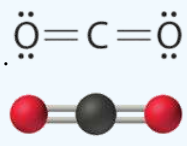
Solution

Steps for Writing Lewis Structures	Example **15.4.2**
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has one valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  Six electrons are used, and 6 are left over.
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 15.4.1

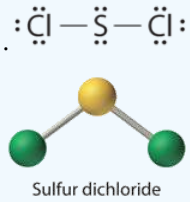
Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO_2



Carbon dioxide

Answer SCl_2



The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

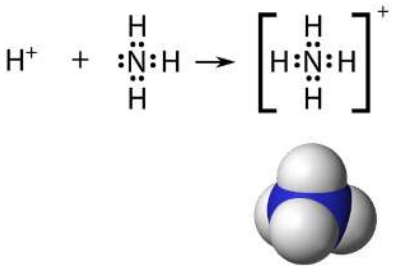


Figure 15.4.3: The ammonium ion.

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$1\ N$ atom $= 5$ valence electrons

$4\ H$ atoms $= 4 \times 1 = 4$ valence electrons

subtract 1 electron for the $1+$ charge of the ion

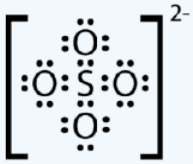
total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside the brackets.

Exercise 15.4.2

Draw the Lewis electron dot structure for the sulfate ion.

Answer



Exceptions to the Octet Rule

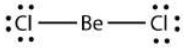
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

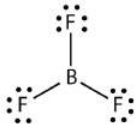


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

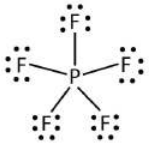
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 15.4.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

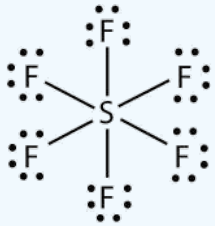
a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 15.4.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer:



The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)
-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.4: Lewis Structures: Counting Valence Electrons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.5: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is **called valence shell electron pair repulsion (VSEPR).** It basically says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from each other as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:



Figure **15.5.1**: *Beryllium hydride and carbon dioxide bonding.*

The two molecules, shown in the figure below in a "ball and stick" model.

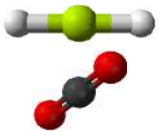


Figure **15.5.2**: *Beryllium hydride and carbon dioxide models.*

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

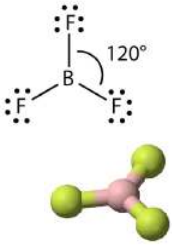


Figure **15.5.3**: *Boron trifluoride bonding.*

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

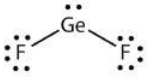


Figure **15.5.4**: *Germanium difluoride bonding.*

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure **15.5.1** Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

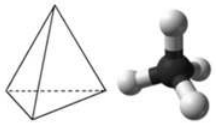


Figure **15.5.5**: *Tetrahedral structure of methane.*

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

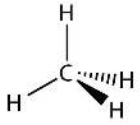


Figure **15.5.3**: *Methane bonding.*

NH_3 is an example of a molecule whose central atom has four electron groups but only three of them are bonded to surrounding atoms.

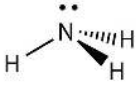


Figure **15.5.3**: *Ammonia bonding.*

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups but only two of them are bonded to surrounding atoms.



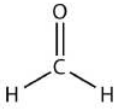
Figure **15.5.3**: *Water bonding.*

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom but only one electron group bonded to another atom is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. CH_2O has the following Lewis electron dot diagram.



The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape:



(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table **15.5.1** summarizes the shapes of molecules based on their number of electron groups and surrounding atoms.

Table 15.5.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom (Steric Number)	Number of Bonding Pairs	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

Example 15.5.1:

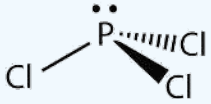
What is the approximate shape of each molecule?

a. PCl_3
b. NOF

Solution

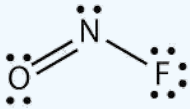
The first step is to draw the Lewis structure of the molecule.

1. For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

2. The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

Exercise 15.5.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

Exercise 15.5.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. Hint, hydrogen is a terminal atom.

Answer

> Trigonal planar about both central C atoms

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.5: Predicting the Shapes of Molecules is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.6: Electronegativity and Polarity - Why Oil and Water Don’t Mix

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Identify whether or not a molecule can exhibit hydrogen bonding.
- List important phenomena which are a result of hydrogen bonding.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur) and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure **15.6.1**.

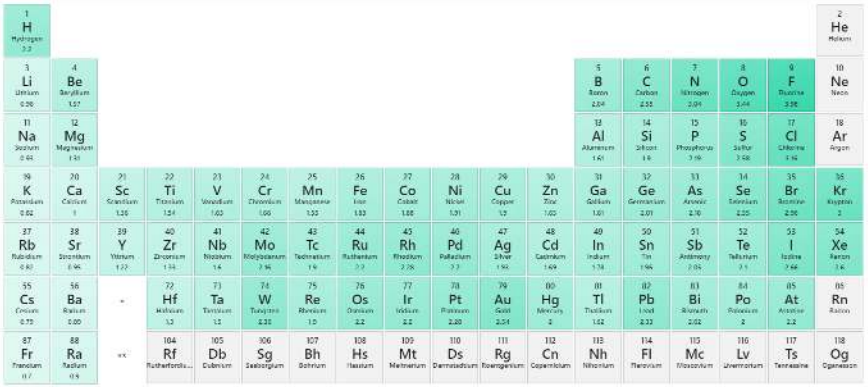


Figure **15.6.1**: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Electronegativity Difference	Bond Type
0	nonpolar covalent
0–0.4	slightly polar covalent
0.5–2.0	definitely polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

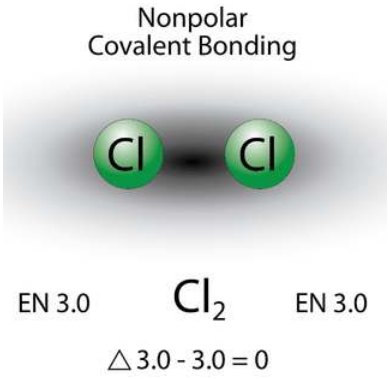


Figure **15.6.2**: *A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.*

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathbf{Cl_2}$ molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta\text{EN} = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

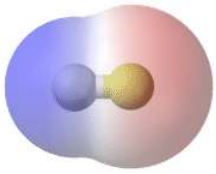


Figure **15.6.3**: In the polar covalent bond of **HF**, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

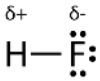


Figure **15.6.4**: Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.



Figure **15.6.5**: Use of crossed arrow to indicate polarity



Electronegativity differences in bonding using Pauling scale. Using differences in electronegativity to classify bonds as covalent, polar covalent, or ionic.

Example 15.6.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 15.6.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 − 2.1 = 0.4. Thus we predict that this bond will be non polar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 − 2.1 = 1.4, so we predict that this bond will be definitely polar covalent.

Exercise 15.6.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is frequently useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical - bonded to the same element with no unshared pairs of electrons. **Polar molecules** are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well - as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming you do, you can look at the structure of each one and decide if it is polar or not - *whether or not you know the individual atom electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as **HF**, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.



Figure **15.6.6**: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide (CO_2) is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the **C** atom to each **O** atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of CO_2 is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the **H** atoms toward the **O** atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

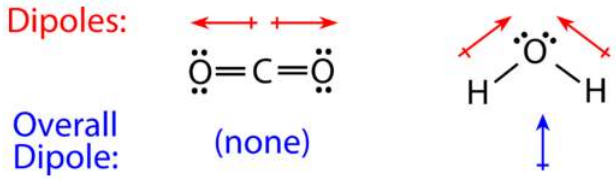


Figure **15.6.7**: *The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.*

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as CH_4 is nonpolar. However, if one of the peripheral **H** atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule (BF_3) may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule (NH_3) is polar.

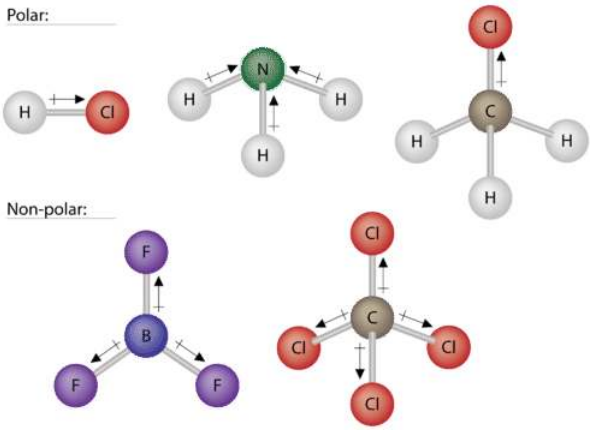


Figure **15.6.8**: Some examples of polar and nonpolar molecules based on molecular geometry.

To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment does not cancel.

Steps to Identify Polar Molecules

1. Draw the Lewis structure
2. Figure out the geometry (using VSEPR theory)
3. Visualize or draw the geometry
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it)
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure **15.6.14**). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

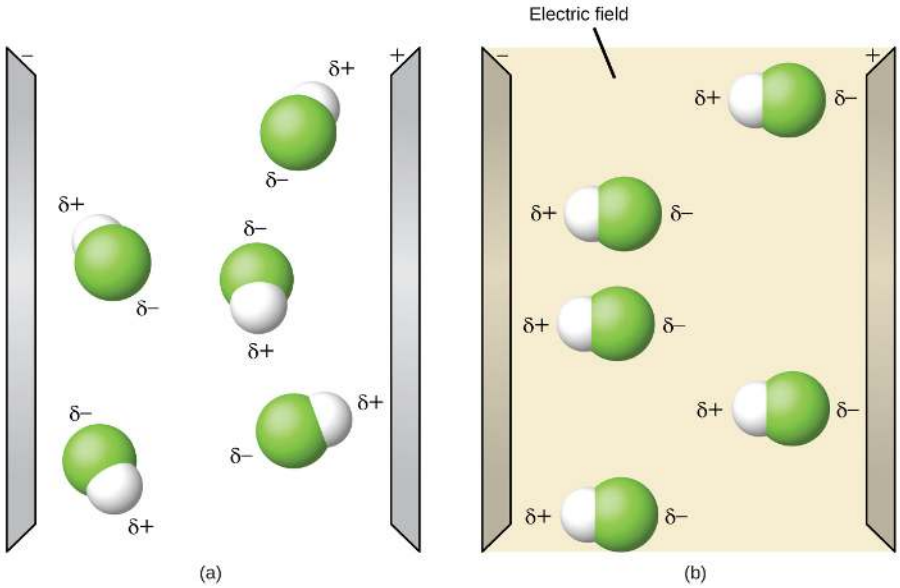


Figure **15.6.9**: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Example 15.6.2:

Label each of the following as polar or nonpolar.

a. Water, H_2O
b. Methanol, CH_3OH
c. Hydrogen Cyanide, HCN
d. Oxygen, O_2
e. Propane, C_3H_8

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.

b. Methanol is polar. This is not a symmetric molecule. The $-OH$ side is different from the other 3 $-H$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

Exercise 15.6.2

Label each of the following as polar or nonpolar.

a. SO_3
b. NH_3

Answer a

non polar

Answer b

polar

Contributors and Attributions

-
- StackExchange (thomij).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

15.6: Electronegativity and Polarity - Why Oil and Water Don't Mix is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

16: Gases

Template:HideTOC

Of the three basic phases of matter—solids, liquids, and gases—only one of them has predictable physical properties: gases. In fact, the study of the properties of gases was the beginning of the development of modern chemistry from its alchemical roots. The interesting thing about some of these properties is that they are independent of the identity of the gas. That is, it doesn't matter if the gas is helium gas, oxygen gas, or sulfur vapors; some of their behavior is predictable and, as we will find, very similar. In this chapter, we will review some of the common behaviors of gases. Gases have no definite shape or volume; they tend to fill whatever container they are in. They can compress and expand, sometimes to a great extent. Gases have extremely low densities, one-thousandth or less the density of a liquid or solid. Combinations of gases tend to mix together spontaneously; that is, they form solutions. Air, for example, is a solution of mostly nitrogen and oxygen. Any understanding of the properties of gases must be able to explain these characteristics.

16.1: Extra-Long Straws

11.2: Kinetic Molecular Theory: Real Gases are Not Ideal

16.3: Pressure: The Result of Constant Molecular Collisions

16.4: Boyle's Law: Pressure and Volume

16.5: Charles's Law: Volume and Temperature

16.6: Gay-Lussac's Law: Temperature and Pressure

16.7: The Combined Gas Law: Pressure, Volume, and Temperature

16.8: Avogadro's Law: Volume and Moles

16.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles

---

16: Gases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.1: Extra-Long Straws

A drinking straw is a tube for transferring a beverage from its container to the mouth of the drinker. A thin tube of plastic (such as polypropylene and polystyrene) or other material, straight or with an angle-adjustable bellows segment, it is used by being held with one end in the mouth and another end in the drink. Muscular action (a combination of the muscles of the tongue and cheeks) reduces air pressure in the mouth and above the liquid in the straw, whereupon atmospheric pressure forces the beverage through the straw.



Eight long drinking straws made from plastic and having flexible heads positioned next to each other in order to look like the color gradation of the rainbow. The are standing against a white background. (CC BY 2.0; Horia Varlan).

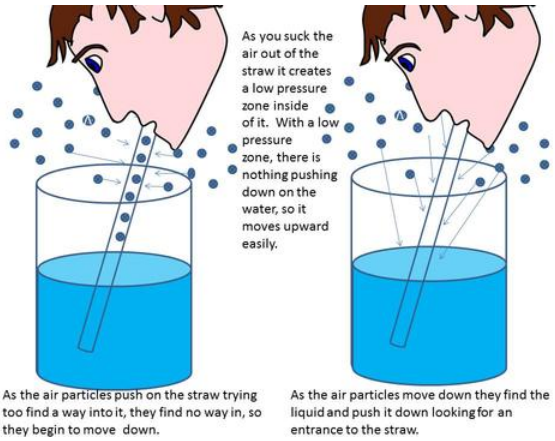
How Straws Work

With the straw just sitting in the glass, the pressure on the surface of the tea is the same all over, including on the little bit of surface inside the straw. When you suck the air out of the straw, you decrease the pressure inside the straw, allowing the higher pressure on the rest of the surface to push the tea up the straw and into your mouth. Because it is really the atmosphere that is doing the pushing, the atmospheric pressure limits how high water will go up a straw. At sea level, the air pressure is enough to support a column of water about thirty feet high.

Many people believe that when they drink a liquid they are sucking the liquid up, however the liquid is really being pushed up. A straw works because when you suck the air out of the straw it creates a vacuum. The atmospheric pressure outside the straw is greater then in the inside pressure, therefore the liquid is pushed up into your mouth

DRINKING STRAW: A drinking straw is used by creating a suction with your mouth. Actually this causes a decrease in air pressure on the inside of the straw. Since the atmospheric pressure is greater on the outside of the straw, liquid is forced into and up the straw.

SIPHON: With a siphon water can be made to flow "uphill". A siphon can be started by filling the tube with water (perhaps by suction). Once started, atmospheric pressure upon the surface of the upper container forces water up the short tube to replace water flowing out of the long tube.



Your misconception seems to be coming from the idea of a vacuum and a straw. The vacuum itself is not what causes the sucking. It is the atmospheric pressure that causes sucking.

How long of a Straw is Possible?

Straws work because sucking creates a pressure difference between the inside of the straw and the outside. If you formed a perfect vacuum within the straw, the pressure outside of the straw at sea level would be enough to push the orange soda (which is mostly water) to a total height of about 10.3 m. • A 10.3-m column of water exerts the same pressure— 101,325 N/m^2 or 14.7 lb/in^2 (psi)—as do the gas molecules in our atmosphere. *At sea level, the air pressure is enough to support a column of water about thirty feet high. This means that even if you could suck all the air out of a forty-foot straw, the water wouldn't rise more than thirty feet.*

Contributors and Attributions

- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Lisa Peck's Conceptual Physics Class

16.1: Extra-Long Straws is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.2: Kinetic Molecular Theory: Real Gases are Not Ideal

Learning Objectives

- State the major concepts behind the kinetic theory of gases.
- Relate the general properties of gases to the kinetic theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas.

Figure **11.2.1** shows a representation of how we mentally picture the gas phase.

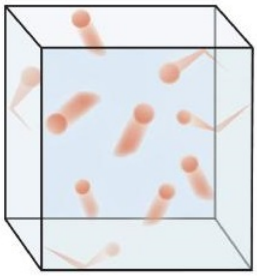


Figure **11.2.1**: The Kinetic Theory of Gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

- The physical behavior of gases is explained by the kinetic theory of gases.
- An ideal gas adheres exactly to the kinetic theory of gases.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

11.2: Kinetic Molecular Theory: Real Gases are Not Ideal is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.3: Pressure: The Result of Constant Molecular Collisions

Learning Objectives

- Define *pressure*.
- Learn the units of pressure and how to convert between them.

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$pressure = \frac{force}{area} \tag{16.3.1}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being "held in" by the earth's gravity, rather than the gas being in a container. The pressure of the atmosphere is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

However, "average atmospheric pressure at sea level" is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mmHg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalences:

$$1\ atm = 760\ mmHg = 760\ torr$$

We can use these equivalences as with any equivalences—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, 1 atm = 101,325 Pa.

Example 16.3.1: Pressure Conversion

	Example **16.3.1** How many atmospheres are there in 595 torr
Steps for Problem Solving	**Unit Conversion**
Identify the "given" information and what the problem is asking you to "find."	Given: 595 torr Find: ?atm
List other known quantities.	1 atm = 760 torr
Prepare a concept map.	torr → atm $\frac{1\ atm}{760\ torr}$
Cancel units and calculate.	$595\ \cancel{torr} \times \frac{1\ atm}{760\ \cancel{torr}} = 0.783\ atm$
Think about your result.	595 torr is less than 760 torr so the final answer should be less than 1 atm

Exercise 16.3.1

How many atmospheres are there in 1,022 torr?

Answer

1.345 atm

Example 16.3.2: Mars

	Example **16.3.2** The atmosphere on Mars is largely CO_2 at a pressure of 6.01 mmHg. What is this pressure in atmospheres?
Steps for Problem Solving	**Unit Conversion**
Identify the "given" information and what the problem is asking you to "find."	Given: 6.01mmHg Find: ?atm
List other known quantities.	1 atm = 760 mmHg
Prepare a concept map.	mmHg → atm $\frac{1\ atm}{760\ mmHg}$
Cancel units and calculate.	$6.01\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}} = 0.00791\ atm = 7.91 \times 10^{-3} atm$
Think about your result.	6.01 is a very small number relative to 760 mmHg just like the value in atmospheres.

Exercise 16.3.2

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer

652 torr

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.3: Pressure: The Result of Constant Molecular Collisions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.4: Boyle's Law: Pressure and Volume

Learning Objectives

- Learn what is meant by the term *gas laws*.
- Learn and apply Boyle's law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Take pressure (P) and volume (V), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [n]), if the temperature (T) of the gas was kept constant, pressure and volume were related: As one increases, the other decreases. As one decreases, the other increases. We say that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{ constant at constant n and T} \tag{16.4.1}$$

If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled P_1 and V_1 and the new conditions are labeled P_2 and V_2, we have

$$P_1V_1 = \text{constant} = P_2V_2 \tag{16.4.2}$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1V_1 = P_2V_2 \text{ at constant n and T} \tag{16.4.3}$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law ia called **Boyle's law**, after the English scientist Robert Boyle, who first announced it in 1662. Figure **16.4.1** shows two representations of how Boyle's law works.

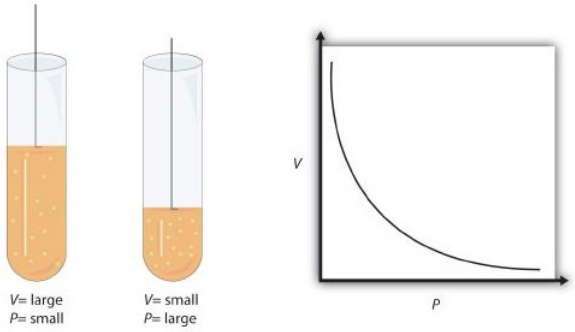


Figure **16.4.1**: Boyle's Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot P versus V for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

Boyle's law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle's law there are two pressure variables; they must have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won't matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

Example 16.4.1

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: P_1 = 2.44 atm and V_1 = 4.01 L P_2 = 1.93 atm Find: V_2 = ? L
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{P_1 \times V_1}{P_2} \quad (16.4.4)$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.44 \cancel{\text{ atm}} \times 4.01 \text{ L}}{1.93 \cancel{\text{ atm}}} = 5.07 \text{ L} \quad (16.4.5)$$
Think about your result.	We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answer makes sense based on Boyle's law.

Exercise 16.4.1

If P_1 = 334 torr, V_1 = 37.8 mL, and P_2 = 102 torr, what is V_2?

Answer

124 mL

As mentioned, you can use any units for pressure or volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

Example 16.4.2:

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: P_1 = 722 torr and V_1 = 88.8 mL V_2 = 0.633 L Find: P_2 = ? torr
List other known quantities.	1 L = 1000 mL to have the same units for volume.
Plan the problem.	1. Perform the conversion of the second volume unit from L to mL. 2. Rearrange the equation algebraically to solve for P_2. $$P_2 = \frac{P_1 \times V_1}{V_2} \quad (16.4.6)$$
Cancel units and calculate.	1. $$0.663 \cancel{L} \times \frac{1000\,ml}{1\,\cancel{L}} = 663\,ml \quad (16.4.7)$$ 2. Substitute the known quantities into the equation and solve. $$P_2 = \frac{722\ \text{torr} \times 88.8\ \cancel{mL}}{663\ \cancel{mL}} = 96.7\ \text{torr} \quad (16.4.8)$$
Think about your result.	When the volume increased, the pressure decreased, which is as expected for Boyle's law.

Exercise 16.4.2

If V_1 = 456 mL, P_1 = 308 torr, and P_2 = 1.55 atm, what is V_2?

Answer

119 mL

Summary

- The behavior of gases can be modeled with gas laws.
- Boyle's law relates a gas's pressure and volume at constant temperature and amount.

16.4: Boyle's Law: Pressure and Volume is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.5: Charles's Law: Volume and Temperature

Learning Objectives

- Learn and apply Charles's Law

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end-result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

French physicist Jacques Charles (1746 - 1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stoppage of molecular motion.

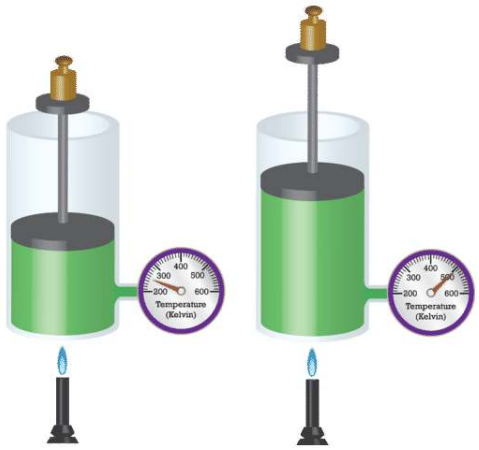


Figure **16.5.1**: *As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.*

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k \tag{16.5.1}$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 16.5.1: Temperature-Volume Data

Temperature (K)	Volume (mL)	$\frac{V}{T} = k \left(\frac{\text{mL}}{\text{K}}\right)$
50	20	0.40
100	40	0.40
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

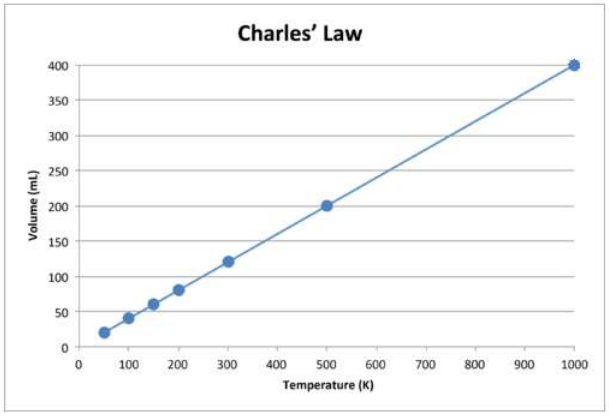


Figure **16.5.2**:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \tag{16.5.2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\mathrm{K} = {}^\circ\mathrm{C} + 273$.

Example 16.5.1:

A balloon is filled to a volume of $2.20\ \mathrm{L}$ at a temperature of $22^\circ\mathrm{C}$. The balloon is then heated to a temperature of $71^\circ\mathrm{C}$. Find the new volume of the balloon.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: $V_1 = 2.20\ \text{L}$ and $T_1 = 22°\text{C} = 295\ \text{K}$ $T_2 = 71°\text{C} = 344\ \text{K}$ Find: V_2 = ? L
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times T_2}{T_1} \quad (16.5.3)$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.20\ \text{L} \times 344\ \cancel{\text{K}}}{295\ \cancel{\text{K}}} = 2.57\ \text{L} \quad (16.5.4)$$
Think about your result.	The volume increases as the temperature increases. The result has three significant figures.

Exercise 16.5.1

If V_1 = 3.77 L and T_1 = 255 K, what is V_2 if T_2 = 123 K?

Answer

1.82 L

Example 16.5.2:

A sample of a gas has an initial volume of 34.8 L and an initial temperature of −67°C. What must be the temperature of the gas for its volume to be 25.0 L?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: Given:T_1 = -27°C and V_1 = 34.8 L V_2 = 25.0 L Find: T_2 = ? K
List other known quantities.	K = -27°C + 273
Plan the problem.	1. Convert the initial temperature to Kelvin 2. Rearrange the equation algebraically to solve for T_2. $$T_2 = \frac{V_2 \times T_1}{V_1} \quad (16.5.5)$$
Cancel units and calculate.	1. −67°C + 273 = 206 K 2. Substitute the known quantities into the equation and solve. $$T_2 = \frac{25.0 \cancel{L} \times 206 \text{ K}}{34.8 \cancel{L}} = 148 \text{ K} \quad (16.5.6)$$
Think about your result.	This is also equal to −125°C. As temperature decreases, volume decreases, which it does in this example.

Exercise 16.5.2

If V_1 = 623 mL, T_1 = 255°C, and V_2 = 277 mL, what is T_2?

Answer

235 K, or −38°C

Summary

- Increasing the temperature of a gas at constant pressure will produce and increase in the volume.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.5: Charles's Law: Volume and Temperature is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.6: Gay-Lussac's Law: Temperature and Pressure

Learning Objectives

- Explain the following laws within the Ideal Gas Law

Propane tanks are widely used with barbeque grills. But it's not fun to find out half-way through your grilling that you've run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778 - 1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.



Figure **16.6.1**: *Joseph Gay-Lussac.*

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \text{ and } \frac{P_1}{T_1} = \frac{P_2}{T_2} \tag{16.6.1}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume its pressure continually decreases until the gas condenses to a liquid.

Example 16.6.1

The gas in an aerosol can is under a pressure of $3.00\ \text{atm}$ at a temperature of 25°C. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of 845°C?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: $P_1 = 3.00\ \text{atm}$ $T_1 = 25°\text{C} = 298\ \text{K}$ $T_2 = 845°\text{C} = 1118\ \text{K}$ Find: $P_2 = ?\ \text{atm}$
List other known quantities	The temperatures have first been converted to Kelvin.
Plan the problem	First, rearrange the equation algebraically to solve for P_2. $$P_2 = \frac{P_1 \times T_2}{T_1} \quad (16.6.2)$$
Calculate	Now substitute the known quantities into the equation and solve. $$P_2 = \frac{3.00\ \text{atm} \times 1118\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.3\ \text{atm} \quad (16.6.3)$$
Think about your result.	The pressure increases dramatically due to large increase in temperature.

Summary

- Pressure and temperature at constant volume are directly proportional.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.6: Gay-Lussac's Law: Temperature and Pressure is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.7: The Combined Gas Law: Pressure, Volume, and Temperature

Learning Objectives

- Learn and apply the combined gas law.

One thing we notice about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that we can propose a gas law that combines pressure, volume, and temperature. This gas law is known as the **combined gas law**, and its mathematical form is

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \; at \; constant \; n \tag{16.7.1}$$

This allows us to follow changes in all three major properties of a gas. Again, the usual warnings apply about how to solve for an unknown algebraically (isolate it on one side of the equation in the numerator), units (they must be the same for the two similar variables of each type), and units of temperature must be in Kelvin.

Example 16.7.1:

A sample of gas at an initial volume of 8.33 L, an initial pressure of 1.82 atm, and an initial temperature of 286 K simultaneously changes its temperature to 355 K and its volume to 5.72 L. What is the final pressure of the gas?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: V_1 = 8.33 L, P_1 = 1.82 atm, and T_1 = 286 K V2 = 5.72 L and T_2 = 355 K Find: P_2 = ? atm
List other known quantities	none
Plan the problem	First, rearrange the equation algebraically to solve for V_2. $P_2 = \frac{P_1V_1T_2}{T_1V_2}$
Calculate	Now substitute the known quantities into the equation and solve. $P_2 = \frac{(1.82\,atm)(8.33\,\cancel{L})(355\,\cancel{K})}{(286\,\cancel{K})(5.72\,\cancel{L})} = 3.22atm$ (16.7.2)
Think about your result.	Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas were changing.

Exercise 16.7.1

If P_1 = 662 torr, V_1 = 46.7 mL, T_1 = 266 K, P_2 = 409 torr, and T_2 = 371 K, what is V_2?

Answer

105 mL

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be in the numerator and all by itself on one side of the equation.

- The combined gas law relates pressure, volume, and temperature of a gas.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.7: The Combined Gas Law: Pressure, Volume, and Temperature is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.8: Avogadro's Law: Volume and Moles

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. How much air should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles (or number of particles) of gas when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n \tag{16.8.1}$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \tag{16.8.2}$$

where n is the number of moles of gas and k is a constant. Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

Example 16.8.1

A balloon has been filled to a volume of $1.90\ \text{L}$ with $0.0920\ \text{mol}$ of helium gas. If $0.0210\ \text{mol}$ of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: $V_1 = 1.90\ \text{L}$ $n_1 = 0.0920\ \text{mol}$ Find: $V_2 = ?\ \text{L}$
List other known quantities	Note that the final number of moles has to be calculated by adding the original number of moles to the moles of added helium. $n_2 = 0.0920 + 0.0210 = 0.1130\ \text{mol}$
Plan the problem	First, rearrange the equation algebraically to solve for V_2. $V_2 = \frac{V_1 \times n_2}{n_1}$ (16.8.3)
Calculate	Now substitute the known quantities into the equation and solve. $V_2 = \frac{1.90\ \text{L} \times 0.1130\ \cancel{\text{mol}}}{0.0920\ \cancel{\text{mol}}} = 2.33\ \text{L}$ (16.8.4)
Think about your result.	Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly.

Exercise 16.8.1

A 12.8 L volume of gas contains .000498 moles of oxygen gas. At constant temperature and pressure, what volume does .0000136 moles of the gas fill?

Answer

0.350 L

Summary

- Calculations are shown for relationships between volume and number of moles of a gas.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.8: Avogadro's Law: Volume and Moles is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles

Learning Objectives

- Explain the Ideal Gas Law

There are a number of chemical reactions that require ammonia. In order to carry out the reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

Ideal Gas Law

The combined gas law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. putting these together leaves us with the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2} \tag{16.9.1}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **ideal gas law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable R for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R \tag{16.9.2}$$

The ideal gas law is conveniently rearranged to look this way, with the multiplication sings omitted:

$$PV = nRT \tag{16.9.3}$$

The variable R in the equation is called the **ideal gas constant**.

Evaluating the Ideal Gas Constant

The value of R, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: $\mathbf{kPa}$, $\mathbf{atm}$, or $\mathbf{mm\ Hg}$. Therefore, R can have three different values.

We will demonstrate how R is calculated when the pressure is measured in $\mathbf{kPa}$. The volume of $\mathbf{1.00\ mol}$ of any gas at STP (Standard temperature, 273.15 K and pressure, 1 atm)is measured to be $\mathbf{22.414\ L}$. We can substitute $\mathbf{101.325\ kPa}$ for pressure, $\mathbf{22.414\ L}$ for volume, and $\mathbf{273.15\ K}$ for temperature into the ideal gas equation and solve for R.

$$\begin{aligned} R &= \frac{PV}{nT} \\ &= \frac{101.325\ \text{kPa} \times 22.414\ \text{L}}{1.000\ \text{mol} \times 273.15\ \text{K}} \\ &= 8.314\ \text{kPa} \cdot \text{L/K} \cdot \text{mol} \end{aligned}$$

This is the value of R that is to be used in the ideal gas equation when the pressure is given in $\mathbf{kPa}$. The table below shows a summary of this and the other possible values of R. It is important to choose the correct value of R to use for a given problem.

Table 16.9.1: Values of the Ideal Gas Constant

Unit of P	Unit of V	Unit of n	Unit of T	Value and Unit of R
kPa	L	mol	K	$8.314\ \text{J/K} \cdot \text{mol}$
atm	L	mol	K	$0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol}$
mm Hg	L	mol	K	$62.36\ \text{L} \cdot \text{mm Hg/K} \cdot \text{mol}$

Notice that the unit for R when the pressure is in kPa has been changed to $\mathrm{J/K \cdot mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

Example 16.9.1 Oxygen Gas

What volume is occupied by 3.76 g of oxygen gas at a pressure of 88.4 kPa and a temperature of 19°C? Assume the oxygen is ideal.

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: • $P = 88.4\ \mathrm{kPa}$ • $T = 19°\mathrm{C} = 292\ \mathrm{K}$ Mass $\mathrm{O_2} = 3.76\ \mathrm{g}$ Find: V = ? L
List other known quantities	$\mathrm{O_2} = 32.00\ \mathrm{g/mol}$ $R = 8.314\ \mathrm{J/K \cdot mol}$
Plan the problem	1. First, determine the number of moles of O_2 from the given mass and the molar mass. 2. Then, rearrange the equation algebraically to solve for V $$V = \frac{nRT}{P}$$
Calculate	1. $$3.76\ \cancel{\mathrm{g}} \times \frac{1\ \mathrm{mol\ O_2}}{32.00\ \cancel{\mathrm{g}}\ \mathrm{O_2}} = 0.1175\ \mathrm{mol\ O_2}$$ 2. Now substitute the known quantities into the equation and solve. $$V = \frac{nRT}{P}$$ $$= \frac{0.1175\ \cancel{\mathrm{mol}} \times 8.314\ \mathrm{J/\cancel{K} \cdot \cancel{mol}} \times 292\ \cancel{\mathrm{K}}}{88.4\ \cancel{\mathrm{kPa}}}$$ $$= 3.23\ \mathrm{L\ O_2}$$
Think about your result.	The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar volume (22.4 L/mol) since the pressure and temperature are reasonably close to standard. The result has three significant figures because of the values for T and P. Since a joule $(\mathrm{J}) = \mathrm{kPa \cdot L}$, the units cancel out correctly, leaving a volume in liters.

Example 16.9.2: Argon Gas

A 4.22 mol sample of Ar has a pressure of 1.21 atm and a temperature of 34°C. What is its volume?

Solution

Steps for Problem Solving	
Identify the "given"information and what the problem is asking you to "find."	Given: n = 4.22 mol P = 1.21 atm T = 34°C Find: V = ? L
List other known quantities	none
Plan the problem	1. The first step is to convert temperature to kelvin. 2. Then, rearrange the equation algebraically to solve for V $$V = \frac{nRT}{P} \quad (16.9.4)$$
Calculate	1. 34 + 273 = 307 K 2. Now substitute the known quantities into the equation and solve. $$V = \frac{(4.22\ \cancel{mol})(0.08205\ \frac{L.\cancel{atm}}{\cancel{mol}.\cancel{K}})(307\ \cancel{K})}{1.21\ \cancel{atm}} \quad (16.9.5)$$ V = 87.9 L
Think about your result.	The number of moles of Ar is large so the expected volume should also be large.

Exercise 16.9.1

A 0.0997 mol sample of O_2 has a pressure of 0.692 atm and a temperature of 333 K. What is its volume?

Answer

3.94 L

Exercise 16.9.2

For a 0.00554 mol sample of H_2, P = 23.44 torr and T = 557 K. What is its volume?

Answer

8.21 L

Summary

- The ideal gas constant is calculated.
- An example of calculations using the ideal gas law is shown.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

16.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

2.4: Solid, Liquid, and Gas

Water can take many forms. At low temperatures (below $0°C$), it is a solid. When at "normal" temperatures (between $0°C$ and $100°C$), it is a liquid. While at temperatures above $100°C$, water is a gas (steam). The state the water is in depends upon the temperature. Each state (solid, liquid, and gas) has its own unique set of physical properties.

Matter and Its States

Matter typically exists in one of three states: **solid**, **liquid**, or **gas**. The state a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states.



Figure 2.4.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From top to bottom: quartz (solid), water (liquid), nitrogen dioxide (gas),

Note

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

Solid

Solids are defined by the following characteristics:

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

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

Liquid

Liquids have the following characteristics:

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

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



Mercury.

If we heat liquid mercury to its boiling point of **357°C**, and under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gas

Gases have the following characteristics:

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

Summary

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

Contributors and Attributions

-

2.4: Solid, Liquid, and Gas is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

3: Introduction to Energy

Template:HideTOC

In previous science classes you may have learned that one way to distinguish chemical changes from physical changes is that physical changes - such as the melting and freezing of water - are reversible but that chemical changes are not. In this unit we will see that this simple answer is not necessarily what it seems.

3.1: Conservation of Mass - There is No New Matter

3.2: Energy

3.3: Energy and Chemical and Physical Change

3.4: Temperature: Random Motion of Molecules and Atoms

3.5: Temperature Changes - Heat Capacity

3: Introduction to Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.1: Conservation of Mass - There is No New Matter

Template:HideTOC

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Look at Figure **3.1.1** below. It shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose you had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose you had been able to measure the oxygen used by the fire and the gases produced by the fire. What would you find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

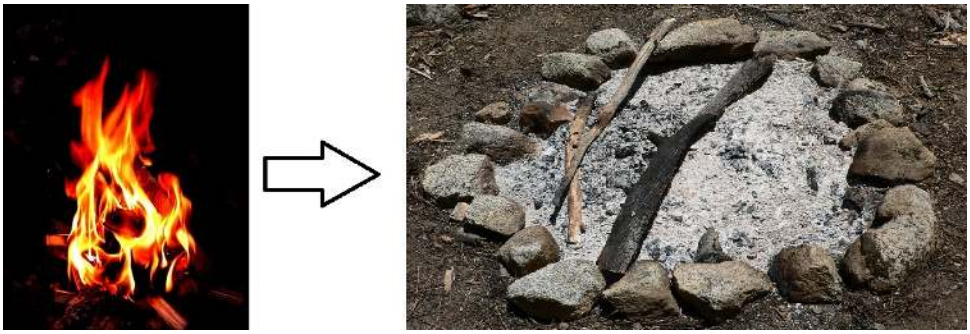


Figure **3.1.1**: Burning is a chemical process. The flames caused as a result of a fuel undergoing combustion (burning). Images used with permission (CC BY-SA 2.5; Einar Helland Berger for fire and Walter Siegmund for ash).

Law of Conservation of Mass

The law of conservation of mass was created in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases, equals the original mass of the charcoal and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. A reactant is when two or more elements chemically interact to make a new substance and a product is the substance that is formed as the result of a chemical reaction (Video **3.1.1**). Mass and matter may not be able to be created or destroyed, but it can change forms to other substances like liquids, gasses, solids, etc.



Video* 3.1.1*: *This is a nice little demonstration showing the Conservation of Mass in action.*

It's important to know the law well because if you're in a situation where there's a 300 kg tree that is burning down, when it is done burning down, there are only ashes left and all of them together weigh 10 kg. Well, it makes you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke so the only thing left that you could see was the 10 kg of ash, and if you knew the law of conservation of mass then you would know that the other 290 kg had to go somewhere because it had to equal the mass of the tree before it burnt down.

Example 3.1.1

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

Solution

$$\text{Mass of the reactants} = \text{Mass of the products}$$

$$10.0\,\text{g of } CaCO_3 = 4.4\,\text{g of } CO_2 + 5.6\,\text{g of CaO}$$

$$10.0\,\text{g of reactant} = 10.0\,\text{g of products}$$

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

Exercise 3.1.1

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

Answer

276.4 g of potassium carbonate

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

Summary

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

Contributors and Attributions

- Binod Shrestha (University of Lorraine)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

3.1: Conservation of Mass - There is No New Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.2: Energy

Template:HideTOC

Learning Objectives

- Define heat and work.
- Distinguish between kinetic energy and potential energy.
- State the law of conservation of matter and energy.

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we're awake or asleep.

Ability to Do Work or Produce Heat

When we speak of *using* energy, we are really referring to *transferring* energy from one place to another. When you use energy to throw a ball, you transfer energy from your body to the ball, and this causes the ball to fly through the air. When you use energy to warm your house, you transfer energy from the furnace to the air in your home, and this causes the temperature in your house to rise. Although energy is used in many kinds of different situations, all of these uses rely on energy being transferred in one of two ways. Energy can be *transferred* as *heat* or as *work*.

When scientists speak of ***heat***, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot – the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure 3.2.1).



Figure 3.2.1: *Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the food.*

Heat is only one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is *force (any push or pull) applied over a distance*. Whenever you push an object and cause it to move, you've done work, and you've transferred some of *your* energy to the object. At this point, it's important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but it isn't true in science. By definition, scientific work requires that force be applied *over a distance*. It does not matter how hard you push or how hard you pull. If you have not moved the object, you haven't done any work.

So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work. But the question still remains – *what IS energy*? We'll try to at least partially tackle that question in the next section.

Kinetic Energy

Machines use energy, our bodies use energy, energy comes from the sun, energy comes from volcanoes, energy causes forest fires, and energy helps us to grow food. With all these seemingly different types of energy, it's hard to believe that there are really only two different *forms* of energy – *kinetic energy* and *potential energy*. **Kinetic energy** is energy associated with motion. When an object is moving, it has kinetic energy. When the object stops moving, it has no kinetic energy. While all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you're from California, you might have driven through the Tehachapi Pass near Mojave or the Montezuma Hills in Solano County and seen the windmills lining the slopes of the mountains (Figure **3.2.2**). These are two of the larger wind farms in North America. As wind rushes along the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.



Figure **3.2.2**: A wind farm in Solano County harnesses the kinetic energy of the wind. Attribution: BDS2006 at the English Wikipedia, CC-BY-SA-3.0 (http://creativecommons.org/licenses/by-sa/3.0/)], via Wikimedia Commons

Potential Energy

Potential energy is *stored* energy. It is energy that remains available until we choose to use it. Think of a battery in a flashlight. If you leave a flashlight on, the battery will run out of energy within a couple of hours, and your flashlight will die. If, however, you only use the flashlight when you need it, and you turn it off when you don't, the battery will last for days or even months. The battery contains a certain amount of energy, and it will power the flashlight for a certain amount of time, but because the battery stores *potential* energy, you can choose to use the energy all at once, or you can save it and only use a small amount at a time.

Any stored energy is potential energy. Unfortunately, there are a lot of different ways in which energy can be stored, and that can make potential energy very difficult to recognize. In general, an object has potential energy because of its *position relative to another object*. For example when you hold a rock above the earth, it has potential energy because of its position relative to the ground. You can tell that this is *potential energy* because the energy is *stored* for as long as you hold the rock in the air. Once you drop the rock, though, the stored energy is released as kinetic energy as the rock falls.

Chemical Energy

There are other common examples of potential energy. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When you hold two magnets next to each other, they store potential energy too. For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

As you learned earlier, the world, and all of the chemicals in it are made up of atoms and molecules. These store potential energy that is dependent on their positions relative to one another. Of course, you can't see atoms and molecules. Nevertheless, scientists do know a lot about the ways in which atoms and molecules interact, and this allows them to figure out how much potential energy is stored in a specific quantity (like a cup or a gallon) of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

> Since different chemicals have different amounts of potential energy, scientists will sometimes say potential energy depends not only on ***position***, but also on ***composition***. Composition affects potential energy because it determines which molecules and atoms end up next to each other. For example, the total potential energy in a cup of pure water is different than the total potential energy in a cup of apple juice, because the cup of water and the cup of apple juice are *composed* of different amounts of different chemicals.

At this point, you might be wondering just how useful chemical potential energy is. If you want to release the potential energy stored in an object held above the ground, you just drop it. But how do you get potential energy out of chemicals? It's actually not that difficult. You use the fact that different chemicals have *different amounts of potential energy*. If you start with chemicals that

have a lot of potential energy and allow them to react and form chemicals with less potential energy, all the extra energy that was in the chemicals at the beginning but not at the end is released.

Units of Energy

Energy 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.

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories} \tag{3.2.1}$$

To say that the snack "contains" 85 Calories means that $85\ kcal$ of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1\ J = 0.2390\ cal \text{ or } 1\ cal = 4.184\ J \tag{3.2.2}$$

We can calculate the amount of heat released in kilojoules when a 400. Calorie hamburger is digested.

$$400\ Cal = 400\ kcal \times \frac{4.184\ kJ}{1\ kcal} = 1.67 \times 10^3\ kJ \tag{3.2.3}$$

Summary

- Any time we use energy, we transfer energy from one object to another.
- Energy can be transferred in one of two ways – as heat, or as work.
- Heat is the term given to energy that is transferred from a hot object to a cooler object due to the difference in their temperatures.
- Work is the term given to energy that is transferred as a result of a force applied over a distance.
- Energy comes in two fundamentally different forms – kinetic energy and potential energy.
- Kinetic energy is the energy of motion.
- Potential energy is stored energy that depends on the position of an object relative to another object.
- Chemical potential energy is a special type of potential energy that depends on the positions of different atoms and molecules relative to one another.
- Chemical potential energy can also be thought of as depending on chemical composition.
- Energy can be converted from one form to another. The total amount of mass and energy in the universe is conserved.

Contributors and Attributions

-
- Wikibooks
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

3.2: Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.3: Energy and Chemical and Physical Change

Template:HideTOC

Learning Objectives

- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

So far we've talked about how energy exists as either *kinetic energy* or *potential energy* and how energy can be transferred as either *heat* or *work*. While it's important to understand the difference between kinetic energy and potential energy and the difference between heat and work, the truth is, energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Likewise, energy that is transferred as work might later end up transferred as heat, while energy that is transferred as heat might later end up being used to do work.

Even though energy can change form, it must still follow one fundamental law – *Energy cannot be created or destroyed, it can only be changed from one form to another*. This law is known as the **Law of Conservation of Energy**. In a lot of ways, energy is like money. You can exchange quarters for dollar bills and dollar bills for quarters, but no matter how often you convert between the two, you will not end up with any more or any less money than you started with. Similarly, you can transfer (or spend) money using cash, or transfer money using a credit card, but you still spend the same amount of money, and the store still makes the same amount of money.

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 just keep warm when it's cold outside.



An image of a campfire with colored flames, made by the burning of a garden hose in a copper pipe. (CC SA-BY 3.0; Jared)

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, called 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 is 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.

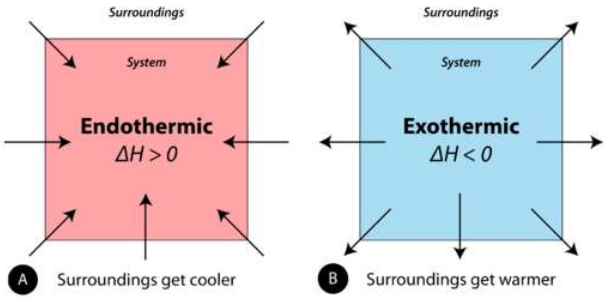


Figure 3.3.1: (A) Endothermic reaction. (B) Exothermic reaction.

During phase changes, energy changes are usually involved. For example, when solid dry ice vaporizes (physical change), carbon dioxide molecules absorb energy. Meanwhile, when liquid water becomes ice energy is released. Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have **potential energy** or "stored energy". Because we are changing the bonding, this means we are also changing how much of this "stored energy" there is in a reaction.

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the stored/hidden energy of the reactants and products as well as the activation energy. If, on an energy diagram, the products have more stored energy than the reactants started with, the reaction is endothermic.. You had to give the reaction energy. If, on the energy diagram, the products have less stored energy than the reactants started with, the reaction is exothermic.

Example 3.3.1

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. gasoline burning
c. ice forming on a pond

Solution

a. endothermic - you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b. exothermic - when you burn something, it feels hot to you because it is giving off heat into the surroundings.
c. exothermic - think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

Exercise 3.3.1

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer (a)

exothermic

Answer (b)

endothermic

Summary

Phase changes involve changes in energy. All chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy. Reactions that absorb energy are endothermic. Reactions that release energy are exothermic.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

3.3: Energy and Chemical and Physical Change is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.4: Temperature: Random Motion of Molecules and Atoms

Template:HideTOC

Learning Objectives

- To identify the different between temperature and heat
- To recognize the different scales used to measuring temperature

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



Figure **3.4.1**: *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 **3.4.2**: *Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right)*

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32°F$ 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.

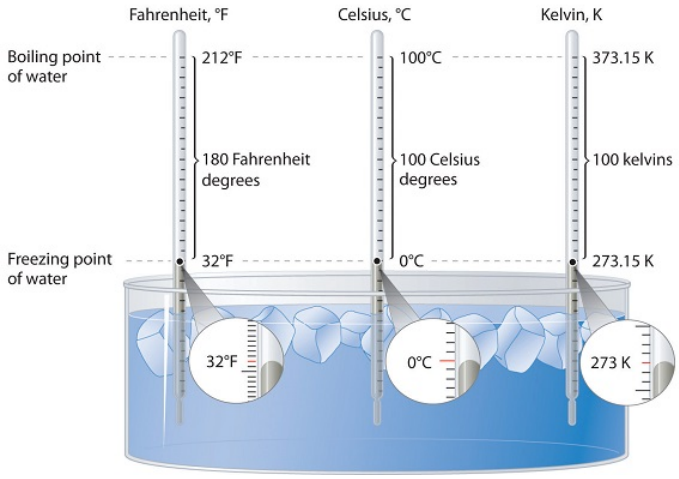


Figure 3.4.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In 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:

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

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

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales are as follows:

$$°C = \frac{5}{9} \times (°F - 32) \tag{3.10.3}$$

$$°F = \frac{9}{5} \times (°C) + 32 \tag{3.10.4}$$

Example 3.4.1: Temperature Conversions

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

Solution

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

$$°C = \frac{5}{9} \times (103.5°F - 32) \tag{3.4.1}$$
$$= 39.7\,°C \tag{3.4.2}$$

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

$$K = 39.7\,°C + 273.15 \tag{3.4.3}$$
$$= 312.9\,K \tag{3.4.4}$$

Exercise 3.4.1

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 K, 9980 °F

Answer (b)

2807 K, 5084 °F

Answer (c)

-195.79 K, -320.42 °F

Summary

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

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

3.4: Temperature: Random Motion of Molecules and Atoms is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

3.5: Temperature Changes - Heat Capacity

Template:HideTOC

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. The heat capacity of an object depends both on its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the wading pool.

Heat Capacity and Specific Heat

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot. We would say that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by $1°C$). Water is very resistant to changes in temperature, while metals in general are not. The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1°C$. The symbol for specific heat is c_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree ($J/g°C$) or calories per gram per degree ($cal/g°C$). This text will use $J/g°C$ for specific heat.

Notice that water has a very high specific heat compared to most other substances. Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see table above). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.



Figure **3.5.1:** *This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.*

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by $1°C$).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1°C$.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

3.5: Temperature Changes - Heat Capacity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

9: Precipitation Reactions

Template:HideTOC

Topic hierarchy

9.1: Aqueous Solutions and Solubility: Compounds Dissolved in Water

9.2: Precipitation Reactions

9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

9: Precipitation Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

9.1: Aqueous Solutions and Solubility: Compounds Dissolved in Water

Learning Objectives

- Define and give examples of electrolytes
- Use the solubility rules to predict if a compound is soluble, insoluble, or slightly soluble

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called electrolytes. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 9.1.1).

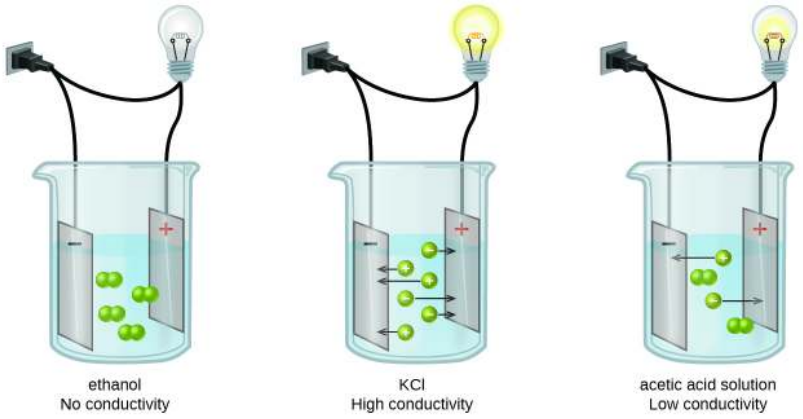


Figure 9.1.1: Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 9.1.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

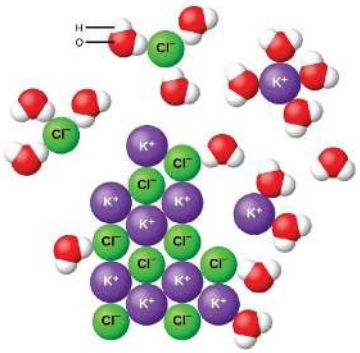


Figure 9.1.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (Table 9.1.1).

***Table 9.1.1**: Solubility Rules*

Solubility of Common Compounds in Water

Negative Ions (Anions)	Positive Ions (Cations)	Compounds
essentially all	alkali ions	soluble
essentially all	hydrogen ion, H^+	soluble
essentially all	ammonium ion, NH_4^+	soluble
nitrate, NO_3^-	essentially all	soluble
acetate, CH_3COO^-	essentially all	soluble
chloride, Cl^- bromide, Br^- iodide, I^-	Ag^+, Pb^{2+}, Hg_2^{2+}, Cu^{1+} all others	*low solubility* soluble
sulfate, SO_4^{2-}	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Ra^{2+}, Ag^+, Hg^{2+} all others	*low solubility* soluble
sulfide, S^{2-}	alkali ions*, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+} all others	soluble *low solubility*
hydroxide, OH^-	alkali ions*, H^+, NH_4^+, Sr^{2+}, Ba^{2+}, Ra^{2+} all others	soluble *low solubility*
phosphate, PO_4^{3-} carbonate, CO_3^{2-} sulfite, SO_3^{2-}	alkali ions*, H^+, NH_4^+ all others	soluble *low solubility*

*Alkali ions = Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+

***Low solubility* means a precipitate will form**

Example 9.1.1: Solubility

Classify each compound as soluble or insoluble

a. $Zn(NO_3)_2$
b. $PbBr_2$
c. $Sr_3(PO_4)_2$

Solution

a. All nitrates are soluble in water so $Zn(NO_3)_2$ is soluble.
b. All bromides are soluble in water, except those combined with Pb^{2+} so $PbBr_2$ is insoluble.
c. All phosphates are insoluble, so $Sr_3(PO_4)_2$ is insoluble

Exercise 9.1.1: Solubility

Classify each compound as soluble or insoluble.

a. $Mg(OH)_2$
b. KBr
c. $Pb(NO_3)_2$

Answer a:
insoluble

Answer b:
soluble

Answer c:
soluble

Summary

Substances that dissolve in water to yield ions are called electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water. Solubility rules allow prediction of what products will be insoluble in water.

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

9.1: Aqueous Solutions and Solubility: Compounds Dissolved in Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

9.2: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate a reddish precipitate of silver dichromate is produced.

$$AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + KNO_3(aq) \tag{9.2.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$AC + BD \rightarrow \underset{insoluble}{AD} + BC \tag{9.2.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called *double-displacement reactions*. Two important uses of precipitation reactions are to isolate metals that have been extracted from their ores and to recover precious metals for recycling.



***Video**: Mixing Potassium dicromate and Silver Nitrate together to initiate a precipitation reaction (Equation ???).*

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $NaCl$ solution is mixed with 500 mL of aqueous KBr solution, the final solution has a volume of 1.00 L and contains $Na^+(aq)$., $Cl^-(aq)$., $K^+(aq)$., and $Br^-(aq)$. As you will see in (Figure 9.2.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

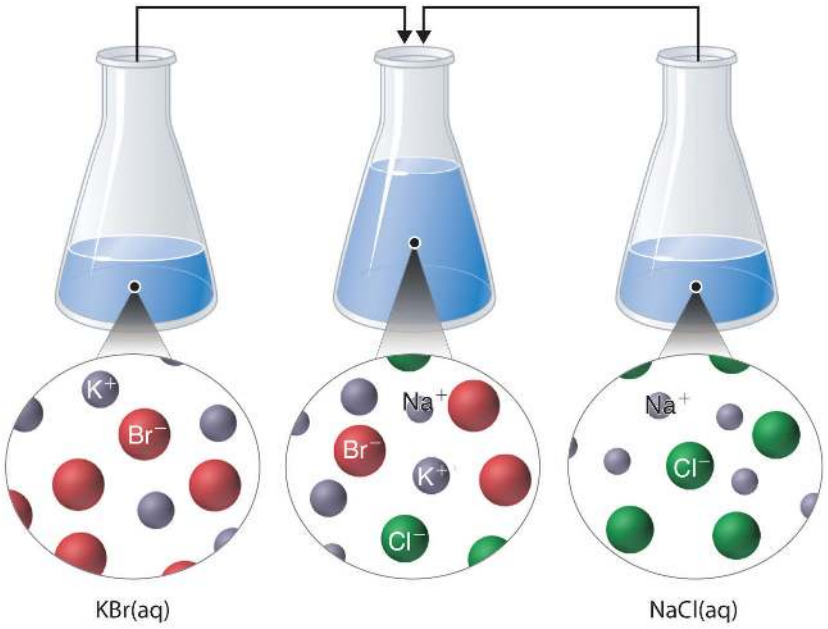


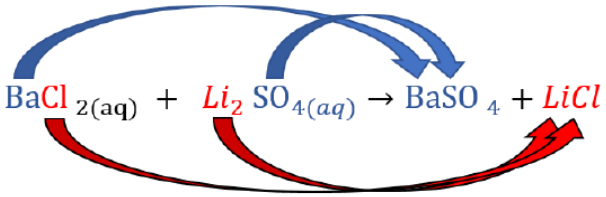
Figure 9.2.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

Change the partners of the anions and cations on the reactant side to form new compounds (products).



Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains Ba^{2+}, Cl^{-}, Li^{+}, and SO_4^{2-} ions. The only possible exchange reaction is to form LiCl and $BaSO_4$.

Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4 + LiCl \tag{9.2.3}$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + LiCl(aq) \tag{9.2.4}$$

Table 9.1.1 shows that LiCl is soluble in water, but $BaSO_4$ is not soluble in water.

Balance the equation

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2LiCl(aq) \tag{9.2.5}$$

Although soluble barium salts are toxic, $BaSO_4$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a

"barium milkshake" or a "barium enema"—a suspension of very fine $BaSO_4$ particles in water.

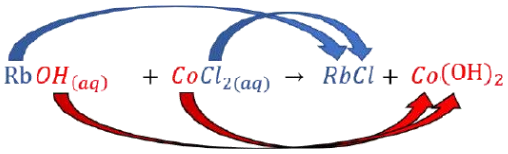


An x-ray of the digestive organs of a patient who has swallowed a "barium milkshake." A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. from Wikipedia.

Example 9.2.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

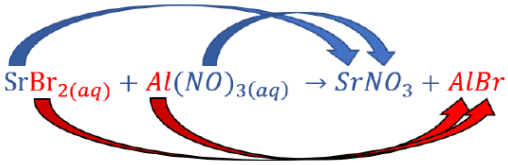
Solution

Change the partners of the anions and cations on the reactant side to form new compounds (products).	
Correct the formulas of the products based on the charges of the ions.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2(aq) + Co(OH)_2(s)$
Balance the equation	Coefficients already balanced $Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$

Example 9.2.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution

Change the partners of the anions and cations on the reactant side to form new compounds (products).	
Correct the formulas of the products based on the charges of the ions.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2 + AlBr_3$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2(aq) + AlBr_3(aq)$ According to Table 7.5.1 , both $AlBr_3$ (rule 4) and $Sr(NO_3)_2$ (rule 2) are soluble.
If all possible products are soluble then there is no net reaction will occur	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow$ **NO REACTION**

Exercise 9.2.1

Using the information in Table 9.1.1, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a:

$Fe(OH)_2$ precipitate is formed.

Answer b:

$Hg_3(PO_4)_2$ precipitate is formed.

Answer c:

No Reaction

Answer d:

$CaCO_3$ is precipitate formed.

Precipitation reactions are a subclass of double displacement reactions.

Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

9.2: Precipitation Reactions is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

https://chem.libretexts.org/@go/page/105732

9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

Template:HideTOC

The independent behavior of each type of ion in solution was illustrated in Chemical Bonding by means of precipitation reactions. **Precipitation** is a process in which a solute separates from a supersaturated solution. In a chemical laboratory it usually refers to a solid crystallizing from a liquid solution, but in weather reports it applies to liquid or solid water separating from supersaturated air.

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations.

$$\underbrace{BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,NaCl(aq)}_{\text{Complete Chemical Equation}} \tag{9.3.1}$$

However, Equation 9.3.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Thus we might write below the complete ionic equation

$$\underbrace{Ba^{2+}(aq) + \overbrace{2Cl^-(aq)}^{\text{spectator}} + \overbrace{2Na^+(aq)}^{\text{spectator}} + SO_4{}^{2-}(aq) \longrightarrow BaSO_4(s) + \overbrace{2Na^+(aq)}^{\text{spectator}} + \overbrace{Cl^-(aq)}^{\text{spectator}}}_{\text{Complete Ionic Equation}} \tag{9.3.2}$$

Equation 9.3.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $BaSO_4(s)$ will form when *any* solution containing $Ba^{2+}(aq)$ is mixed with *any* solution containing $SO_4{}^{2-}(aq)$ (provided concentrations are not extremely small). This happens independently of the $Cl^-(aq)$ and $Na^+(aq)$ ions in Equation 9.3.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $BaSO_4$ the net ionic equation is

$$\underbrace{Ba^{2+}(aq) + SO_4{}^{2-}(aq) \longrightarrow BaSO_4(s)}_{\text{Net Ionic Equation}} \tag{9.3.3}$$

Example 9.3.1

a. When a solution of $AgNO_3$ is added to a solution of $CaCl_2$, insoluble $AgCl$ precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.

b. Write balanced net ionic equation to describe any reaction which occur when the solutions of Na_2SO_4 and NH_4I are mixed.

Solution

	Example 9.3.1*a*	Example 9.3.1*b*
Complete Chemical Equation	$2\,AgNO_3(aq) + CaCl_2(aq) \longrightarrow 2\,AgCl(s) + Ca(NO_3)_2(aq)$ The proper states and formulas of all products are written and the chemical equation is balanced.	$Na_2SO_4(aq) + NH_4I_2(aq) \longrightarrow 2\,NaI(aq) + (NI$ Both products are aqueous so there is no net ionic equation that can be written.
Complete Ionic Equation	$2\,Ag^+(aq) + 2\,NO_3{}^-(aq) + Ca^{2+}(aq) + Cl^-(aq) \longrightarrow 2\,AgCl(s) + Ca^{2+}(aq) + 2\,NO_3{}^-(aq)$ AgCl is a solid so it does not break up into ions in solution.	
Net Ionic Equation	$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$ All spectator ions are removed.	NaI and $(NH_4)_2SO_4$ are both soluble. **No net ionic equation**

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $BaCl_2$ solution, for instance, is often used as a test for the presence of $SO_4{}^{2-}(aq)$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $BaSO_4$. Thus, if $BaCl_2$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the $SO_4{}^{2-}$ ion.



Figure 9.3.1: The three common silver halide precipitates: AgI, $AgBr$ and $AgCl$ (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$AgNO_3$ solutions are often used in a similar way to test for halide ions. If $AgNO_3$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of Cl^- ions, a cream-colored precipitate indicates the presence of Br^- ions, and a yellow precipitate indicates the presence of I^- ions (Figure 9.3.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $AgNO_3$ is added to tap water, a white precipitate is almost always formed. The Cl^- ions in tap water usually come from the Cl_2 which is added to municipal water supplies to kill microorganisms.

Exercise 9.3.1

Write balanced net ionic equations to describe any reaction which occur when the following solutions below are mixed.

a. $K_2CO_3 + SrCl_2$
b. $FeSO_4 + Ba(NO_3)_2$

Answer a:

$$Sr^{2+}(aq) + CO_3{}^{2-}(aq) \longrightarrow SrCO_3(s)$$

Answer b:

$$Ba^{2+}(aq) + SO_4{}^{2-}(aq) \longrightarrow Ba(SO_4)(s)$$

Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then the rules of stoichiometry may be applied.

Contributors and Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

Chapter 3: Introduction to Energy

Your page has been created!

Remove this content and add your own.

- Edit page

  Click the Edit page button in your user bar. You will see a suggested structure for your content. Add your content and hit Save.

Tips:

- Drag and drop

  Drag one or more image files from your computer and drop them onto your browser window to add them to your page.

- Classifications

  Tags are used to link pages to one another along common themes. Tags are also used as markers for the dynamic organization of content in the CXone Expert framework.

- Working with templates

  CXone Expert templates help guide and organize your documentation, making it flow easier and more uniformly. Edit existing templates or create your own.

- Visit for all help topics.

Chapter 3: Introduction to Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

Index

A

acid–base titration
- 13.9: Acid–Base Titration

anion
- 5.7: Ions - Losing and Gaining Electrons

atomic mass unit
- 5.4: The Properties of Protons, Neutrons, and Electrons

Aufbau principle
- 14.2: Electron Configurations

B

Balancing a Chemical Equation
- 8.3: How to Write Balanced Chemical Equations

Bends
- 13.4: Solutions of Gases in Water

buoyancy
- 2.4: Density and its Applications

C

cation
- 5.7: Ions - Losing and Gaining Electrons

chemical change
- 1.7: Physical and Chemical Changes
- 1.8: Classification of Matter
- 2.7: Changes in Matter: Physical and Chemical Changes
- 1.7: Classification of Matter

chemical property
- 2.6: Differences in Matter: Physical and Chemical Properties

Coefficients and Subscripts
- 8.3: How to Write Balanced Chemical Equations

combination reaction
- 11.4: Oxidation States & Redox Reactions

complete ionic equation
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

compound
- 1.8: Classification of Matter
- 2.5: Classifying Matter According to Its Composition
- 1.7: Classification of Matter

conservation of energy
- 3.3: Energy and Chemical and Physical Change

conservation of mass
- 3.1: Conservation of Mass - There is No New Matter

crystallization
- 1.8: Classification of Matter
- 1.7: Classification of Matter

D

decomposition reaction
- 11.4: Oxidation States & Redox Reactions

density
- 2.4: Density and its Applications

disproportionation reaction
- 11.4: Oxidation States & Redox Reactions

distillation
- 1.8: Classification of Matter
- 1.7: Classification of Matter

E

electron configuration
- 14.2: Electron Configurations

element
- 1.8: Classification of Matter
- 2.5: Classifying Matter According to Its Composition
- 1.7: Classification of Matter

empirical formula
- 7.9: Calculating Molecular Formulas for Compounds

empirical formula mass
- 7.9: Calculating Molecular Formulas for Compounds

endothermic process
- 3.3: Energy and Chemical and Physical Change

exothermic process
- 3.3: Energy and Chemical and Physical Change

G

gas
- 1.8: Classification of Matter
- 1.7: Classification of Matter

H

heat capacity
- 3.5: Temperature Changes - Heat Capacity

Heat Capacity Calculations
- 4.10: Energy and Heat Capacity Calculations

heterogeneous mixture
- 1.8: Classification of Matter
- 1.7: Classification of Matter

homogeneous mixture
- 1.8: Classification of Matter
- 1.7: Classification of Matter

hund's rule
- 14.2: Electron Configurations

hydrometer
- 2.4: Density and its Applications

hypothesis
- 1.3: Hypothesis, Theories, and Laws

I

ions
- 5.7: Ions - Losing and Gaining Electrons

isotopes
- 5.8: Isotopes - When the Number of Neutrons Varies

L

law
- 1.3: Hypothesis, Theories, and Laws

liquid
- 1.8: Classification of Matter
- 1.7: Classification of Matter

M

mass
- 1.8: Classification of Matter
- 1.7: Classification of Matter

matter
- 1.8: Classification of Matter
- 2.2: What is Matter?
- 1.7: Classification of Matter

mixture
- 2.5: Classifying Matter According to Its Composition

MO
- 14.4: Electron Configurations of Ions

molarity
- 13.6: Solution Concentration: Molarity

molecular formula
- 7.9: Calculating Molecular Formulas for Compounds

N

net ionic equation
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

O

OpenChem
- 14.4: Electron Configurations of Ions

oxidation number
- 11.4: Oxidation States & Redox Reactions

oxidizing agent
- 11.4: Oxidation States & Redox Reactions

P

pauli exclusion principle
- 14.2: Electron Configurations

physical change
- 1.7: Physical and Chemical Changes
- 1.8: Classification of Matter
- 2.7: Changes in Matter: Physical and Chemical Changes
- 1.7: Classification of Matter

physical property
- 2.6: Differences in Matter: Physical and Chemical Properties

potential energy
- 3.2: Energy

pressure
- 1.8: Classification of Matter
- 1.7: Classification of Matter

R

redox reaction
- 11.4: Oxidation States & Redox Reactions
- 11.5: Half-Reactions

reducing agent
- 11.4: Oxidation States & Redox Reactions

relative abundances
- 5.9: Atomic Mass: The Average Mass of an Element's Atoms

S

Separation of Mixtures
- 2.7: Changes in Matter: Physical and Chemical Changes

significant figures
- 4.4: Significant Figures in Calculations

single replacement reaction
- 11.4: Oxidation States & Redox Reactions

solid
- 1.8: Classification of Matter
- 1.7: Classification of Matter

Solubility of gases
- 13.4: Solutions of Gases in Water

specific gravity
- 2.4: Density and its Applications

specific heat capacity
- 3.5: Temperature Changes - Heat Capacity

specific volume
- 2.4: Density and its Applications

spectator ions
- 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations

Stability of Isotopes
- 5.8: Isotopes - When the Number of Neutrons Varies

Stock system
- 6.7: Naming Ionic Compounds

T

temperature
- 3.4: Temperature: Random Motion of Molecules and Atoms

theory
- 1.3: Hypothesis, Theories, and Laws

U

units of energy
- 3.2: Energy

V

valence electrons
- 14.2: Electron Configurations

W

weight
- 1.8: Classification of Matter
- 1.7: Classification of Matter

work
- 3.2: Energy

Glossary

Sample Word 1 | Sample Definition 1

Detailed Licensing

Overview

Title: PC: CHEM100 - Fundamentals of Chemistry

Webpages: 150

Applicable Restrictions: Noncommercial

All licenses found:

- CC BY-NC-SA 4.0: 90% (135 pages)
- Undeclared: 10% (15 pages)

By Page

- PC: CHEM100 - Fundamentals of Chemistry - *CC BY-NC-SA 4.0*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1.7: Classification of Matter - *CC BY-NC-SA 4.0*
  - 1: Introduction to Chemistry - *CC BY-NC-SA 4.0*
    - Front Matter - *Undeclared*
      - TitlePage - *Undeclared*
      - InfoPage - *Undeclared*
    - 1.1: The Scope of Chemistry - *CC BY-NC-SA 4.0*
    - 1.2: Chemicals Compose Ordinary Things - *CC BY-NC-SA 4.0*
    - 1.3: Hypothesis, Theories, and Laws - *CC BY-NC-SA 4.0*
    - 1.4: The Scientific Method - How Chemists Think - *CC BY-NC-SA 4.0*
    - 1.5: A Beginning Chemist - How to Succeed - *CC BY-NC-SA 4.0*
    - 1.6: Solid, Liquid, and Gas - *CC BY-NC-SA 4.0*
    - 1.7: Physical and Chemical Changes - *CC BY-NC-SA 4.0*
    - 1.8: Classification of Matter - *CC BY-NC-SA 4.0*
    - Back Matter - *Undeclared*
      - Index - *Undeclared*
  - 2.4: Solid, Liquid, and Gas - *CC BY-NC-SA 4.0*
  - 4: Measurements, Conversion Factors, and Calculations - *CC BY-NC-SA 4.0*
    - 4.1: Taking Measurements - *CC BY-NC-SA 4.0*
    - 4.2: Scientific Notation: Writing Large and Small Numbers - *CC BY-NC-SA 4.0*
    - 4.3: Significant Figures: Writing Numbers to Reflect Precision - *CC BY-NC-SA 4.0*
    - 4.4: Significant Figures in Calculations - *CC BY-NC-SA 4.0*
    - 4.5: The Basic Units of Measurement - *CC BY-NC-SA 4.0*
    - 4.6: Problem Solving and Unit Conversions - *CC BY-NC-SA 4.0*
    - 4.7: Solving Multistep Conversion Problems - *CC BY-NC-SA 4.0*
    - 4.8: Units Raised to a Power - *CC BY-NC-SA 4.0*
    - 4.9: Density - *CC BY-NC-SA 4.0*
    - 4.10: Energy and Heat Capacity Calculations - *CC BY-NC-SA 4.0*
  - 3: Introduction to Energy - *CC BY-NC-SA 4.0*
    - 3.1: Conservation of Mass - There is No New Matter - *CC BY-NC-SA 4.0*
    - 3.2: Energy - *CC BY-NC-SA 4.0*
    - 3.3: Energy and Chemical and Physical Change - *CC BY-NC-SA 4.0*
    - 3.4: Temperature: Random Motion of Molecules and Atoms - *CC BY-NC-SA 4.0*
    - 3.5: Temperature Changes - Heat Capacity - *CC BY-NC-SA 4.0*
  - 2: Matter - *CC BY-NC-SA 4.0*
    - 2.4: Density and its Applications - *CC BY-NC-SA 4.0*
    - 2.1: In Your Room - *CC BY-NC-SA 4.0*
    - 2.2: What is Matter? - *CC BY-NC-SA 4.0*
    - 2.3: Classifying Matter According to Its State: Solid, Liquid, and Gas - *CC BY-NC-SA 4.0*
    - 2.5: Classifying Matter According to Its Composition - *CC BY-NC-SA 4.0*
    - 2.6: Differences in Matter: Physical and Chemical Properties - *CC BY-NC-SA 4.0*

- 2.7: Changes in Matter: Physical and Chemical Changes - *CC BY-NC-SA 4.0*

- 5: Atoms and Elements - *CC BY-NC-SA 4.0*
  - 5.1: Cutting Alumimun until you get Atoms - *CC BY-NC-SA 4.0*
  - 5.2: Indivisible: The Atomic Theory - *CC BY-NC-SA 4.0*
  - 5.3: The Nuclear Atom - *CC BY-NC-SA 4.0*
  - 5.4: The Properties of Protons, Neutrons, and Electrons - *CC BY-NC-SA 4.0*
  - 5.5: Elements: Defined by Their Number of Protons - *CC BY-NC-SA 4.0*
  - 5.6: Looking for Patterns: The Periodic Table - *CC BY-NC-SA 4.0*
  - 5.7: Ions - Losing and Gaining Electrons - *CC BY-NC-SA 4.0*
  - 5.8: Isotopes - When the Number of Neutrons Varies - *CC BY-NC-SA 4.0*
  - 5.9: Atomic Mass: The Average Mass of an Element's Atoms - *CC BY-NC-SA 4.0*
- 6: Molecules and Compounds - *CC BY-NC-SA 4.0*
  - 6.1: Sugar and Salt - *CC BY-NC-SA 4.0*
  - 6.2: Compounds Display Constant Composition - *CC BY-NC-SA 4.0*
  - 6.3: Chemical Formulas: How to Represent Compounds - *CC BY-NC-SA 4.0*
  - 6.4: A Molecular View of Elements and Compounds - *CC BY-NC-SA 4.0*
  - 6.5: Writing Formulas for Ionic Compounds - *CC BY-NC-SA 4.0*
  - 6.6: Nomenclature: Naming Compounds - *CC BY-NC-SA 4.0*
  - 6.7: Naming Ionic Compounds - *CC BY-NC-SA 4.0*
  - 6.8: Naming Molecular Compounds - *CC BY-NC-SA 4.0*
  - 6.9: Naming Acids - *CC BY-NC-SA 4.0*
  - 6.10: Nomenclature Summary - *CC BY-NC-SA 4.0*
  - 6.11: Formula Mass: The Mass of a Molecule or Formula Unit - *CC BY-NC-SA 4.0*
- 7: Quantitative Composition of Compounds - *CC BY-NC-SA 4.0*
  - 7.1: How Much Sodium? - *Undeclared*
  - 7.2: Counting Nails by the Pound - *CC BY-NC-SA 4.0*
  - 7.3: Counting Atoms by the Gram - *CC BY-NC-SA 4.0*
  - 7.4: Counting Molecules by the Gram - *CC BY-NC-SA 4.0*
  - 7.5: Chemical Formulas as Conversion Factors - *CC BY-NC-SA 4.0*
  - 7.6: Mass Percent Composition of Compounds - *CC BY-NC-SA 4.0*
  - 7.7: Mass Percent Composition from a Chemical Formula - *CC BY-NC-SA 4.0*
  - 7.8: Calculating Empirical Formulas for Compounds - *CC BY-NC-SA 4.0*
  - 7.9: Calculating Molecular Formulas for Compounds - *CC BY-NC-SA 4.0*
- 8: Chemical Reactions - *CC BY-NC-SA 4.0*
  - 8.1: Evidence of a Chemical Reaction - *CC BY-NC-SA 4.0*
  - 8.2: Chemical Equations - *CC BY-NC-SA 4.0*
  - 8.3: How to Write Balanced Chemical Equations - *CC BY-NC-SA 4.0*
  - 8.4: Classifying Chemical Reactions - *CC BY-NC-SA 4.0*
- 12: Quantities in Chemical Reactions - *CC BY-NC-SA 4.0*
  - 12.1: Climate Change - Too Much Carbon Dioxide - *CC BY-NC-SA 4.0*
  - 12.2: Stoichiometry - *CC BY-NC-SA 4.0*
  - 12.3: Mole-to-Mole Conversions - *CC BY-NC-SA 4.0*
  - 12.4: Making Molecules: Mole to Mass (or vice versa) and Mass-to-Mass Conversions - *CC BY-NC-SA 4.0*
  - 12.5: Limiting Reactant and Theoretical Yield - *CC BY-NC-SA 4.0*
  - 12.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants - *CC BY-NC-SA 4.0*
  - 12.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed - *CC BY-NC-SA 4.0*
- 9: Precipitation Reactions - *CC BY-NC-SA 4.0*
  - 9.1: Aqueous Solutions and Solubility: Compounds Dissolved in Water - *CC BY-NC-SA 4.0*
  - 9.2: Precipitation Reactions - *CC BY-NC-SA 4.0*
  - 9.3: Writing Chemical Equations for Reactions in Solution: Molecular, Complete Ionic, and Net Ionic Equations - *CC BY-NC-SA 4.0*
- 14: Electrons in Atoms and the Periodic Table - *CC BY-NC-SA 4.0*
  - 14.1: The Perioidic Table - *CC BY-NC-SA 4.0*
  - 14.2: Electron Configurations - *CC BY-NC-SA 4.0*
  - 14.3: Electron Configurations and the Periodic Table (The Block Method) - *CC BY-NC-SA 4.0*
  - 14.4: Electron Configurations of Ions - *CC BY-NC-SA 4.0*
  - 14.5: Periodic Trends: Atomic Size, Ionization Energy, and Metallic Character - *CC BY-NC-SA 4.0*

- 10: Acid-Base and Gas Evolution Reactions - *CC BY-NC-SA 4.0*
  - 10.1: Acids: Properties and Examples - *CC BY-NC-SA 4.0*
  - 10.2: Bases: Properties and Examples - *CC BY-NC-SA 4.0*
  - 10.3: Molecular Definitions of Acids and Bases - *CC BY-NC-SA 4.0*
  - 10.4: Reactions of Acids and Bases - *CC BY-NC-SA 4.0*
  - 10.5: Acid–Base and Gas Evolution Reactions - *CC BY-NC-SA 4.0*
- 11: Redox Reactions - *CC BY-NC-SA 4.0*
  - 11.1: Oxidation and Reduction - Some Definitions - *CC BY-NC-SA 4.0*
  - 11.2: Electronegativity - *CC BY-NC-SA 4.0*
  - 11.3: Oxidation States: Electron Bookkeeping - *CC BY-NC-SA 4.0*
  - 11.4: Oxidation States & Redox Reactions - *CC BY-NC-SA 4.0*
  - 11.5: Half-Reactions - *CC BY-NC-SA 4.0*
  - 11.6: The Activity Series: Predicting Spontaneous Redox Reactions - *CC BY-NC-SA 4.0*
  - 11.7: Combustion Reactions - *CC BY-NC-SA 4.0*
- 13: Solutions - *CC BY-NC-SA 4.0*
  - 13.1: Prelude - Tragedy in Cameroon - *CC BY-NC-SA 4.0*
  - 13.2: Solutions: Homogeneous Mixtures - *CC BY-NC-SA 4.0*
  - 13.3: Solutions of Solids Dissolved in Water: How to Make Rock Candy - *CC BY-NC-SA 4.0*
  - 13.4: Solutions of Gases in Water - *CC BY-NC-SA 4.0*
  - 13.5: Solution Concentration: Mass Percent - *CC BY-NC-SA 4.0*
  - 13.6: Solution Concentration: Molarity - *CC BY-NC-SA 4.0*
  - 13.7: Solution Dilution - *CC BY-NC-SA 4.0*
  - 13.8: Solution Stoichiometry - *CC BY-NC-SA 4.0*
  - 13.9: Acid–Base Titration - *CC BY-NC-SA 4.0*
  - 13.10: Water - Acid and Base in One - *CC BY-NC-SA 4.0*
  - 13.11: The pH and pOH Scales - Ways to Express Acidity and Basicity - *CC BY-NC-SA 4.0*
- 15: Chemical Bonding - *CC BY-NC-SA 4.0*
  - 15.1: Representing Valence Electrons with Dots - *CC BY-NC-SA 4.0*
  - 15.2: Electrons are Transferred in Ionic Compounds - *CC BY-NC-SA 4.0*
  - 15.3: Lewis Structures: Electrons Shared - *CC BY-NC-SA 4.0*
  - 15.4: Lewis Structures: Counting Valence Electrons - *CC BY-NC-SA 4.0*
  - 15.5: Predicting the Shapes of Molecules - *CC BY-NC-SA 4.0*
  - 15.6: Electronegativity and Polarity - Why Oil and Water Don’t Mix - *CC BY-NC-SA 4.0*
- 16: Gases - *CC BY-NC-SA 4.0*
  - 16.1: Extra-Long Straws - *CC BY-NC-SA 4.0*
  - 11.2: Kinetic Molecular Theory: Real Gases are Not Ideal - *CC BY-NC-SA 4.0*
  - 16.3: Pressure: The Result of Constant Molecular Collisions - *CC BY-NC-SA 4.0*
  - 16.4: Boyle’s Law: Pressure and Volume - *CC BY-NC-SA 4.0*
  - 16.5: Charles’s Law: Volume and Temperature - *CC BY-NC-SA 4.0*
  - 16.6: Gay-Lussac's Law: Temperature and Pressure - *CC BY-NC-SA 4.0*
  - 16.7: The Combined Gas Law: Pressure, Volume, and Temperature - *CC BY-NC-SA 4.0*
  - 16.8: Avogadro’s Law: Volume and Moles - *CC BY-NC-SA 4.0*
  - 16.9: The Ideal Gas Law: Pressure, Volume, Temperature, and Moles - *CC BY-NC-SA 4.0*
- Chapter 3: Introduction to Energy - *CC BY-NC-SA 4.0*
- Back Matter - *Undeclared*
  - Index - *Undeclared*
  - Glossary - *Undeclared*
  - Detailed Licensing - *Undeclared*