INTRODUCTORY CHEMISTRY



Commodore St Germain
Solano College



Solano College

Introductory_Chemistry_at_Solano_College

Commodore St Germain

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 borne 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 were 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.

The LibreTexts name, logo, and book covers are trademarked by LibreTexts, Inc. (a 501(c)(3) not-for-profit organization) and protected against unauthorized use.

This text was compiled on 10/13/2026

TABLE OF CONTENTS

Licensing

1: Introduction To Chemistry

1.1: Scope of Chemistry

1.2: Areas of Chemistry

1.3: Research

1.4: Scientific Problem Solving

2: Numbers From Measurements

2.1: Accuracy and Precision

2.2: Percent Error

2.3: Measurement Uncertainty

2.4: Significant Figures

2.5: Rounding

2.6: Significant Figures in Addition and Subtraction

2.7: Significant Figures in Multiplication and Division

2.8: Scientific Notation

3: Unit Systems and Dimensional Analysis

3.1: SI Base Units

3.2: Metric Prefixes

3.3: Length and Volume

3.4: Mass and Weight

3.5: Metric Unit Conversions

3.6: Derived Units

3.7: Dimensional Analysis

3.8: Density

3.9: Kinetic Energy

3.10: Temperature and Temperature Scales

4: Matter and Change

4.1: Matter, Mass, and Volume

4.2: States of Matter

4.3: Physical Properties

4.4: Chemical Properties and Chemical Reactions

4.5: Extensive and Intensive Properties

4.6: Physical Change

4.7: Chemical Change

4.8: Recognizing Chemical Reactions

4.9: Pure Substances

4.10: Mixture
4.11: Heterogeneous Mixtures
4.12: Homogeneous Mixture
4.13: Separating Mixtures
4.14: Elements
4.15: Compounds
4.16: Chemical Symbols and Formulas
4.17: A Molecular View of Elements and Compounds

5: Atoms, Molecules, and Subatomic Particles

5.1: Atom
5.2: Protons
5.3: Neutrons
5.4: Electrons
5.5: Atomic Nucleus
5.6: Atomic Number
5.7: Mass Number
5.8: Isotopes
5.9: Atomic Mass Unit
5.10: Calculating Average Atomic Mass
5.11: Ion

6: Electronic Structure and Chemical Periodicity

6.1: Periodic Law
6.2: Modern Periodic Table- Periods and Groups
6.3: Blocks of the Periodic Table
6.4: Energy Level
6.5: Orbitals
6.6: Electron Configurations
6.7: Noble Gas Configuration
6.8: Hund's Rule and Orbital Filling Diagrams
6.9: Aufbau Principle
6.10: Periodic Trends- Atomic Radius
6.11: Metals
6.12: Nonmetals
6.13: Metalloids

7: Chemical Bonds

7.1: Chemical Bond
7.2: Electron Dot Diagrams
7.3: Octet Rule
7.4: Cation Formation
7.5: Anion Formation

7.6: Ionic Bonding
7.7: Polyatomic Ions
7.8: Covalent Bond
7.9: Lewis Electron-Dot Structures
7.10: Single Covalent Bonds
7.11: Multiple Covalent Bonds
7.12: Valence Electrons
7.13: Coordinate Covalent Bond
7.14: Resonance
7.15: VSEPR Theory
7.16: Molecular Shapes- No Lone Pairs on Central Atoms
7.17: Molecular Shapes - Lone Pair(s) on Central Atom
7.18: Bond Polarity
7.19: Polar Molecules

8: Chemical Nomenclature

8.1: Molecular Formula
8.2: Empirical Formula
8.3: Cations
8.4: Anions
8.5: Transition Metal Ions
8.6: The Stock System of Nomenclature
8.7: Naming Binary Ionic Compounds
8.8: Formulas for Binary Ionic Compounds
8.9: Polyatomic Ions
8.10: Ternary Ionic Compounds- Naming and Formulas
8.11: Binary Molecular Compounds- Naming and Formulas
8.12: Acids - Naming and Formulas
8.13: Bases- Naming and Formulas

9: The Mole

9.1: Law of Definite Proportions
9.2: Avogadro's Number
9.3: Conversions Between Moles and Atoms
9.4: Molar Mass
9.5: Conversions Between Moles and Mass
9.6: Conversions Between Mass and Number of Particles
9.7: Avogadro's Hypothesis and Molar Volume
9.8: Conversions Between Moles and Gas Volume
9.9: Gas Density
9.10: Mole Road Map
9.11: Percent Composition

9.12: Percent of Water in a Hydrate
9.13: Determining Empirical Formulas
9.14: Determining Molecular Formulas

10: Chemical Calculations Involving Chemical Equations

10.1: Law of Conservation of Mass
10.2: Word Equations
10.3: Chemical Equations
10.4: Balancing Equations
10.5: Combination Reactions
10.6: Decomposition Reactions
10.7: Combustion Reactions
10.8: Single Replacement Reactions
10.9: Activity Series
10.10: Double Replacement Reactions
10.11: Everyday Stoichiometry
10.12: Mole Ratios
10.13: Mass-Mole Stoichiometry
10.14: Mass-Mass Stoichiometry
10.15: Volume-Volume Stoichiometry
10.16: Mass-Volume Stoichiometry
10.17: Limiting Reactant
10.18: Determining the Limiting Reactant
10.19: Theoretical Yield and Percent Yield

11: States of Matter

11.1: Physical Properties and Intermolecular Forces
11.2: Hydrogen Bonding
11.3: Kinetic Molecular Theory
11.4: Gas Pressure
11.5: Atmospheric Pressure
11.6: Pressure Units and Conversions
11.7: Average Kinetic Energy and Temperature
11.8: Physical Properties and Intermolecular Forces
11.9: Surface Tension
11.10: Evaporation
11.11: Vapor Pressure
11.12: Boiling
11.13: Vapor Pressure Curves
11.14: Melting
11.15: Sublimation
11.16: Heating and Cooling Curves

11.17: General Phase Diagram
11.18: Phase Diagram for Water

12: Gas Laws

13: Solutions

13.1: Solute-Solvent Combinations
13.2: Rate of Dissolution
13.3: Saturated and Unsaturated Solutions
13.4: How Temperature Influences Solubility
13.5: Supersaturated Solutions
13.6: Predicting Precipitates Using Solubility Rules
13.7: Percent Solutions
13.8: Molarity
13.9: Preparing Solutions
13.10: Dilution
13.11: Molality
13.12: Calculating Molar Mass
13.13: Electrolytes and Nonelectrolytes
13.14: Strong and Weak Electrolytes
13.15: Molecular and Ionic Equations
13.16: Net Ionic Equations

14: Acids and Bases

14.1: Properties of Acids
14.2: Properties of Bases
14.3: Arrhenius Acids
14.4: Arrhenius Bases
14.5: Brønsted-Lowry Acids and Bases
14.6: Brønsted-Lowry Acid-Base Reactions
14.7: Ion-Product of Water
14.8: The pH Scale
14.9: Calculating pH of Acids and Bases
14.10: The pOH Concept
14.11: Strong and Weak Acids and Acid Ionization Constant (K_a)
14.12: Strong and Weak Bases and Base Ionization Constant
14.13: Calculating Acid and Base Dissociation Constants
14.14: Calculating pH of Weak Acid and Base Solutions
14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions
14.16: Titration Experiment
14.17: Titration Calculations
14.18: Titration Curves
14.19: Indicators

14.20: Hydrolysis of Salts - Equations
14.21: Calculating pH of Salt Solutions

Index

Glossary

Detailed Licensing

This page titled Introductory Chemistry at Solano College - Updated 2023/04/03 is shared under a license and was authored, remixed, and/or curated by LibreTexts.

Licensing

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

This page titled Licensing is shared under a license and was authored, remixed, and/or curated by LibreTexts.

1: INTRODUCTION TO CHEMISTRY

1.1: Scope of Chemistry
1.2: Areas of Chemistry
1.3: Research
1.4: Scientific Problem Solving

This page titled 1: Introduction To Chemistry is shared under a license and was authored, remixed, and/or curated by LibreTexts.

1.1: Scope of Chemistry



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

How does soap relate to chemistry?

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

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

What is Chemistry?

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

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



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

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

YouTube Video

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

View this YouTube Video online

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

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org



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

Summary

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

Review

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

Explore More

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

This page titled 1.1: Scope of Chemistry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 1.1: Scope of Chemistry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

1.2: 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
- Organic chemistry
- Inorganic chemistry
- Analytical chemistry
- Biochemistry

Physical Chemistry

Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. As examples, a physical chemist may study 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 chemicals found in every living organism are based on carbon.

Inorganic Chemistry

Inorganic chemistry is the study of chemicals that do not, in general, contain 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.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

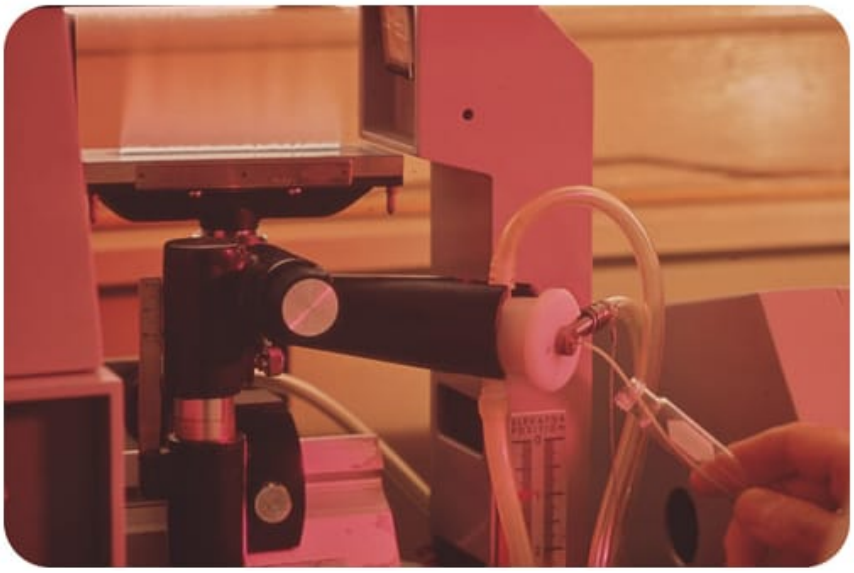


Figure 1.2.1: Measurement of trace metals using atomic spectroscopy. (Bill Shrout, Environmental Protection Agency;n Source: http://commons.wikimedia.org/wiki/File:EPA_GULF_BREEZE_LABORATORY,_THE_MICROBIOLOGY_LAB._AN_ATOMIC_ABSORPTION_SYSTEM,_WHICH_MEASURES_MINUTE_QUAN_NARA_-_546269.tif; License: Public Domain)

Biochemistry

Biochemistry is the study of chemical processes that occur in living things. Research may cover basic cellular processes up to understanding disease states, so that better treatments can be developed.



Figure 1.2.2: Measuring hormone concentrations. (Fred W. Baker III, U.S. Department of Defense; Source: http://commons.wikimedia.org/wiki/File:Defense.gov_photo_essay_090506-D-1852B-019.jpg; License: Public Domain)

In practice, chemical research is often not limited to just one of the five major disciplines. A 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.

Summary

- Five areas of chemistry are described:
  - Physical chemistry
  - Organic chemistry
  - Inorganic chemistry
  - Analytical chemistry
  - Biochemistry

Review

Match the project with the proper chemistry discipline.

Match the project with the proper chemistry discipline.

a. measuring mercury in seawater	1. biochemistry
b. studying enzymes in cells	2. organic chemistry
c. measuring the electrical properties of solutions	3. inorganic chemistry
d. synthesizing new carbon compounds	4. physical chemistry
e. making new compounds for energy processes	5. analytical chemistry

Explore More

1. Do an internet search using one of the five chemistry areas as a search term. List two significant contributions made to chemistry by that area.

This page titled 1.2: Areas of Chemistry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 1.4: Areas of Chemistry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

1.3: Research

What can be accomplished through team work?

Most jobs to day involve a fair amount of meetings and team involvement. It doesn't matter what you do, you will spend a certain amount of time meeting with other people, sharing what you have done, planning projects, and organizing the work. Hardly anybody works by him or herself these days – everybody is part of a larger group.

The Nature of Science

Today's scientists rarely work alone. Rather, most scientists collaborate with one another as part of a group effort, no matter the setting. The majority of research scientists work either for a company such as DuPont Chemical Company in Wilmington, Delaware or for one of many universities, such as the California Institute of Technology. Working as part of a group has many advantages. Most scientific problems are so complex and time-consuming, that one person could not hope to address all of the issues by himself or herself. Instead, different members of a research group are each tasked with a particular small aspect of a larger research problem. Collaboration between members of the group is frequent. This occurs informally in the laboratory on an everyday basis. Research groups typically have regular meetings where one or more members of the group may give a presentation to the others on the status of the research that they are doing. Progress normally occurs in small steps rather than grand, sweeping discoveries, and that progress is helped along by the teamwork that comes from working as part of a group.

Modern scientific research is usually expensive. Lab equipment, chemicals, research space, and the upkeep of technical instrumentation all costs money. So research groups need to raise money in order to continue their research. Much of that money comes from government sources, such as the National Science Foundation or the National Institute of Health, especially in the case of research being done at universities. Private companies can fund their own research, but may also seek outside funding as well. Scientists write grants explaining the goals of their research along with projected costs, and funding agencies make decisions on which research projects they would like to fund. The long-term viability of most research labs depends on the ability to get and maintain funding.

YouTube Video

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

View this YouTube Video online

Communicating Results

Suppose that your research is a success. What now? Scientists communicate their results to one another and to the public at large in several ways. One is to publish their research findings in one of many publications called scientific **journals**. There are many hundreds of scientific journals covering every field of science imaginable. In chemistry, there is the "Journal of the American Chemical Society," the "Journal of Physical Chemistry," and the "Inorganic Chemistry," to name just a few. Some journals have a very narrow scope while others publish articles from many different sciences and appeal to a wider audience. Examples of the latter include the journals "Science" and "Nature." Journal articles are often very complex and detailed. They must be accurate, since the research field as a whole uses these journal articles as a way to make scientific progress. Therefore, journal articles are only published after having been extensively reviewed by other professional scientists in the same field. Reviewers have the power to make suggestions about the research or possibly question the validity of the author's conclusions. Only when the reviewers are satisfied that the research is correct, will the journal publish the article. In this way, all scientists can trust the research findings that they read about in journals.

Scientists also communicate with one another by presenting their findings at international conferences. Some scientists are

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

chosen to give a lecture at a conference, typically about research that has already been published. Many other scientists at the same conference will present their work at **poster sessions**. These poster sessions are more informal and may often represent research that is still in progress.



Figure 1.3.1: Scientific presentation. (Public Domain; Journalist 1st Class Dave Kaylor, U.S. Navy via Wikipedia)

YouTube Video

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

View this YouTube Video online

Science Friday

A Shot in the Dark: Alternative Uses for Squid Ink

Squid ink is commonly thought of as a defensive mechanism used by squids, but could there be other uses? In this video by Science Friday, Dr. Stephanie Bush talks about possible other functions of squid ink.

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

View this YouTube Video online

Snowflake Safari

Next snowstorm, grab a magnifying glass and look carefully at snowflakes. Bullet rosettes, stellar plates and capped columns are just a few of the many varieties of snow crystals. In this video by Science Friday, physicist Kenneth Libbrecht shares secrets about snowflakes.

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

View this YouTube Video online

Summary

- Scientific research today is a team effort.
- Support for research usually comes from grants.
- Findings are shared in meetings, conferences, and scientific publications.

Review

1. How is most scientific research carried out today?
2. How is research usually supported?
3. How do scientists share their findings?

This page titled 1.3: Research is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

1.4: Scientific Problem Solving

How can we use problem solving in our everyday routines?

One day you wake up and realize your clock radio did not turn on to get you out of bed. You are puzzled, so you decide to find out what happened. You list three possible explanations:

1. There was a power failure and your radio cannot turn on.
2. Your little sister turned it off as a joke.
3. You did not set the alarm last night.

Upon investigation, you find that the clock is on, so there is no power failure. Your little sister was spending the night with a friend and could not have turned the alarm off. You notice that the alarm is not set—your forgetfulness made you late. You have used the scientific method to answer a question.

Scientific Problem Solving

Humans have always wondered about the world around them. One of the questions of interest was (and still is): *what is this world made of?* Chemistry has been defined in various ways as the study of matter. What matter consists of has been a source of debate over the centuries. One of the key areas for this debate in the Western world was Greek philosophy.

The basic approach of the Greek philosophers was to discuss and debate the questions they had about the world. There was no gathering of information to speak of, just talking. As a result, several ideas about matter were put forth, but never resolved. The first philosopher to carry out the gathering of data was Aristotle (384-322 B.C.). He recorded many observations on the weather, on plant and animal life and behavior, on physical motions, and a number of other topics. Aristotle could probably be considered the first "real" scientist, because he made systematic observations of nature and tried to understand what he was seeing.



Figure 1.4.1: Aristotle. (Public Domain; Raphael via Wikipedia)

Inductive and Deductive Reasoning

Two approaches to logical thinking developed over the centuries. These two methods are **inductive reasoning** and **deductive reasoning**. Inductive reasoning involves getting a collection of specific examples and drawing a general conclusion from them. Deductive reasoning takes a general principle and then draws a specific conclusion from the general concept. Both are used in the development of scientific ideas.

Inductive reasoning first involves the collection of data: "If I add sodium metal to water, I observe a very violent reaction. Every time I repeat the process, I see the same thing happen." A general conclusion is drawn from these observations: the addition of sodium to water results in a violent reaction.

In deductive reasoning, a specific prediction is made based on a general principle. One general principle is that acids turn blue

litmus paper red. Using the deductive reasoning process, one might predict: "If I have a bottle of liquid labeled 'acid', I expect the litmus paper to turn red when I immerse it in the liquid."

YouTube Video

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

View this YouTube Video online

The Idea of the Experiment

Inductive reasoning is at the heart of what is now called the "**scientific method**." In European culture, this approach was developed mainly by Francis Bacon (1561-1626), a British scholar. He advocated the use of inductive reasoning in every area of life, not just science. The scientific method, as developed by Bacon and others, involves several steps:

1. Ask a question - identify the problem to be considered.
2. Make observations - gather data that pertains to the question.
3. Propose an explanation (a hypothesis) for the observations.
4. Make new observations to test the hypothesis further.



Figure 1.4.2: Sir Francis Bacon. (Public Domain; Paul van Somer via Wikipedia)

Note that this should not be considered a "cookbook" for scientific research. Scientists do not sit down with their daily "to do" list and write down these steps. The steps may not necessarily be followed in order. But this does provide a general idea of how scientific research is usually done.

When a hypothesis is confirmed repeatedly, it eventually becomes a theory—a general principle that is offered to explain natural phenomena. Note a key word—*explain*, or *explanation*. A theory offers a description of why something happens. A law, on the other hand, is a statement that is always true, but offers no explanation as to why. The law of gravity says a rock will fall when dropped, but does not explain why (gravitational theory is very complex and incomplete at present). The kinetic molecular theory of gases, on the other hand, states what happens when a gas is heated in a closed container (the pressure increases), but also explains why (the motions of the gas molecules are increased due to the change in temperature). Theories do not get "promoted" to laws, because laws do not answer the "why" question.

YouTube Video

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

View this YouTube Video online

Summary

- The early Greek philosophers spent their time talking about nature, but did little or no actual exploration or investigation.
- Inductive reasoning - to develop a general conclusion from a collection of observations.
- Deductive reasoning - to make a specific statement based on a general principle.
- Scientific method - a process of observation, developing a hypothesis, and testing that hypothesis.

Review

1. What was the basic shortcoming of the Greek philosophers approach to studying the material world?
2. How did Aristotle improve the approach?
3. Define “inductive reasoning” and give an example.
4. Define “deductive reasoning” and give an example.
5. What is the difference between a hypothesis and a theory?
6. What is the difference between a theory and a law?

This page titled 1.4: Scientific Problem Solving is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 1.12: Scientific Problem Solving by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2: NUMBERS FROM MEASUREMENTS

2.1: Accuracy and Precision
2.2: Percent Error
2.3: Measurement Uncertainty
2.4: Significant Figures
2.5: Rounding
2.6: Significant Figures in Addition and Subtraction
2.7: Significant Figures in Multiplication and Division
2.8: Scientific Notation

This page titled 2: Numbers From Measurements is shared under a license and was authored, remixed, and/or curated by LibreTexts.

2.1: Accuracy and Precision



Figure 2.1.1 (Public Domain; Daniel Arizpe via Commons Wikimedia, High School Basketball Game [commons.wikimedia.org])

How do professional basketball players improve their shooting accuracy?

Basketball is one of those sports where you need to hit the target. A football field goal kicker might have room for some deviation from a straight line – for college and pro football there is an 18 foot 6 inch space for the ball to go through. In basketball, the basket is only 18 inches across and the ball is a little less than 10 inches across – not much room for error. The ball has to be on target in order to go into the basket and score.

Accuracy and Precision

In everyday speech, the terms **accuracy** and **precision** are frequently used interchangeably. However, their scientific meanings are quite different. Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured. Precision is a measure of how close a series of measurements are to one another. Precise measurements are highly reproducible, even if the measurements are not near the correct value. Darts thrown at a dartboard are helpful in illustrating accuracy and precision.

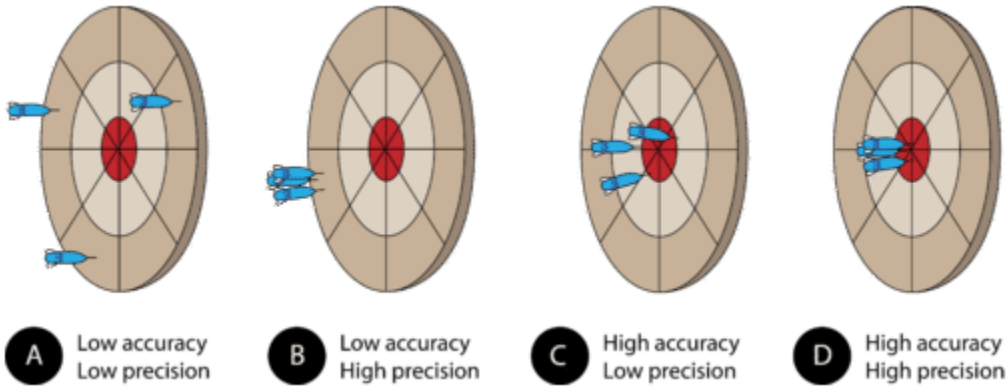


Figure 2.1.2: The distribution of darts on a dartboard shows the difference between accuracy and precision. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being

measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. The figure above demonstrates four possible outcomes:

a. The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate nor precise.
b. The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly, or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.
c. The darts are not grouped very near to each other, but are generally centered around the bulls-eye. This demonstrates poor precision, but fairly high accuracy. This situation is not desirable in a lab situation because the "high" accuracy may simply be random chance and not a true indicator of good measuring skill.
d. The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements.



Figure 2.1.3: Students in a chemistry lab are making careful measurements with a series of volumetric flasks. Accuracy and precision are critical in every experiment. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Interactive Content

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

View this Interactive Content online

Interactive Content

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

View this Interactive Content online

Simulation

How is your accuracy and precision? Try out your skills with this simulation.

Summary

- Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured.
- Precision is a measure of how close a series of measurements are to one another.

Review

1. Define accuracy.
2. Define precision.
3. What can be said about the reproducibility of precise values?

This page titled 2.1: Accuracy and Precision is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.12: Accuracy and Precision by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.2: Percent Error

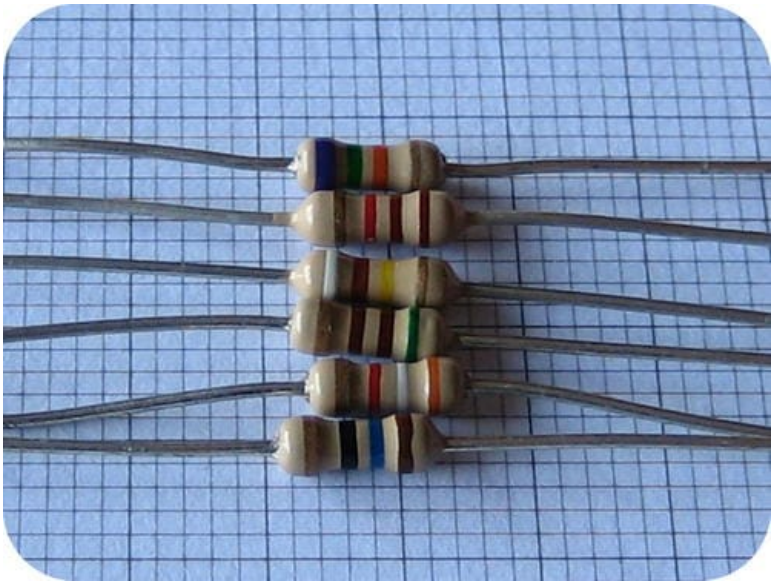


Figure 2.2.1 (Public Domain; Vald Kliper via Wikipedia)

How does an electrical circuit work?

A complicated piece of electronic equipment may contain several resistors that have the role of controlling the voltage and current in the electrical circuit. If there is too much current, the apparatus malfunctions. Conversely, too little current means the system simply will not perform. The resistor values are always given with an error range. A resistor may have a stated value of 200 ohms, but a 10% error range, meaning the resistance could be anywhere between 195 and 205 ohms. By knowing these values, an electronics person can design and service the equipment to make sure that it functions properly.

Percent Error

An individual measurement may be accurate or inaccurate, depending on how close it is to the true value. Suppose that you are doing an experiment to determine the density of a sample of aluminum metal. The **accepted value** of a measurement is the true or correct value based on general agreement with a reliable reference. For aluminum, the accepted density is $2.70\ \text{g/cm}^3$. The **experimental value** of a measurement is the value that is measured during the experiment. Suppose that in your experiment, you determine an experimental value for the aluminum density to be $2.42\ \text{g/cm}^3$. The **error** of an experiment is the difference between the experimental and accepted values.

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

If the experimental value is less than the accepted value, the error is negative. If the experimental value is larger than the accepted value, the error is positive. Often, error is reported as the absolute value of the difference in order to avoid the confusion of a negative error. The **percent error** is the absolute value of the error, divided by the accepted value, and multiplied by 100%.

$$\%\ \text{Error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

To calculate the percent error for the aluminum density measurement, we can substitute the given values of $2.45\ \text{g/cm}^3$ for the experimental value and $2.70\ \text{g/cm}^3$ for the accepted value.

$$\%\ \text{Error} = \frac{\left|2.45\ \text{g/cm}^3 - 2.70\ \text{g/cm}^3\right|}{2.70\ \text{g/cm}^3} \times 100\% = 9.26\%$$

If the experimental value is equal to the accepted value, the percent error is equal to 0. As the accuracy of a measurement decreases, the percent error of that measurement rises.

YouTube Video

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

View this YouTube Video online

Summary

- Definitions of accepted value and experimental values are given.
- Calculations of error and percent error are demonstrated.

Review

1. Define accepted value.
2. Define experimental value
3. What happens as the accuracy of the measurement decreases?

This page titled 2.2: Percent Error is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.13: Percent Error by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.3: Measurement Uncertainty



Figure 2.3.1 (CC BY-SA 4.0; Tony Webster via Commons Wikimedia, Police Line Crime Scene [commons.wikimedia.org])

How do police officers identify criminals?

After a bank robbery has been committed, police will ask witnesses to describe the thieves. They will usually get an answer such as "medium height". Others may say something along the lines of "between 5 foot 8 inches and 5 foot 10 inches". In both cases, there is a significant amount of uncertainty about the height of the criminals.

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. The figure below shows two rulers making the same measurement of an object (indicated by the blue arrow).

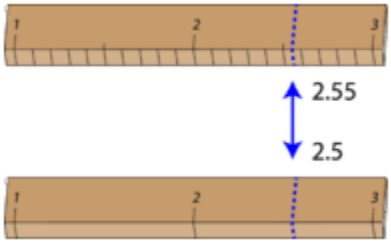


Figure 2.3.2: Uncertainty in measurement. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

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). With this ruler, 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.

YouTube Video

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

View this YouTube Video online

Summary

- Uncertainty exists in all measurements.
- The degree of uncertainty is affected in part by the quality of the measuring tool.

Review

1. What is uncertainty in measurements?
2. Why is the top ruler more reliable in measuring length than the bottom ruler?
3. How could the top ruler be made more accurate?

This page titled 2.3: Measurement Uncertainty is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.14: Measurement Uncertainty by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.4: Significant Figures



Figure 2.4.1 (Public Domain; User:WhisperToMe/Wikimedia Commons via Wikipedia)

How fast do you drive?

As you enter the town of Jacinto City, Texas, a sign tells you that the speed limit is 30 miles per hour. But what if you happen to be driving 31 miles an hour? Are you in trouble? Probably not, because there is a certain amount of leeway built into enforcing the regulation. Most speedometers do not measure the vehicle speed very accurately and could easily be off by a mile or so (on the other hand, radar measurements are much more accurate). So, a couple of miles per hour difference won't matter that much. However, if you stretch the limits any further, you may end up receiving a traffic ticket.

Significant Figures

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

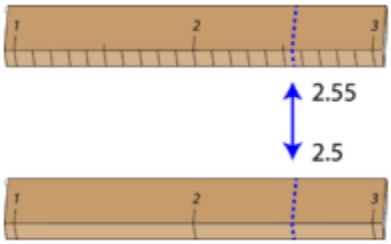


Figure 2.4.2: Measurement with two different rulers. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

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.

Figure 2.4.1*: 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 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 left-end zeros. Left-end 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 right-end zeros. Right-end zeros in a number that lacks a decimal point are not significant.	• 140 has two significant figures. • 75,210 has four significant figures.
5. Right-end 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.

It needs to be emphasized that to say a certain digit is not significant, does not mean that it is not important or can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2 with two significant figures in the coefficient. For a number with left-end zeros, such as 0.000416, it can be written as 4.16×10^{-4} with 3 significant figures. In some cases, scientific notation is the only way to correctly indicate the correct number of significant figures. In order to report a value of 15,000,000 with four significant figures, it would need to be written as 1.500×10^7. The right-end zeros after the 5 are significant. The original number of 15,000,000 only has two significant figures.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Interactive Element

Practice using significant figures in this simulation.

Summary

- 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.

Review

1. What does a significant figure tell us?
2. What is a left-end zero?
3. What is a right-end zero?
4. What does an insignificant zero do?

This page titled 2.4: Significant Figures is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.16: Significant Figures by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.5: Rounding



Figure 2.5.1 (Public Domain; Nancy Heise via Wikipedia)

Have you ever been fishing

People who fish often are a little unreliable when it comes to describing what fish they caught and how much it weighed. It's easier to say that the fish weighed ten pounds than it is to accurately describe the weight of 8 pounds 11 ounces. Ten pounds is "close enough" when a fisherman is talking about their catch.

Rounding

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

Consider the measurement $207.518\ \text{m}$. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table 2.5.1.

Table 2.5.1: 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.

YouTube Video

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

View this YouTube Video online

Summary

- Rounding involves the adjustment of a value to account for the proper number of significant digits.
- Rules exist for rounding numbers.

Review

1. Why do we round numbers?
2. What do we need to know before we round a number?
3. What is “rounding up”?
4. What is “rounding down”?

This page titled 2.5: Rounding is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.15: Rounding by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.6: Significant Figures in Addition and Subtraction



Figure 2.6.1 (Public Domain; User:Septagram/Wikipedia via Wikipedia)

How old do you think this calculator is?

Calculators are great devices. Their invention has allowed for quick computation at work, school, and other places where manipulation of numbers needs to be done rapidly and accurately. But they are only as good as the numbers put into them. The calculator cannot determine how accurate each set of numbers is, and the answer given on the screen must be assessed by the user for reliability.

Uncertainty in Addition and Subtraction

Consider two separate mass measurements: $16.7\ \text{g}$ and $5.24\ \text{g}$. The first mass measurement, $(16.7\,\text{g})$, is known only to the tenths place, or to one digit after the decimal point. There is no information about its hundredth place and so that digit cannot be assumed to be zero. The second measurement, $(5.24\,\text{g})$, is known to the hundredths place, or to two digits after the decimal point.

When these masses are added together, the result on a calculator is $16.7 + 5.24 = 21.94\ \text{g}$. Reporting the answer as $21.94\ \text{g}$ suggests that the sum is known all the way to the hundredths place. However, that cannot be true because the hundredths place of the first mass was completely unknown. The calculated answer needs to be rounded in such a way as to reflect the certainty of each of the measured values that contribute to it. For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places. The sum of the above masses would be properly rounded to a result of $21.9\ \text{g}$.

Example 2.6.1

Determine the combined molecular mass of a glucose molecule and a maltose molecule.

Glucose molecule = $180.156\frac{g}{mol}$

Maltose molecule = $342.3\frac{g}{mol}$

Solution

$180.156 + 342.4 = 522.456$

When adding and subtracting we know to look at the least number of decimals in our starting values; in this case 342.3 has only 1 digit after the decimal, so we need to round our answer to the same place.

\(522.456\to 522.5\frag{g}{mol}

When working with whole numbers, pay attention to the last significant digit that is to the left of the decimal point, and round your answer to that same point. For example, consider the subtraction: $78,500\ \text{m} - 362\ \text{m}$. The calculated result is $78,138\ \text{m}$. However, the first measurement is known only to the hundreds place, as the 5 is the last significant digit. Rounding the result to that same point means that the correct result is $78,100\ \text{m}$.

Example 2.6.2

What is $4200 + 540$ = ?

Solution

$4200 + 540 = 4740$

To determine where to round our answer, we look at our starting numbers to see which has the fewest number of decimal places. They both have 0 so we round to the nearest whole number, 4740.

YouTube Video

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

View this YouTube Video online

Summary

- For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places.

Review

1. What is the basic principle to use in working with addition and subtraction?
2. What do you pay attention to when working with whole numbers?

This page titled 2.6: Significant Figures in Addition and Subtraction is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.17: Significant Figures in Addition and Subtraction by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.7: Significant Figures in Multiplication and Division



Figure 2.7.1 (Public Domain; Adrian Pingstone (User:Arpingstone/Wikimedia Commons) via Wikipedia)

Who should report the numbers - you or your calculator?

Calculators do just what is asked of them—no more and no less. However, they sometimes can get a little out of hand. If you multiply 2.48 times 6.3, you get an answer of 15.687, a value that ignores the number of significant figures in either number. Division with a calculator is even worse. When you divide 12.2 by 1.7, the answer you obtain is 7.176470588. Neither piece of data is accurate to nine decimal places, but the calculator does not know that. The human being operating the instrument has to make the decision about how many places to report.

Uncertainty in Multiplication and Division

The density of a certain object is calculated by dividing the mass by the volume. Suppose that a mass of $37.46\ \text{g}$ is divided by a volume of $12.7\ \text{cm}^3$. The result on a calculator would be:

$$D = \frac{m}{V} = \frac{37.46\ \text{g}}{12.7\ \text{cm}^3} = 2.949606299\ \text{g/cm}^3$$

The value of the mass measurement has four significant figures, while the value of the volume measurement has only three significant figures. For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures. Applying this rule results in a density of $2.95\ \text{g/cm}^3$, with three significant figures—the same as the volume measurement.

YouTube Video

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

View this YouTube Video online

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

Example 2.7.1

Perform the following calculations, rounding the answers to the appropriate number of significant figures.

A. $0.048\ \text{m} \times 32.97\ \text{m}$

B. $14,570\ \text{kg} \div 5.81\ \text{L}$

Solution

Step 1: Plan the problem.

Analyze each of the measured values to determine how many significant figures should be in the result. Perform the calculation and round appropriately. Apply the correct units to the answer. When multiplying or dividing, the units are also multiplied or divided.

Step 2: Calculate.

A. $0.048\ \text{m} \times 32.97\ \text{m} = 1.6\ \text{m}^2$ Round to two significant figures because 0.048 has two.

B. $14,570\ \text{kg} \div 5.81\ \text{L} = 2510\ \text{kg/L}$ Round to three significant figures because 5.81 has three.

Exercise 2.7.1

How many significant figures should the answer contain?

$10.61 \times 12.133 \times 3.25$

Answer available at the end of this section.

Summary

- For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures.

Review

1. Perform the calculation and round your answer with the correct number of significant figures.
   a. $78.2 \div 32\ \text{cm}^3$
   b. 3.0 m/s × 9.21 s
2. What happens to units in multiplication and division problems?

Answers

Exercise 2.7.1

=418; 3 significant figures

This page titled 2.7: Significant Figures in Multiplication and Division is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.18: Significant Figures in Multiplication and Division by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

2.8: Scientific Notation

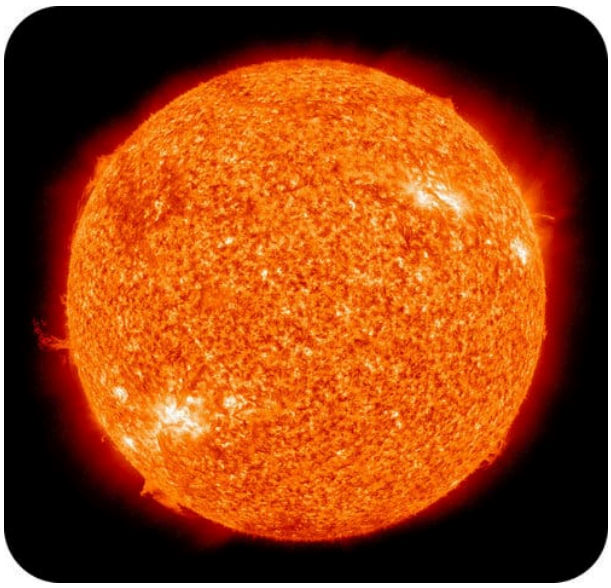


Figure 2.8.1 (Public Domain; NASA via Wikipedia)

How far is the Sun from Earth?

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

Scientific Notation

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

Table 2.8.1: SI Prefixes

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
giga	G	10^9	1,000,000,000	1 gigameter $(\text{Gm}) = 10^9\ \text{m}$
mega	M	10^6	1,000,000	1 megameter $(\text{Mm}) = 10^6\ \text{m}$
kilo	k	10^3	1,000	1 kilometer $(\text{km}) = 1,000\ \text{m}$
hecto	h	10^2	100	1 hectometer $(\text{hm}) = 100\ \text{m}$
deka	da	10^1	10	1 dekameter $(\text{dam}) = 10\ \text{m}$
		10^0	**1**	**1 meter** (m)
deci	d	10^{-1}	1/10	1 decimeter $(\text{dm}) = 0.1\ \text{m}$
centi	c	10^{-2}	1/100	1 centimeter $(\text{cm}) = 0.01\ \text{m}$

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
milli	m	10^{-3}	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001\ \mathrm{m}$
micro	μ	10^{-6}	1/1,000,000	1 micrometer $(\mu\mathrm{m}) = 10^{-6}\ \mathrm{m}$
nano	n	10^{-9}	1/1,000,000,000	1 nanometer $(\mathrm{nm}) = 10^{-9}\ \mathrm{m}$
pico	p	10^{-12}	1/1,000,000,000,000	1 picometer $(\mathrm{pm}) = 10^{-12}\ \mathrm{m}$

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

Interactive Content

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

View this Interactive Content online

Interactive Content

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

View this Interactive Content online

Summary

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

Review

1. What is scientific notation?
2. What do we use scientific notation for?
3. What is a leading zero?
4. Express 150,000,000 in scientific notation.
5. Express 0.0000043 in scientific notation.

This page titled 2.8: Scientific Notation is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.3: Scientific Notation by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3: UNIT SYSTEMS AND DIMENSIONAL ANALYSIS

3.1: SI Base Units

3.2: Metric Prefixes

3.3: Length and Volume

3.4: Mass and Weight

3.5: Metric Unit Conversions

3.6: Derived Units

3.7: Dimensional Analysis

3.8: Density

3.9: Kinetic Energy

3.10: Temperature and Temperature Scales

This page titled 3: Unit Systems and Dimensional Analysis is shared under a license and was authored, remixed, and/or curated by LibreTexts.

3.1: SI Base Units

How long is a yard?

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

SI Base Units

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

Table 3.1.1: SI Base Units of Measurement

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

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



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



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

YouTube Video

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

View this YouTube Video online

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

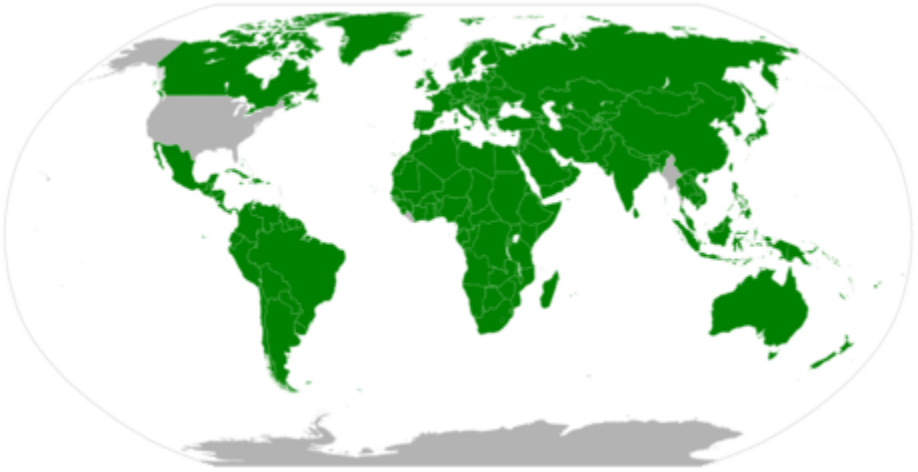


Figure 3.1.3: Areas of the world that use the metric system (green). (Public Domain; User:Canuckguy/Wikipedia and User:AzaToth/Wikimedia Commons via Wikipedia)

Summary

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

Review

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

This page titled 3.1: SI Base Units is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.1: SI Base Units by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.2: Metric Prefixes



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

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

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

Metric Prefixes

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

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

Table 3.2.1: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\mathrm{Gm}) = 10^9\ \mathrm{m}$
mega	M	1,000,000	1 megameter $(\mathrm{Mm}) = 10^6\ \mathrm{m}$
kilo	k	1,000	1 kilometer $(\mathrm{km}) = 1,000\ \mathrm{m}$
hecto	h	100	1 hectometer $(\mathrm{hm}) = 100\ \mathrm{m}$
deka	da	10	1 dekameter $(\mathrm{dam}) = 10\ \mathrm{m}$
		1	**1 meter** (m)

Prefix	Unit Abbreviation	Meaning	Example
deci	d	1/10	1 decimeter $(\mathrm{dm}) = 0.1\ \mathrm{m}$
centi	c	1/100	1 centimeter $(\mathrm{cm}) = 0.01\ \mathrm{m}$
milli	m	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001\ \mathrm{m}$
micro	μ	1/1,000,000	1 micrometer $(\mu\mathrm{m}) = 10^{-6}\ \mathrm{m}$
nano	n	1/1,000,000,000	1 nanometer $(\mathrm{nm}) = 10^{-9}\ \mathrm{m}$
pico	p	1/1,000,000,000,000	1 picometer $(\mathrm{pm}) = 10^{-12}\ \mathrm{m}$

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

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

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 $6.5\ \mathrm{kg}$, this is easier than saying that it weighs $6500\ \mathrm{g}$ or $0.65\ \mathrm{dag}$. 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.

YouTube Video

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

View this YouTube Video online

Simulations

Practice using metric prefixes with this simulation.

Summary

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

Review

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

This page titled 3.2: Metric Prefixes is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.2: Metric Prefixes by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.3: Length and Volume



Figure 3.3.1 (Public Domain; User:Żeglarz/Wikimedia Commons via Wikipedia)

How were sailors able to measure the depths of seas?

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

Length and Volume

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

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

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 3.3.2: A typical water bottle is 1 liter in volume. (Public Domain; User:Kenyon/Wikimedia Commons via Wikipedia)

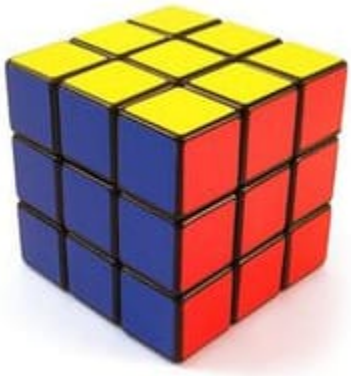


Figure 3.3.3: This Rubik's cube is $5.7\ \text{cm}$ on each side and has a volume of $185.2\ \text{cm}^3$ or $185.2\ \text{mL}$. (CC by 2.0; Bram Van Damme (Flickr: Bramus!) via Flickr)



Figure 3.3.4: A graduated cylinder is often used to measure volume in the laboratory, and comes in a variety of sizes. (Public Domain; User:Darrien/Wikipedia via Wikipedia)

Summary

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

Review

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

4. How big is a mL?

This page titled 3.3: Length and Volume is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.4: Length and Volume by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.4: Mass and Weight



Figure 3.4.1 (Public Domain; NASA via Wikipedia)

How is he floating?

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

Mass and Weight

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



Figure 3.4.2: An analytical balance makes very sensitive mass measurements in a laboratory, usually in grams. (Public Domain; US Drug Enforcement Administration via Wikipedia)

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are $1000\ \mathrm{g}$ in

$1\ \text{kg}$. A milligram is 1/1000th of a gram, so there are $1000\ \text{mg}$ in $1\ \text{g}$.

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

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 3.4: Mass and Weight is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.5: Mass and Weight by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.5: Metric Unit Conversions

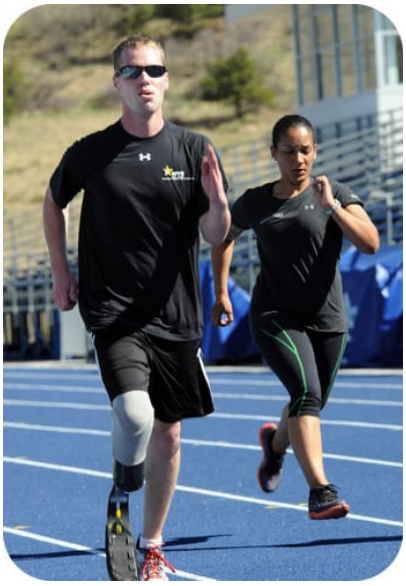


Figure 3.5.1 (Public Domain; the US Army via Wikipedia)

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

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

Metric Unit Conversions

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

Example 3.5.1

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

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

Step 3:Think about your result.

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

Two-Step Metric Unit Conversions

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

Example 3.5.2

Convert $4.3\ \text{cm}$ to μm.

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

Step 3: Think about your result.

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

YouTube Video

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

View this YouTube Video online

cK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

Summary

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

Review

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

This page titled 3.5: Metric Unit Conversions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.9: Metric Unit Conversions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.6: Derived Units



Figure 3.6.1 (Public Domain; User:Joegrimes/Wikipedia via Commons Wikimedia, Farm landly sander [commons.wikimedia.org])

How has farming evolved?

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

Dimensional Analysis and Derived Units

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

Derived SI Units

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

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

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

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

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

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

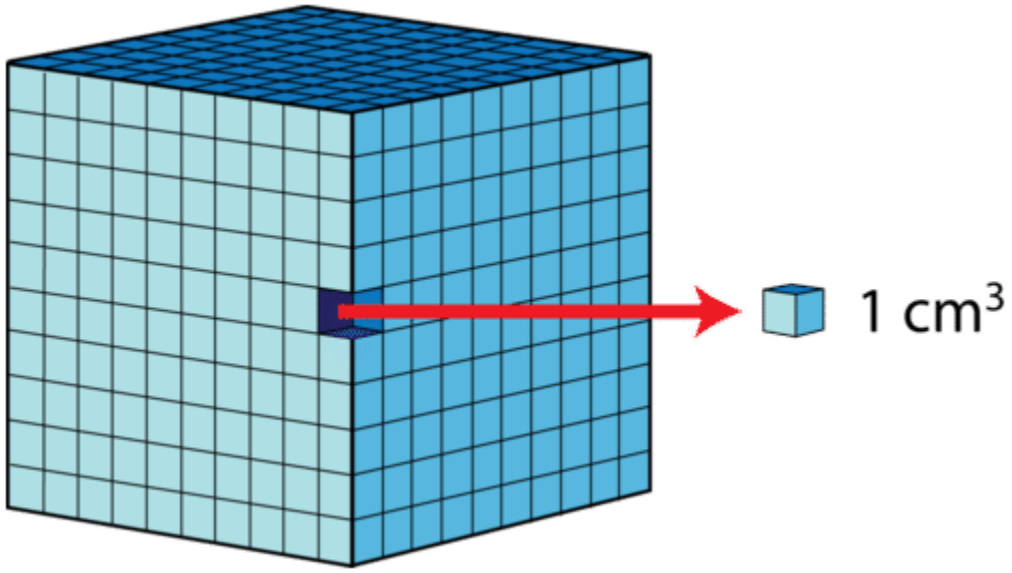


Figure 3.6.2: There are $1000\ \text{cm}^3$ in $1\ \text{dm}^3$. Since a cm^3 is equal to a mL and a dm^3 is equal to a L, we can say that there are $1000\ \text{mL}$ in $1\ \text{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 3.6.1

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

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

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

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

Step 3: Think about your result.

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 3.6: Derived Units is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.10: Derived Units by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.7: Dimensional Analysis

Conversion Factors

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

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

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

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

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

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

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

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

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

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

Scientific Dimensional Analysis

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

Example 3.7.1

How many seconds are in a day?

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

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

Step 3: Think about your result.

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

Summary

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Review

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

Explore More

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

1. What do we always need to express measurements correctly?
2. What does dimensional analysis tell you?
3. How do you know that you have set the problem up incorrectly?
4. How do you know that you have set the problem up correctly?

This page titled 3.7: Dimensional Analysis is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.8: Dimensional Analysis by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.8: Density



Figure 3.8.1 (CC BY 2.0; Tony Hisgett (Flickr:ahisgett) via Flickr)

How do logs stay afloat in water?

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

Density

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

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

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

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

Table 3.8.1: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084

Liquids and Solids	Density at $20^{\circ}\mathrm{C}$ $(\mathrm{g/mL})$	Gases	Density at $20^{\circ}\mathrm{C}$ $(\mathrm{g/L})$
Ice $(0^{\circ}\mathrm{C})$	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water $(4^{\circ}\mathrm{C})$	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

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

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

Example 3.8.1

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

Solution

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

Known

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

Unknown

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

Step 2: Calculate.

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

Step 3: Think about your result.

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

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

Example 3.8.2

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

Solution

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

Known

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

Unknown

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

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

Step 2: Calculate.

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

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

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

Step 3: Think about your results.

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

Interactive Content

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

View this Interactive Content online

Interactive Content

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

View this Interactive Content online

Simulation

Go fishing and explore the variables that affect density in this simulation.

Summary

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

Review

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

This page titled 3.8: Density is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.11: Density by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.9: Kinetic Energy



Figure 3.9.1 (Public Domain; User:Dlls/Wikimedia Commons via Wikipedia)

Have you ever watched a cat in action?

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

SI Kinetic Energy Units

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

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

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

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



Figure 3.9.2: James Prescott Joule. (Public Domain; John Collier via Wikipedia)

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 3.9: Kinetic Energy is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.6: Kinetic Energy by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

3.10: Temperature and Temperature Scales



Figure 3.10.1 (CC0; Tantien3 via Pixabay / Public Domain)

An ice cube in your hand feels cold, but why?

The particles of matter in a hot object are moving much faster than the particles of matter in a cold object. An object's **kinetic energy** is the energy due to motion. The particles of matter that make up the hot stove have a greater amount of kinetic energy than those in the ice cube.

Temperature and Temperature Scales

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 an 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 3.10.2: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy. (Charcoal: CC-BY 2.0; Snow: Public Domain; Charcoal: Serge Melki; Snow: User:Haymanj/Wikimedia Commons via Charcoal: http://www.flickr.com/photos/sergemelki/3106924114; Snow: http://commons.wikimedia.org/wiki/File:Snow_Gum1.JPG)

Temperature Scales

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 problems with this scale are the arbitrary definitions of temperature. The freezing point of water was defined as $32^{\circ}\mathrm{F}$ and the boiling point as $212^{\circ}\mathrm{F}$. The Fahrenheit scale is typically not used for scientific purposes.



Figure 3.10.3: Daniel Gabriel Fahrenheit. (Public Domain; Anonymous via Wikipedia)

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



Figure 3.10.4: Anders Celsius. (Public Domain; Olof Arenius via Wikipedia)

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



Figure 3.10.5: Lord Kelvin. (Public Domain; Anonymous via Wikipedia)

YouTube Video

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

View this YouTube Video online

As can be seen by the 100 kelvin difference between the two, a change of one degree on the Celsius scale is equivalent to the change of one kelvin on the Kelvin scale. Converting from one scale to another is easy, as you simply add or subtract 273.

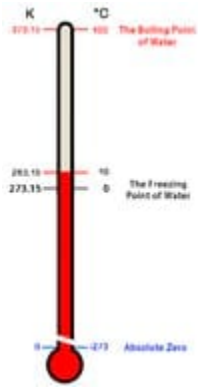


Figure 3.10.6: Comparison between Kelvin and Celsius temperature Scales. (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Simulation

No matter the temperature scale you use, the temperature at which a liquid boils depends on a few different variables. Try boiling water in a few places around the world to see what we mean.

Summary

- Temperature is a measure of the average kinetic energy of the particles in matter.
- The Fahrenheit scale defines the freezing point of water as 32°F and the boiling point as 212°F.
- The Celsius scale sets the freezing point and boiling point of water at 0°C and 100°C, respectively.
- The Kelvin scale is based on molecular motion, with the temperature of $0\ \text{K}$—also known as absolute zero—being the point where all molecular motion ceases.

Review

1. What is absolute zero on the Celsius temperature scale?
2. What are the freezing and boiling points of water in the Celsius scale?
3. Convert the following Kelvin temperatures to degrees Celsius.
   a. 188 K
   b. 631 K
4. Temperature in degrees Fahrenheit can be converted to Celsius by first subtracting 32, then dividing by 1.8. What is the Celsius temperature outside on a warm day (88°F)?
5. Why is the Celsius scale sometimes called “centigrade”?

This page titled 3.10: Temperature and Temperature Scales is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 3.7: Temperature and Temperature Scales by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4: MATTER AND CHANGE

4.1: Matter, Mass, and Volume

4.2: States of Matter

4.3: Physical Properties

4.4: Chemical Properties and Chemical Reactions

4.5: Extensive and Intensive Properties

4.6: Physical Change

4.7: Chemical Change

4.8: Recognizing Chemical Reactions

4.9: Pure Substances

4.10: Mixture

4.11: Heterogeneous Mixtures

4.12: Homogeneous Mixture

4.13: Separating Mixtures

4.14: Elements

4.15: Compounds

4.16: Chemical Symbols and Formulas

4.17: A Molecular View of Elements and Compounds

This page titled 4: Matter and Change is shared under a license and was authored, remixed, and/or curated by LibreTexts.

4.1: Matter, Mass, and Volume

Can you guess what this colorful image shows?

Believe it or not, it actually depicts individual atoms of cesium (reddish-orange) on a surface of gallium arsenide molecules (blue). The image was created with an extremely powerful microscope, called a scanning tunneling microscope. This is one of a few types of microscope that can make images of atoms, the basic building blocks of matter.

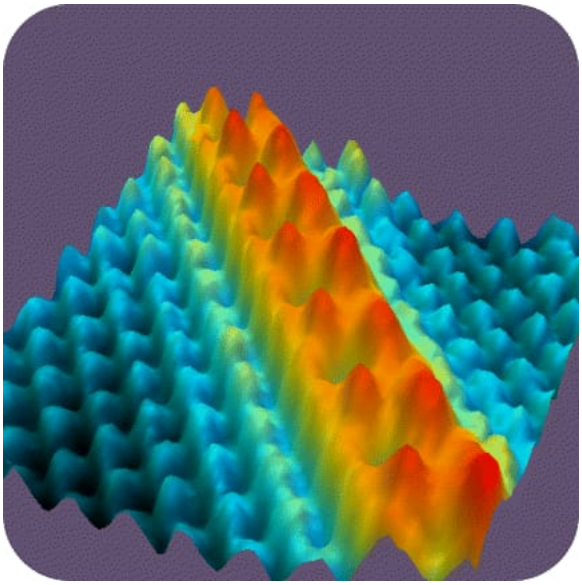


Figure 4.1.1 (Public Domain; L.J. Whitman, J.A. Stroscio, R.A. Dragoset, R.J. Celotta, and the National Institute of Standards and Technology via NIST, Physical Measurement Laboratory, Scanning Tunneling Microscope [www.nist.gov])

What's the Matter?

Matter is all the "stuff" that exists in the universe. Everything you can see and touch is made of matter, including you! The only things that aren't matter are forms of energy, such as light and sound. In science, **matter** is defined as anything that has mass and volume. Mass and volume measure different aspects of matter.

Mass

Mass is a measure of the amount of matter in a substance or an object. The basic SI unit for mass is the kilogram (kg), but smaller masses may be measured in grams (g). To measure mass, you would use a balance. In the lab, mass may be measured with a triple beam balance or an electronic balance, but the old-fashioned balance pictured below may give you a better idea of what mass is. If both sides of this balance were at the same level, it would mean that the fruit in the left pan has the same mass as the iron object in the right pan. In that case, the fruit would have a mass of 1 kg, the same as the iron. As you can see, however, the fruit is at a higher level than the iron. This means that the fruit has less mass than the iron, that is, the fruit's mass is less than 1 kg.

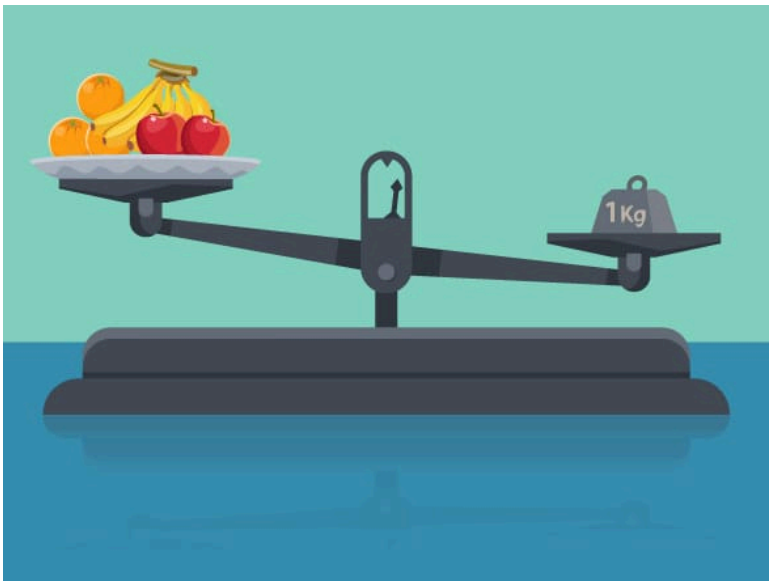


Figure 4.1.2 (CK-12 Curriculum Materials License ; CK-12 Foundation via CK-12 Foundation)

Example 4.1.1

Refer to the picture above. If the fruit were at a lower level than the iron object, what would be the mass of the fruit?

Solution

The mass of the fruit would be greater than 1 kg.

Mass vs. Weight

Mass is commonly confused with weight. The two are closely related, but they measure different things. Whereas mass measures the amount of matter in an object, weight measures the force of gravity acting on an object. The force of gravity on an object depends on its mass but also on the strength of gravity. If the strength of gravity is held constant (as it is all over Earth), then an object mass is directly proportional to the objects weight, so a greater mass also has a greater weight.

Example 4.1.2

With Earth's gravity, an object with a mass of 1 kg has a weight of 2.2 lb. How much does a 10 kg object weigh on Earth?

Solution

A 10 kg object weighs ten times as much as a 1 kg object: 10 × 2.2 lb = **22 lb**

Volume

Volume is a measure of the amount of space that a substance or an object takes up. The basic SI unit for volume is the cubic meter (m^3), but smaller volumes may be measured in cm^3, and liquids may be measured in liters (L) or milliliters (mL). How the volume of matter is measured depends on its state.

- The volume of a liquid is measured with a measuring container, such as a measuring cup or graduated cylinder.
- The volume of a gas depends on the volume of its container: gases expand to fill whatever space is available to them.
- The volume of a regularly shaped solid can be calculated from its dimensions. For example, the volume of a rectangular solid is the product of its length, width, and height.

CK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

- The volume of an irregularly shaped solid can be measured by the displacement method. You can read below how this method works.

Example 4.1.3: Calculating Volume from Dimensions

How could you find the volume of air in an otherwise empty room?

Solution

If the room has a regular shape, you could calculate its volume from its dimensions. For example, the volume of a rectangular room can be calculated with the formula:

$$\text{Volume} = \text{length} \times \text{width} \times \text{height}$$

If the length of the room is 5.0 meters, the width is 3.0 meters, and the height is 2.5 meters, then the volume of the room is:

$$\text{Volume} = 5.0 \text{ m} \times 3.0 \text{ m} \times 2.5 \text{ m} = 37.5 \text{ m3}$$

Measuring Volume Using the Displacement Method

The following video shows how the volume of an irregular shaped object, like your science teacher can be measured by the displacement method.

YouTube Video

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

View this YouTube Video online

Example 4.1.4

What is the volume of the dinosaur in the diagram below?

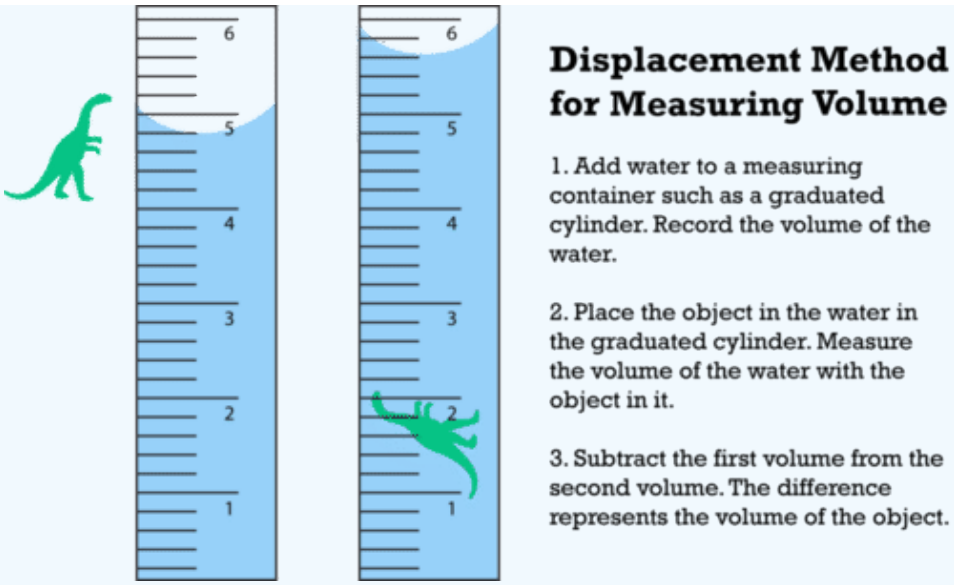


Figure 4.1.3 (CK-12 Curriculum Materials License; Christopher AuYeung via CK-12 Foundation)

Solution

The volume of the water alone is 4.8 mL. The volume of the water and dinosaur together is 5.6 mL. Therefore, the volume of the dinosaur alone is 5.6 mL – 4.8 mL = 0.8 mL.

Summary

- Matter is all the “stuff” that exists in the universe. It has both mass and volume.
- Mass measures the amount of matter in a substance or an object. The basic SI unit for mass is the kilogram (kg).
- Volume measures the amount of space that a substance or an object takes up. The basic SI unit for volume is the cubic meter (m^3).

Review

1. How do scientists define matter?
2. What is mass? What is the basic SI unit of mass?
3. What does volume measure? Name two different units that might be used to measure volume.
4. Explain how to use the displacement method to find the volume of an irregularly shaped object.

This page titled 4.1: Matter, Mass, and Volume is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

4.2: States of Matter

Why is the state of water different in each picture?

Water can take many forms. At low temperatures (below $0^{\circ}\mathrm{C}$), water is a solid. When at "normal" temperatures (between $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$), it is a liquid. At temperatures above $100^{\circ}\mathrm{C}$, water is a gas (steam).



Figure $4.2.1$ (Public Domain; Iceberg: Rear Admiral Harley D. Nygren, NOAA; Beach: User:Wicki/Wikimedia Commons; Volcano: E. Klett, US Fish and Wildlife Service via Iceberg: http://commons.wikimedia.org/wiki/File:Ice_berg.jpg; Beach: http://commons.wikimedia.org/wiki/File:Ocean_Spokojny.JPG; Volcano: http://commons.wikimedia.org/wiki/File:Dds40-097_large.jpeg)

The state of water depends on 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**. There is a fourth state of matter called **plasma**, which rarely exists on earth, but we will omit this from our current discussion. 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.

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), which is a property that all liquids exhibit. Mercury has a relatively high surface tension, and this makes it very unique. Here you can see mercury in its common liquid form.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 4.2.2: Mercury. (the EPA; Source: http://commons.wikimedia.org/wiki/File:Mercury-element.jpg; License: Public Domain)

If we heat liquid mercury to its boiling point of $357^{\circ}\mathrm{C}$, and contain it 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.

Solid

Solids are defined by the following characteristics:

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

If we were to cool a sample of liquid mercury to its freezing point of $-39^{\circ}\mathrm{C}$, and had it contained under the right pressure conditions, we would notice all of the liquid particles would go into the solid state.

YouTube Video

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

View this YouTube Video online

As you can see in the video, 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.

YouTube Video

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

View this YouTube Video online

Plasma

Plasma is a state of matter that resembles a gas but has certain properties that gases do not have. Like a gas, plasma consists of particles of matter that can pull apart and spread out, so it lacks a fixed volume and a fixed shape. Unlike a gas, plasma can conduct electricity and respond to a magnetic field. That's because plasma consists of electrically charged particles called ions, instead of uncharged particles such as atoms or molecules.

Plasma are defined by the following characteristics:

- particles are charged ions and free electrons
- no definite shape
- no definite volume
- conducts electricity
- responds to magnetic field

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 their container.
- Gases have no definite shape or volume.

Review

1. How many states of matter are there?
2. What is a solid?
3. What is a liquid?
4. What is a gas?

This page titled 4.2: States of Matter is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.5: States of Matter by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.3: Physical Properties

Why are drag car standards constantly reinforced?

Drag racing is a highly competitive (and expensive) sport. There are a variety of classes of vehicles, ranging from stock classes (dependent on car weight, engine size, and degree of engine modification) all the way up to the Top Fuel class. Cars in the Top Fuel class may weigh over two thousand pounds, and are capable of top speeds of well over 300 miles per hour at the end of the quarter-mile. The standards for each class are well-defined, and frequent checks are made of engine dimensions and components to ensure that the rules are followed.

Physical Properties

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, brittle, and conducts electricity when it has been dissolved into water (which it does quite easily). Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, **melting point**, and **boiling point**.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red.

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.

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 4.3.1: Pencil (Public Domain; Public Domain via Wikipedia)



Figure 4.3.2: Diamond ring (Public Domain; User:Ikkyu2/Wikimedia Commons via Wikipedia)

Melting point and boiling point are somewhat unique identifiers, especially in the case 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.

YouTube Video

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

View this YouTube Video online

Is that Really Gold? Archimedes and Density

How do you determine the density of an irregularly shaped object? Learn how Archimedes did this thousands of years ago with just a beaker, water, and some simple math.

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

View this YouTube Video online

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, melting point, and boiling point.

Review

1. What is a physical property?
2. What color are most metals?
3. Is titanium harder or softer than gold?

This page titled 4.3: Physical Properties is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.3: Physical Properties by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.4: Chemical Properties and Chemical Reactions

Have you ever left your bicycle out in the rain?

It's not a good idea because the rain can cause the bicycle to rust. You start to get a reddish-orange build-up of a rough deposit on the metal. It may start with the chain, but can spread to other parts of the bicycle, especially if there are scratches that create a bare metal surface. The formation of rust is a chemical process that takes place when iron is exposed to water and oxygen. It is estimated that damage due to rust costs U.S. businesses, military, and government over 276 billion dollars a year—a very expensive chemical process.

Chemical Properties

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. The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used to describe 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.

Chemical Change

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). The figure below shows the substances involved in this reaction.

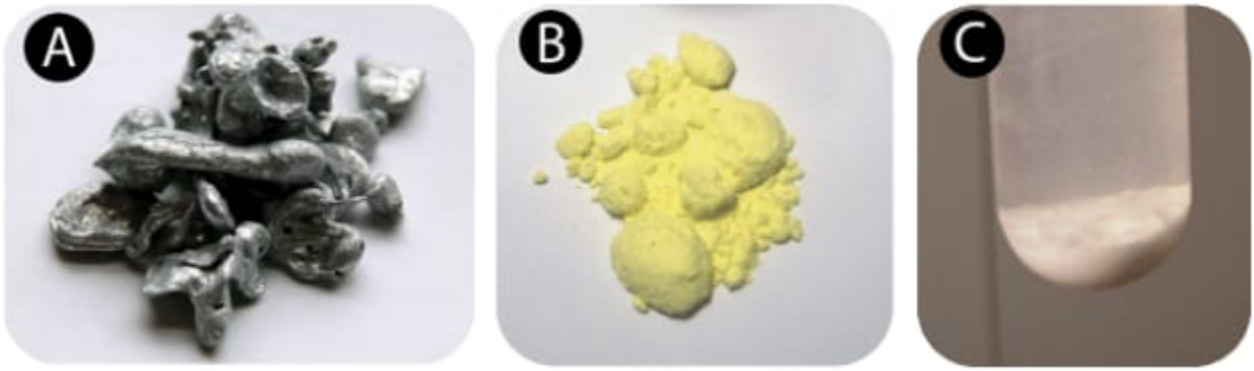


Figure $4.4.1$: Zinc and sulfur are two elements that undergo a chemical reaction when heated to form the compound zinc sulfide. ((A; (A) Ben Mills (Benjah-bmm27); (B) User:Vineyard/Wikimedia Commons; (C) A. Eisen/NOAA via (A) http://commons.wikimedia.org/wiki/File:Zinc-sample.jpg; (B) http://commons.wikimedia.org/wiki/File:SulfurReagent.jpg; (C) http://oceanexplorer.noaa.gov/explorations/03windows/logs/jul30/media/zincsulfideprecipitate.html) Public Domain; (B) CC BY 3.0; (C) Public Domain)

Alkaline metals are highly reactive with water. Watch the video to see what happens when different alkaline metals are exposed to water.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Summary

- A chemical property describes the ability of a substance to undergo a specific chemical change.
- A chemical reaction is a process that occurs when one or more substances are changed into one or more new substances.

Review

1. What is a chemical property?
2. What is a chemical reaction?
3. When can we observe a chemical property?
4. Is freezing water a chemical property?

This page titled 4.4: Chemical Properties and Chemical Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.16: Chemical Properties and Chemical Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.5: Extensive and Intensive Properties

How much is twenty dollars really worth?

You agree to mow someone's lawn for twenty dollars (it's a fairly large yard). When they pay you, they give you a $20 bill. It doesn't matter whether the bill is brand new or old, dirty, and wrinkled—all of these bills have the same value of $20. If you want more $20 bills, you have to mow more lawns. You can't say, "this particular bill is actually worth more than $20." To have more money, you have to put in more work.

Extensive Properties

Some properties of matter depend on the size of the sample, while some do not. An **extensive** property is a property that depends on the amount of matter in a sample. The mass of an object is a measure of the amount of matter that an object contains. A small sample of a certain type of matter will have a small mass, while a larger sample will have a greater mass. Another extensive property is volume. The volume of an object is a measure of the space that is occupied by that object.

The figure below illustrates the extensive property of volume. The pitcher and glass both contain milk. The pitcher holds approximately two quarts and the glass will hold about 8 ounces of milk. The same milk is in each container. The only difference is the amount of milk contained in the glass and in the pitcher.



Figure 4.5.1: Milk pitcher and glass. (Public Domain; Zenon Niewada (Wikimedia: Pitcherman) via Wikipedia)

Intensive Properties

The electrical conductivity of a substance is a property that depends only on the type of substance. Silver, gold, and copper are excellent conductors of electricity, while glass and plastic are poor conductors. A larger or smaller piece of glass will not change this property. An **intensive** property is a property of matter that depends only on the type of matter in a sample and not on the amount. Other intensive properties include color, temperature, density, and solubility.



Figure 4.5.1: Copper wire. (Credit: User:Inductiveload/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Tinned_Copper_Wire_anaglyph.jpg)

The copper wire shown in the picture below has a certain electrical conductivity. You could cut off the small end that sticks out, and it would have the same conductivity as the entire long roll of wire shown here. The conductivity is a property of the copper metal itself, not of the length of the wire.

YouTube Video

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

View this YouTube Video online

Summary

- An extensive property is a property that depends on the amount of matter in a sample.
- Mass and volume are examples of extensive properties.
- An intensive property is a property of matter that depends only on the type of matter in a sample and not on the amount.
- Color, temperature, and solubility are examples of intensive properties.

Review

1. Define extensive property.
2. Give two examples of extensive properties.
3. Define intensive property.
4. Give two examples of intensive properties.

This page titled 4.5: Extensive and Intensive Properties is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.4: Extensive and Intensive Properties by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.6: Physical Change

How does a car go from new and shiny to old and beaten up?

Want to buy a car – cheap? Notice there is no specification such as “in good condition” or “needs a little work.” The car above is pretty beat up. The body is damaged, the windows are broken, and the interior is probably torn up. But this is still a car. It has all the components of a car, even though you would not want to buy it in the present condition. But change that condition and you have a (possibly) useable car.



Figure 4.6.1 (CC BY 2.0; User: Paul / Wikimedia Commons via Wikipedia)

Physical Change

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 4.6.2: Melting ice in the Beaufort Sea. (Public Domain; Rear Admiral Harley D. Nygren, NOAA via Wikipedia)

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



Figure 4.6.3: Chopping wood is an irreversible physical change. (Pixabay License; stevepb via Pixabay)

Physical Changes vs. Chemical Changes

What is the difference between physical and chemical changes? Watch the video to find out.

YouTube Video

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

View this YouTube Video online

Practice

Explore a campout to see if you can determine which things are undergoing physical changes and which are undergoing chemical changes.

Summary

- 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.
- A physical change can be reversible where the original form of the matter can be restored, or irreversible where the original form cannot be restored.
- Melting ice and grinding wood into sawdust are examples of physical changes.

Review

1. Define physical change.
2. Why is melting an ice cube a reversible physical change?
3. Give an example of an irreversible physical change.

This page titled 4.6: Physical Change is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.6: Physical Change by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.7: Chemical Change



Figure 4.7.1 (Public Domain; Jon Sullivan via Wikipedia)

Do you like to cook?

Cooking is a valuable skill that can benefit everyone. Whether it is fixing a simple grilled cheese sandwich or preparing an elaborate meal, cooking demonstrates some basic concepts of chemistry. When you bake bread, you mix flour, sugar, yeast, and water together. After baking, this mixture has been changed to form bread, another substance that has different characteristics and qualities from the original materials. The process of baking has produced chemical changes in the ingredients that result in bread being made.

Chemical Change

Most of the elements we know about do not exist freely in nature. Sodium cannot be found by itself (unless we prepare it in the laboratory) because it interacts too easily with other materials. On the other hand, the element helium does not interact with other elements to any extent. We can isolate helium from natural gas during the process of drilling for oil.

A chemical change produces different materials than the ones we started with. One aspect of the science of chemistry is the study of the changes that matter undergoes. If chemistry was simply a study of elements that did nothing, life would be very boring (in fact, life would not exist since the elements are what make up our bodies and sustain us). But the processes of change that take place when different chemicals are combined produce all the materials that we use daily.

One type of chemical change (already mentioned) is when two elements combine to form a compound. Another type involves the breakdown of a compound to produce the elements that make it up. If we pass an electric current through bauxite (aluminum oxide, the raw material for aluminum metal), we get metallic aluminum as a product.



Figure 4.7.1: Electrolytic production of aluminum. (Public Domain; Ronny Jacques via Wikipedia)

However, the vast majority of chemical changes involve one compound being transformed into another compound. There are literally millions of possibilities when we take this approach to chemical change. New compounds can be made to produce better fabrics that are easier to clean and maintain; they can help preserve food so it doesn't spoil as quickly; we can make new medicines to treat diseases; even products, such as airbags (as seen in the simulation below), use a chemical change to improve the world around us.

Simulation

How does a chemical reaction inflate an airbag?

Summary

A chemical change produces different materials than the ones we started with.

Review

1. What is a chemical change?
2. List three types of chemical changes.

This page titled 4.7: Chemical Change is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.14: Chemical Change by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.8: Recognizing Chemical Reactions



Figure 4.8.1 (Public Domain; Jon Sullivan via Wikipedia)

Have you ever cooked a pizza?

Making a pizza can be as easy as buying a "take and bake" from a store and putting it in the oven, or as complicated as mixing the dough and loading it up with your favorite toppings before baking it. How do you know when it is done? The most obvious sign is that the crust turns light brown. The dough is no longer flexible, but much more rigid. Maybe the cheese has melted. You certainly want the pizza to be cooked, not half-raw.

Recognizing Chemical Reactions

How can a scientist tell if a **chemical reaction** is taking place? There are four visual clues that indicate that a chemical reaction is likely occurring:

1. A change in 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 transfer of energy occurs as a result of the reaction.

Mercury (II) oxide is a red solid. When it is heated to a temperature above $500^{\circ}\mathrm{C}$, it easily decomposes into mercury and oxygen gas. The red color of the mercury oxide reactant becomes the silver color of mercury. The color change is a sign that the reaction is occurring.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 4.8.1: Mercuric oxide. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)



Figure 4.8.2: Mercury metal. (Public Domain; User:A/Wikimedia Commons via Wikipedia)

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced. The production of a gas is also an indication that a chemical reaction is occurring.



Figure 4.8.3: Zinc reacting with hydrochloric acid produces bubbles of hydrogen gas. (Public Domain; User:Chemicalinterest/Wikimedia Commons via Wikipedia)

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. A precipitate is a solid product that forms from a reaction and settles out of a liquid solution. The formation of a precipitate is an indication of a chemical reaction.



Figure 4.8.4: A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed. (CC by 2.0; Paige Powers (Flickr: paigggeyy) via Flickr)

All chemical changes involve a transfer of energy. When zinc reacts with hydrochloric acid, the test tube becomes very warm as energy is released during the reaction. Some other reactions absorb energy. While energy changes are a potential sign of a chemical reaction, care must be taken to ensure that a chemical reaction is indeed taking place. Physical changes also involve a transfer of energy. The melting of a solid absorbs energy, while the condensation of a gas releases energy. The only way to be certain that a chemical reaction has taken place is to test the composition of the substances after the change has taken place, to see if they are different from the starting substances.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- There are four visual clues that indicate that a chemical reaction is likely occurring:
  - A change of color occurs during the reaction.
  - A gas is produced during the reaction.
  - A solid product called a precipitate is produced in the reaction.
  - A transfer of energy occurs as a result of the reaction.

Review

1. What was the color change when mercury (II) oxide was heated?
2. What happened when zinc metal was mixed with hydrochloric acid?
3. What happens when lead nitrate and potassium iodide are mixed?

This page titled 4.8: Recognizing Chemical Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.18: Recognizing Chemical Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.9: Pure Substances

Have you ever had the experience of opening the refrigerator and seeing something that you don't recognize?

Not an unfamiliar food, but something that has changed over the time period it was in the refrigerator. You are not sure just what it is, so you call it "stuff". Likely, after looking at it, you know it's not a good idea to eat it, so you throw it out. If you're not sure just what it is, you certainly do not want to put it in your body.

Substances

Silver and ordinary table salt are two examples of chemical **substances**. A substance is matter that has a uniform and definite composition. All samples of substances, sometimes called pure substances, have identical properties. When chemists run a chemical reaction, they want to use pure materials so they know exactly what they are dealing with. They know that the reaction involves a specific substance, so they expect the same reaction to give the same results each time that it is run.

One characteristic of a substance is that it cannot be separated into components by physical separation techniques. The **compound** acetylsalicylic acid (the active ingredient in aspirin) is a substance because it is a chemical compound that cannot be further reduced to other materials by physical means. However, the aspirin that contains the acetylsalicylic acid is a mixture. In addition to the active pain-relieving ingredient, aspirin also may contain waxes, corn starch, and cellulose, among other ingredients. These added materials bind the tablet components into a solid pill and influence how rapidly the material dissolves.



Figure 4.9.1: Aspirin tablets. (Pixabay License; Heung Soon via Pixabay)

Substances may be either **elements** or pure compounds. They may also be solid, liquid, or gas. Some of the elements (such as silver) exist in the solid form. Other elements (mercury) exist as liquids at room temperature. Elements such as helium and radon exist as gases at room temperature. A compound such as water will take the form of a solid, a liquid, or a gas, depending on the temperature.

Summary

- A substance is a pure material with a uniform and definite composition.
- A substance cannot be separated into components using physical means.
- A substance can be a solid, a liquid, or a gas, dependent on temperature.

Review

1. Define substance.
2. Can salt be a substance?
3. Is aspirin a substance?

This page titled 4.9: Pure Substances is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.2: Pure Substances by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.10: Mixture

Ahhhh! A tall glass of ice-cold lemonade is really refreshing on a hot day. Lemonade is a combination of lemon juice, water, and sugar. Do you know what kind of matter lemonade is? It's obviously not an element because it consists of more than one substance. Is it a compound? Not all combined substances are compounds. Some—including lemonade—are mixtures.



Figure 4.10.1 (Pixabay License; JillWellington via Pixabay)

What Is a Mixture?

A **mixture** is a combination of two or more substances in any proportion. This is different from a compound, which consists of substances in fixed proportions. The substances in a mixture also do not combine chemically to form a new substance, as they do in a compound. Instead, they just intermingle and keep their original properties. The lemonade pictured above is a mixture because it doesn't have fixed proportions of ingredients. It could have more or less lemon juice, for example, or more or less sugar, and it would still be lemonade.

Example 4.10.1

What are some other examples of mixtures?

Solution

Other examples of liquid mixtures include salt water and salad dressing. Air is a mixture of gases, mainly nitrogen and oxygen. The rock pictured in the figure below is a solid mixture.

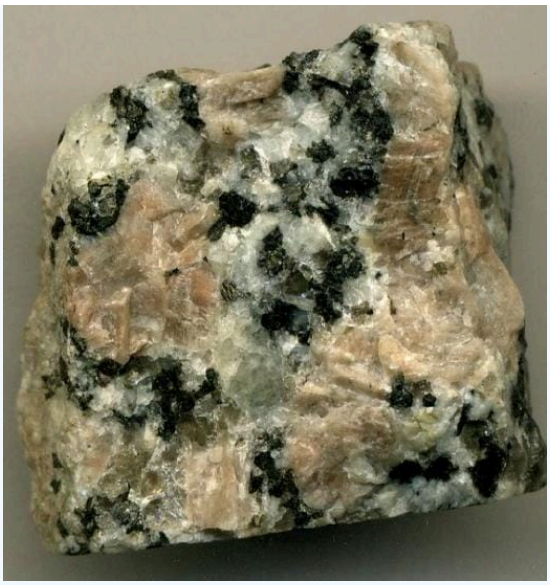


Figure 4.10.2: This rock is a mixture of smaller rocks and minerals. (CC BY 2.0; James St. John (Flickr: jsj1771) via Flickr)

Homogeneous or Heterogeneous?

The lemonade in the opening picture is an example of a homogeneous mixture. A homogeneous mixture has the same composition throughout. Another example of a homogeneous mixture is salt water. If you analyzed samples of ocean water in different places, you would find that the proportion of salt in each sample is the same: 3.5 percent.

The rock in figure above is an example of a heterogeneous mixture. A heterogeneous mixture varies in its composition. The black nuggets, for example, are not distributed evenly throughout the rock.

Types of Mixtures

Mixtures have different properties depending on the size of their particles. Three types of mixtures based on particle size are solutions, suspensions, and colloids, all of which are described in the table below.

YouTube Video

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

View this YouTube Video online

Table 4.10.1: Solutions, Suspensions, and Colloids

Type of Mixture	Description
Solutions  Figure 4.10.3 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)	A **solution** is a homogeneous mixture with tiny particles. The particles are too small to see and also too small to settle or be filtered out of the mixture. When the salt is thoroughly mixed into the water in this glass, it will form a solution. The salt will no longer be visible in the water, and it won't settle to the bottom of the glass.
Colloids  Figure 4.10.4 (CC BY 2.0; Ray Bouknight via Flickr)	A **colloid** is a homogeneous mixture with medium-sized particles. The particles are large enough to see but not large enough to settle or be filtered out of the mixture. The gelatin in this dish is a colloid. It looks red because you can see the red gelatin particles in the mixture. However, the particles are too small to settle to the bottom of the dish.

Type of Mixture	Description
Suspensions  Figure 4.10.5 (Pixabay License; jlarbig via Pixabay)	A **suspension** is a heterogeneous mixture with large particles. The particles are large enough to see and also to settle or be filtered out of the mixture. The salad dressing in this bottle is a suspension. It contains oil, vinegar, herbs, and spices. If the bottle sits undisturbed for very long, the mixture will separate into its component parts. That's why you should shake it before you use it.

Example 4.10.2

If you buy a can of paint at a paint store, a store employee may put the can on a shaker machine to mix up the paint in the can. What type of mixture is the paint?

Solution

The paint is a suspension. Some of the components of the paint settle out of the mixture when it sits undisturbed for a long time. This explains why you need to shake (or stir) the paint before you use it.

Example 4.10.3

The milk you buy in the supermarket has gone through a process called homogenization. This process breaks up the cream in the milk into smaller particles. As a result, the cream doesn't separate out of the milk no matter how long it sits on the shelf. Which type of mixture is homogenized milk?

Solution

Homogenized milk is a colloid. The particles in the milk are large enough to see—that's why milk is white instead of clear like water, which is the main component of milk. However, the particles are not large enough to settle out of the mixture.

Separating Mixtures

The components of a mixture keep their own identity when they combine, so they retain their physical properties. Examples of physical properties include boiling point, ability to dissolve, and particle size. When components of mixtures vary in physical properties such as these, processes such as boiling, dissolving, or filtering can be used to separate them.

Look at the figure below of the Great Salt Lake in Utah. The water in the lake is a solution of salt and water. Do you see the white salt deposits near the shore? How did the salt separate from the salt water? Water has a lower boiling point than salt, and it

evaporates in the heat of the sun. With its higher boiling point, the salt doesn't get hot enough to evaporate, so it is left behind.



Figure 4.10.6: "Great Salt Lake" (CC BY 2.0; John Morgan via Flickr)

Example 4.10.4

Suppose you have a mixture of salt and pepper. What properties of the salt and pepper might allow you to separate them?

Solution

Salt dissolves in water but pepper does not. If you mix salt and pepper with water, only the salt will dissolve, leaving the pepper floating in the water. You can separate the pepper from the water by pouring the mixture through a filter, such as a coffee filter.

Example 4.10.5

After you separate the pepper from the salt water, how could you separate the salt from the water?

Solution

You could heat the water until it boils and evaporates. The salt would be left behind.

Summary

- A mixture is a combination of two or more substances in any proportions. The substances in a mixture do not combine chemically, so they retain their physical properties.
- A homogeneous mixture has the same composition throughout. A heterogeneous mixture varies in its composition.
- Mixtures can be classified on the basis of particle size into three different types: solutions, suspensions, and colloids.
- The components of a mixture retain their own physical properties. These properties can be used to separate the components by filtering, boiling, or other physical processes.

Review

1. What is a mixture?
2. What is the difference between a homogeneous and a heterogeneous mixture?

3. Make a table to compare and contrast solutions, colloids, and suspensions. Include an example of each type of mixture in your table.
4. Iron filings are attracted by a magnet. This is a physical property of iron but not of most other materials, including sand. How could you use this difference in physical properties to separate a mixture of iron filings and sand?

This page titled 4.10: Mixture is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

4.11: Heterogeneous Mixtures

What is the best way to eat a bag of jelly beans? Many people open the bag and eat all the candy, no matter what flavor each piece is. Others pick through the collection. They may say, "I don't like the orange ones." Or maybe they only care for the lemon ones. There are different kinds of jelly beans in the mixture and people will eat what they want and get rid of the rest.

Heterogeneous Mixtures

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.

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.



Figure 4.11.1: Oil and water do not mix, instead forming two distinct layers called phases. The oil phase is less dense than the water phase, and so the oil floats on top of the water. (CC BY 2.0; Flickr: Yortw via Flickr)

In the case of vegetable soup, one phase would be the liquid soup itself. This phase has vitamins, minerals, and other components dissolved in the water. This phase would be homogeneous. The carrots, peas, corn, or other vegetables represent other phases of the soup.The various vegetables are not mixed evenly in the soup, but are spread around at random.

There are a large number of heterogeneous mixtures around us. Soil is composed of a variety of substances, and is often of different composition depending on the sample taken. One shovelful may come up with dirt and grass, while the next shovelful could contain an earthworm.

Smog is another example of a heterogeneous mixture. This murky collection of pollutants can be a mixture of water and contaminants from burning gasoline or plastics, mixed with nitric oxide derivatives and ozone. You can see that the smog distribution in the air illustrated below is not evenly spread out, but varies from one part of the atmosphere to another.



Figure 4.11.2: Smog in New York City. (Public Domain; Dr. Edwin P. Ewing, Jr./CDC via Wikipedia)

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Simulation

Whether the air is clean or smoggy, it's a collection of molecules. See what it's made of in this simulation.

Summary

- A heterogeneous mixture is a mixture in which the composition is not uniform throughout the mixture.
- A phase is a separate layer in a heterogeneous mixture.

Review

Use the link below to answer the following questions:

1. Define a heterogeneous mixture.
2. Why is vegetable soup a heterogeneous mixture?
3. How many phases are in a heterogeneous mixture?

This page titled 4.11: Heterogeneous Mixtures is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.9: Heterogeneous Mixtures by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.12: Homogeneous Mixture

How do you like your coffee?

Many people enjoy a cup of coffee at some point during the day. Some may drink it black, while others may put cream (or some dairy substitute) and sugar in their coffee. High-end coffee drinks can be purchased at espresso stands (either sit-down or drive-through). Whatever the coffee drinker's preference, they want the coffee to be the same at the beginning as it is at the end of the drink. They don't want the components to separate out, but want their drink to be uniform from top to bottom.

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 or large amount of salt 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.

YouTube Video

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

View this YouTube Video online

Homogenous Mixtures

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.

One characteristic of mixtures is that they can be separated into their components. Since each part of the mixture has not reacted with another part of the mixture, the identities of the different materials are unchanged.

YouTube Video

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

View this YouTube Video online

Simulation

Take a deep breath...you just breathed in a mixture. This simulation will help you explore the mixture that surrounds all of us.

Summary

- A homogeneous mixture is a mixture in which the composition is uniform throughout the mixture.
- All solutions would be considered homogeneous.

Review

1. What is a mixture?
2. What is a homogeneous mixture?
3. Are all solutions homogeneous mixtures?
4. Can homogeneous mixtures be separated into their components?

This page titled 4.12: Homogeneous Mixture is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.8: Homogeneous Mixture by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.13: Separating Mixtures



Figure 4.13.1 (CC by 2.0; Marcin Chady via Flicker photo, Marcin Chady [www.flickr.com])

How did gold miners search for gold?

Beginning in the late 1840s, thousands of prospectors rushed to California in search of gold. One of the approaches taken to isolate the gold from the soil was called "panning". Dirt would be placed in the pan and covered with water. After thorough mixing, the pan is gently swirled to remove dissolved material while the heavier gold settles to the bottom of the pan. The gold is then separated from the mixture of soil and water.

Separation of Mixtures

Not everyone is out searching for gold (and not many of those searchers is going to get much gold, either). In a chemical reaction, it is important to isolate the component(s) of interest from all the other materials so they can be further characterized. Studies of biochemical systems, environmental analysis, pharmaceutical research – these and many other areas of research require reliable separation methods.

Here are a number of common separation techniques:

Chromatography

Chromatography is the separation of a mixture by passing it in solution or suspension, or as a vapor (as in gas chromatography), through a medium in which the components move at different rates. Thin-layer chromatography is a special type of chromatography used for separating and identifying mixtures that are or can be colored, especially pigments.

Distillation

Distillation is an effective method to separate mixtures that are comprised of two or more pure liquids. Distillation is a purification process where the components of a liquid mixture are vaporized and then condensed and isolated. In simple distillation, a mixture is heated, and the most volatile component vaporizes at the lowest temperature. The vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state. The condensate that is collected is called distillate.

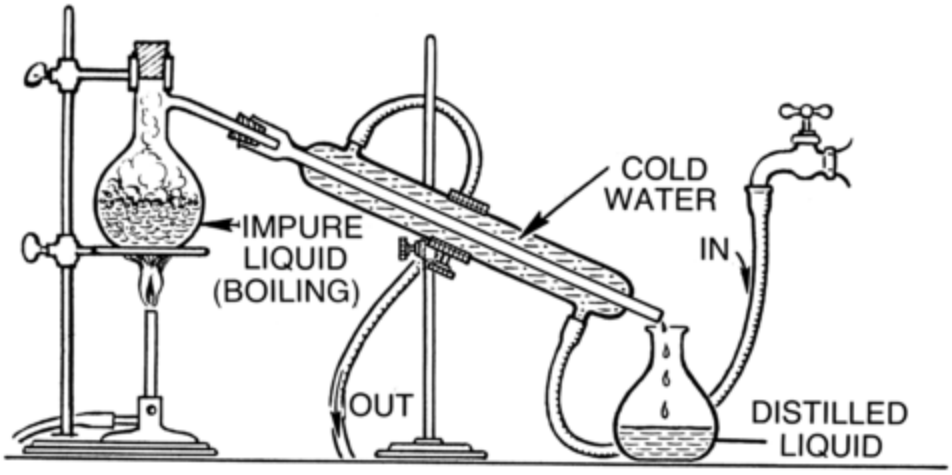


Figure 4.13.2: Distillation apparatus. (Public Domain; Pearson Scott Foresman via Commons Wikimedia, Distillation Apparatus [commons.wikimedia.org])

In the figure above, we see several important pieces of equipment. There is a heat source, a test tube with a one-hole stopper attached to a glass elbow, and rubber tubing. The rubber tubing is placed into a collection tube, which is submerged in cold water. There are other more complicated assemblies for distillation that can also be used, especially to separate mixtures which are comprised of pure liquids, with boiling points that are close to one another.

Evaporation

Evaporation is a technique used to separate out homogeneous mixtures that contain one or more dissolved salts. The method drives off the liquid components from the solid components. The process typically involves heating the mixture until no more liquid remains. Prior to using this method, the mixture should only contain one liquid component, unless it is not important to isolate the liquid components. This is because all liquid components will evaporate over time. The evaporation method is suitable to separate a soluble solid from a liquid.

In many parts of the world, table salt is obtained from the evaporation of sea water. The heat for the process comes from the sun.



Figure 4.13.3: Once the sea water in these evaporation ponds has evaporated, the salt can be harvested. (CC BY 3.0; Sandip Dey (Wikimedia: Dey.sandip) via Commons Wikimedia, Salt pans in Marakkanam, India [commons.wikimedia.org])

Filtration

Filtration is a separation method used to separate out pure substances in mixtures comprised of particles—some of which are large enough in size to be captured with a porous material. Particle size can vary considerably, given the type of mixture. For instance, stream water is a mixture that contains naturally occurring biological organisms like bacteria, viruses, and protozoa. Some water filters can filter out bacteria, the length of which are on the order of 1 micron. Other mixtures, like soil, have relatively large particle sizes, which can be filtered through something like a coffee filter.

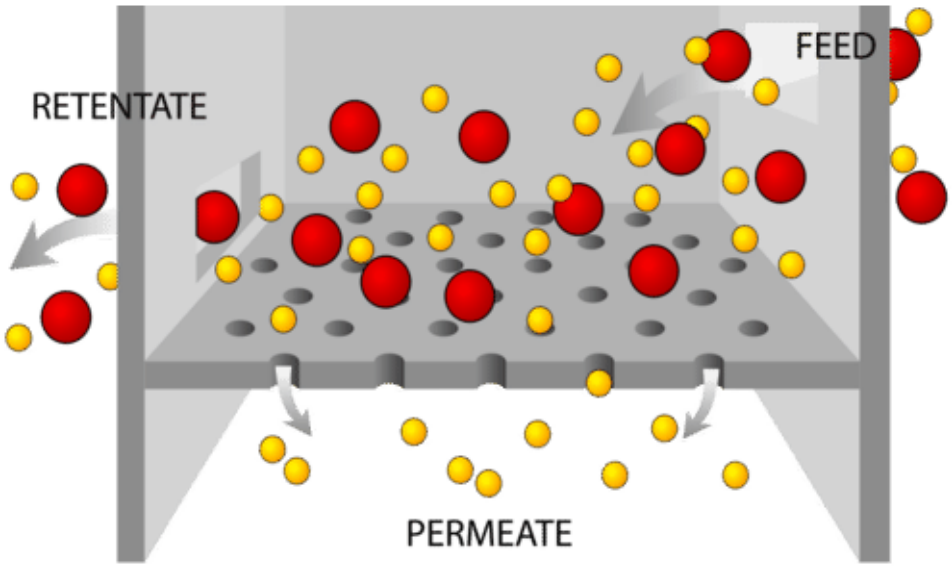


Figure 4.13.3: Filtration. (Public Domain; Mariana Ruiz Villarreal (Wikimedia: LadyofHats) via Commons Wikimedia, Filtration Diagram [commons.wikimedia.org])

YouTube Video

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

View this YouTube Video online

Summary

- Mixtures can be separated using a variety of techniques.
- Chromatography involves solvent separation on a solid medium.
- Distillation takes advantage of differences in boiling points.
- Evaporation removes a liquid from a solution to leave a solid material.
- Filtration separates solids of different sizes.

Review

1. Why is it important to separate materials from a mixture?
2. What is chromatography?
3. What is distillation
4. What is filtration?
5. What is evaporation?
6. What technique would you use to separate sand from water? There are two possibilities.

7. What technique would you use to separate alcohol from water?

This page titled 4.13: Separating Mixtures is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.10: Separating Mixtures by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.14: Elements



Figure 4.14.1 (Public Domain; Sidney Paget via Wikipedia)

Who is Sherlock Holmes?

The famous fictional British detective Sherlock Holmes supposedly often made the statement "elementary, my dear Watson". In reality, the closest he ever came to that line was an exchange with Watson in the short story "The Crooked Man". Holmes demonstrated shrewd insight into Watson's activities of the day. When asked how he knew what Watson was doing, Holmes simply replies "Elementary". Regardless of exactly how stated it, Sherlock Holmes was simply referring to what the Free Dictionary defines as "relating to, or constituting the basic, essential, or fundamental part".

Elements

An **element** is the simplest form of matter that has a unique set of properties. It is a pure substance that cannot be broken down into different types of substances. As of 2010, there are 118 known elements. Ninety eight of the elements are naturally occurring on earth, while the other 20 elements have been made synthetically in a lab. The names of all the elements can be found on the periodic table.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

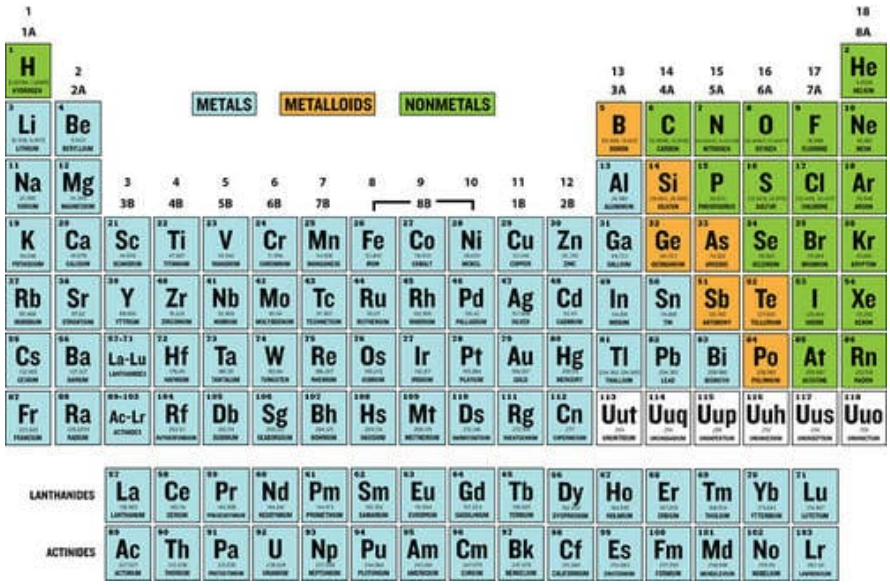


Figure 4.14.2: The Periodic Table. (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Examples of well-known elements include oxygen, iron, and gold (see the figure below). Elements cannot be broken down into a simpler substance. Likewise, one element cannot be chemically converted into a different element.



Figure 4.14.3: Chemical elements are the simplest of substances. (A) An oxygen tank like this is used by people who need breathing assistance. (B) A simple skillet can be made from cast iron. (C) Gold bars are formed and used for monetary purposes. (Credit: (A) Leon Brocard (Flickr: acme); (B) User:Evan-Amos/Wikimedia Commons; (C) Agnico-Eagle Mines Limited; Source: (A) http://www.flickr.com/photos/acme/8532536; (B) http://commons.wikimedia.org/wiki/File:Cast-Iron-Pan.jpg; (C) http://commons.wikimedia.org/wiki/File:Gold_Bars.jpg; http://commons.wikimedia.org/wiki/File:Coypel_Democritus.jpg License: (A) CC-BY 2.0; (B) Public Domain; (C) Public Domain)

Some elements have been known for centuries (gold, silver, iron, and copper, for example), while others have been created in a lab only within the last several years. Most elements do not exist as such in nature. They are so reactive that they can be found only in combination with other materials.

Some elements are very valuable, while others are quite inexpensive. Gold is currently worth almost $1700 per ounce. Aluminum, on the other hand, only sells for about 90 cents per pound, considerably lower than gold. Copper is worth somewhat more, selling for approximately $3.50 per pound. Platinum is very valuable at about $1650 an ounce, though not quite as expensive as gold.

Simulation

There are elements all around you, including in the air you breathe. See what elements are in the air in this simulation.

Video: The Chemical Elements Song

The Chemical Elements song, sung by Tom Lehrer. At the time, there were only 102 elements discovered.

YouTube Video

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

View this YouTube Video online

Summary

- An element is the simplest form of matter that has a unique set of properties.
- One element cannot be chemically converted to another element.

Review

1. What is an element?
2. Give the names of any two elements.
3. What is the name of the table that has the names of the elements?
4. How many elements are naturally occurring?

This page titled 4.14: Elements is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.11: Elements by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.15: Compounds



Figure 4.15.1 (Public Domain; User:JohnABerring27A via Wikipedia)

What do you need to start building a house?

Houses take on a variety of shapes. Some are large mansions while others are simpler dwellings. No matter what the house looks like there will be certain common components – wood, brick, plaster, nails, plumbing, wiring, etc. These components go together in numerous ways to make up the unique final product.

Compounds in chemistry are composed of different types and numbers of atoms. But atoms are common building blocks for a wide variety of compounds, just as brick, wood, and other materials are common building blocks for many types of houses.

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).

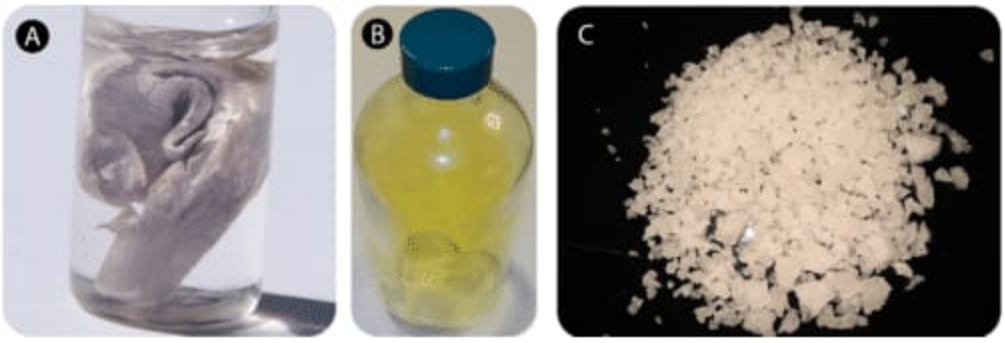


Figure 4.15.1: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine. (Credit: (A) Hi-Res Images of Chemical Elements; (B) User: Greenhorn1/Wikimedia Commons; (C) User:Miansari66/Wikimedia Commons; Source: (A) http://images-of-elements.com/sodium.php; (B) http://commons.wikimedia.org/wiki/File:Chlorine2.jpg; (C) http://commons.wikimedia.org/wiki/File:Salt_,_Lake_salt.JPG License: (A) CC by 3.0 (B) Public Domain; (C) Public Domain)

Science Friday: Celebrating Explosive Chemistry

July Fourth: A day for picnics, parades and chemistry. In this video by Science Friday, Professor Bassam Shakhashiri explains some of the science of fireworks.

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

View this YouTube Video online

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.

Review

1. What is a compound?
2. How is a compound different from an element?
3. What is a chemical change?

Explore More

Use the resource below to answer the following questions.

YouTube Video

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

View this YouTube Video online

1. What lists all the elements we know about?
2. What compound is made of two hydrogen atoms and one oxygen atom?
3. What compound is made of one carbon atom and two oxygen atoms?

This page titled 4.15: Compounds is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.12: Compounds by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.16: Chemical Symbols and Formulas

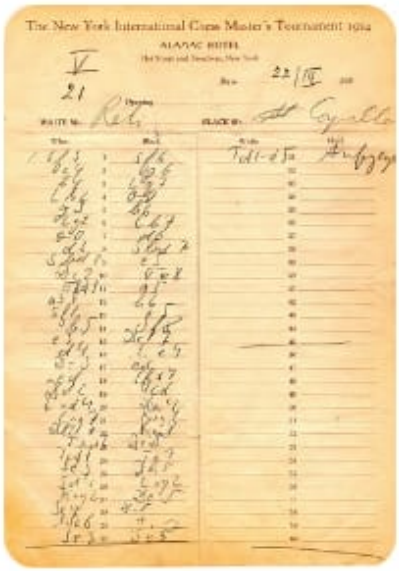


Figure 4.16.1 (Author Unknown; Source: Commons Wikimedia, Capablanca [commons.wikimedia.org]; License: Public Domain)

How do chess players monitor their moves in a game?

Suppose you were walking along and noticed a piece of paper on the ground with markings on it. You pick it up and see the paper in the picture above.

To most people, these notes are meaningless (maybe they're a secret spy code). But to a chess player, these symbols tell the story of a chess game. Each abbreviation describes a chess piece or a move during the game. The use of special symbols allows chess players to "see" the game without having to read a wordy and possibly incomplete description of what happened.

Chemical Symbols and Formulas

Chemical Formula

In order to illustrate chemical reactions and the elements and **compounds** involved in them, chemists use symbols and formulas. 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 below.

Table 1: 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

Chemical Symbol	Name	Latin Name
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Watch the following video to learn more about the origin of different element's symbols.

YouTube Video

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

View this YouTube Video online

Chemical Formulas

Compounds are combinations of two or more elements. 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 two to one ratio. The chemical formula for water is $\mathrm{H_2O}$. 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 $\mathrm{H_2SO_4}$.

Summary

- A chemical symbol is a one- or two-letter designation of an element.
- Compounds are combinations of two or more elements.
- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- Some elements have symbols that derive from the Latin name for the element.

Review

1. What is a chemical symbol?
2. What is a chemical formula?
3. How many hydrogen atoms are in one molecule of the compound $\mathrm{C_{12}H_{22}O_{11}}$?
4. What is the Latin name for the element potassium?

This page titled 4.16: Chemical Symbols and Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 2.15: Chemical Symbols and Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

4.17: A Molecular View of Elements and Compounds

Learning Objectives

- Classify substances as atomic elements, molecular elements, molecular compounds, or ionic compounds.

Atomic Elements

Most elements exist with **individual atoms** as their basic unit. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

Molecular Elements

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

Table 4.17.1*: 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 4.17.1). 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 4.17.1).

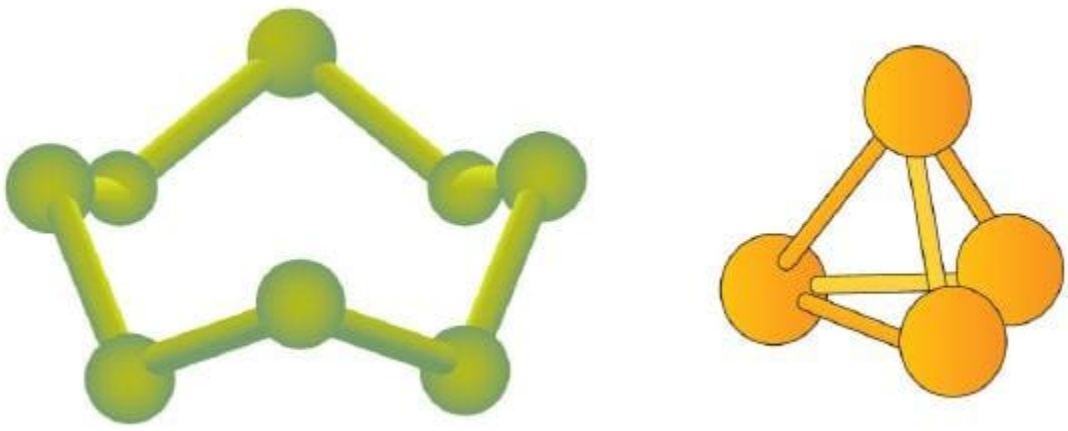


Figure 4.17.1: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 4.17.1 shows two examples of how molecules will be represented in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

CK-12 LICENSED UNDER CK-12 Science
© CK-12 Foundation • Visit us at ck12.org

Ionic Compounds

The elements in the periodic table are divided into specific groupings; the metals, the non-metals, the semi-metals, and so on. These groupings are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). The basic unit of ionic compounds is the **formula unit**.

Molecular Compounds

Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds. Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals) and silicon dioxide (SiO_2) will also be a covalently bound molecule (a semi-metal and a non-metal). The basic unit of molecular compounds is the **molecule**.

Example 4.17.1

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 4.17.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

a. I_2
b. He
c. H_2O
d. Al
e. CuCl

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

Answer e: available at the end of this section.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

Answers

Exercise 4.17.1
molecular element

Exercise 4.17.1
atomic element

Exercise 4.17.1
molecular compound

Exercise 4.17.1
atomic element

Exercise 4.17.1
ionic compound

This page titled 4.17: A Molecular View of Elements and Compounds is shared under a CK-12 license and was authored, remixed, and/or curated by Henry Agnew & Marisa Alviar-Agnew.

- 5.4: A Molecular View of Elements and Compounds by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

5: ATOMS, MOLECULES, AND SUBATOMIC PARTICLES

5.1: Atom

5.2: Protons

5.3: Neutrons

5.4: Electrons

5.5: Atomic Nucleus

5.6: Atomic Number

5.7: Mass Number

5.8: Isotopes

5.9: Atomic Mass Unit

5.10: Calculating Average Atomic Mass

5.11: Ion

This page titled 5: Atoms, Molecules, and Subatomic Particles is shared under a license and was authored, remixed, and/or curated by LibreTexts.

5.1: Atom

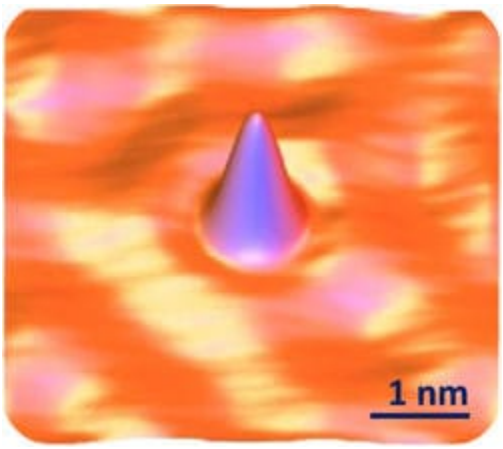


Figure 5.1.1 (Public Domain; NIST, Joseph Stroscio et. al. via Commons Wikimedia, STM Image of Single Co Atom [commons.wikimedia.org])

What could this hilly blue surface possibly be? Do you have any idea? The answer is a single atom of the element cobalt. The picture was created using a scanning tunneling microscope. No other microscope can make images of things as small as atoms. How small are atoms? You will find out in this lesson.

What Are Atoms?

Atoms are the building blocks of matter. They are the smallest particles of an element that still have the element's properties. Elements, in turn, are pure substances—such as nickel, hydrogen, and helium—that make up all kinds of matter. All the atoms of a given element are identical in that they have the same number of protons, one of the building blocks of atoms (see below). They are also different from the atoms of all other elements, as atoms of different elements have different number of protons.

Size of Atoms

Unlike bricks, atoms are extremely small. The radius of an atom is well under 1 nanometer, which is one-billionth of a meter. If a size that small is hard to imagine, consider this: trillions of atoms would fit inside the period at the end of this sentence. Although all atoms are very small, elements vary in the size of their atoms. The figure below compares the sizes of atoms of more than 40 different elements. The elements in the figure are represented by chemical symbols, such as H for hydrogen and He for helium. Of course, real atoms are much smaller than the circles representing them in the figure below.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

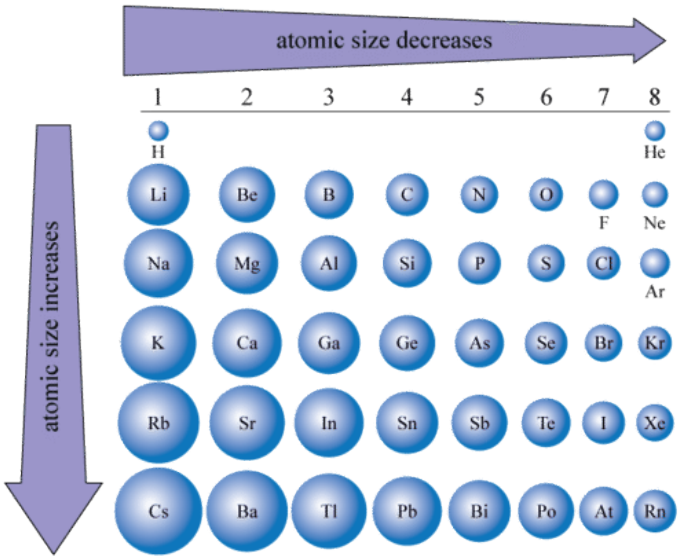


Figure 5.1.2 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: Which element in the figure above has the biggest atoms?

A: The element in the figure with the biggest atoms is cesium (Cs).

Subatomic Particles

Although atoms are very tiny, they consist of even smaller particles. Three main types of particles that make up all atoms are:

- protons, which have a positive electric charge.
- electrons, which have a negative electric charge.
- neutrons, which are neutral in electric charge.

The model in the figure below shows how these particles are arranged in an atom. The particular atom represented by the model is helium, but the particles of all atoms are arranged in the same way. At the center of the atom is a dense area called the nucleus, where all the protons and neutrons are clustered closely together. The electrons constantly move around the nucleus. Helium has two protons and two neutrons in its nucleus and two electrons moving around the nucleus. Atoms of other elements have different numbers of subatomic particles, but the number of protons always equals the number of electrons. This makes atoms neutral in charge because the positive and negative charges "cancel out."

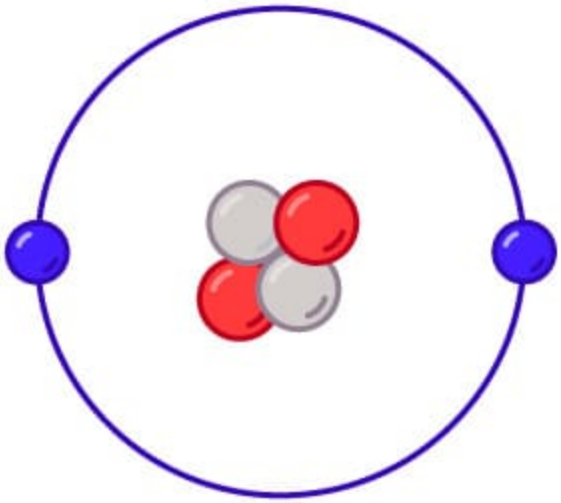


Figure 5.1.3: Model of a helium atom. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Q: Lithium has three protons, four neutrons, and three electrons. Sketch a model of a lithium atom, similar to the model above for helium.

A: Does your sketch resemble the model in the figure below? The model has three protons (blue) and four neutrons (gray) in the nucleus, with three electrons (red) moving around the nucleus.

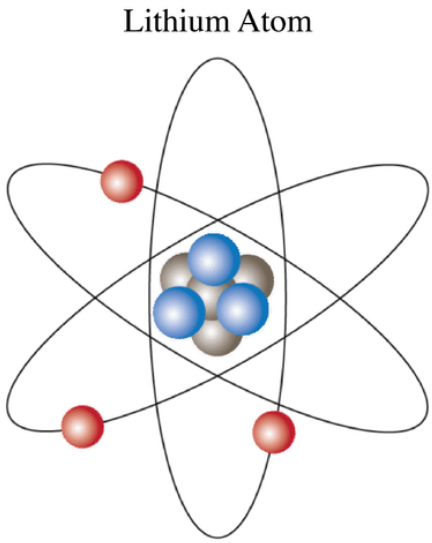


Figure 5.1.4: Copy and Paste Caption here. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: All atoms of carbon have six protons. How many electrons do carbon atoms have?

A: Carbon atoms must have six electrons to "cancel out" the positive charges of the six protons. That's because atoms are always neutral in electric charge.

YouTube Video

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

View this YouTube Video online

Summary

- Atoms are the building blocks of matter. They are the smallest particles of an element that still have the element's properties.
- All atoms are very small, but atoms of different elements vary in size.
- Three main types of particles that make up all atoms are protons, neutrons, and electrons.

Review

1. What is an atom?
2. Which of the following statement(s) are true about the atoms of any element?
   a. The number of protons in an atom of an element is unique to each element.
   b. The number of protons and neutrons in an atom of an element is unique to each element
   c. A proton in an atom of one element is identical to a proton in an atom of another element.
   d. The number of protons in an atom of an element is the same for all elements.
3. Which of the following statements explains why atoms are always neutral in charge
   a. They have the same number of protons as the atoms of all other elements.
   b. They have protons that are identical to the protons of all other elements.
   c. They have the same size as the atoms of all other elements.
   d. They have the same number of protons as electrons.

This page titled 5.1: Atom is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

5.2: Protons



Figure 5.2.1 (Public Domain; Adrian Pingstone (User:Arpingstone/Wikimedia Commons) via Wikipedia)

Can you name this car?

Describing what we can see is a fairly easy matter. If we are asked to describe the sports car illustrated below, we could all quickly come up with a fairly accurate description. A person knowledgeable about cars would include more details, but everyone would have the basic information in their description.

What makes the description easy to come up with? We can see it, we have a common language to describe it (size, color, construction), and we have a basic idea of what it is (a car, not a house or a tree). Scientists have far more difficulty in describing things they cannot see. There is no way to look directly at an atom and see its detailed structure. When a discovery is first made, there is often no language to use to tell others exactly what it is. This was the problem in talking about the atom and its structure.

Putting the Puzzle Pieces Together

Research builds upon itself – one piece connects to another. Sometimes the puzzle doesn't seem to make sense because some of the pieces are missing at the moment. Each finding gives a clearer picture of the whole and also raises new questions. The detective work that led to the discovery of the proton was built upon finding pieces to the puzzle and putting them together in the right way.

The electron was discovered using a cathode ray tube. An electric current was passed from the cathode (the negative pole) to the anode (positive pole). Several experiments showed that particles were emitted at the cathode and that these particles had a negative charge. These experiments demonstrated the presence of electrons.

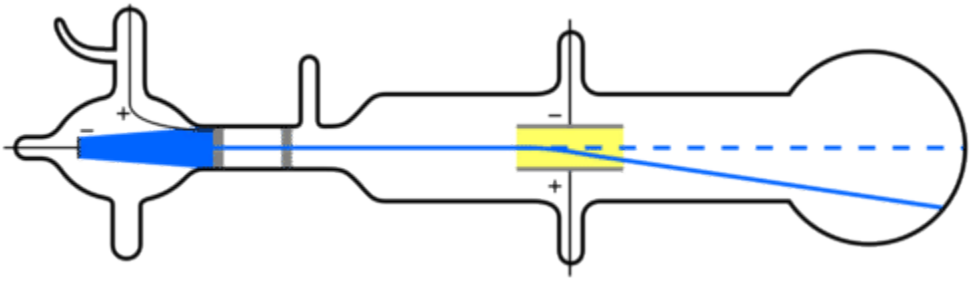


Figure 5.2.2: JJ Thomson's experiment with cathode rays. (Public Domain; User:Kurzon/Wikimedia Commons via Wikipedia)

If cathode rays are electrons that are given off by the metal atoms of the cathode, then what remains of the atoms that have lost those electrons? We know several basic things about electrical charges. They are carried by particles of matter. Millikan's experiment showed that they exist as whole-number multiples of a single basic unit. Atoms have no overall electrical charge, meaning that each and every atom contains an exactly equal number of positively and negatively charged particles. A hydrogen atom is the simplest kind of atom with only one electron. When that electron is removed, a positively charged particle should remain.

Discovery of the Proton

In 1886, Eugene Goldstein (1850-1930) discovered evidence for the existence of this positively charged particle. Using a cathode ray tube with holes in the cathode, he noticed that there were rays traveling in the opposite direction from the cathode rays. He called these **canal rays** and showed that they were composed of positively charged particles. The **proton** is the positively charged subatomic particle present in all atoms. The mass of the proton is about 1840 times the mass of the electron.

YouTube Video

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

View this YouTube Video online

Summary

- When an electron is removed from a hydrogen atom, a proton remains.
- Goldstein observed rays traveling in the opposite direction of the cathode rays in a cathode ray tube.
- He demonstrated that these rays were positive particles and called the canal rays.

Review

1. Why is it easy to describe things we can see?
2. Why did researchers believe that the particle left after electrons were emitted as cathode rays had to be positive?
3. Atoms, which are always neutral in electric charge, contain electrons as well as protons and neutrons. An electron has an electrical charge of -1. If an atom has three electrons, infer how many protons it has.
4. How many electrons does it take to weight the same as one proton?

This page titled 5.2: Protons is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

5.3: Neutrons



Figure 5.3.1 (Public Domain; Sidney Paget via Wikipedia)

Was Sherlock Holmes real or the product of someone's imagination?

The most famous detective in history and literature never existed. Sherlock Holmes was the creation of the British author Sir Arthur Conan Doyle. This mythical person had capabilities far beyond those of mere mortals. Holmes was capable of spotting the tiniest clue, the smallest piece of evidence to solve the crime. He could link all sorts of seemingly irrelevant data into a coherent whole to clear up whatever mystery he was dealing with.

The Quest for the Neutron

Clues are generally considered to involve the presence of something – a footprint, a piece of fabric, a bloodstain, something tangible that we can measure directly. The discoveries of the electron and the proton were accomplished with the help of those kinds of clues. Cathode ray tube experiments showed both the negatively charged electrons emitted by the cathode and the positively charged proton (also emitted by the cathode). The **neutron** was initially found not by a direct observation, but by noting what was not found.

Research had shown the properties of the electron and the proton. Scientists learned that approximately 1,837 electrons weighed the same as one proton. There was evidence to suggest that electrons went around the heavy nucleus composed of protons. Charge was balanced with equal numbers of electrons and protons which made up an electrically neutral atom. But there was a problem with this model – the atomic number (number of protons) did not match the atomic weight. In fact, the atomic number was usually about half the atomic weight. This indicated that something else must be present. That something must weigh about the same as a proton, but could not have a charge – this new particle had to be electrically neutral.

In 1920, Ernest Rutherford tried to explain this phenomenon through the presence of another particle in the nucleus. He proposed that the "extra" particles were protons and electrons that had combined into a new particle in the nucleus (this did not turn out to be the case). These new particles would have a mass very similar to a proton, but would be electrically neutral since the positive charge of the proton and the negative charge of the electron would cancel each other out.

In 1930, German researchers bombarded the element beryllium with alpha particles (helium nuclei containing two protons and two neutrons with a charge of +2). The particles produced in this process had strong penetrating power, which suggested they were fairly large. In addition, they were not affected by a magnetic field, so they were electrically neutral. The French husband-wife research team of Frederic and Irene Joliot-Curie used these new "rays" to bombard paraffin, which was rich in protons. The unknown particles produced a large emission of protons from the paraffin.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

The English physicist James Chadwick (1891-1974) repeated these experiments and studied the energy of these particles. By measuring velocities, he was able to show that the new particle has essentially the same mass as a proton. So we now have a third subatomic particle with a mass equal to that of a proton, but with no charge. This particle is called the neutron. Chadwick won the Nobel Prize in Physics in 1935 for his research.

YouTube Video

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

View this YouTube Video online

Neutron Applications

Neutrons can be used in a variety of ways. One important use is in nuclear **fission** to produce new isotopes. A neutron will collide with a large atom (such as uranium) and cause it to split into smaller atoms, such as in figure below.

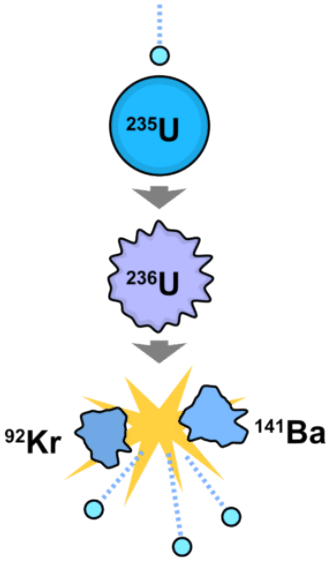


Figure 5.3.2: A neutron collides with a large atom, splitting it into smaller atoms. (Public Domain; User:Fastfission/Wikimedia Commons via Wikipedia)

Nuclear reactors utilize **chain reactions** involving neutrons to heat water which drive turbines for the generation of electricity. When a neutron collides with a large atom, the atom splits with the release of more neutrons and also a large amount of energy. The energy converts water to steam for the turbine, while the neutrons serve to continue the chain reaction (see figure below).

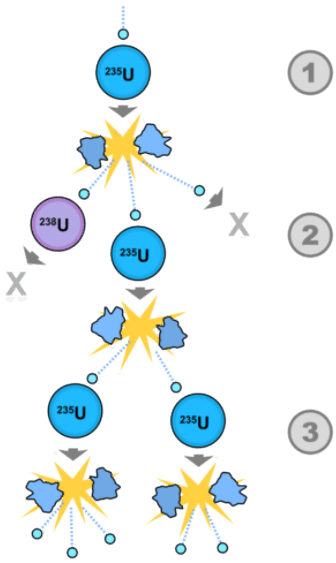


Figure 5.3.3: How nuclear fission produces new isotopes. (Public Domain; User: Fastfission/Wikimedia Commons via Wikipedia)

Summary

- Rutherford proposed that "extra" particles in nucleus were combinations of protons and electrons.
- Bombardment of beryllium with alpha particles produced large, neutral particles.
- Chadwick determined the mass of the neutron.
- Nuclear fission produces new elements.
- Nuclear reactors use chain reactions to produce heat.

Review

1. How did Rutherford try to explain the differences between the number of protons in the nucleus and the atomic weight?
2. What did German researchers find when they bombarded beryllium with alpha particles?
3. What did Chadwick determine about these new particle (observed by the German scientist and the Curies)?

This page titled 5.3: Neutrons is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

5.4: Electrons

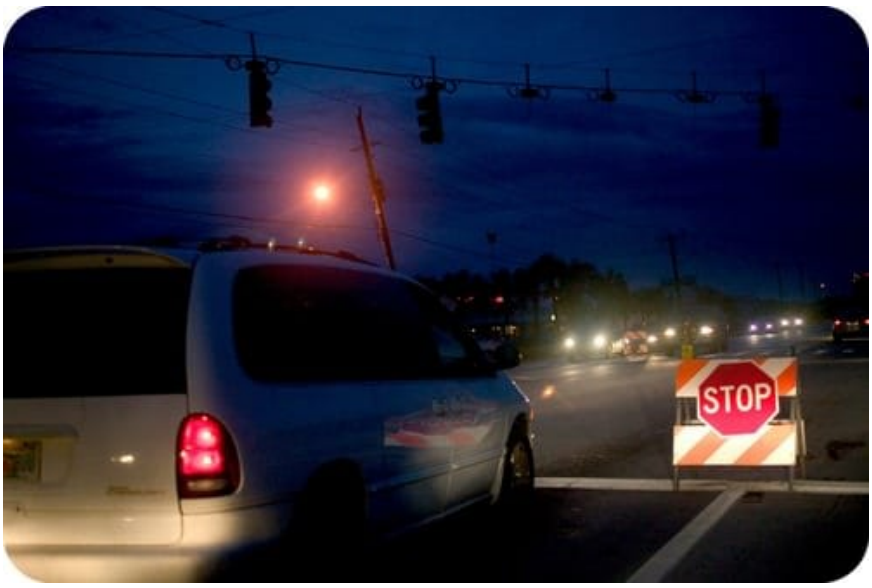


Figure 5.4.1 (Public Domain; Andrea Booher, FEMA via Wikipedia)

What causes a power outage?

In a power outage, all your electrical equipment suddenly stops working. The radio was on just a minute ago and now it is silent. What happened? Somewhere between a power generator and your electrical device was an interruption. Power stopped flowing through the wires and into your radio. That “power” turns out to be electrons that move through the wires and cause an electrical current to flow.

Is There Anything Inside an Atom?

As the nineteenth century began to draw to a close, the concept of atoms was well-established. We could determine the mass of different atoms and had some good ideas about the atomic composition of many compounds. Dalton’s atomic theory held that atoms were indivisible, so scientists did not ask questions about what was inside the atom – it was solid and could not be broken down further. But then things began to change.

The Electron

In 1897, English physicist J.J. Thomson (1856-1940) experimented with a device called a cathode ray tube, in which an electric current was passed through gases at low pressure. A cathode ray tube consists of a sealed glass tube fitted at both ends with metal disks called electrodes. The electrodes are then connected to a source of electricity. One electrode, called the anode, becomes positively charged while the other electrode, called the cathode, becomes negatively charged. A glowing beam (the cathode ray) travels from the cathode to the anode.

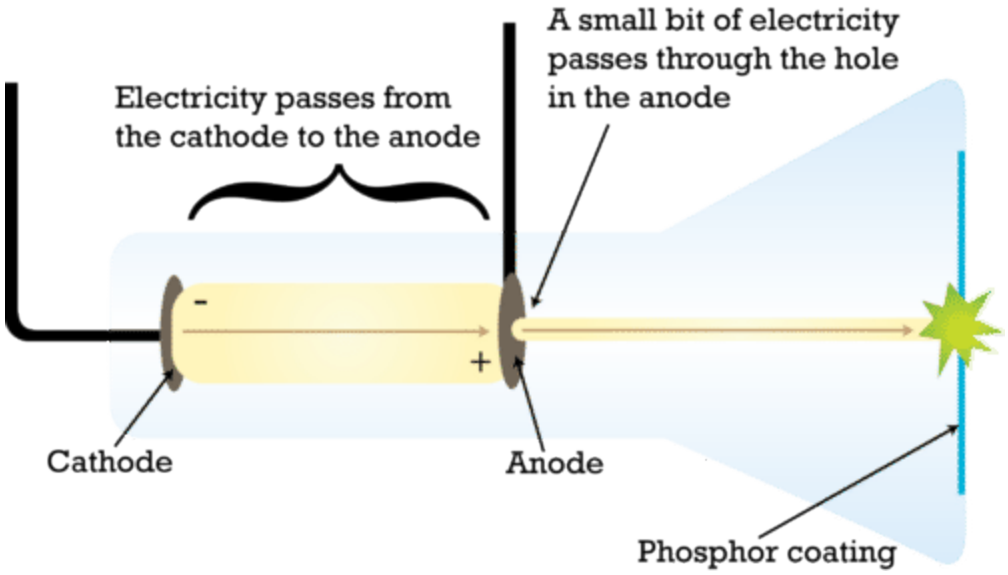


Figure 5.4.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Earlier investigations by Sir William Crookes and others had been carried out to determine the nature of the cathode ray. Thomson modified and extended these experiments in an effort to learn about these mysterious rays. He discovered two things, which supported the hypothesis that the cathode ray consisted of a stream of particles.

- When an object was placed between the cathode and the opposite end of the tube, it cast a shadow on the glass.
- A cathode ray tube was constructed with a small metal rail between the two electrodes. Attached to the rail was a paddle wheel capable of rotating along the rail. Upon starting up the cathode ray tube, the wheel rotated from the cathode towards the anode. This proved that the cathode ray was made of particles which must have mass. Crooke had first observed this phenomenon and attributed it to pressure by these particles on the wheel. Thomson correctly surmised that these particles were producing heat, which caused the wheel to turn.

In order to determine if the cathode ray consisted of charged particles, Thomson used magnets and charged plates to deflect the cathode ray. He observed that cathode rays were deflected by a **magnetic field** in the same manner as a wire carrying an electric current, which was known to be negatively charged. In addition, the cathode ray was deflected away from a negatively charged metal plate and towards a positively charged plate.

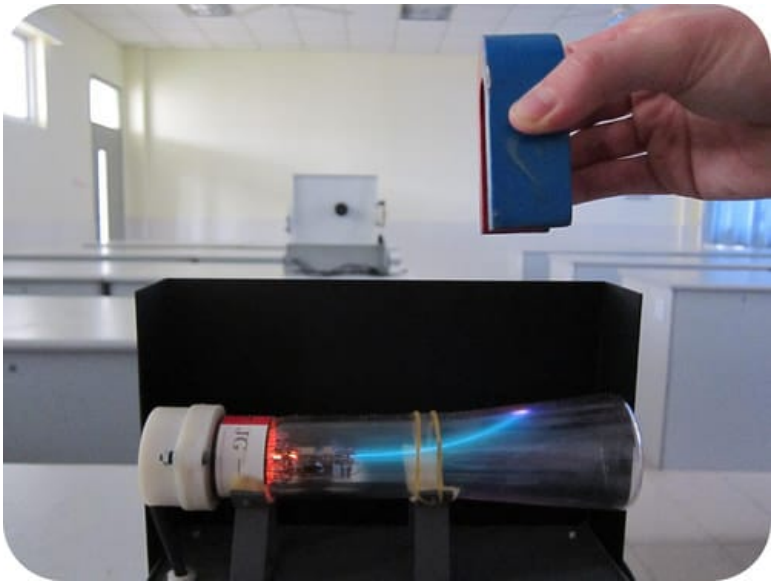


Figure 5.4.3 (CC BY 2.0; Micah Sittig via Flickr)

Thomson knew that opposite charges attract one another, while like charges repel one another. Together, the results of the

cathode ray tube experiments showed that cathode rays are actually streams of tiny negatively charged particles moving at very high speeds. While Thomson originally called these particles corpuscles, they were later named electrons.

Thomson conducted further experiments, which allowed him to calculate the charge-to-mass ratio $\frac{e}{m_e}$ of the electron. In units of coulombs to grams, this value is 1.8×10^8 Coulombs/gram. He found that this value was a constant and did not depend on the gas used in the cathode ray tube or on the metal used as the electrodes. He concluded that electrons were negatively charged subatomic particles present in atoms of all elements.

YouTube Video

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

View this YouTube Video online

Summary

- Cathode rays are deflected by a magnetic field.
- The rays are deflected away from a negatively charged electrical field and toward a positively charge field.
- The charge/mass ratio for the electron is 1.8×10^8 Coulombs/gram.

Review

1. What subatomic particle creates electric power, and how does it do it?
2. Whose work did Thomson repeat and revise?
3. What experiment did Thomson perform that showed cathode rays to be particles?
4. How did he show that these particles had a charge on them?
5. Did the cathode ray have positive or negative charge?

This page titled 5.4: Electrons is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

5.5: Atomic Nucleus



Figure 5.5.1 (CC by 3.0; Charles Hamm via Wikipedia)

How is science like a jigsaw puzzle?

Many people enjoy putting jigsaw puzzles together. As the different pieces go together, the picture begins to become clearer. When the puzzle is completed, you see that what had been a confused collection of individual components that made little or no sense by themselves fit together to give a clear picture.

Science works the same way as a jigsaw puzzle. Different researchers make individual discoveries that answer a specific question or questions. When enough data are gathered, we get a better understanding of a process or a structure. The experiments of Crooke, Millikan, Rutherford, and many others gave us pieces of the puzzle that is the atomic nucleus.

Different Models of the Nucleus

After the electron and proton were discovered, people began to try to build a picture of the atom. The Thomson model had electrons mixed in with a positive core of some sort that provided electrical neutrality. Rutherford showed that this model was incomplete. His picture of the atom involved a small solid core that alpha particles could zoom past with very few collisions. So the picture of the atom became a little clearer – electrons surrounded a very small core nucleus. The discovery of the neutron helped fill out the picture even more. We now have protons and neutrons in a concentrated center of the atom with electrons surrounding the nucleus.

One problem still existed. We have a number of positively charged protons in the nucleus. Why don't they push each other apart? Physicists postulate a strong nuclear force that acts at very short distances. At these distances the attraction between protons is greater than the force causing the positive charges to push each other away. Neutrons are also involved in this process somehow. So the strong force holds protons together, it holds neutrons together, and it causes protons and neutrons to be attracted to one another.

YouTube Video

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

View this YouTube Video online

Summary

- Rutherford proposed a model of the atomic nucleus which had a solid core.
- Physicists postulate a strong nuclear force holds the protons and neutrons together in the nucleus of the atom.

Review

1. How did Rutherford change our thinking about atomic structure?
2. What is our current picture of the atom?
3. Why is the presence of positively charged protons a problem with current models of the atom?
4. How do we explain why the nucleus does not fall apart?

This page titled 5.5: Atomic Nucleus is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.15: Atomic Nucleus by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.6: Atomic Number



Figure 5.6.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What is unique about each one of us?

For the vast majority of people, it is not their name, because it is quite possible for others in the world to have the same name (check it out by doing an internet search for your name and see how many other of "you" there are). It is not your physical description. Eye-witnesses to crime scenes often pick the wrong person when trying to identify the criminal.

There may be some unique identifiers for us. If you have a cell phone in your name, nobody else in the world has that number. Email addresses are different for each of us, which is a good thing since we can email almost anywhere in the world. Our DNA is unique, but getting a DNA analysis is expensive and time-consuming, so we really don't want to have to explore that.

Organizing the Elements

One of the goals of science is to discover the order in the universe and to organize information that reflects that order. As information about the different elements was made known, efforts were made to see if there were patterns in all of the data. An early attempt to organize data was made by Mendeleev, who developed the first periodic table. His data set was based on **atomic weights** and was instrumental in providing clues as to the possible identity of new elements. Once we learned the details of the atomic nucleus, the table was based on the number of protons in the nucleus, called the **atomic number** of the element.

Atomic Number

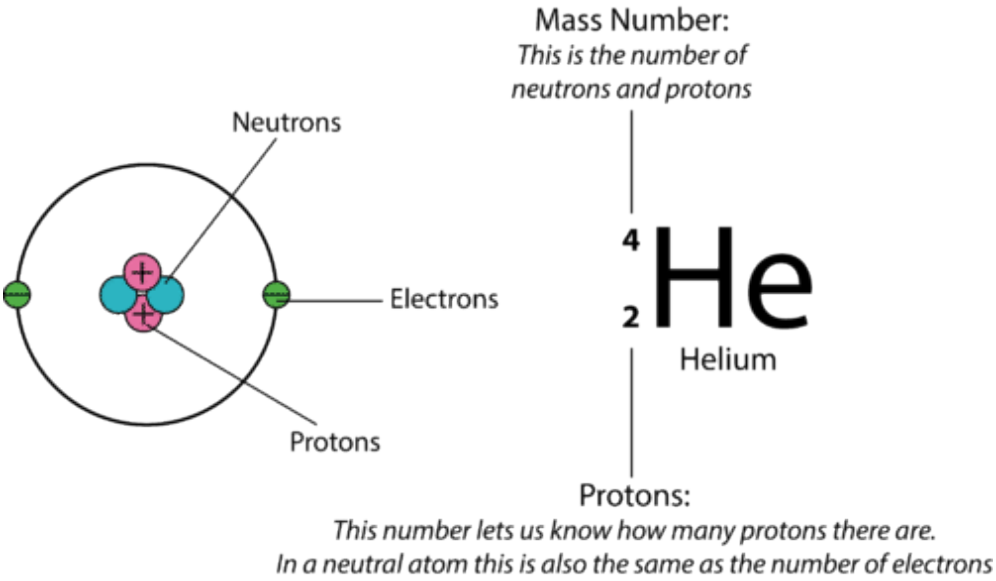


Figure 5.6.2: How can you determine the atomic number of an element. (Public Domain; User:Materialscientist/Wikimedia Commons via Wikipedia)

The atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element. This means that the number of protons is the characteristic which makes each element unique compared to all other elements. Elements are different because of their atomic number. The **periodic table** displays all of the known elements and is arranged in order of increasing atomic number. In this table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Following on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, and so forth.

Since atoms are neutral, the number of electrons is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus. Manganese (atomic number 25) would have twenty-five protons and twenty-five electrons.

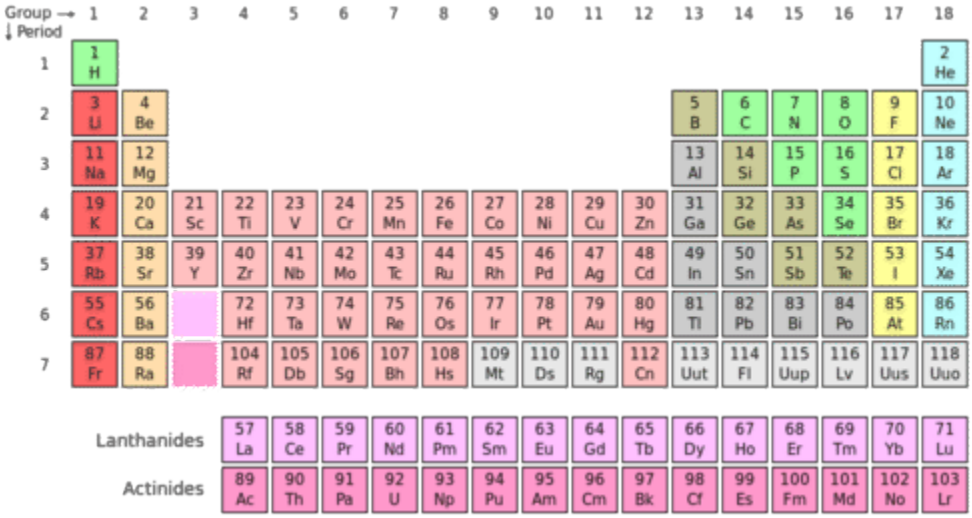


Figure 5.6.3: The periodic table classifies elements by atomic number. (Public Domain; User:Cepheus/Wikimedia Commons via Wikipedia)

The classification of elements by atomic number allows us to understand many properties of the atom and makes it possible to predict behaviors instead of just having to memorize everything.

YouTube Video

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

View this YouTube Video online

Summary

- The atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element
- The number of electrons is equal to the number of protons in an atom of an element.

Review

1. What is the atomic number of an atom? Why is this number important?
2. Using a periodic table, what is the atomic number of helium have
3. How many protons are in the following elements:?
   a. Ne
   b. Ca
   c. Pt
4. Write the symbol for the element with the following atomic number:
   a. 18
   b. 41
   c. 82
   d. 12

This page titled 5.6: Atomic Number is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.16: Atomic Number by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.7: Mass Number

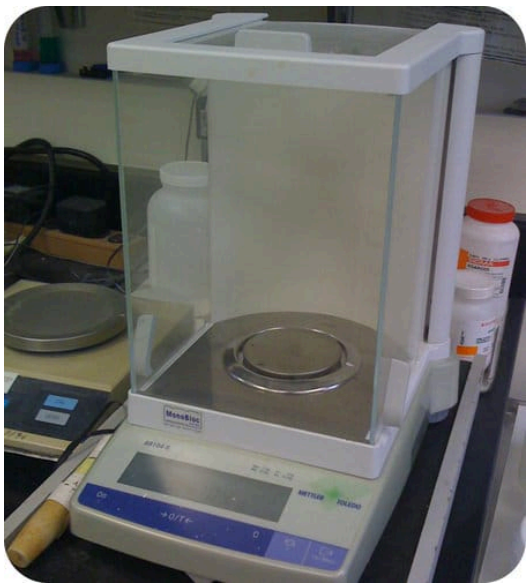


Figure 5.7.1 (Public Domain; User:Ebultoof/Wikipedia via Wikipedia)

How can you determine the mass of a chemical?

Often a student will need to weigh out a chemical for an experiment. If he or she uses a watch glass (a small, round piece that will hold the solid chemical), the weight of the watch glass must be determined first. Then the solid is added to the glass and the weight of the glass plus the solid is measured. The balance reading will be the total of the glass plus the chemical.

History of Atomic Weight Determinations

As a part of his research on atoms, John Dalton determined a number of atomic weights of elements in the early 1800s. Atomic weights were the basis for the periodic table that Mendeleev developed. Originally all atomic weights were based on a comparison to hydrogen, which has an atomic weight of one. After the discovery of the proton, scientists assumed that the weight of an atom was essentially that of the protons—electrons were known to contribute almost nothing to the atomic weight of the element.

This approach worked until we learned how to determine the number of protons in an element. We then saw that the atomic weight for an element was often twice the number of protons (or more). The discovery of the neutron provided the missing part of the picture. The atomic mass is now known to be the sum of the protons and neutrons in the nucleus.

Mass Number

Rutherford showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number** is defined as the total number of protons and neutrons in an atom. It can be calculated by adding the number of neutrons and the number of protons (atomic number) together.

Mass number = atomic number + number of neutrons

Consider Table 5.7.1 below that shows data from the first six elements of the periodic table.

Table 5.7.1: Atoms of the First Six Elements

Name	Symbol	Protons	Neutrons	Electrons	Atomic Number	Mass Number
Hydrogen	H	1	0	1	1	1

Name	Symbol	Protons	Neutrons	Electrons	Atomic Number	Mass Number
Helium	He	2	2	2	2	4
Lithium	Li	3	4	3	3	7
Beryllium	Be	4	5	4	4	9
Boron	B	5	6	5	5	11
Carbon	C	6	6	6	6	12

Consider the element helium. Its atomic number is 2, so it has two protons in its nucleus. Its nucleus also contains two neutrons. Since $2 + 2 = 4$, we know that the mass number of the helium atom is 4. Finally, the helium atom also contains two electrons since the number of electrons must equal the number of protons. This example may lead you to believe that atoms have the same number of protons and neutrons, but further examination of the table above will show that this is not the case. Lithium, for example, has three protons and four neutrons, leaving it with a mass number of 7.

Knowing the mass number and the atomic number of an atom allows you to determine the number of neutrons present in that atom by subtraction.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

Example 5.7.1

Atoms of the element chromium (Cr) have an atomic number of 24 and a mass number of 52. How many neutrons are in the nucleus of a chromium atom?

Solution

To determine this, you subtract the atomic number from the mass number, as shown above:

$$52 - 24 = 28 \text{ neutrons in a chromium atom}$$

The composition of any atom can be illustrated with a shorthand notation using the atomic number and the mass number. Both are written before the chemical symbol, with the mass number written as a superscript and the atomic number written as a subscript. The chromium atom discussed above would be written as:

$$^{52}_{24}\mathrm{Cr}$$

Another way to refer to a specific atom is to write the mass number of the atom after the name, separated by a hyphen. The above atom would be written as chromium-52.

YouTube Video

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

View this YouTube Video online

Summary

- The mass number is defined as the total number of protons and neutrons in an atom.
- The mass number = number of neutrons + atomic number.

Review

1. Who first determined atomic weights for elements?
2. What were the original atomic weights based on?
3. Why were calculations based on numbers of protons not valid for determining atomic weights?
4. A tin atom has an atomic number of 50 and a mass number of 118. How many neutrons are present in this atom?
5. What is the mass number of a cobalt atom that has 27 protons and 30 neutrons?

This page titled 5.7: Mass Number is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.17: Mass Number by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.8: Isotopes



Figure 5.8.1 (Public Domain; Chief Photographer's Mate Johnny Bivera, U.S. Navy via Wikipedia)

Are all the members of the football team shown above identical?

They are on the same team and are all known by the same team name, but there are individual differences among the players. We do not expect the kicker to be as big as the quarterback. The tight end is very likely to weigh less than the defensive tackle on the other side of the ball. They play as a unit, but they have different weights and heights.

What are Isotopes?

The history of the atom is full of some of these differences. Although John Dalton stated in his atomic theory of 1804 that all atoms of an element are identical, the discovery of the neutron began to show that this assumption was not correct. The study of radioactive materials (elements that spontaneously give off particles to form new elements) by Frederick Soddy (1877-1956) gave important clues about the internal structure of atoms. His work showed that some substances with different radioactive properties and different atomic masses were in fact the same element. He coined the term **isotope** from the Greek roots isos (ίσος "equal") and topos (τόπος "place"). He described isotopes as, "Put colloquially, their atoms have identical outsides but different insides." Soddy won the Nobel Prize in Chemistry in 1921 for his work.

As stated earlier, not all atoms of a given element are identical. Specifically, the number of neutrons can be variable for many elements. As an example, naturally occurring carbon exists in three forms. Each carbon atom has the same number of protons (6), which is its atomic number. Each carbon atom also contains six electrons in order to maintain electrical neutrality. However the number of neutrons varies as six, seven, or eight. Isotopes are atoms that have the same number atomic number, but different mass numbers due to a change in the number of neutrons.

The three isotopes of carbon can be referred to as carbon-12 ($^{12}_{6}C$), carbon-13 ($^{13}_{6}C$), and carbon-14 ($^{14}_{6}C$) refers to the nucleus of a given isotope of an element. A carbon atom is one of three different nuclides. Most elements naturally consist of mixtures of isotopes. Carbon has three natural isotopes, while some heavier elements can have many more. Tin has ten stable isotopes, the most of any element.

While the presence of isotopes affects the mass of an atom, it does not affect its chemical reactivity. Chemical behavior is governed by the number of electrons and the number of protons. Carbon-13 behaves chemically in exactly the same way as the more plentiful carbon-12.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Summary

- Isotopes are atoms that have the same atomic number, but different mass numbers due to a change in the number of neutrons.
- The term nuclide refers to the nucleus of a given isotope of an element.
- The atomic mass of an atom equals the sum of the protons and the neutrons.

Review

1. What are isotopes?
2. Why do different isotopes of an element generally have the same physical and chemical properties?
3. How would the nucleus of the hydrogen-1 and hydrogen-2 differ?
4. Relate the concepts of isotope and mass number.
5. All oxygen atoms have eight protons, and most have eight neutrons as well. What is the mass number of an oxygen isotope that has nine neutrons? What is the name of this isotope?
6. An isotope of yttrium has 39 protons and 59 neutrons. What is the mass number of that isotope?
7. An isotope with a mass number of 193 has 116 neutrons. What is the atomic number of this isotope?
8. An isotope of barium (atomic number 56) has an mass of 138. How many neutrons are in the nucleus of this isotope?
9. Relate the concepts of isotope and mass number.

This page titled 5.8: Isotopes is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.18: Isotopes by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.9: Atomic Mass Unit



Figure $5.9.1$ (Public Domain; Derek Jensen (User:Tysto/Wikimedia Commons) via Wikipedia)

The current system of measurement in the United States is a hodge-podge of different units, many of which are hard to interconvert. We at least have standardized units these days, unlike centuries past. At one time, measurements of length were defined as the distance from the end of one appendage to another. For example, the yard was defined as the distance from the king's nose to the tip of his thumb, when his arm was stretched out.

Standardized measurements make it possible for people everywhere to get the same amount of something. Note the red and white labels on the gas pump pictured below. These labels certify that the gas pump has been checked, and is pumping an accurate gallon of gas. Standard measurements in science are very important so that we can compare experimental data from one lab to another, and make sure that we are all talking about the same thing.

Atomic Mass

Masses of individual atoms are very, very small. Using a modern device called a mass spectrometer, it is possible to measure such miniscule masses. An atom of oxygen-16, for example, has a mass of 2.66×10^{-23} g. While comparisons of masses measured in grams would have some usefulness, it is far more practical to have a system that will allow us to more easily compare relative atomic masses. Scientists decided on using the carbon-12 nuclide as the reference standard by which all other masses would be compared. By definition, one atom of carbon-12 is assigned a mass of 12 **atomic mass units** (amu). An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12. The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of 4.0026 amu. An atom of sulfur-32 has a mass of 31.972 amu.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

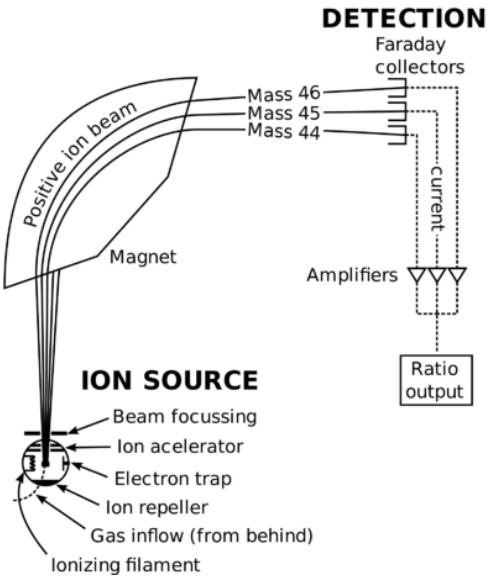


Figure 5.9.2: Mass spectrometer schematic. (Public Domain; Devon Fyson, based on image created by the U.S. Geological Survey via Wikipedia)

The carbon-12 atom has six protons and six neutrons in its nucleus for a mass number of 12. Since the nucleus accounts for nearly all of the mass of the atom, a single proton or single neutron has a mass of approximately $1\ \text{amu}$. However, as seen by the helium and sulfur examples, the masses of individual atoms are not whole numbers. This is because an atom's mass is affected very slightly by the interactions of the various particles within the nucleus, and the small mass of the electron is taken into account.

YouTube Video

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

View this YouTube Video online

Summary

- Carbon-12 is the reference for all atomic mass calculations.
- An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12.
- The mass of an atom is affected by the interactions of particles within the nucleus.

Review

1. What instrument is used to measure the mass of atoms?
2. How much does a single oxygen-16 atom weigh in grams?
3. What is the reference standard for atomic mass units?
4. How is an atomic mass unit defined?
5. Why are the numbers for atomic mass of individual atoms not whole numbers?

This page titled 5.9: Atomic Mass Unit is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.19: Atomic Mass Unit by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.10: Calculating Average Atomic Mass



Figure 5.10.1 (CC BY 2.0; George M. Groutas via Flickr Boulders beach, Cape Peninsula [www.flickr.com])

Have you ever tried to move a boulder?

You have a pile of rocks to move and need to decide what equipment you want to rent to move them. If the rocks are fairly small, you can get a shovel to pick them up. Larger rocks could be moved by hand, but big boulders will need some sort of mechanical scoop. The amount of each kind of rock will also determine how much time you will need to get the job done. Knowing the relative amounts of large, medium, and small rocks can be very useful in deciding how to approach the job.

Percent Natural Abundance

Most elements occur naturally as a mixture of two or more isotopes. The table below shows the natural isotopes of several elements, along with the **percent natural abundance** of each.

Table 5.10.1: Atomic Masses and Percents of Abundance of Some Natural Isotopes

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Hydrogen	$^{1}_{1}\mathrm{H}$	99.985	1.0078	1.0080
	$^{2}_{1}\mathrm{H}$	0.015	2.0141	
	$^{3}_{1}\mathrm{H}$	negligible	3.0160	
Carbon	$^{12}_{6}\mathrm{C}$	98.89	12.000	12.011
	$^{13}_{6}\mathrm{C}$	1.11	13.003	
	$^{14}_{6}\mathrm{C}$	trace	14.003	
Oxygen	$^{16}_{8}\mathrm{O}$	99.759	15.995	15.999
	$^{17}_{8}\mathrm{O}$	0.037	16.995	
	$^{18}_{8}\mathrm{O}$	0.204	17.999	
Chlorine	$^{35}_{17}\mathrm{Cl}$	75.77	34.969	35.453
	$^{37}_{17}\mathrm{Cl}$	24.23	36.966	

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Copper	$^{63}_{29}\mathrm{Cu}$	69.17	62.930	63.546
	$^{65}_{29}\mathrm{Cu}$	30.83	64.928	

For some elements, one particular isotope predominates greatly over the other isotopes. Naturally occurring hydrogen is nearly all hydrogen-1 and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in more substantial quantities. Chlorine (atomic number 17) is a yellowish-green toxic gas. About three quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise **atomic masses**, you would get 36.

$$\frac{(34.969 + 36.966)}{2} = 35.968 \text{ amu}$$

Clearly the actual average atomic mass from the last column of the table is significantly lower. Why? We need to take into account the percent natural abundance of each isotope, in order to calculate the weighted average. The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element. The sample problem below demonstrates how to calculate the atomic mass of chlorine.

Example 5.10.1

Use the atomic masses of each of the two isotopes of chlorine along with their respective percent abundances to calculate the average atomic mass of chlorine.

Solution

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

Known

- Chlorine-35: atomic mass $= 34.969$ amu and percent abundance $= 75.77\%$
- Chlorine-37: atomic mass $= 36.966$ amu and percent abundance $= 24.23\%$

Unknown

- Average atomic mass of chlorine

Change each percent abundance into decimal form by dividing by 100. Multiply this value by the atomic mass of that isotope. Add together for each isotope to get the average atomic mass.

Step 2: Calculate.

$$\begin{array}{ll} \text{chlorine-35} & 0.7577 \times 34.969 = 26.50 \text{ amu} \\ \text{chlorine-37} & 0.2423 \times 36.966 = 8.957 \text{ amu} \\ \text{average atomic mass} & 26.50 + 8.957 = 35.46 \text{ amu} \end{array}$$

Note: Applying significant figure rules results in the 35.45 amu result without excessive rounding error. In one step:

$$(0.7577 \times 34.969) + (0.2423 \times 36.966) = 35.46 \text{ amu}$$

Step 3: Think about your result.

The calculated average atomic mass is closer to 35 than to 37 because a greater percentage of naturally occurring chlorine atoms have the mass number of 35. It agrees with the value from the table above.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Summary

- The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element.
- Calculations of atomic mass use the percent abundance of each isotope.

Review

1. Define atomic mass.
2. What information do you need to calculate atomic mass for an element?
3. Calculate the atomic mass for carbon using the data provided in the table below.

Isotope	Atomic Mass	Percent Abundance
carbon-12	12.000000	98.90
carbon-13	13.003355	1.100

This page titled 5.10: Calculating Average Atomic Mass is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.20: Calculating Average Atomic Mass by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

5.11: Ion



Figure 5.11.1 (CC by 2.0; Nick Russill via Flickr, Northern Lights, Greenland [www.flickr.com])

The incredible green lights in this cold northern sky consist of charged particles known as ions. Their swirling pattern is caused by the pull of Earth's magnetic field. Called the northern lights, this phenomenon of nature shows that ions respond to a magnetic field. Do you know what ions are? Read on to find out.

Atoms Are Neutral

The northern lights aren't caused by atoms, because atoms are not charged particles. An atom always has the same number of electrons as protons. Electrons have an electric charge of -1 and protons have an electric charge of +1. Therefore, the charges of an atom's electrons and protons "cancel out." This explains why atoms are neutral in electric charge.

Q: What would happen to an atom's charge if it were to gain extra electrons?

A: If an atom were to gain extra electrons, it would have more electrons than protons. This would give it a negative charge, so it would no longer be neutral.

Atoms to Ions

Atoms cannot only gain extra electrons. They can also lose electrons. In either case, they become **ions**. Ions are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons. If atoms lose electrons, they become positive ions, or cations. If atoms gain electrons, they become negative ions, or anions. Consider the example of fluorine (see figure below). A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with an electric charge of -1.

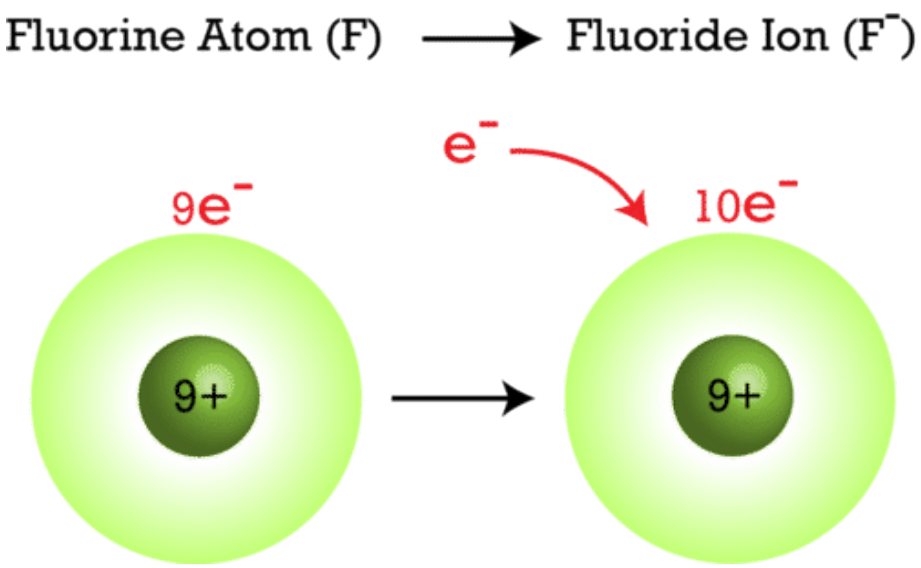


Figure 5.11.2 (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Names and Symbols

Like fluoride, other negative ions usually have names ending in *–ide*. Positive ions, on the other hand, are just given the element name followed by the word *ion*. For example, when a sodium atom loses an electron, it becomes a positive sodium ion. The charge of an ion is indicated by a plus (+) or minus sign (-), which is written to the right of and just above the ion's chemical symbol. For example, the fluoride ion is represented by the symbol F^-, and the sodium ion is represented by the symbol Na^+. If the charge is greater than one, a number is used to indicate it. For example, iron (Fe) may lose two electrons to form an ion with a charge of plus two. This ion would be represented by the symbol Fe^{2+}. This and some other common ions are listed with their symbols in the table below.

Table $PageIndex1$: Some Common Ions

Cations		Anions	
Name of Ion	Chemical Symbol	Name of Ion	Chemical Symbol
Calcium ion	Ca^{2+}	Chloride	Cl^-
Hydrogen ion	H^+	Fluoride	F^-
Iron(II) ion	Fe^{2+}	Bromide	Br^-
Iron(III) ion	Fe^{3+}	Oxide	O^{2-}

Q: How does the iron(III) ion differ from the iron(II) ion?

A: The iron(III) ion has a charge of +3, so it has one less electron than the iron(II) ion, which has a charge of +2.

Q: What is the charge of an oxide ion? How does its number of electrons compare to its number of protons?

A: An oxide ion has a charge of -2. It has two more electrons than protons.

How Ions Form

The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation. The radiation may give their outer electrons enough energy to escape from the attraction of the positive nucleus. However, most ions form when atoms transfer electrons to or from other atoms or molecules. For example, sodium atoms may

transfer electrons to chlorine atoms. This forms positive sodium ions (Na^+) and negative chloride ions (Cl^-).

YouTube Video

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

View this YouTube Video online

Q: Why do you think atoms lose electrons to, or gain electrons from, other atoms?

A: Atoms form ions by losing or gaining electrons because it makes them more stable and this state takes less energy to maintain. The most stable state for an atom is to have its outermost energy level filled with the maximum possible number of electrons. In the case of metals such as lithium, with just one electron in the outermost energy level, a more stable state can be achieved by losing that one outer electron. In the case of nonmetals such as fluorine, which has seven electrons in the outermost energy level, a more stable state can be achieved by gaining one electron and filling up the outer energy level.

YouTube Video

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

View this YouTube Video online

Properties of Ions

Ions are highly reactive, especially as gases. They usually react with ions of opposite charge to form neutral compounds. For example, positive sodium ions and negative chloride ions react to form the neutral compound sodium chloride, commonly known as table salt. This occurs because oppositely charged ions attract each other. Ions with the same charge, on the other hand, repel each other. Ions are also deflected by a magnetic field, as you saw in the opening image of the northern lights.

Summary

- Atoms have equal numbers of positive protons and negative electrons, so they are neutral in electric charge.
- Atoms can gain or lose electrons and become ions, which are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons.
- The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation or when atoms transfer electrons to or from other atoms.
- Ions are reactive, attracted or repulsed by other charged particles, and deflected by a magnetic field.

Review

1. Why are atoms neutral in electric charge?
2. Define ion.
3. Compare and contrast cations and anions, and give an example of each.
4. Describe how ions form.
5. List properties of ions.
6. The model in the illustration below represents an atom of lithium (Li). If the lithium atom becomes an ion, which type of ion

will it be, a cation or an anion? What will be the electric charge of this ion? What will the ion be named? What symbol will be used to represent it?

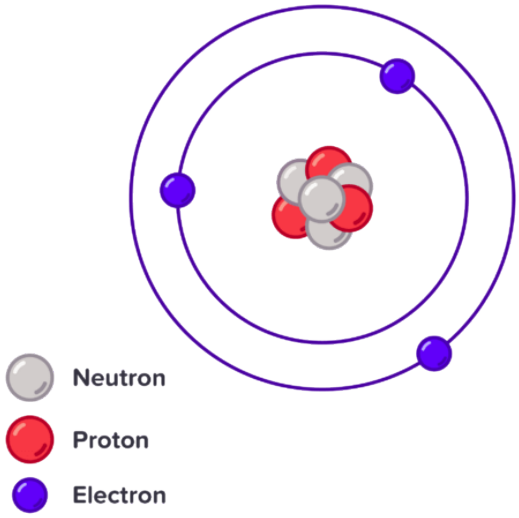


Figure 5.11.3 (CC BY-NC 3.0; CK-12 Foundation Employee via CK-12 Foundation)

This page titled 5.11: Ion is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

6: ELECTRONIC STRUCTURE AND CHEMICAL PERIODICITY

6.1: Periodic Law

6.2: Modern Periodic Table- Periods and Groups

6.3: Blocks of the Periodic Table

6.4: Energy Level

6.5: Orbitals

6.6: Electron Configurations

6.7: Noble Gas Configuration

6.8: Hund's Rule and Orbital Filling Diagrams

6.9: Aufbau Principle

6.10: Periodic Trends- Atomic Radius

6.11: Metals

6.12: Nonmetals

6.13: Metalloids

This page titled 6: Electronic Structure and Chemical Periodicity is shared under a license and was authored, remixed, and/or curated by LibreTexts.

6.1: Periodic Law



Figure 6.1.1: Copy and Paste Caption here. (CC-BY 2.0; Christopher Fynn via Flickr, Mysore - Fruit vendor [www.flickr.com])

How are these items related to one another?

We have all enjoyed looking around a market for delicious foods to eat later at home. When you get to the market you know you need to get fruits, vegetables and grains to ensure you eat a balanced diet. In the market, these items are all grouped together to make it easier to find the type of fruit, vegetable or grain you are looking for; it would be inconvenient if the different fruits were all scattered in different places. The periodic table is organized in a similar way, ensuring similar elements are found in the same group or period.

The Periodic Law

When Mendeleev put his periodic table together, nobody knew about the existence of the nucleus. It was not until 1911 that Rutherford conducted his gold foil experiment that demonstrated the presence of the nucleus in the atom. Just two years later, in 1913, English physicist Henry Moseley (1887-1915) examined x-ray spectra of a number of chemical elements. He would shoot x-rays through crystals of the element and study the wavelengths of the radiation that he detected. Moseley found that there was a relationship between wavelength and atomic number. His results led to the definition of atomic number as the number of protons contained in the nucleus of each atom. He then realized that the elements of the periodic table should be arranged in order of increasing atomic number, rather than increasing atomic mass.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

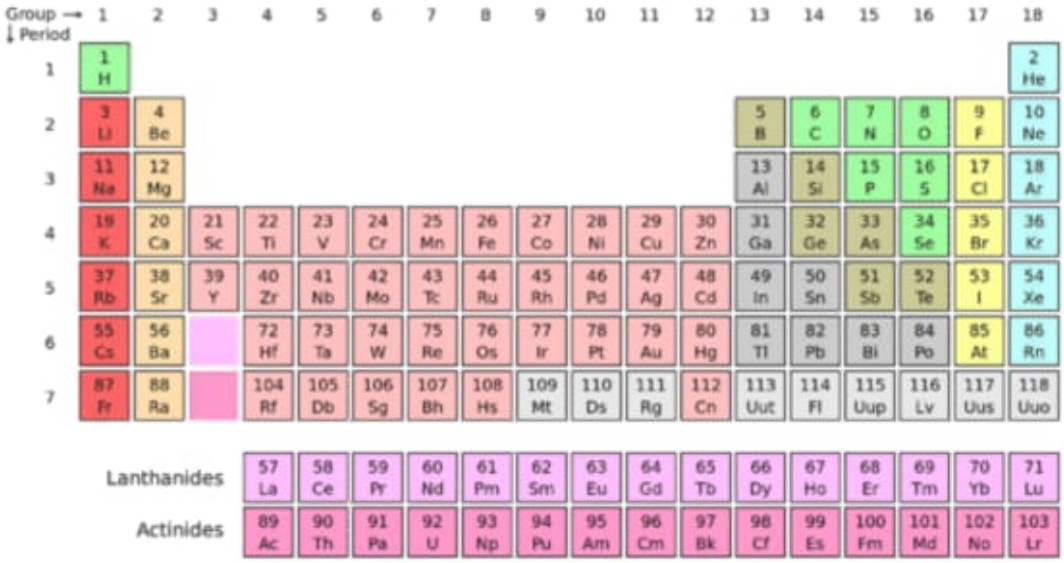


Figure 6.1.2 (Public Domain; User:Cepheus/Wikimedia Commons via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

When ordered by atomic number, the discrepancies within Mendeleev's table disappeared. Tellurium has an atomic number of 52, while iodine has an atomic number of 53. So even though tellurium does indeed have a greater atomic mass than iodine, it is properly placed before iodine in the periodic table. Mendeleev and Moseley are credited with being most responsible for the modern **periodic law**: When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties. The result is the periodic table as we know it today. Each new horizontal row of the periodic table corresponds to the beginning of a new **period** because a new principal energy level is being filled with electrons. Elements with similar chemical properties appear at regular intervals, within the vertical columns called **groups**.

YouTube Video

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

View this YouTube Video online

Summary

- Elements of the periodic table are arranged in order of increasing atomic number.
- The periodic law states: "When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties."

Review

1. Did Mendeleev know about the nucleus of an atom?
2. Who discovered the relationship between wavelength of X-rays and atomic number?
3. What did Moseley conclude from his research?
4. What is the “periodic law”?
5. What do the vertical columns (groups) in the periodic table represent?

This page titled 6.1: Periodic Law is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.3: Periodic Law by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.2: Modern Periodic Table- Periods and Groups

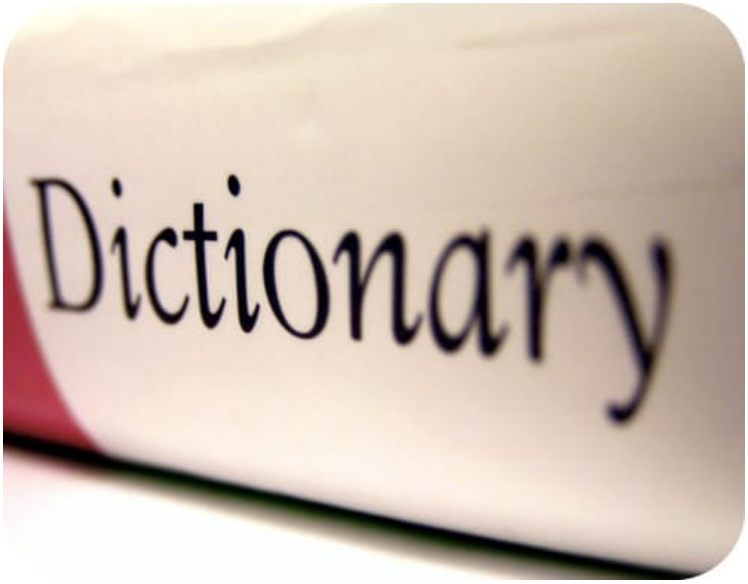


Figure 6.2.1 (Copyright; author via source)

How has the English dictionary evolved over time?

Language changes with time. New words enter the language and old words often disappear from lack of use. Dictionaries are published so that people can keep up with changes in language and know how to use words properly. These publications may be in print, as is the law dictionary below, or they may be electronic. Dictionaries can be found on the internet and apps are available for smartphones. Dictionaries are invaluable for good, reliable communication.

The Modern Periodic Table

The periodic table has undergone extensive changes in the time since it was originally developed by Mendeleev and Moseley. Many new elements have been discovered, while others have been artificially synthesized. Each fits properly into a **group** of elements with similar properties. The periodic table is an arrangement of the elements in order of their atomic numbers, so that elements with similar properties appear in the same vertical column or group.

The figure below shows the most commonly used form of the periodic table. Each square shows the chemical symbol of the element along with its name. Notice that several of the symbols seem to be unrelated to the name of the element: Fe for iron, Pb for lead, etc. Most of these are the elements that have been known since ancient times and have symbols based on their Latin names. The atomic number of each element is written above the symbol.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

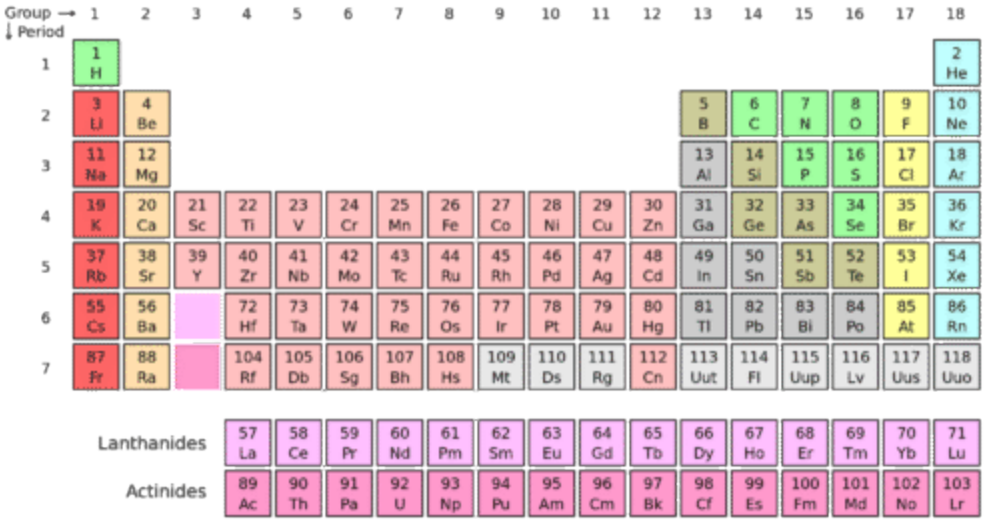


Figure 6.2.2 (Public Domain; User:Cepheus/Wikimedia Commons via Wikipedia)

A **period** is a horizontal row of the periodic table. There are seven periods in the periodic table, with each one beginning at the far left. A new period begins when a new principal energy level begins filling with electrons. Period 1 has only two elements (hydrogen and helium), while periods 2 and 3 have 8 elements. Periods 4 and 5 have 18 elements. Periods 6 and 7 have 32 elements, because the two bottom rows that are separate from the rest of the table belong to those periods. These two rows are pulled out in order to make the table itself fit more easily onto a single page.

A group is a vertical column of the periodic table, based on the organization of the outer shell electrons. There are a total of 18 groups. There are two different numbering systems that are commonly used to designate groups, and you should be familiar with both. The traditional system used in the United States involves the use of the letters A and B. The first two groups are 1A and 2A, while the last six groups are 3A through 8A. The middle groups use B in their titles. Unfortunately, there was a slightly different system in place in Europe. To eliminate confusion, the International Union of Pure and Applied Chemistry (IUPAC) decided that the official system for numbering groups would be a simple 1 through 18 from left to right. Many periodic tables show both systems simultaneously.

Most recent depictions of the periodic table show an incomplete seventh period. In fact, many of those seventh period elements were not known before the early twentieth century. Many elements have been synthesized by bombarding known elements with sub-atomic particles such as neutrons or alpha particles. Uranium has been used in this manner to produce elements 93-100. Larger elements (atomic numbers 101 and above) are formed by fusing nuclei of smaller elements together.

These synthetic elements tend to be very unstable, often existing for less than a second, so little is known about them. Elements up to atomic number 112 are known (not in any detail, however), and some evidence for elements 113 and above had been put forth, with elements 114 and 116 being added to the table in 2011. In December 2015, the IUPAC verified the existence of four new elements 113, 115, 117, and 118 and approved their addition to the periodic table.

The following are the new element names and their origins:

- Element 113 was named Nihonium, symbol Nh, proposed by Japanese researchers after the Japanese word Nihon, which means Japan.
- A team of scientists from Russia and the United States named element 115, Moscovium, symbol Mc, after Moscow and element 117, Tennessine, symbol Ts, after Tennessee.
- The Russian team that discovered element 118 named it Oganesson, symbol Og, after Yuri Oganessian, a prolific element hunter.

These elements complete the seventh period or row of the periodic table.

The video below reviews groups and trends in the periodic table:

YouTube Video

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

View this YouTube Video online

Watch the additional videos below to learn more about the newly identified and named chemical elements:

Element Name and Symbol: Nihonium (Nh)

Atomic Number: 113

YouTube Video

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

View this YouTube Video online

Element Name and Symbol: Moscovium (Mc)

Atomic Number: 115

YouTube Video

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

View this YouTube Video online

Element Name and Symbol: Tennessine (Ts)

Atomic Number: 117

YouTube Video

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

View this YouTube Video online

Element Name and Symbol: Oganesson (Og)

Atomic Number: 118

YouTube Video

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

View this YouTube Video online

Summary

- The periodic table is arranged in order of atomic number.
- A period is a horizontal row of the periodic table.
- A group is a vertical row of the periodic table.

Review

1. How is today's periodic table different from the one that Mendeleev published?
2. Are all the elements in today's periodic table naturally occurring? Explain your answer.
3. What is a "period"? What does it represent?
4. What is a "group"? What does it represent?
5. Why are there two different numbering systems for groups?

This page titled 6.2: Modern Periodic Table- Periods and Groups is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.4: Modern Periodic Table- Periods and Groups by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.3: Blocks of the Periodic Table
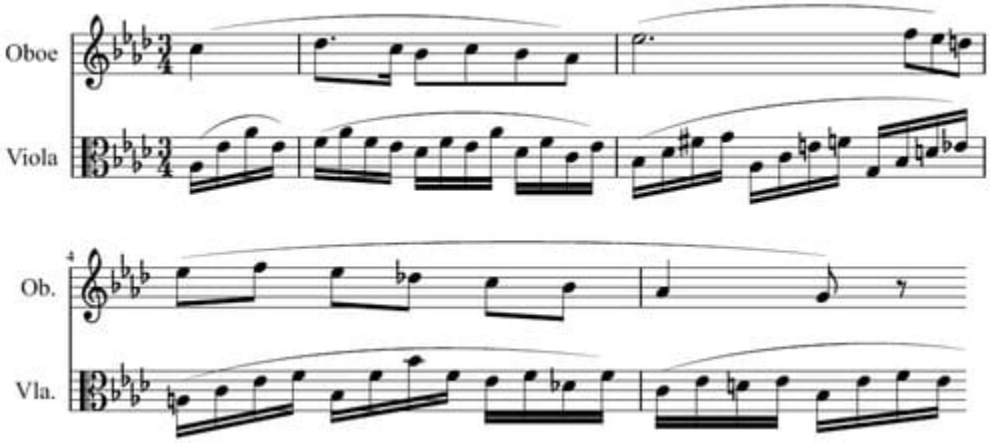


Figure 6.3.1 (Public Domain; User:Szalax/Wikimedia Commons via Wikipedia)

What makes these music notes unique?

We all enjoy music of some sort. Some people like classical music, others like jazz or country. Music styles change from one period of time to the next, and from one region to another. Each type of music has its language that describes it. Classical music has a certain structure, style, and content. There are different expressions of classical music – the symphony, concerto, sonata. We have ballet and opera as well as choral music. Jazz has a different set of characteristics from classical and different styles of performance. Each type of music can be described and compared to other types on the basis of certain common qualities like notes, chords, and melodic styles.

The elements in the periodic table could be considered to be similar to types of music. Each set of elements has its unique set of properties, with different sets of elements having some common characteristics in terms of electron arrangements. We can see patterns of electronic structure and reactivity in the periodic table that allow us to understand better the behavior of individual elements.

Periods and Blocks

There are seven horizontal rows of the periodic table, called **periods**. The length of each period is determined by the number of electrons that are capable of occupying the **sublevels** that fill during that period, as seen in the table below.

Table 6.3.1: Period Length and Sublevels in the Periodic Table

Period	Number of Elements in Period	Sublevels in Order of Fill
1	2	$1s$
2	8	$2s\ 2p$
3	8	$3s\ 3p$
4	18	$4s\ 3d\ 4p$
5	18	$5s\ 4d\ 5p$
6	32	$6s\ 4f\ 5d\ 6p$
7	32	$7s\ 5f\ 6d\ 7s$

Recall that the four different sublevels each consist of a different number of orbitals. The s sublevel has one orbital, the p sublevel has three orbitals, the d sublevel has five orbitals, and the f sublevel has seven orbitals. In the first period, only the $1s$ sublevel is being filled. Since all orbitals can hold two electrons, the entire first period consists of just two elements. In the second period, the

$2s$ sublevel, with two electrons, and the $2p$ sublevel, with six electrons, are being filled. Consequently, the second period contains eight elements. The third period is similar to the second, filling the $3s$ and $3p$ sublevels. Notice that the $3d$ sublevel does not actually fill until after the $4s$ sublevel. This results in the fourth period containing 18 elements due to the additional 10 electrons that are contributed by the d sublevel. The fifth period is similar to the fourth. After the $6s$ sublevel fills, the $4f$ sublevel with its 14 electrons fills. This is followed by the $5d$ and the $6p$. The total number of elements in the sixth period is 32. The later elements in the seventh period are still being created. So while there is a possible number of 32 elements in the period, the current number is slightly less.

The period to which a given element belongs can be easily determined by its electron configuration. For example, consider the element nickel (Ni). Its electron configuration is $[\mathrm{Ar}]\ \ 3d^8\ \ 4s^2$. The highest occupied principal energy level is the fourth, indicated by the 4 in the $4s^2$ portion of the configuration. Therefore, nickel can be found in the fourth period of the periodic table.

Based on electron configurations, the periodic table can be divided into **blocks** denoting which sublevel is in the process of being filled. The s, p, d, and f blocks are illustrated below.

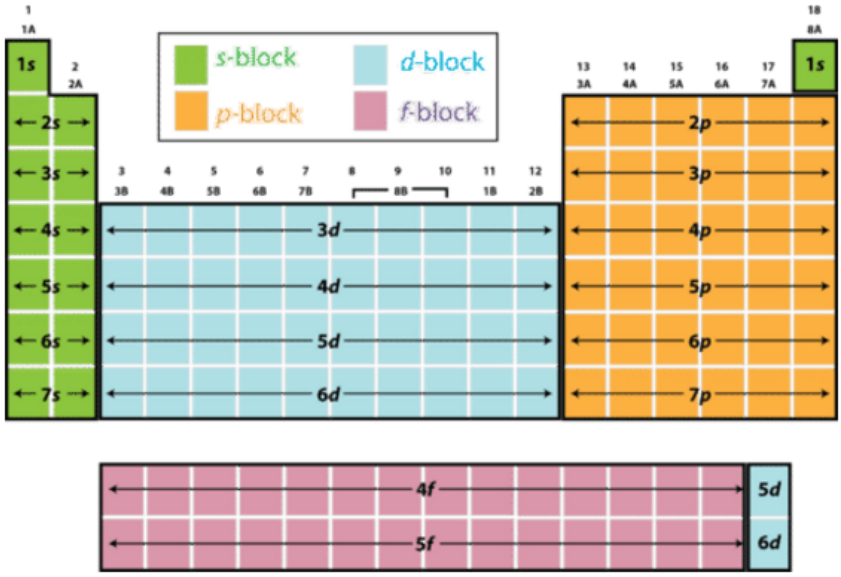


Figure 6.3.2 (CC BY-NC 3.0; Christopher Auyeung and Joy Sheng via CK-12 Foundation)

The figure also illustrates how the d sublevel is always one principal level behind the period in which that sublevel occurs. In other words, the $3d$ sublevel fills during the fourth period. The f sublevel is always two levels behind. The $4f$ sublevel belongs to the sixth period.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- The horizontal rows of the periodic table are called periods.
- The length of a period depends on how many electrons are needed to occupy the sublevels that fill the period.
- Blocks indicate which sublevel is being filled.

Review

1. What are the horizontal rows of the periodic table called?
2. Which sublevel is being filled in period 1?
3. Which sublevel is being filled in period 7?
4. How does the electron configuration of an element give information about the period it is in?
5. What block of elements has the d sublevels being filled?

This page titled 6.3: Blocks of the Periodic Table is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.8: Blocks of the Periodic Table by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.4: Energy Level



Figure 6.4.1 (Pixabay License; PublicDomainPictures via Pixabay)

Fireworks are a great way to celebrate happy events. Do you know what causes the brilliant, colored lights of fireworks? The lights are bursts of energy given off by atoms in the fireworks. What do you suppose causes these bursts of light? The answer has to do with energy levels of atoms.

What Are Energy Levels?

Energy levels (also called electron shells) are fixed distances from the nucleus of an atom where electrons may be found. Electrons are tiny, negatively charged particles in an atom that move around the positive nucleus at the center. Energy levels are a little like the steps of a staircase. You can stand on one step or another but not in between the steps. The same goes for electrons. They can occupy one energy level or another but not the space between energy levels.

The model in the figure below shows the first four energy levels of an atom. Electrons in energy level I (also called energy level K) have the least amount of energy. As you go farther from the nucleus, electrons at higher levels have more energy, and their energy increases by a fixed, discrete amount. Electrons can jump from a lower to the next higher energy level if they absorb this amount of energy. Conversely, if electrons jump from a higher to a lower energy level, they give off energy, often in the form of light. This explains the fireworks pictured above. When the fireworks explode, electrons gain energy and jump to higher energy levels. When they jump back to their original energy levels, they release the energy as light. Different atoms have different arrangements of electrons, so they give off light of different colors.

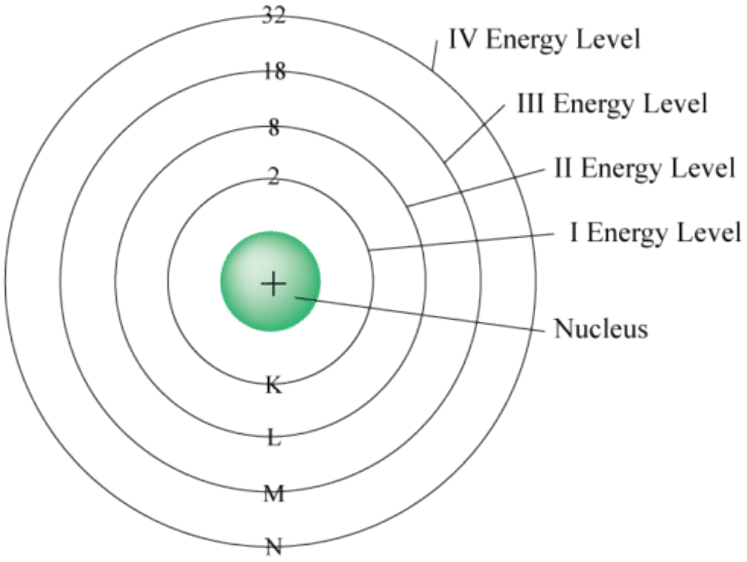


Figure 6.4.2 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: In the atomic model Figure above, where would you find electrons that have the most energy?

A: Electrons with the most energy would be found in energy level IV.

Energy Levels and Orbitals

The smallest atoms are hydrogen atoms. They have just one electron. That one electron is in the first energy level. Bigger atoms have more electrons. Electrons are always added to the lowest energy level first until it has the maximum number of electrons possible. Then electrons are added to the next higher energy level until that level is full, and so on.

How many electrons can a given energy level hold? The maximum numbers of electrons possible for the first four energy levels are shown in the figure above. For example, energy level I can hold a maximum of two electrons, and energy level II can hold a maximum of eight electrons. The maximum number depends on the number of orbitals at a given energy level. An orbital is a volume of space within an atom where an electron is most likely to be found. As you can see by the images in the figure below, some orbitals are shaped like spheres (S orbitals) and some are shaped like dumbbells (P orbitals). There are other types of orbitals as well.

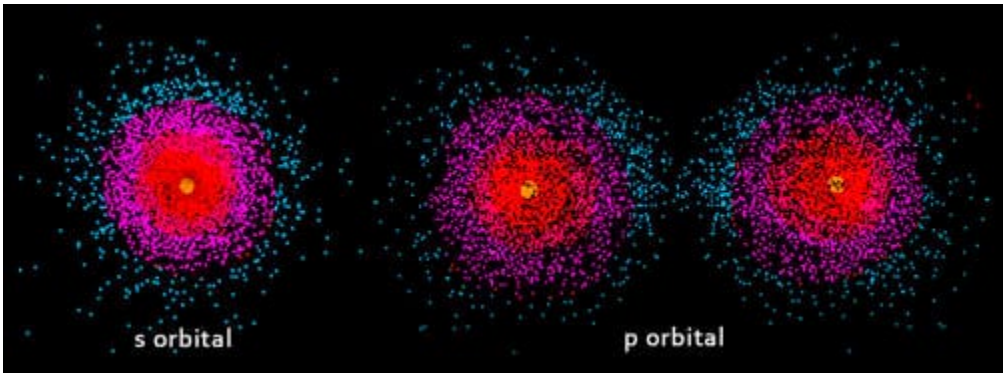


Figure 6.4.3 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

Regardless of its shape, each orbital can hold a maximum of two electrons. Energy level I has just one orbital, so two electrons will fill this energy level. Energy level II has four orbitals, so it takes eight electrons to fill this energy level.

Q: Energy level III can hold a maximum of 18 electrons. How many orbitals does this energy level have?

A: At two electrons per orbital, this energy level must have nine orbitals.

The Outermost Level

Electrons in the outermost energy level of an atom have a special significance. These electrons are called valence electrons, and they determine many of the properties of an atom. An atom is most stable if its outermost energy level contains as many electrons as it can hold. For example, helium has two electrons, both in the first energy level. This energy level can hold only two electrons, so helium's only energy level is full. This makes helium a very stable element. In other words, its atoms are unlikely to react with other atoms.

Consider the elements fluorine and lithium, modeled in the figure below. Fluorine has seven of eight possible electrons in its outermost energy level, which is energy level II. It would be more stable if it had one more electron because this would fill its outermost energy level. Lithium, on the other hand, has just one of eight possible electrons in its outermost energy level (also energy level II). It would be more stable if it had one less electron because it would have a full outer energy level (now energy level I).

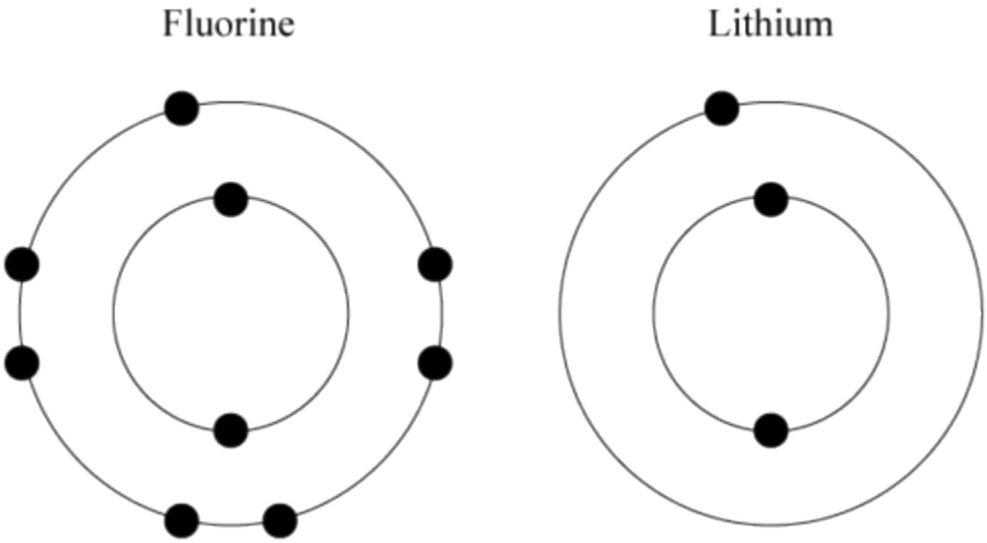


Figure 6.4.4 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Both fluorine and lithium are highly reactive elements because of their number of valence electrons. Fluorine will readily gain one electron and lithium will just as readily give up one electron to become more stable. In fact, lithium and fluorine will react together as shown in the figure below. When the two elements react, lithium transfers its one "extra" electron to fluorine.

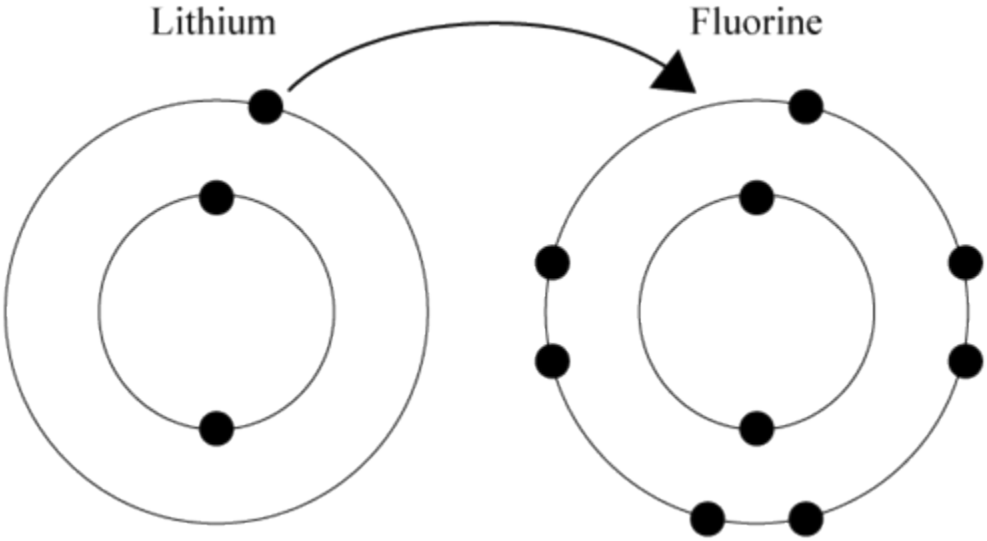


Figure 6.4.5 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: A neon atom has ten electrons. How many electrons does it have in its outermost energy level? How stable do you think a neon atom is?

A: A neon atom has two electrons in energy level I and its remaining eight electrons in energy level II, which can hold only eight electrons. This means that is outermost energy level is full. Therefore, a neon atom is very stable.

Summary

- Energy levels (also called electron shells) are fixed distances from the nucleus of an atom where electrons may be found. As you go farther from the nucleus, electrons at higher energy levels have more energy.
- Electrons are always added to the lowest energy level first until it has the maximum number of electrons possible, and then electrons are added to the next higher energy level until that level is full, and so on. The maximum number of electrons at a given energy level depends on its number of orbitals. There are at most two electrons per orbital.
- Electrons in the outermost energy level of an atom are called valence electrons. They determine many of the properties of an atom, including how reactive it is.

Review

1. What are energy levels?
2. Relate energy levels to the amount of energy their electrons have.
3. What must happen for an electron to jump to a different energy level?
4. How many electrons can the fourth energy level have? How many orbitals are there at this energy level?
5. An atom of sodium has 11 electrons. Make a sketch of a sodium atom, showing how many electrons it has at each energy level. Infer how reactive sodium atoms are.

Resources

YouTube Video

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

View this YouTube Video online

This page titled 6.4: Energy Level is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

6.5: Orbitals



Figure 6.5.1 (Public Domain; User:Baseball Bugs/Wikimedia Commons via Commons Wikimedia, Plane over Lake Harriet [commons.wikimedia.org])

How is it that so many planes are able to fly without running into each other?

The flight path of a commercial airliner is carefully regulated by the Federal Aviation Administration. Each airplane must maintain a distance of five miles from another plane flying at the same altitude, and be 2,000 feet above and below another aircraft (1,000 feet if the altitude is less than 29,000 feet). So, each aircraft only has certain positions it is allowed to maintain while it flies. Quantum mechanics demonstrates that electrons have similar restrictions on their locations.

Orbitals

We can apply our knowledge of quantum numbers to describe the arrangement of electrons for a given atom. We do this with something called **electron configurations**. They are effectively a map of the electrons for a given atom. We look at the four quantum numbers for a given electron and then assign that electron to a specific orbital.

s Orbitals

For any value of n, a value of $l = 0$ places that electron in an s **orbital**. This orbital is spherical in shape:

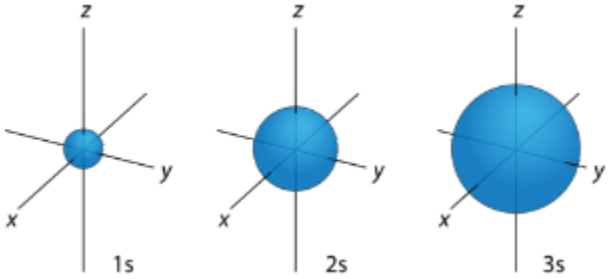


Figure 6.5.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

p Orbitals

From the table below, it is evident that there are three possible orbitals when $l = 1$. These are designated as p **orbitals** and have dumbbell shapes. Each of the p orbitals has a different orientation in three-dimensional space.

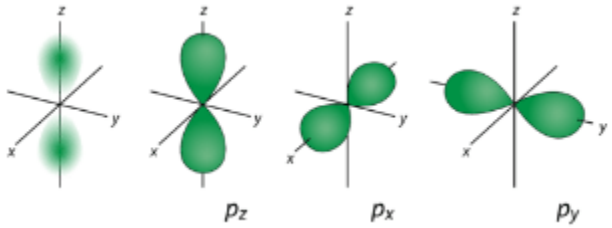


Figure 6.5.3 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

d Orbitals

When $l = 2$, m_l values can be $-2,\ -1,\ 0,\ +1,\ +2$ for a total of five d **orbitals**. Note that all five of the orbitals have specific three-dimensional orientations.

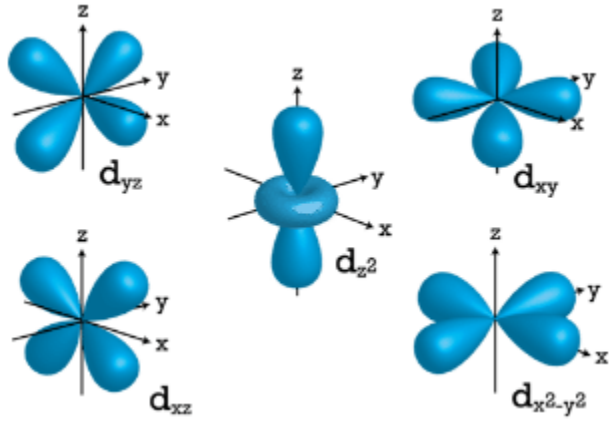


Figure 6.5.4 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

f Orbitals

The most complex set of orbitals are the f **orbitals**. When $l = 3$, m_l values can be $-3,\ -2,\ -1,\ 0,\ +1,\ +2,\ +3$ for a total of seven different orbital shapes. Again, note the specific orientations of the different f orbitals.

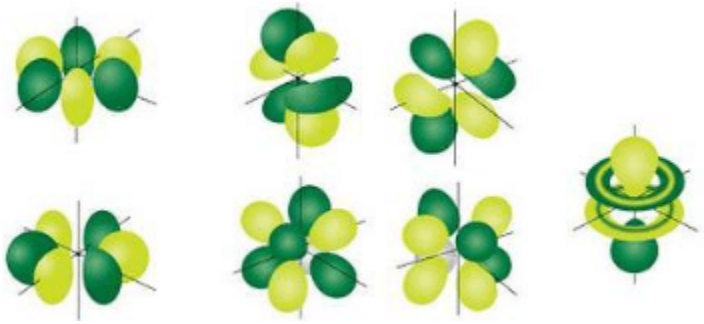


Figure 6.5.5 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Table 6.5.1: Electron Arrangement Within Energy Levels

Principal Quantum Number (n)	Allowable Sublevels	Number of Orbitals per Sublevel	Number of Orbitals per Principal Energy Level	Number of Electrons per Sublevel	Number of Electrons per Principal Energy Level
1	s	1	1	2	2
2	s	1	4	2	8
	p	3		6	
3	s	1	9	2	18
	p	3		6	
	d	5		10	
4	s	1	16	2	32
	p	3		6	
	d	5		10	
	f	7		14	

YouTube Video

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

View this YouTube Video online

Summary

- There are four different classes of electron orbitals.
- Electron orbitals are determined by the value of the angular momentum quantum number l.

Review

1. What is an electron configuration?
2. How many electrons are in the $n = 1$ orbital?
3. What is the total number of electrons in a p orbital?
4. How many electrons does it take to completely fill a d orbital?

This page titled 6.5: Orbitals is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.13: Orbitals by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.6: Electron Configurations

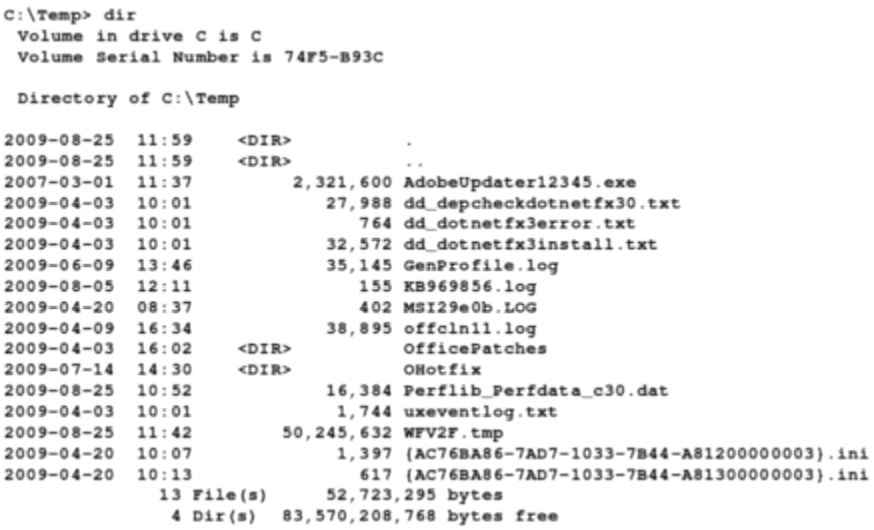


Figure 6.6.1 (Public Domain; David R. Tribble (User:Loadmaster/Wikimedia Commons) via Commons wikimedia, Dir Command in Windows Command Prompt [commons.wikimedia.org])

How big is a file?

If you keep your papers in manila folders, you can pick up a folder and see how much it weighs. If you want to know how many different papers (articles, bank records, or whatever else you keep in a folder), you have to take everything out and count. A computer directory, on the other hand, tells you exactly how much you have in each file. We can get the same information on atoms. If we use an orbital filling diagram, we have to count arrows. When we look at electron configuration data, we simply add up the numbers.

Electron Configurations

Electron configuration notation eliminates the boxes and arrows of orbital filling diagrams. Each occupied sublevel designation is written followed by a superscript that is the number of electrons in that sublevel. For example, the hydrogen configuration is $1s^1$, while the helium configuration is $1s^2$. Multiple occupied sublevels are written one after another. The electron configuration of lithium is $1s^2 2s^1$. The sum of the superscripts in an electron configuration is equal to the number of electrons in that atom, which is in turn equal to its atomic number.

Example 6.6.1

Draw the orbital filling diagram for carbon and write its electron configuration.

Solution

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

Known

- atomic number of carbon, Z=6

Use the order of fill diagram to draw an orbital filling diagram with a total of six electrons. Follow Hund's rule. Write the electron configuration.

Step 2: Construct diagram.

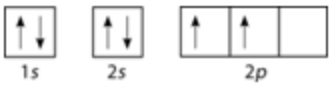


Figure 6.6.2: Orbital filling diagram for carbon. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Electron configuration $1s^2 2s^2 2p^2$

Step 3: Think about your result.

Following the $2s$ sublevel is the $2p$, and p sublevels always consist of three orbitals. All three orbitals need to be drawn even if one or more is unoccupied. According to Hund's rule, the sixth electron enters the second of those p orbitals, with the same spin as the fifth electron.

Second Period Elements

Periods refer to the horizontal rows of the periodic table. The first period of the periodic table contains only the elements hydrogen and helium. This is because the first principal energy level consists of only the s sublevel, and so only two electrons are required in order to fill the entire principal energy level. Each time a new principal energy level begins, as with the third element lithium, a new period is started on the periodic table. As one moves across the second period, electrons are successively added. With beryllium $(Z = 4)$, the $2s$ sublevel is complete and the $2p$ sublevel begins with boron $(Z = 5)$. Since there are three $2p$ orbitals and each orbital holds two electrons, the $2p$ sublevel is filled after six elements. Table 6.6.1 shows the electron configurations of the elements in the second period.

Table 6.6.1: Electron Configurations of Second-Period Elements

Element Name	Symbol	Atomic Number	Electron Configuration
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Electron configuration notation simplifies the indication of where electrons are located in a specific atom.
- Superscripts are used to indicate the number of electrons in a given sublevel.

Review

1. What does electron configuration notation eliminate?
2. How do we know how many electrons are in each sublevel?
3. An atom has the electron configuration of $1s^2 2s^2 2p^5$. How many electrons are in that atom?
4. Which element has the electron configuration of $1s^2 2s^2 2p^6 3s^2$?

This page titled 6.6: Electron Configurations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.18: Electron Configurations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.7: Noble Gas Configuration



Figure 6.7.1 (Public Domain; User:Danbold/Wikipedia via Wikipedia)

How does it feel to be full after a meal?

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

Noble Gas Configuration

Sodium, element number 11, is the first element in the third period of the periodic table. Its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon $(Z = 10)$. This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. The elements that are found in the last column of the periodic table are an important group of elements called the noble gases. They are helium, neon, argon, krypton, xenon, and radon. A **noble gas configuration** of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons. So for sodium, we make the substitution of $[\mathrm{Ne}]$ for the $1s^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes $[\mathrm{Ne}]\, 3s^1$. Table 6.7.1 shows the noble gas configurations of the third period elements.

Table 6.7.1: Electron Configurations of Third-Period Elements

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Sodium	Na	11	$[\mathrm{Ne}]\, 3s^1$
Magnesium	Mg	12	$[\mathrm{Ne}]\, 3s^2$
Aluminum	Al	13	$[\mathrm{Ne}]\, 3s^2 3p^1$
Silicon	Si	14	$[\mathrm{Ne}]\, 3s^2 3p^2$
Phosphorus	P	15	$[\mathrm{Ne}]\, 3s^2 3p^3$
Sulfur	S	16	$[\mathrm{Ne}]\, 3s^2 3p^4$
Chlorine	Cl	17	$[\mathrm{Ne}]\, 3s^2 3p^5$
Argon	Ar	18	$[\mathrm{Ne}]\, 3s^2 3p^6$

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

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

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 6.7: Noble Gas Configuration is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.20: Noble Gas Configuration by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.8: Hund's Rule and Orbital Filling Diagrams



Figure 6.8.1 (CC BY 3.0; Laura Guerin via CK-12 Foundation)

Have you ever wondered what those load limit signs mean on a bridge?

The sign above signifies that nothing over five tons is allowed because it will do damage to the structure. There are limits to the amount of weight that a bridge can support, there are limits to the number of people that can safely occupy a room, and there are limits to what can go into an electron orbital.

Hund's Rule

The last of the three rules for constructing electron arrangements requires electrons to be placed one at a time in a set of orbitals within the same sublevel. This minimizes the natural repulsive forces that one electron has for another. **Hund's rule** states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and that each of the single electrons must have the same spin. The figure below shows how a set of three p orbitals is filled with one, two, three, and four electrons.

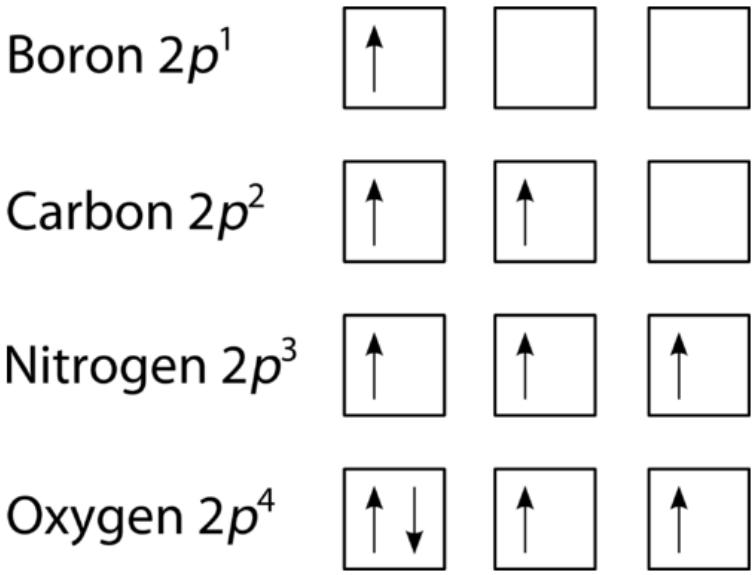


Figure $6.8.2$: The $2p$ sublevel, for the elements boron $(Z = 5)$, carbon $(Z = 6)$, nitrogen $(Z = 7)$, and oxygen $(Z = 8)$. According to Hund's rule, as electrons are added to a set of orbitals of equal energy, one electron enters each orbital before any orbital receives a second electron. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Orbital Filling Diagrams

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

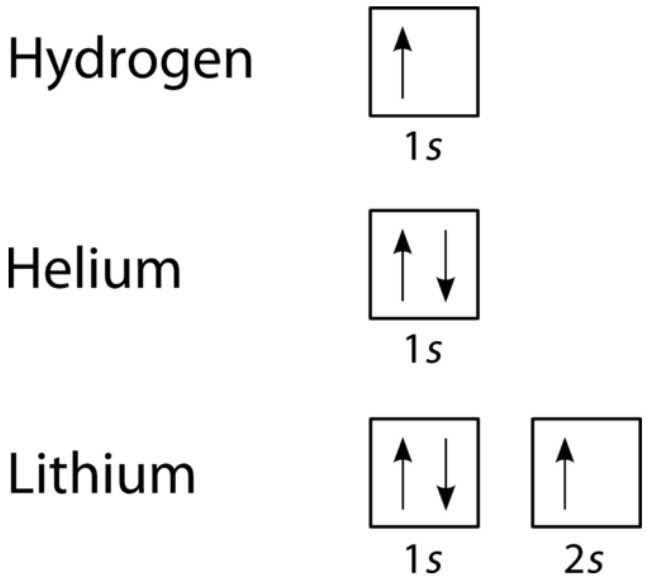


Figure $6.8.3$: Orbital filling diagrams for hydrogen, helium, and lithium. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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

The filling diagram for carbon is shown in the figure below. There are two $2p$ electrons for carbon and each occupies its own $2p$ orbital.

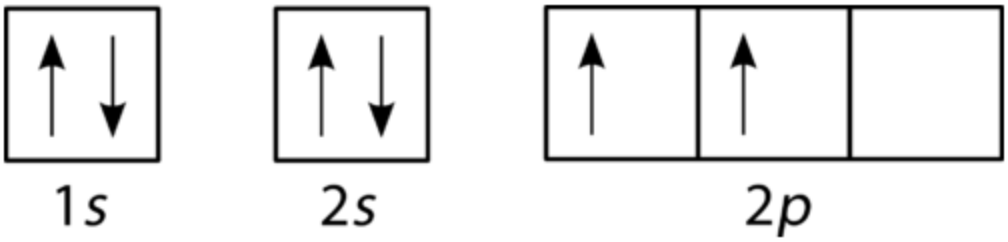


Figure 6.8.4: Orbital filling diagram for carbon. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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

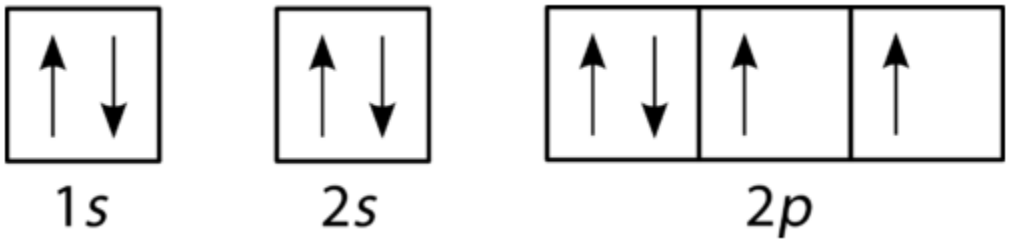


Figure 6.8.5: Orbital filling diagram for oxygen. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Summary

- Hund's rule specifies the order of electron filling within a set of orbitals.
- Orbital filling diagrams are a way of indicating electron locales in orbitals.

Review

1. State Hund's rule.
2. What is an orbital filling diagram?
3. Is the diagram in figure below correct? Explain your answer.

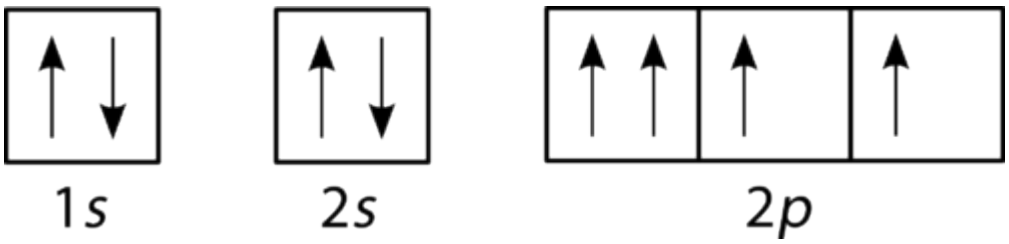


Figure 6.8.6 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

4. Is the diagram in figure below correct? Explain your answer.

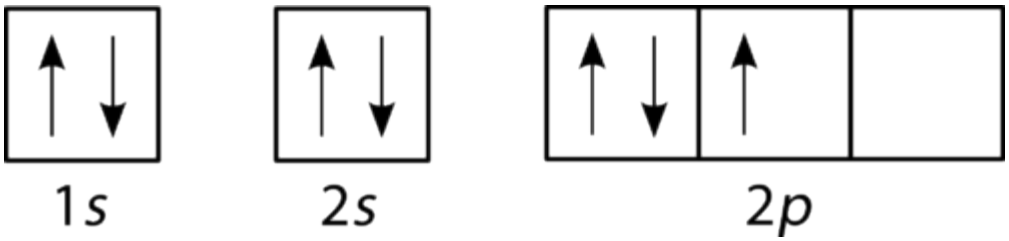


Figure 6.8.7 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

This page titled 6.8: Hund's Rule and Orbital Filling Diagrams is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.17: Hund's Rule and Orbital Filling Diagrams by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.9: Aufbau Principle



Figure 6.9.1 (Public Domain; Gary Minnaert via Commons Wikimedia, File:LACMA BCAM02.jpg [commons.wikimedia.org])

How are buildings constructed?

Construction of a building begins at the bottom. The foundation is laid, and the building goes up step by step. You obviously cannot start with the roof, since there is no place to hang it. The building goes from the lowest level to the highest level in a systematic way.

Aufbau Principle

In order to create ground state electron configurations for any element, it is necessary to know the way in which the atomic sublevels are organized in order of increasing energy. Figure 6.9.2 shows the order of increasing energy of the sublevels.

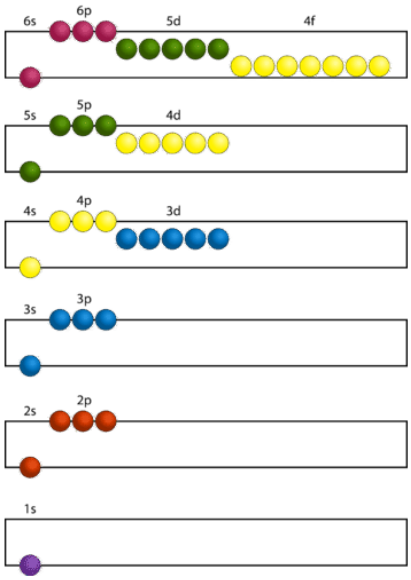


Figure 6.9.2: Electrons are added to atomic orbitals in order from low energy (bottom of the graph) to high (top of the graph), according to the Aufbau principle. Principle energy levels are color coded, while sublevels are grouped together, and each circle represents an orbital capable of holding two electrons. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The lowest energy sublevel is always the $1s$ sublevel, which consists of one orbital. The single electron of the hydrogen atom will occupy the $1s$ orbital when the atom is in its ground state. As we proceed to atoms with multiple electrons, those electrons are added to the next lowest sublevel: $2s$, $2p$, $3s$, and so on. The **Aufbau principle** states that an electron occupies orbitals in order from lowest energy to highest. The Aufbau (German for *building up, construction*) principle is sometimes referred to as the "building up" principle. It is worth noting that in reality, atoms are not built by adding protons and electrons one at a time, and that this method is merely an aid to understand the end result.

As seen in the figure above, the energies of the sublevels in different principal energy levels eventually begin to overlap. After the $3p$ sublevel, it would seem logical that the $3d$ sublevel should be the next lowest in energy. However, the $4s$ sublevel is slightly lower in energy than the $3d$ sublevel and thus fills first. Following the filling of the $3d$ sublevel is the $4p$, then the $5s$ and the $4d$. Note that the $4f$ sublevel does not fill until just after the $6s$ sublevel. Figure $6.9.2$ is a useful and simple aid for keeping track of the order of fill of the atomic sublevels.

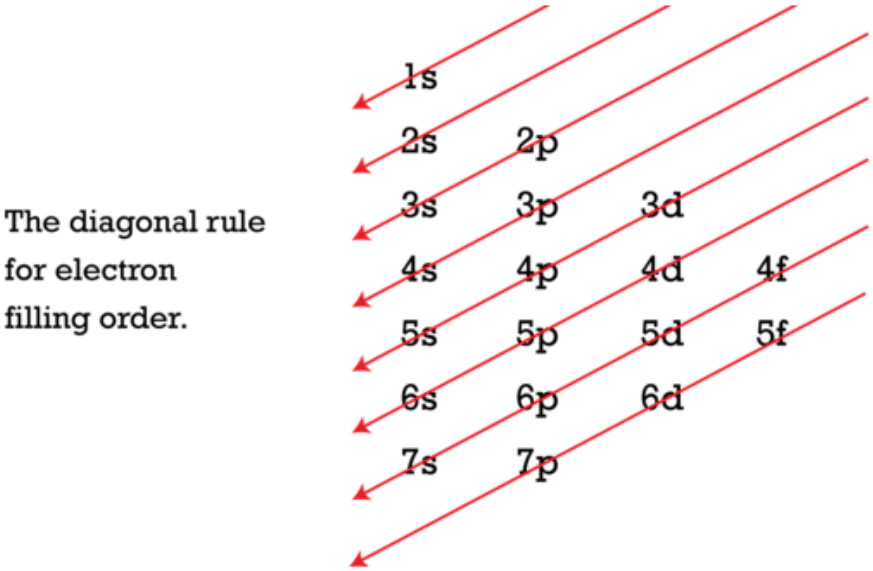


Figure 6.9.3: The Aufbau principle is illustrated in the diagram by following each red arrow in order from top to bottom: $1s$, $2s$, $2p$, $3s$, etc. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Summary

- The Aufbau principle gives the order of electron filling in an atom.
- It can be used to describe the locations and energy levels of every electron in a given atom.

Review

1. What is the Aufbau principle?
2. Which orbital is filled after the $2p$?
3. Which orbital is filled after $4s$?
4. Which orbital is filled after $6s$?

This page titled 6.9: Aufbau Principle is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.15: Aufbau Principle by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.10: Periodic Trends- Atomic Radius



Figure 6.10.1 (CC by 2.0; James Cridland via Flickr)

How can all of these people fit in such a small space?

Events draw large numbers of people to them. Even an outdoor event can fill up so that there is no room for more people. The crowd capacity depends on the amount of space in the venue, and the amount of space depends on the size of the objects filling it. We can get more people into a given space than can elephants, because elephants are larger than people. We can get more squirrels into that same space than we can people for the same reason. Knowing the sizes of objects to be dealt with can be important in deciding how much space is needed.

Atomic Radius

The size of atoms is important to explanations of the behavior of atoms or compounds. One way to express the size of atoms is by use of **atomic radius**. This data helps us understand why some molecules fit together and why other molecules have parts that get too crowded under certain conditions.

The size of an atom is defined by the edge of its orbital. However, orbital boundaries are fuzzy, and variable under different conditions. In order to standardize the measurement of atomic radii, the distance between the nuclei of two identical atoms bonded together is measured. The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together.

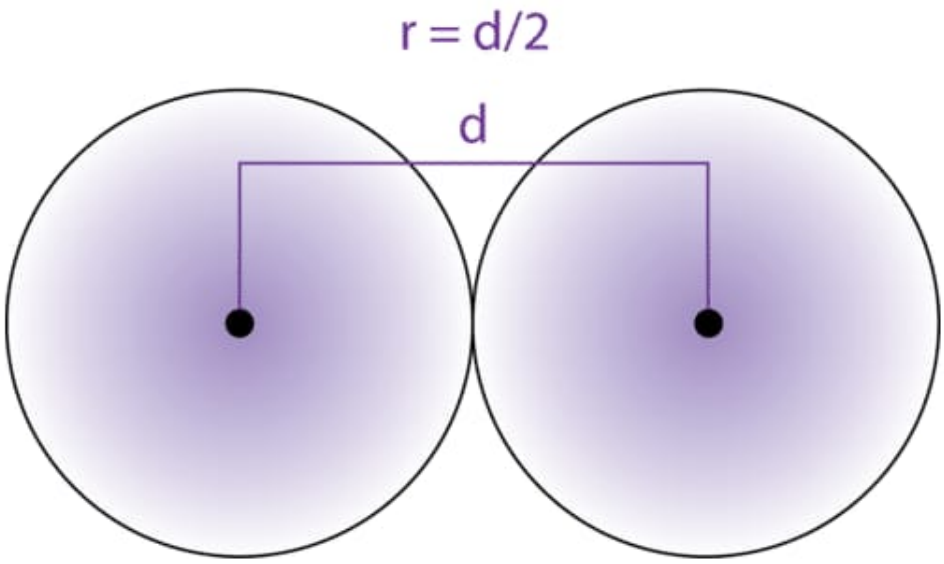


Figure 6.10.2: The atomic radius (r) of an atom can be defined as one half the distance (d) between two nuclei in a diatomic molecule. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Atomic radii have been measured for elements. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be $74\ \text{pm}$. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = 37\ \text{pm}$.

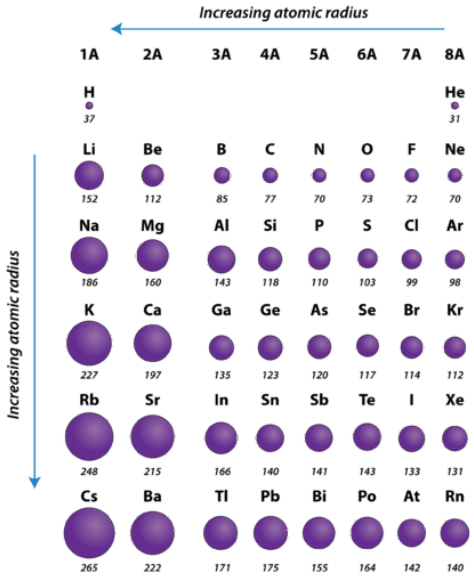


Figure 6.10.3: Atomic radii of the representative elements measured in picometers. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Periodic Trend

The atomic radius of atoms generally decreases from left to right across a period. There are some small exceptions, such as the oxygen radius being slightly greater than the nitrogen radius. Within a period, protons are added to the nucleus as electrons are being added to the same principal energy level. These electrons are gradually pulled closer to the nucleus because of its increased positive charge. Since the force of attraction between nuclei and electrons increases, the size of the atoms decreases. The effect lessens as one moves further to the right in a period, because of electron-electron repulsions that would otherwise cause the atom's size to increase.

Group Trend

The atomic radius of atoms generally increases from top to bottom within a group. As the atomic number increases down a group, there is again an increase in the positive nuclear charge. However, there is also an increase in the number of occupied principal energy levels. Higher principal energy levels consist of orbitals which are larger in size than the orbitals from lower energy levels. The effect of the greater number of principal energy levels outweighs the increase in nuclear charge, and so atomic radius increases down a group.

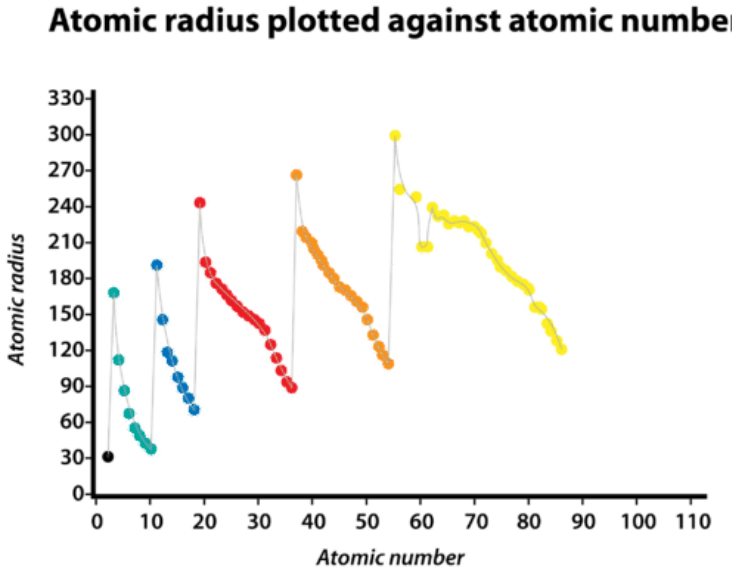


Figure 6.10.4: A graph of atomic radius plotted versus atomic number. Each successive period is shown in a different color. As the atomic number increases within a period, the atomic radius decreases. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Interactive Content

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

View this Interactive Content online

Summary

- Atomic radius is determined as half the distance between the nuclei of two identical atoms bonded together.
- The atomic radius of atoms generally decreases from left to right across a period.
- The atomic radius of atoms generally increases from top to bottom within a group.

Review

1. Define “atomic radius.”
2. What are the units of measurement for atomic radius?
3. How does the atomic radius of different elements change across a period?
4. How does atomic radius change from top to bottom within a group?
5. Explain why the atomic radius of hydrogen is so much smaller than the atomic radius of potassium.

This page titled 6.10: Periodic Trends- Atomic Radius is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.15: Periodic Trends- Atomic Radius by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.11: Metals



Figure 6.11.1 (Public Domain; User:Brianiac/Wikipedia via Wikipedia)

Can you guess what types of metals these screws are made of?

Screws come in all sizes and shapes. They are all (well, almost all) made of some kind of metal. But they have differences in size, shape, and type of metal. Physical characteristics also differ. Some screws are long, and others are short. One screw may have a flat-head slot while another screw may have a Phillips-head. Some of the screws in the picture below are used to fasten things together, and others are used to hang heavy objects on a wall.

Chemists classify materials in many ways. We can sort elements on the basis of their electron arrangements. The way the electrons are distributed determines the chemical properties of the element. Another way is to classify elements based on physical properties. Some common physical properties are color, volume, and density. Other properties that allow us to sort on the basis of behavior are conduction of heat and electricity, malleability (the ability to be hammered into very thin sheets), ductility (the ability to be pulled into thin wires), melting point, and boiling point. Three broad classes of elements based on physical properties are metals, nonmetals, and metalloids.

Metals

A metal is an element that is a good conductor of heat and electricity. Metals are also malleable, which means that they can be hammered into very thin sheets without breaking. They are ductile, which means that they can be drawn into wires. When a fresh surface of any metal is exposed, it will be very shiny because it reflects light well. This is called luster. All metals are solid at room temperature with the exception of mercury (Hg), which is a liquid. Melting points of metals display a very wide variance. The melting point of mercury is -39°C, while the highest melting metal is tungsten (W), with a melting point of 3422°C. The elements in blue in the periodic table below are metals. About 80 percent of the elements are metals.

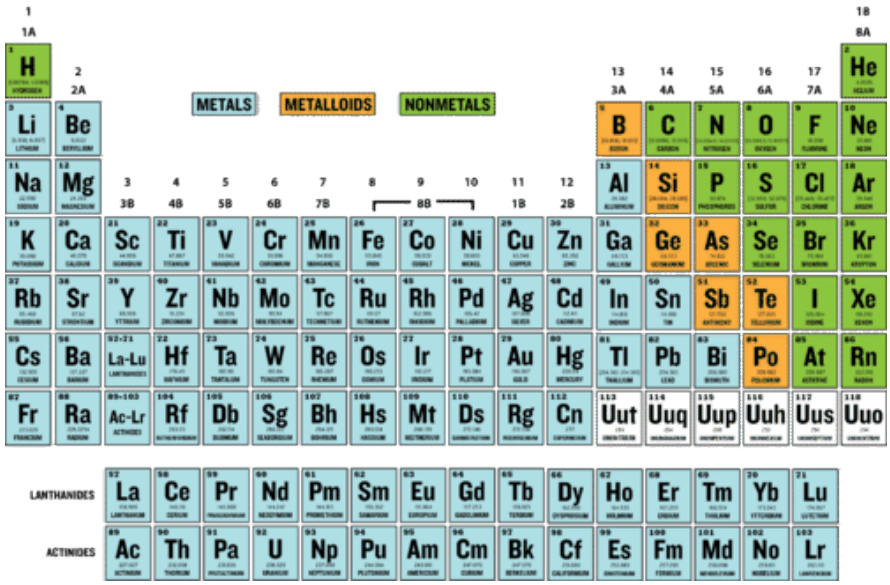


Figure 6.11.2 (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

Gold has been used by many civilizations for making jewelry (see figure below). This metal is soft and easily shaped into a variety of items. Since gold is very valuable and often used as currency, gold jewelry has also often represented wealth.



Figure 6.11.3: Gold jewelry. (Public Domain; User:Dorieo21/Wikimedia Commons via Wikipedia)

Copper is a good conductor of electricity and is very flexible and ductile. This metal is widely used to conduct electric current in a variety of appliances, from lamps to stereo systems to complex electronic devices (see figure below).



Figure 6.11.4: Copper wire exposed. (Public Domain; Scott Ehardt via Wikipedia)

Mercury is the only metal to exist as a liquid at room temperature (see figure below). This metal was extensively used in thermometers for decades until information about its toxicity became known. Mercury switches were once common, but are no longer used. However, new federally-mandated energy-efficient light bulbs that are now used contain trace amounts of mercury and represent a hazardous waste.



Figure 6.11.5: Pouring Mercury. (CC BY 3.0; User:Bionerd/Wikimedia Commons and User:Materialscientist/Wikimedia Commons via Wikipedia)

YouTube Video

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

View this YouTube Video online

Summary

- Metals are good conductors of heat and electricity.
- Metals are malleable and ductile
- All metals are solids at room temperature with the exception of mercury
- Gold, silver, iron, and mercury are typical metals.

Review

1. What properties of an element are affected by electron distribution?
2. Define malleability.
3. Define ductility.
4. State one reason gold is used in jewelry.
5. Why is mercury no longer used in many devices?

This page titled 6.11: Metals is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.5: Metals by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.12: Nonmetals



Figure 6.12.1 (Public Domain; Neville Street via Wikipedia)

When we sort parts in our shop or garage, we often classify them in terms of common properties. One container might hold all the screws (possibly sub-divided by size and type). Another container would be for nails. Maybe there is a set of drawers for plumbing parts.

When you get finished, you could also have a collection of things that don't nicely fit a category. You define them in terms of what they are not. They are not electrical components, or sprinkler heads for the yard, or parts for the car. These parts may have some common properties, but are a variety of items.

Nonmetals

In the chemical world, these "spare parts" would be considered nonmetals, loosely defined as not having the properties of metals. A **nonmetal** is an element that is generally a poor conductor of heat and electricity. Most properties of nonmetals are the opposite of metals. There is a wider variation in properties among the nonmetals than among the metals. Nonmetals exist in all three states of matter. The majority are gases, such as nitrogen and oxygen. Bromine is a liquid. A few are solids, such as carbon and sulfur. In the solid state, nonmetals are **brittle**, meaning that they will shatter if struck with a hammer. The solids are not lustrous. Melting points are generally much lower than those of metals. The green elements in the table below are nonmetals.

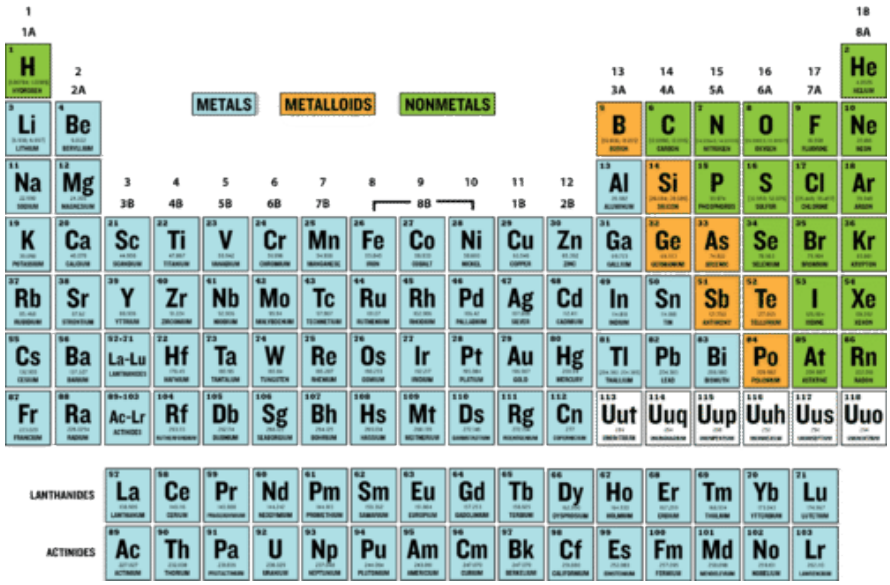


Figure 6.12.2: Copy and Paste Caption here. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

Nonmetals have a wide variety of uses. Sulfur can be employed in gunpowder, fireworks, and matches to facilitate ignition (see Figure below). This element is also widely used as an insecticide, a fumigant, or a means of eliminating certain types of fungus. An important role for sulfur is the manufacture of rubber for tires and other materials. First discovered in 1839 by Charles Goodyear, the process of vulcanization makes the rubber more flexible and elastic as well as being more resistant to changes in temperature. A major use of sulfur is for the preparation of sulfur-containing compounds such as sulfuric acid.



Figure 6.12.3: Sulfur. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Bromine is a versatile compound, used mainly in manufacture of flame-retardant materials, especially important for children's clothing (see Figure below). For treatment of water in swimming pools and hot tubs, bromine is beginning to replace chlorine as a disinfectant because of its higher effectiveness. When incorporated into compounds, bromine atoms play important roles in pharmaceuticals for treatment of pain, cancer, and Alzheimer's disease.



Figure 6.12.4: Bromine. (CC BY 3.0; User:Jurii/Wikimedia Commons via Wikipedia)

Helium is one of the many nonmetals that is a gas. Other nonmetal gases include hydrogen, fluorine, chlorine, and all the group eighteen noble (or inert) gases. Helium is chemically non-reactive, so it is useful for applications such as balloons (see figure below) and lasers, where non-flammability is extremely important. Liquid helium exists at an extremely low temperature and can be used to cool superconducting magnets for imaging studies (MRI, magnetic resonance imaging). Leaks in vessels and many types of high-vacuum apparatus can be detected using helium. Inhaling helium changes the speed of sound, producing a higher pitch in your voice. This is definitely an unsafe practice and can lead to physical harm and death.



Figure 6.12.5: Blimp. (Public Domain; Derek Jensen (Wikimedia: Tysto) via Wikipedia)

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Nonmetals are generally poor conductors of heat and electricity.
- Properties of nonmetals are usually the opposite of properties of metals
- Nonmetals can be solid, liquid, or gas at room temperature depending upon the element.
- Sulfur, bromine, and helium are typical nonmetals.

Review

1. What are the properties of nonmetals?
2. List the states of matter in which nonmetals can exist and give one example of each state.
3. What are the physical properties and uses of sulfur?
4. What are the physical properties and uses of bromine?
5. What are the physical properties and uses of helium?

This page titled 6.12: Nonmetals is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 6.6: Nonmetals by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

6.13: Metalloids

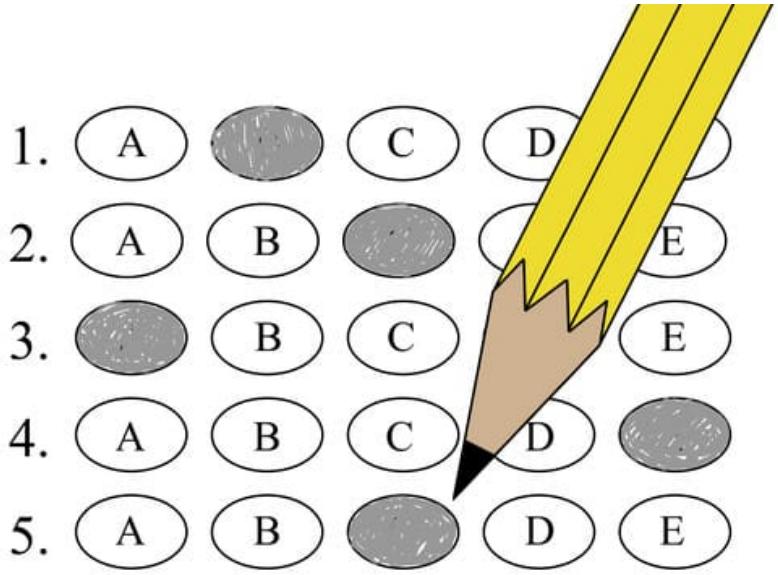


Figure 6.13.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

Have you ever taken a multiple-choice test?

Most of the time the answers are specific choices – is the answer possibility a or possibility b? Quite often you can "think through" the choices to come up with the correct answer. More frustrating is the choice "none of the above." You feel very uncertain checking that possibility.

Metalloids

Some elements are "none of the above." They don't fit neatly into the categories of metal or non-metal because of their characteristics. A **metalloid** is an element that has properties that are intermediate between those of metals and nonmetals. Metalloids can also be called semimetals. On the periodic table, the elements colored yellow, which generally border the stair-step line, are considered to be metalloids. Notice that aluminum borders the line, but it is considered to be a metal since all of its properties are like those of metals.

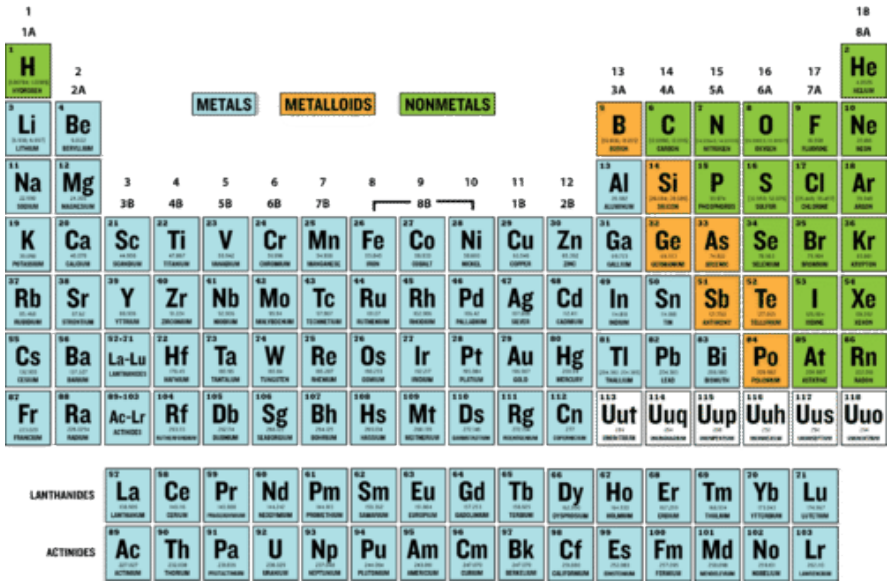


Figure 6.13.2 (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Examples of Metalloids

Silicon is a typical metalloid (see figure below). It has luster like a metal, but is brittle like a nonmetal. Silicon is used extensively in computer chips and other electronics because its electrical conductivity is in between that of a metal and a nonmetal.



Figure 6.13.3: Silicon. (Public Domain; User:Enricoros/Wikipedia via Wikipedia)

Boron is a versatile element that can be incorporated into a number of compounds (see figure below). Borosilicate glass is extremely resistance to thermal shock. Extreme changes in the temperature of objects containing borosilicates will not create any damage to the material, unlike other glass compositions, which would crack or shatter. Because of their strength, boron filaments are used as light, high-strength materials for airplanes, golf clubs, and fishing rods. Sodium tetraborate is widely used in fiberglass as insulation and also is employed in many detergents and cleaners.

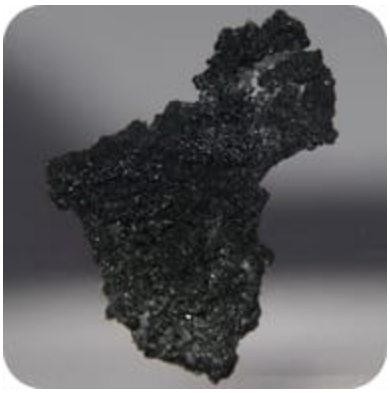


Figure 6.13.4: Boron. (CCSA;CC-BY 3.0; User:Jurii/Wikimedia Commons via Wikipedia)

Arsenic has long played a role in murder mysteries, being used to commit the foul deed (see figure below). This use of the material is not very smart since arsenic can be easily detected on autopsy. We find arsenic in pesticides, herbicides, and insecticides, but the use of arsenic for these applications is decreasing due to the toxicity of the metal. Its effectiveness as an insecticide has led arsenic to be used as a wood preservative.

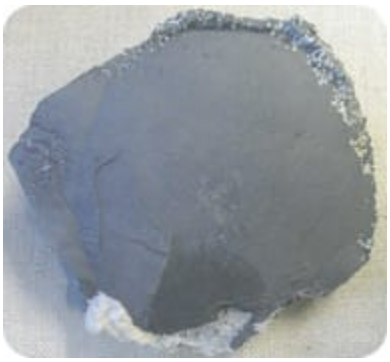


Figure 6.13.5: Arsenic. (Public Domain; Aram Dulyan (Wikimedia: Aramgutang) via Wikipedia)

Antimony is a brittle, bluish-white metallic material that is a poor conductor of electricity (see figure below). Used with lead, antimony increases the hardness and strength of the mixture. This material plays an important role in the fabrication of electronic and semiconductor devices. About half of the antimony used industrially is employed in the production of batteries, bullets, and alloys.



Figure 6.13.6: Antimony. (Public Domain; Aram Dulyan (User:Aramgutang/Wikimedia Commons) via Wikipedia)

Summary

- Metalloids are elements with properties intermediate between those of metals and non-metals
- Silicon is a metalloid because it has luster, but is brittle.
- Boron, arsenic, and antimony are metalloids with a variety of uses.

Review

1. Define “metalloid.”
2. Why would it be difficult to decide whether or not an element was a metalloid based on its properties?
3. Why is silicon used extensively in electronics?
4. What are borosilicates used for?
5. Why is the use of arsenic as an insecticide decreasing?
6. What is a main application of antimony?

This page titled 6.13: Metalloids is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

7: CHEMICAL BONDS

7.1: Chemical Bond
7.2: Electron Dot Diagrams
7.3: Octet Rule
7.4: Cation Formation
7.5: Anion Formation
7.6: Ionic Bonding
7.7: Polyatomic Ions
7.8: Covalent Bond
7.9: Lewis Electron-Dot Structures
7.10: Single Covalent Bonds
7.11: Multiple Covalent Bonds
7.12: Valence Electrons
7.13: Coordinate Covalent Bond
7.14: Resonance
7.15: VSEPR Theory
7.16: Molecular Shapes- No Lone Pairs on Central Atoms
7.17: Molecular Shapes - Lone Pair(s) on Central Atom
7.18: Bond Polarity
7.19: Polar Molecules

This page titled 7: Chemical Bonds is shared under a license and was authored, remixed, and/or curated by LibreTexts.

7.1: Chemical Bond



Figure 7.1.1 (CC by 2.0; Flickr: Amelia Wells via Flickr)

Have you ever made cupcakes from scratch?

You mix together flour, sugar, eggs, and other ingredients to make the batter, put the batter into cupcake papers, and then put them into the oven to bake. The cupcakes that come out of the oven after baking are different from any of the individual ingredients that went into the batter. Like the ingredients that join together to make cupcakes, atoms of different elements can join together to form entirely different substances called compounds. In cupcakes, the eggs and other wet ingredients cause the dry ingredients to stick together. What causes elements to stick together in compounds? The answer is chemical bonds.

What Is a Chemical Bond?

A chemical bond is a force of attraction between atoms or ions. Bonds form when atoms share or transfer valence electrons. Valence electrons are the electrons in the outer energy level of an atom that may be involved in chemical interactions. Valence electrons are the basis of all chemical bonds.

Q: Why do you think that chemical bonds form?

A: Chemical bonds form because they give atoms a more stable arrangement of electrons.

Why Bonds Form

To understand why chemical bonds form, consider the common compound known as water, or H_2O. It consists of two hydrogen (H) atoms and one oxygen (O) atom. As you can see in the on the left side of the figure below, each hydrogen atom has just one electron, which is also its sole valence electron. The oxygen atom has six valence electrons. These are the electrons in the outer energy level of the oxygen atom.

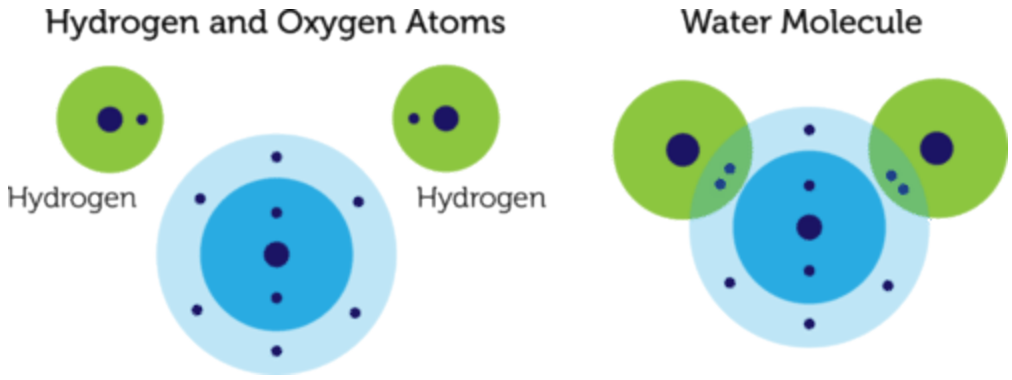


Figure 7.1.2: Copy and Paste Caption here. (CC by-NC 3.0; Christopher AuYeung via CK-12 Foundation)

In the water molecule on the right in the figure above, each hydrogen atom shares a pair of electrons with the oxygen atom. By sharing electrons, each atom has electrons available to fill its sole or outer energy level. The hydrogen atoms each have a pair of shared electrons, so their first and only energy level is full. The oxygen atom has a total of eight valence electrons, so its outer energy level is full. A full outer energy level is the most stable possible arrangement of electrons. It explains why elements form chemical bonds with each other.

Types of Chemical Bonds

Not all chemical bonds form in the same way as the bonds in water. There are actually three different types of chemical bonds, called covalent, ionic, and metallic bonds. Each type of bond is described below.

YouTube Video

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

View this YouTube Video online

- A covalent bond is the force of attraction that holds together two nonmetal atoms that share a pair of electrons. One electron is provided by each atom, and the pair of electrons is attracted to the positive nuclei of both atoms. The water molecule represented in Figure 7.1.2 contains covalent bonds.
- An ionic bond is the force of attraction that holds together oppositely charged ions. Ionic bonds form crystals instead of molecules. Table salt contains ionic bonds.
- A metallic bond is the force of attraction between a positive metal ion and the valence electrons that surround it—both its own valence electrons and those of other ions of the same metal. The ions and electrons form a lattice-like structure. Only metals, such as the copper pictured in Figure 7.1.3, form metallic bonds.



Figure 7.1.3: Metallic bonds explain many of the properties of metals. This coil of wire is made of the metal copper. Like other metals, copper is shiny, can be formed into wires, and conducts electricity. (CC BY 2.0; Flavio~ - "Industrial Photography: Copper Wires-003" via Flickr)

Summary

- A chemical bond is a force of attraction between atoms or ions. Bonds form when atoms share or transfer valence electrons.
- Atoms form chemical bonds to achieve a full outer energy level, which is the most stable arrangement of electrons.
- There are three different types of chemical bonds: covalent, ionic, and metallic bonds.

Review

1. What is a chemical bond?
2. Explain why hydrogen and oxygen atoms are more stable when they form bonds in a water molecule.
3. How do ionic bonds and covalent bonds differ?

Explore More

Watch this video about covalent bonds, and then answer the questions below.

YouTube Video

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

View this YouTube Video online

1. Which types of elements can form covalent bonds?
2. How can you tell the number of covalent bonds an element can form?
3. Why does one atom of nitrogen form bonds with three atoms of hydrogen?

This page titled 7.1: Chemical Bond is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

7.2: Electron Dot Diagrams

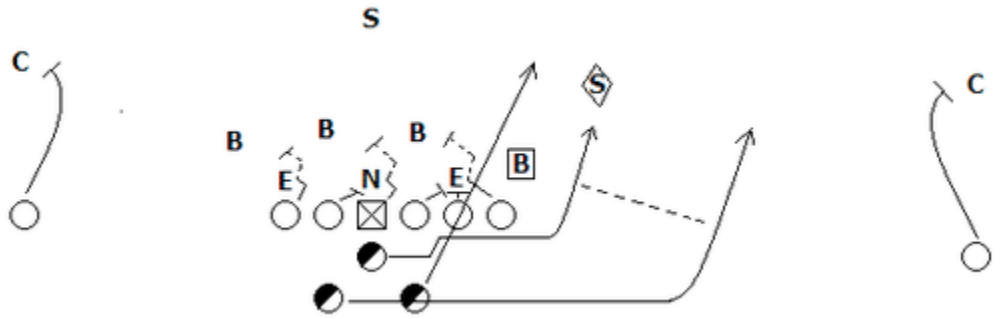


Figure 7.2.1 (Public Domain; User:Veatchw/Wikipedia via Commons Wikimedia, Veer vs 34 [commons.wikimedia.org])

How do we show electrons in atoms?

Diagrams contain a lot of useful information in a compact format. What does the diagram above tell us? The football play diagrammed above describes the lineup of each player on the team and describes how they will move when the ball is snapped. Diagrams of electrons give similar information about where certain electrons are. We can mark these electrons and indicate what happens to them when an element reacts.

Electron Dot Diagrams

Recall that the valence electrons of an atom are the electrons located in the highest occupied principal energy level. Valence electrons are primarily responsible for the chemical properties of elements. The number of valence electrons can be easily determined from the electron configuration. Several examples from the second period elements are shown in the table below.

Table 7.2.1:

Element	Electron Configuration	Number of Valence Electrons
lithium	$1s^2\ 2s^1$	1
beryllium	$1s^2\ 2s^2$	2
nitrogen	$1s^2\ 2s^2\ 2p^3$	5
neon	$1s^2\ 2s^2\ 2p^6$	8

In each case, valence electrons are those in the second principal energy level. As one proceeds left to right across a period, the number of valence electrons increases by one. In the s block, Group 1 elements have one valence electron, while Group 2 elements have two valence electrons. In the p block, the number of valence electrons is equal to the group number minus ten. Group 13 has three valence electrons, Group 14 has four, and so on...up through Group 18 with eight. The eight valence electrons, a full outer s and p sublevel, give the noble gases their special stability.

When examining chemical bonding, it is necessary to keep track of the valence electrons of each atom. **Electron dot diagrams** are diagrams in which the valence electrons of an atom are shown as dots distributed around the element's symbol. A beryllium atom, with two valence electrons, has the electron dot diagram below:



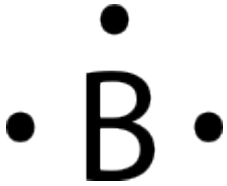
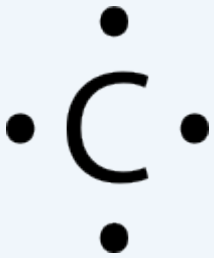
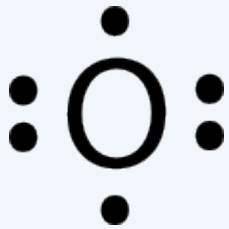
Figure 7.2.2 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

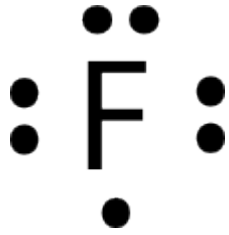
Since electrons repel each other, the dots for a given atom are distributed evenly around the symbol before they are paired. The

table below shows the electron dot diagrams for the entire second period.

Table 7.2.2: Electron Dot Diagrams for the Second Period Elements

Electron Dot Diagrams for the Second Period Elements

Group Number	Electron Dot Diagram
1	 Figure 7.2.3 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
2	 Figure 7.2.4 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
13	 Figure 7.2.5 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
14	 Figure 7.2.6 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
15	 Figure 7.2.7 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
16	 Figure 7.2.8 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

Group Number	Electron Dot Diagram
17	 Figure 7.2.9 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
18	 Figure 7.2.10 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

Electron dot diagrams would be the same for each element in the representative element groups. Most transition elements have two valence electrons, though some that have unusual electron configurations have only one.

YouTube Video

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

View this YouTube Video online

Summary

- Electron dot diagrams show the valence electrons for an atom.
- The dot diagrams are the same for each element in the representative element groups.

Review

1. What are valence electrons primarily responsible for?
2. Calcium would have the same electron dot structure as which element pictured in the table?
3. What is the symbol for an element that would have the same electron dot structure as carbon?
4. Would you expect the group 18 elements to have the same electron dot diagram as neon?

This page titled 7.2: Electron Dot Diagrams is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 8.1: Electron Dot Diagrams by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.3: Octet Rule



Figure 7.3.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How are electrons organized in atoms?

Graduations are exciting events in one's life. The processional, the ceremonies, even the speeches can be handled because this marks the end of a stage of life and the beginning of another one. The ceremony itself is challenging for those organizing it. There are just enough seats on the stage for the graduates. How do you avoid getting too many people in a row and not enough in the next row? Someone is stationed at the end of the row to count the students as they enter. Only so many are allowed to go into a row, and then you begin to fill the next row. Electrons in atoms behave the same way. There are rules that determine where electrons go in compounds.

Octet Rule

The noble gases are unreactive because of their electron configurations. The noble gas neon has the electron configuration of $1s^2\ 2s^2\ 2p^6$. It has a full outer shell and cannot incorporate any more electrons into the valence shell. The other noble gases have the same outer shell electron configuration, even though they have different numbers of inner-shell electrons.



Figure 7.3.2: Electron configuration of neon atom. (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

American chemist Gilbert Lewis (1875-1946) used this observation to explain the types of ions and molecules that are formed by other elements. He called his explanation the **octet rule**. The octet rule states that atoms tend to form compounds in ways that give them eight valence electrons, and thus the electron configuration of a noble gas. An exception to an octet of electrons is in the case of the first noble gas, helium, which only has two valence electrons. This primarily affects the element hydrogen, which forms stable compounds by achieving two valence electrons. Lithium, an alkali metal with three valence electrons, is also an exception to the octet rule. Lithium tends to lose one electron to take on the electron configuration of the nearest noble gas, helium, leaving it with two valence electrons.

There are two ways in which atoms can satisfy the octet rule. One way is by sharing their valence electrons with other atoms. The second way is by transferring valence electrons from one atom to another. Atoms of metals tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level. Atoms of nonmetals tend to gain electrons in order to fill their outermost principal energy level with an octet.

Watch the following video to learn how to use the octet rule to predict the charge of an ion.

YouTube Video

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

View this YouTube Video online

Summary

- Atoms form compounds in ways that give them eight valence electrons.
- Metals tend to lose electrons to achieve this configuration.
- Nonmetals tend to gain electrons to achieve this configuration.

Review

1. What is the electron configuration of a noble gas?
2. Why is this configuration important?
3. What are the two ways in which atoms can satisfy the octet rule?
4. How do metals change to obey the octet rule?
5. How do nonmetals change to obey the octet rule?

This page titled 7.3: Octet Rule is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 8.2: Octet Rule by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.4: Cation Formation



Figure 7.4.1 (Public Domain; User:Bugman/Wikipedia via Wikipedia)

How are cations formed?

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

Cation Formation

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

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

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

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

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

In this case, the magnesium atom loses its two valence electrons in order to achieve the same noble gas configuration. The aluminum atom loses its three valence electrons. The $\mathrm{Mg^{2+}}$ ion, the $\mathrm{Al^{3+}}$ ion, the $\mathrm{Na^+}$ ion, and the $ceNe$ atom are all isoelectronic. For representative elements under typical conditions, three electrons are the maximum number that will be lost.

We can also show the loss of valence electron(s) with an electron dot diagram:

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

$$\mathrm{Na}\cdot \rightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$
$$\cdot\mathrm{Mg}\cdot \rightarrow \blacksquare + 2\mathrm{e}^{-}$$

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Cations form when an atom loses one or more electrons.
- The resulting cation has the electron configuration of the noble gas atom in the row above it in the periodic table.

Review

1. What is a cation?
2. How many valence electrons does the sodium atom have?
3. Which atom is the sodium ion isoelectronic with?
4. How many electrons does magnesium lose to form the magnesium ion?

This page titled 7.4: Cation Formation is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 8.3: Cation Formation by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.5: Anion Formation



Figure 7.5.1: A pressed glass open salt dish made in the early 1830s. (Public Domain; User:Daderot/Wikimedia Commons via Wikipedia)

How do you make chlorine safe to eat?

How do you transform a deadly gas into something you can sprinkle on your eggs and eat for breakfast? Chlorine in its free form is very dangerous if you breathe the fumes or come in contact with the gas. However, in reaction with sodium, sodium chloride forms as the sodium atom gives up an electron to chlorine, which accepts the electron to form the chloride anion.

Anions

Anions are the negative ions formed from the gain of one or more electrons. When nonmetal atoms gain electrons, they often do so until their outermost principal energy level achieves an octet. This process is illustrated below for the elements fluorine, oxygen, and nitrogen:

$$\begin{array}{lcl} \mathrm{F} + \mathrm{e}^- & \rightarrow & \mathrm{F}^- \\ 1s^2\ 2s^2\ 2p^5 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{lcl} \mathrm{O} + 2\mathrm{e}^- & \rightarrow & \blacksquare \\ 1s^2\ 2s^2\ 2p^4 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

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

All of these anions are isoelectronic with each other and with neon. They are also isoelectronic with the three cations from the previous section. Under typical conditions, three electrons are the maximum that will be gained in the formation of anions.

Outer electron configurations are constant within a group, so this pattern of ion formation repeats itself for Periods 3, 4, and following (see below).

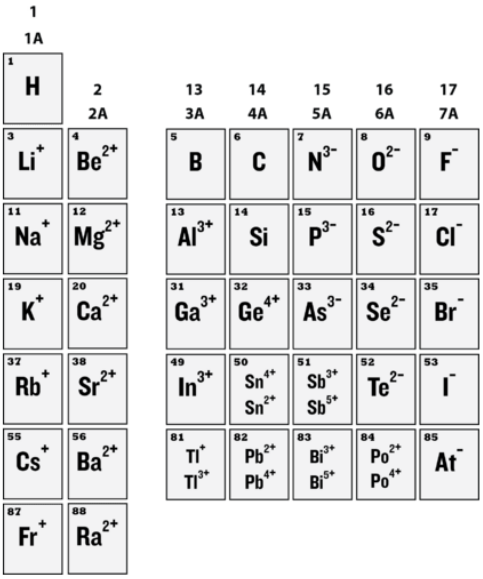


Figure 7.5.2: Ion charges. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

It is important not to misinterpret the concept of being isoelectronic. A sodium ion is very different from a neon atom, because the nuclei of the two contain different numbers of protons. One is an essential ion that is a part of table salt, while the other is an unreactive gas that is a very small part of the atmosphere. Likewise, sodium ions are very different than magnesium ions, fluoride ions, and all the other members of this isoelectronic series ([redacted], [redacted], F^-, Ne, Na^+, [redacted], [redacted]).

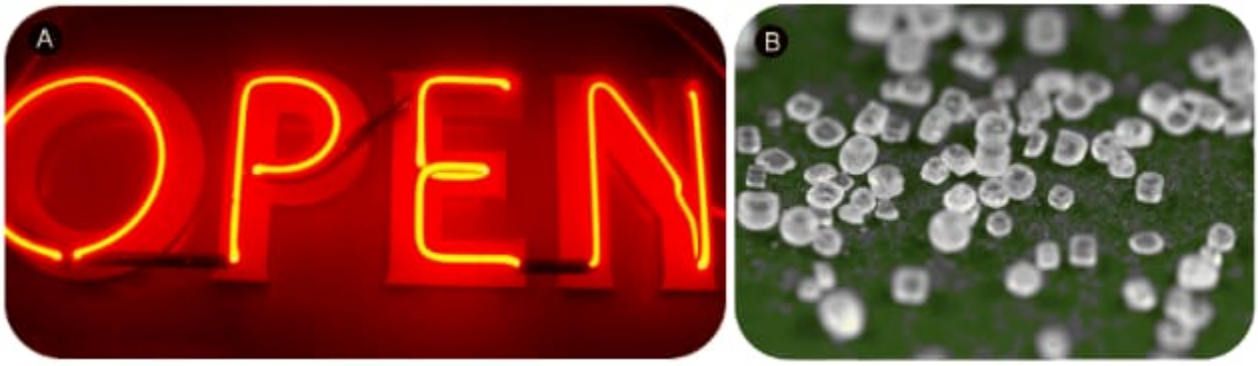


Figure 7.5.3: Neon gas (A) and sodium chloride crystals (B). Neon atoms and sodium ions are isoelectronic. Neon is a colorless and unreactive gas that glows a distinctive red-orange color in a gas discharge tube. Sodium ions are most commonly found in crystals of sodium chloride—ordinary table salt. (CC BY 2.0; (A) Andy Wright (Flickr: rightee); (B) Kevin Dooley;User:Bugman/Wikipedia via (A) http://www.flickr.com/photos/rightee/4356950; (B) http://www.flickr.com/photos/pagedooley/2769134850; http://commons.wikimedia.org/wiki/File:Ion_exchange_resin_beads.jpg)

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Anions are negative ions formed by accepting electrons.
- The outermost principal energy level of an anion is usually an octet.

Review

1. What is an anion?
2. Write the electronic configurations for the chlorine atom and the chloride anion.
3. What does isoelectronic mean?

This page titled 7.5: Anion Formation is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 8.4: Anion Formation by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.6: Ionic Bonding



Figure 7.6.1 (Public Domain; Leon Brooks via Commons Wikimedia, Waneroo Beach [commons.wikimedia.org])

Does the sea really have salt in it?

We can get common table salt from several sources. It can be mined in the solid form in salt mines, or found as a solid in deposits. We can also get salt from the ocean, but it really does not exist as a salt when in solution. The sodium ions and chloride ions are dissolved, but not combined into a structure until all the water is removed.

Most of the rocks and minerals that make up the Earth's crust are composed of positive and negative ions held together by ionic bonding. An ionic compound is an electrically neutral compound consisting of positive and negative ions. You are very familiar with some ionic compounds, such as sodium chloride (NaCl). A sodium chloride crystal consists of equal numbers of positive sodium ions (Na^+) and negative chloride ions (Cl^-).

Ionic Bonds

Oppositely charged particles attract each other. This attractive force is often referred to as an **electrostatic force**. An **ionic bond** is the electrostatic force that holds ions together in an **ionic compound**. The strength of the ionic bond is directly dependent upon the quantity of the charges and inversely dependent on the distance between the charged particles. A cation with a $2+$ charge will make a stronger ionic bond than a cation with a $1+$ charge. A larger ion makes a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion.

YouTube Video

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

View this YouTube Video online

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

Electron Dot Diagrams

We will use sodium chloride as an example to demonstrate the nature of the ionic bond and how it forms. As you know, sodium is a metal and loses its one valence electron to become a cation. Chlorine is a nonmetal and gains one electron in becoming an anion. Both achieve a noble-gas electron configuration. However, electrons cannot be simply "lost" to nowhere in particular. A more accurate way to describe what is happening is that a single electron is transferred from the sodium atom to the chlorine atom, as shown below.

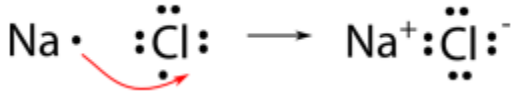


Figure 7.6.2 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The ionic bond is the attraction of the Na^{+} ion for the Cl^{-} ion. It is conventional to show the cation without dots around the symbol to emphasize that the original energy level that contained the valence electron is now empty. The anion is now shown with a complete octet of electrons.

For a compound such as magnesium chloride, it is not quite as simple. Because magnesium has two valence electrons, it needs to lose both to achieve the noble-gas configuration. Therefore, two chlorine atoms will be needed.

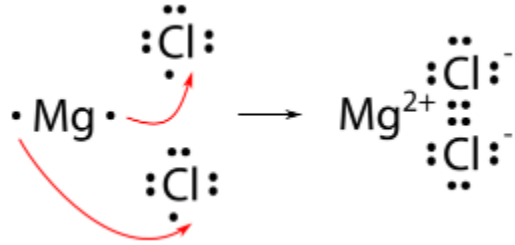


Figure 7.6.3 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The final formula for magnesium chloride is MgCl_2.

Summary

- An ionic compound contains positive and negative ions.
- An ionic bond is electrostatic in nature.
- Electron dot diagrams can be used to illustrate electron movements and ion formation.

Review

1. What is an ionic compound?
2. What is an ionic bond?
3. Which cation (Na^+ or Ca^{2+}) would form a stronger ionic bond with Cl^-?

This page titled 7.6: Ionic Bonding is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 8.6: Ionic Bonding by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.7: Polyatomic Ions



Figure 7.7.1 (Public Domain; Huhu via Wikipedia)

Have you ever read the story of Romeo and Juliet?

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

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

Polyatomic Ions

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

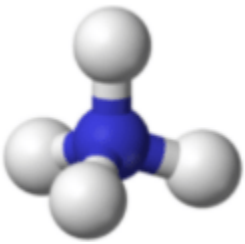


Figure 7.7.2: The ammonium ion $(\mathrm{NH_4}^+)$ is a nitrogen atom (blue) bonded to four hydrogen atoms (white). (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)



Figure 7.7.3: The hydroxide ion (OH^-) is an oxygen atom (red) bonded to a hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

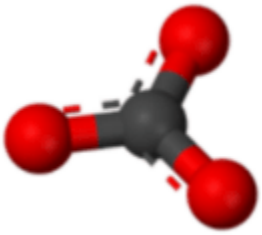


Figure 7.7.4: The carbonate ion (CO_3^{2-}) is a carbon atom (black) bonded to three oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

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

Table 7.7.1: Common Polyatomic Ions

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

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

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

CK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

- ClO_4^-, perchlorate

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

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 7.7: Polyatomic Ions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.9: Polyatomic Ions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.8: Covalent Bond



Figure 7.8.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

In a tennis match, two players keep hitting the ball back and forth. The ball bounces from one player to the other, over and over again. The ball keeps the players moving together on the court. What if the two players represented the nuclei of two atoms and the ball represented valence electrons? What would the back and forth movement of the ball represent? The answer is a covalent bond.

Sharing Electrons

A covalent bond is the force of attraction that holds together two atoms that share a pair of valence electrons. The shared electrons are attracted to the nuclei of both atoms. This forms a molecule consisting of two or more atoms. Covalent bonds form only between atoms of nonmetals.

Covalent Compounds and Diatomic Elements

The two atoms that are held together by a covalent bond may be atoms of the same element or different elements. When atoms of different elements form covalent bonds, a new substance, called a covalent compound, results. Water is an example of a covalent compound. A water molecule is modeled in the figure below. A molecule is the smallest particle of a covalent compound that still has the properties of the compound.

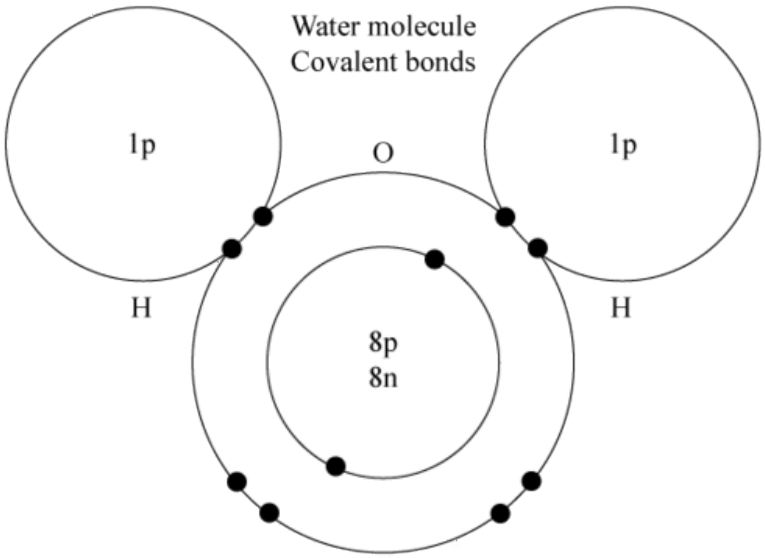


Figure 7.8.2 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: How many valence electrons does the oxygen atom (O) share with each hydrogen atom (H)? How many covalent bonds hold the water molecule together?

A: The oxygen atom shares one pair of valence electrons with each hydrogen atom. Each pair of shared electrons represents one covalent bond, so two covalent bonds hold the water molecule together.

The diagram in the figure below shows an example of covalent bonds between two atoms of the same element, in this case two atoms of oxygen. The diagram represents an oxygen molecule, so it's not a new compound. Oxygen normally occurs in diatomic ("two-atom") molecules. Several other elements also occur as diatomic molecules: hydrogen, nitrogen, and all but one of the halogens (fluorine, chlorine, bromine, and iodine).

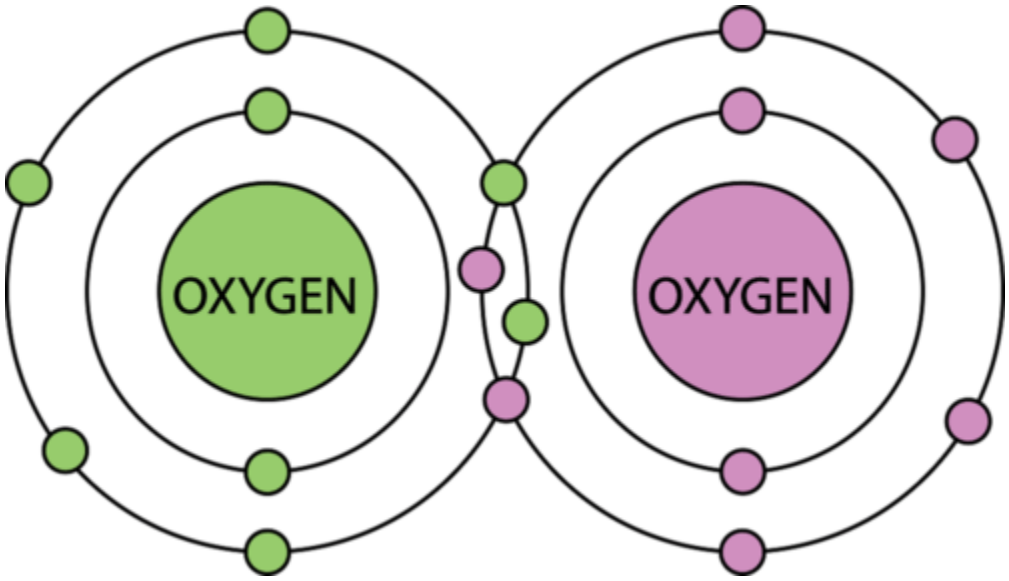


Figure 7.8.3 (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Q: How many electrons do these two oxygen atoms share? How many covalent bonds hold the oxygen molecule together?

A: The two oxygen atoms share two pairs of electrons, so two covalent bonds hold the oxygen molecule together.

Why Covalent Bonds Form

Covalent bonds form because they give atoms a more stable arrangement of electrons. Look at the oxygen atoms in the figure above. Alone, each oxygen atom has six valence electrons. By sharing two pairs of valence electrons, each oxygen atom has a total of eight valence electrons. This fills its outer energy level, giving it the most stable arrangement of electrons. The shared electrons are attracted to both oxygen nuclei, and this force of attraction holds the two atoms together in the oxygen molecule.

Summary

- A covalent bond is the force of attraction that holds together two atoms that share a pair of valence electrons. Covalent bonds form only between atoms of nonmetals.
- The two atoms that are held together in a covalent bond may be atoms of the same element or different elements. When atoms of different elements bond together, it forms a covalent compound.
- Covalent bonds form because the shared electrons fill each atom's outer energy level and this is the most stable arrangement of electrons.

Review

1. What is a covalent bond?
2. Nitrogen is a diatomic element with five valence electrons. Create a model of a molecule of nitrogen.
3. Which of the following represents a covalent compound?
   a. O_2
   b. CO_2
   c. Cl_2
   d. NaCl

Explore More

Watch the video about covalent bonding and then answer the questions below.

YouTube Video

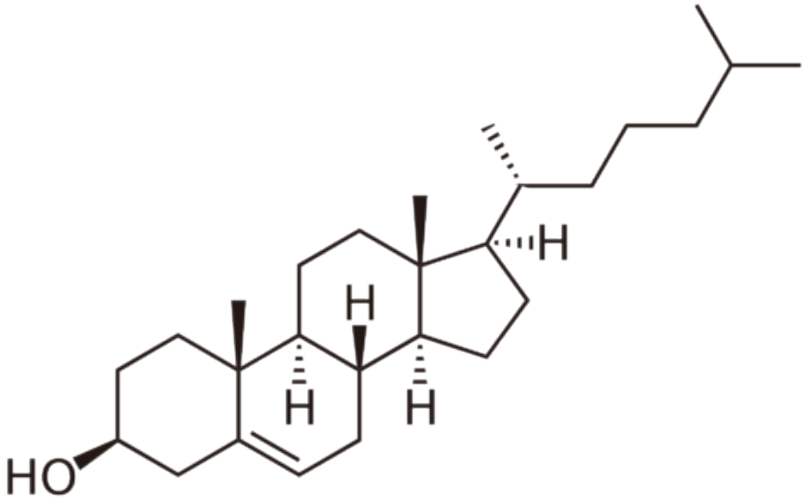
This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

1. How can you tell the number covalent bonds the atoms of an element can form?
2. How many covalent bonds can nitrogen form? How many covalent bonds can chlorine form?
3. Why are covalent bonds stronger than ionic bonds?
4. Why can't molecules of a covalent compound conduct electricity?

This page titled 7.8: Covalent Bond is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

7.9: Lewis Electron-Dot Structures



Figure

(Public Domain; User:BorisTM/Wikimedia Commons via Wikipedia).

What does cholesterol really look like?

We can write the structure of the cholesterol molecule a couple of different ways. The simplest approach is to simply write $C_{27}H_{46}O$. This "structure" is not very useful because it does not tell us how the carbons, hydrogens, and oxygen are connected to one another. The structure in the figure above is much more helpful—we see how the different atoms are connected together to form the molecule.

Lewis Electron-Dot Structures

In a previous chapter, you learned that the valence electrons of an atom can be shown in a simple way with an electron dot diagram. A hydrogen atom is shown as $H\cdot$ because of its one valence electron. The structures of molecules that are held together by covalent bonds can be diagrammed by **Lewis electron-dot structures**. The hydrogen molecule is shown in the figure below.

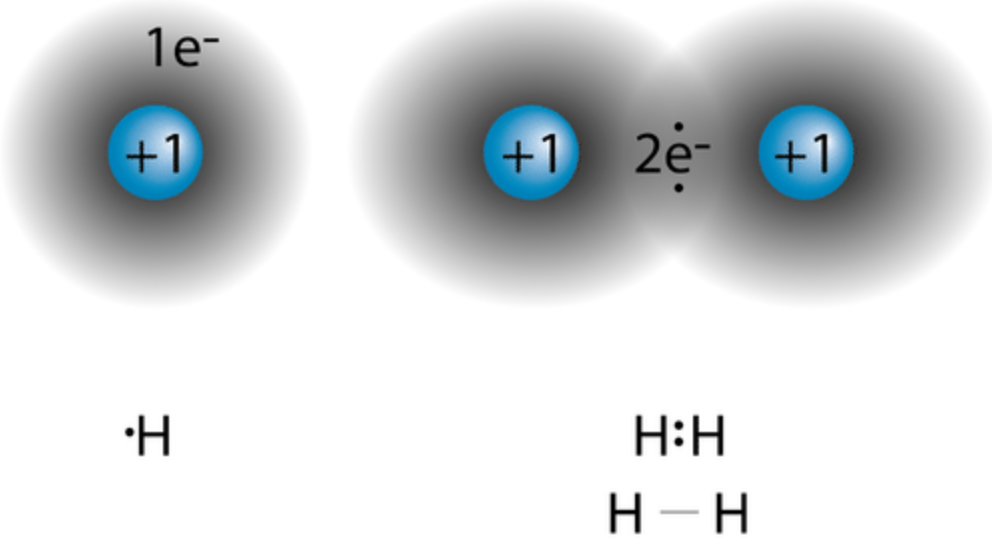


Figure 7.9.2: On the left is a single hydrogen atom with one electron. On the right is an H_2 molecule showing the electron cloud overlap. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The shared pair of electrons is shown as two dots in between the two H symbols $(H : H)$. This is called a **single covalent bond**, when two atoms are joined by the sharing of one pair of electrons. The single covalent bond can also be shown by a dash in between the two symbols $(H—H)$. A **structural formula** is a formula that shows the arrangement of atoms in a molecule and represents covalent bonds between atoms by dashes.

The Octet Rule and Covalent Bonds

When ions form, they conform to the **octet rule** by either losing or gaining electrons in order to achieve the electron configuration of the nearest noble gas. In a similar way, nonmetal atoms share electrons in the formation of a covalent bond in such a way that each of the atoms involved in the bond can attain a noble-gas electron configuration. The shared electrons are "counted" for each of the atoms involved in the sharing. For hydrogen (H_2), the shared pair of electrons means that each of the atoms is able to attain the electron configuration of helium, the noble gas with two electrons. For atoms other than hydrogen, the sharing of electrons will usually provide each of the atoms with eight valence electrons.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Lewis electron-dot structures show the bonding in covalent molecules.
- Covalent bonds between atoms can be indicated either with dots $(:)$ or a dash $(-)$.

Review

1. What is a single covalent bond?
2. How can covalently-bound atoms obey the octet rule?
3. Does the hydrogen molecule obey the octet rule?

This page titled 7.9: Lewis Electron-Dot Structures is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.5: Lewis Electron-Dot Structures by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.10: Single Covalent Bonds

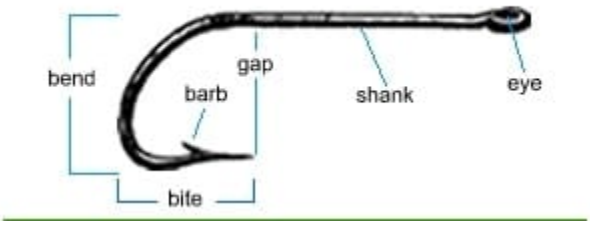


Figure 7.10.1 (Public Domain; the US Department of the Interior via Commons Wikimedia, Fishhook [commons.wikimedia.org])

What holds molecules together?

In one form or another, the idea of atoms connecting to form larger substances has been with us for a long time. The Greek philosopher Democritus (460-370 BCE) believed that atoms had hooks that allowed them to connect with one another. Today, we believe that atoms are held together by bonds formed when two atoms share a set of electrons—a much more complicated picture than the simple hooks that Democritus believed in.

Single Covalent Bonds

A covalent bond forms when two orbitals with one electron each overlap one another. For the hydrogen molecule, this can be shown as:

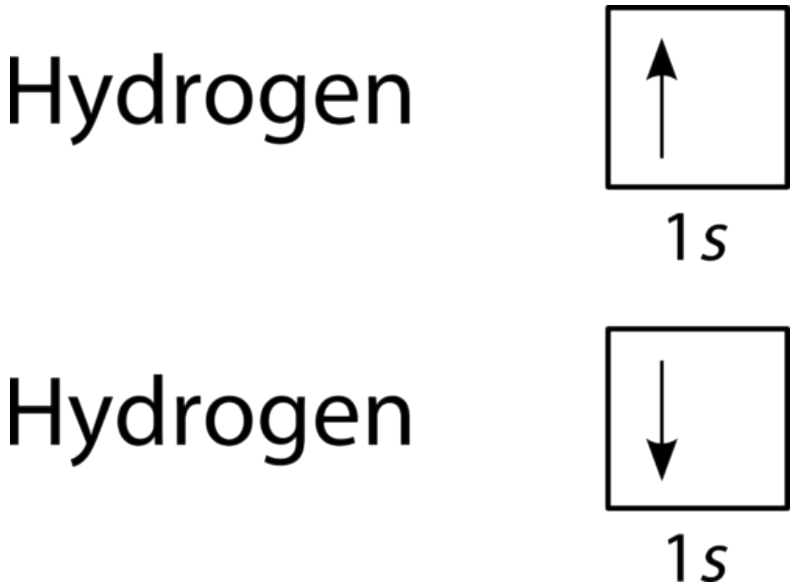


Figure 7.10.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Upon formation of the H_2 molecule, the shared electrons must have opposite spin, so they are shown with opposite spin in the atomic $1s$ orbital.

The halogens also form single covalent bonds in their diatomic molecules. An atom of any halogen, such as fluorine, has seven valence electrons. Its unpaired electron is located in the $2p$ orbital.

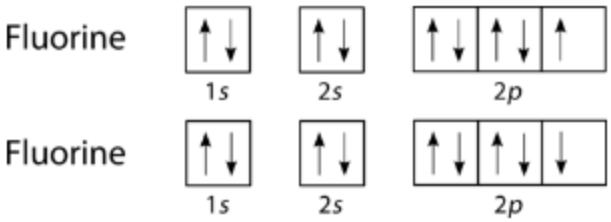


Figure 7.10.3 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The single electrons in the third $2p$ orbital combine to form the covalent bond:

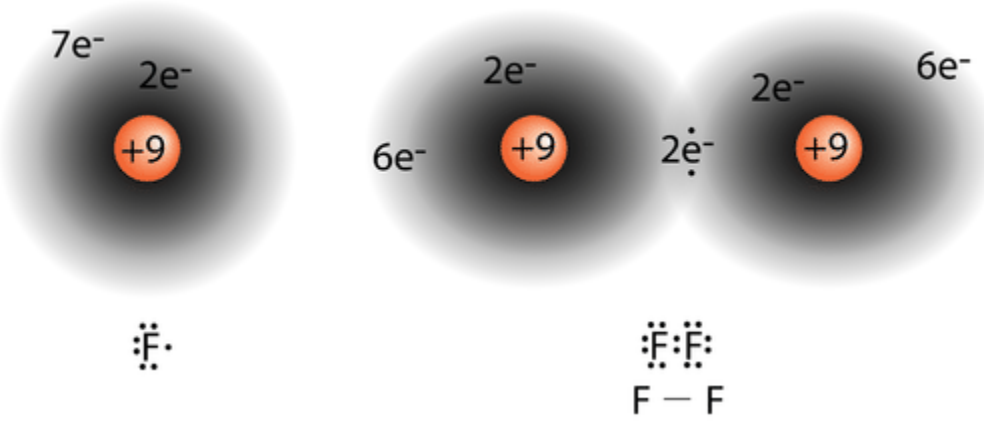


Figure 7.10.4: On the left is a fluorine atom with seven valence electrons. On the right is the F_2 molecule. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The diatomic fluorine molecule (F_2) contains a single shared pair of electrons. Each F atom also has three pair of electrons that are not shared with the other atom. A **lone pair** is a pair of electrons in a Lewis electron-dot structure that is not shared between atoms. The oxygen atom in the water molecule shown below has two lone pair sets of electrons. Each F atom has three lone pairs. Combined with the two electrons in the covalent bond, each F atom follows the octet rule.

YouTube Video

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

View this YouTube Video online

Example 7.10.1

Draw the Lewis electron dot structure for water.

Solution

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

Known

- Molecular formula of water $= H_2O$
- 1 O atom $= 6$ valence electrons
- 2 H atoms $= 2 \times 1 = 2$ valence electrons
- Total number of valence electrons $= 8$

Use the periodic table to determine the number of valence electrons for each atom and the total number of valence electrons. Arrange the atoms and distribute the electrons so that each atom follows the octet rule. The oxygen atom will have 8 electrons, while the hydrogen atoms will each have 2.

Step 2: Solve.

Electron dot diagrams for each atom are:



Figure 7.10.5 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Each hydrogen atom with its single electron will form a covalent bond with the oxygen atom where it has a single electron. The resulting Lewis electron dot structure is:

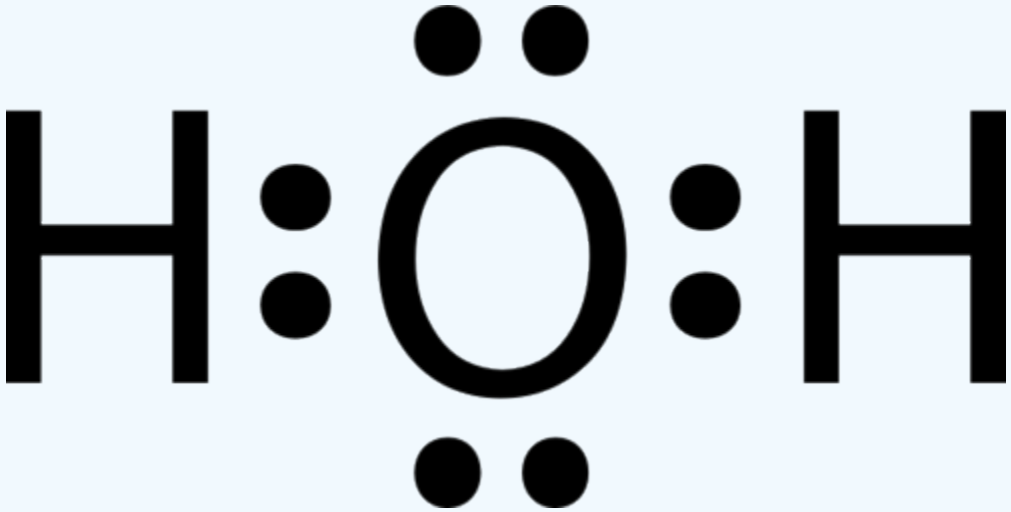


Figure 7.10.6 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Step 3: Think about your result.

The oxygen atom follows the octet rule with two pairs of bonding electrons and two lone pairs. Each hydrogen atom follows the octet rule with one bonding pair of electrons.

YouTube Video

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

View this YouTube Video online

Summary

- Covalent bonds form when electrons in two atoms form overlapping orbitals.
- Lone pair electrons in an atom are not shared with any other atom.

Review

1. How does a covalent bond form?
2. What relationship do the spins of shared electrons have with each other?
3. Do lone pair electrons form covalent bonds?

This page titled 7.10: Single Covalent Bonds is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.6: Single Covalent Bonds by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.11: Multiple Covalent Bonds

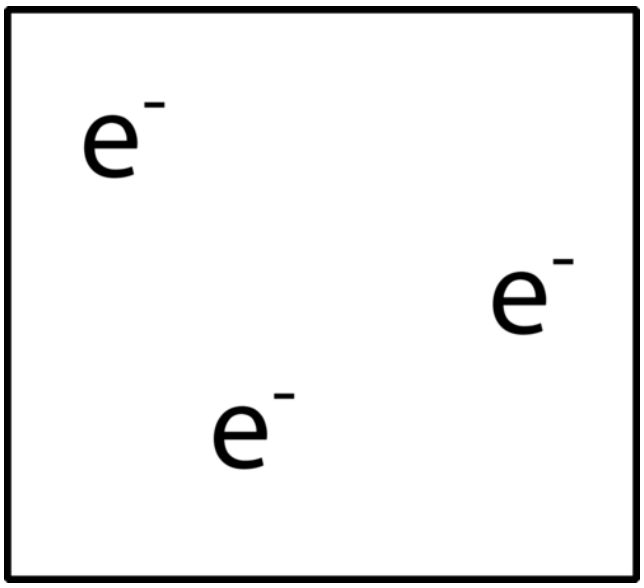


Figure 7.11.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

What do you do with your leftovers?

When working with covalent structures, it sometimes looks like there are leftover electrons. Maybe you have applied the rules you learned so far, but 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 $\mathrm{C_2H_4}$. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two } \mathrm{C} \text{ atoms } = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four } \mathrm{H} \text{ atoms } = 4 \times 1 = 4 \text{ valence electrons}$$

$$= \text{total of 12 valence electrons in the molecule}$$

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:

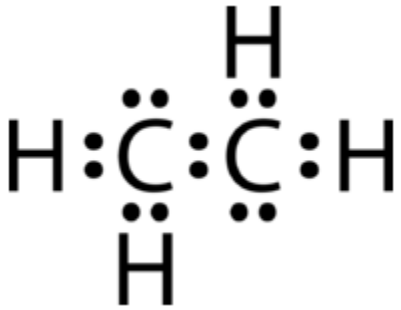


Figure 7.11.2: Incorrect dot structure of ethene. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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 the carbon atoms share two pairs, instead of only one pair, of electrons.

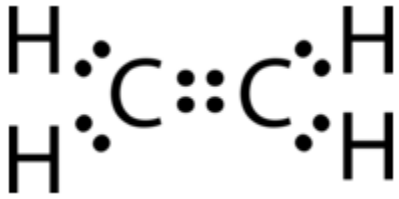


Figure 7.11.3: Correct dot structure for ethene. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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.

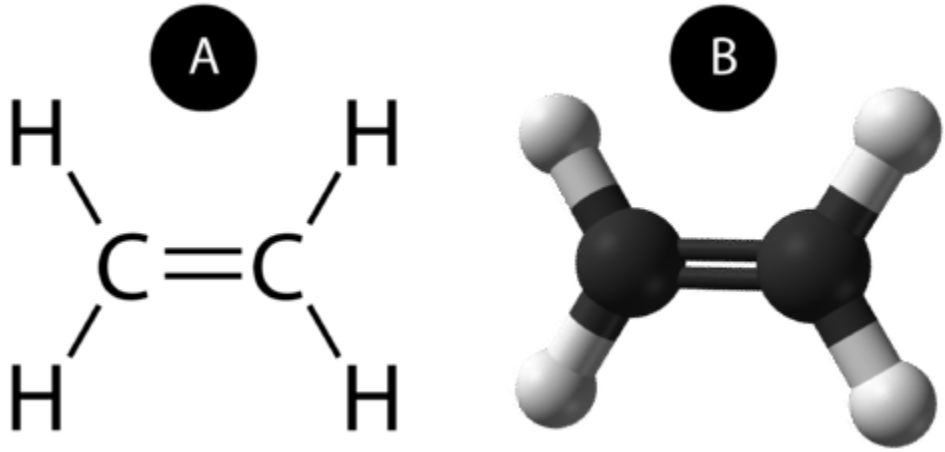


Figure 7.11.4: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-dimensions-2D-Vector.svg; (B) http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png)

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.

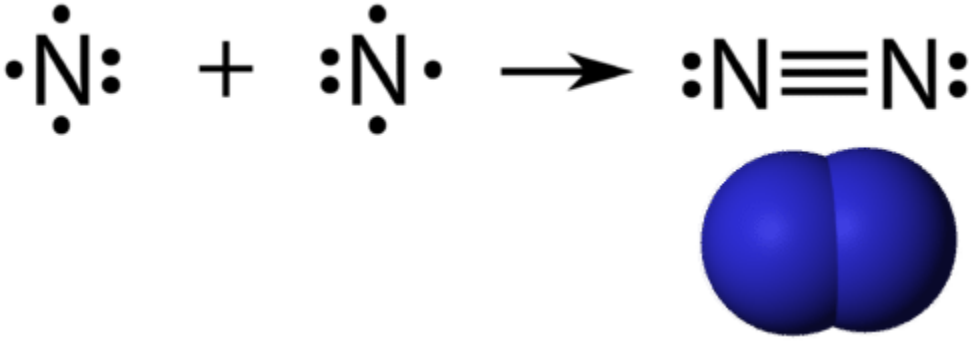


Figure 7.11.5: Triple bond in N_2. (CC BY-NC-SA 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Nitrogen-3D-vdW.png)

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.

YouTube Video

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

View this YouTube Video online

Summary

- Lewis structures can be drawn for molecules that share multiple pairs of electrons.
- A double covalent bond is a covalent bond formed by atoms that share two pairs of electrons.
- A triple covalent bond is a covalent bond formed by atoms that share three pairs of electrons.

Review

1. Why is the first ethene Lewis structure incorrect?
2. What do the single electrons in nitrogen do to form a triple bond?
3. Draw the Lewis structure for ethyne C_2H_2.

This page titled 7.11: Multiple Covalent Bonds is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.7: Multiple Covalent Bonds by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.12: Valence Electrons



Figure 7.12.1 (Public Domain; User:Chemicalinterest/Wikipedia via Wikipedia)

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

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

Valence Electrons

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 7.12: Valence Electrons is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 5.19: Valence Electrons by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.13: Coordinate Covalent Bond



Figure 7.13.1 (Public Domain; P. Periyannan (Wikimedia: TRYPPN) via Wikipedia)

Is sharing a good thing?

Remember when you were younger, and were told to share your favorite toy with your brother, sister, or friend? You probably didn't want to share, but did anyway. It likely turned out that you had more fun playing with the toy together than if you had kept it to yourself. Atoms often share electrons with other atoms that have nothing to contribute to the situation; the end result is a new structure.

Coordinate Covalent Bonds

Each of the covalent bonds that we have looked at so far has involved each of the atoms that are bonding contributing one of the electrons to the shared pair. There is an alternate type of covalent bond in which one of the atoms provides both of the electrons in a shared pair. Carbon monoxide, CO, is a toxic gas that is released as a byproduct during the burning of fossil fuels. The bonding between the C atom and the O atom can be thought of in the following procession:

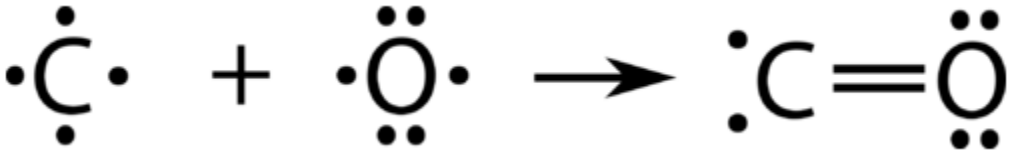


Figure 7.13.2: Formation of a CO double bond (incorrect structure). (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

At this point, a double bond has formed between the two atoms, with each atom providing one of the electrons to each bond. The oxygen atom now has a stable octet of electrons, but the carbon atom only has six electrons and is unstable. This situation is resolved if the oxygen atom contributes one of its lone pairs in order to make a third bond with the carbon atom.

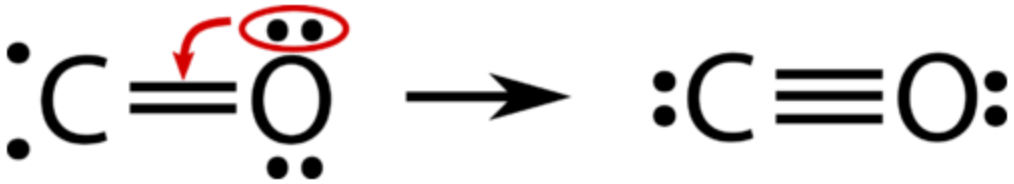


Figure 7.13.3: Correct CO structure. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The carbon monoxide molecule is correctly represented by a triple covalent bond between the carbon and oxygen atoms. One of the bonds is a **coordinate covalent bond**, a covalent bond in which one of the atoms contributes both of the electrons in the shared pair.

Once formed, a coordinate covalent bond is the same as any other covalent bond. It is not as if the two conventional bonds in the CO molecule are stronger or different in any other way than the coordinate covalent bond.

YouTube Video

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

View this YouTube Video online

Summary

- Coordinate covalent bonds can form when one atom provides a lone pair of electrons to the bond.
- Coordinate covalent bonds are as strong as other covalent bonds.

Review

1. Where does the third covalent bond in the CO molecule come from?
2. Why is the incorrect structure for CO above wrong?
3. Are coordinate covalent bonds stronger or weaker than regular covalent bonds?

This page titled 7.13: Coordinate Covalent Bond is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.8: Coordinate Covalent Bond by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.14: Resonance



Figure 7.14.1 (Public Domain; User:Gouwenaar/Nl.Wikipedia via Wikipedia)

Is this a reflection or a doorway?

You may look at the picture below and think you are looking at the image of a room as reflected in a mirror (and you probably are). But the picture can be cropped in such a way as to give the impression of looking at the real room through a door. You would see the same thing and receive the same information, but it would be from a different perspective. There are molecules that can be represented in different ways, such that reality becomes a matter of interpretation.

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone (O_3) molecule in the figure below. There are a total of 18 electrons in the structure, and so the following two structures are possible:

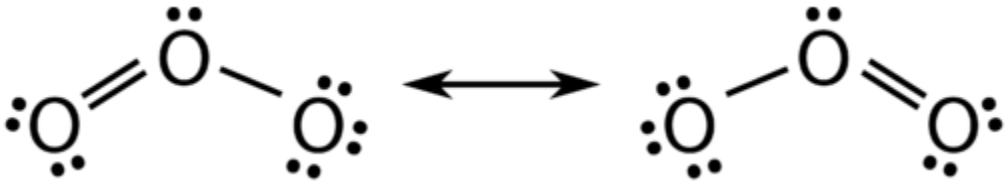


Figure 7.14.2: Resonance forms of ozone. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The structure on the left (see figure above) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $O-O$ single bond and a double bond.

Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (see figure below).



Figure 7.14.3: "Half-bond" model of ozone molecule. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Ozone-CRC-MW-3D-balls.png)

YouTube Video

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

View this YouTube Video online

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- (see figure below).

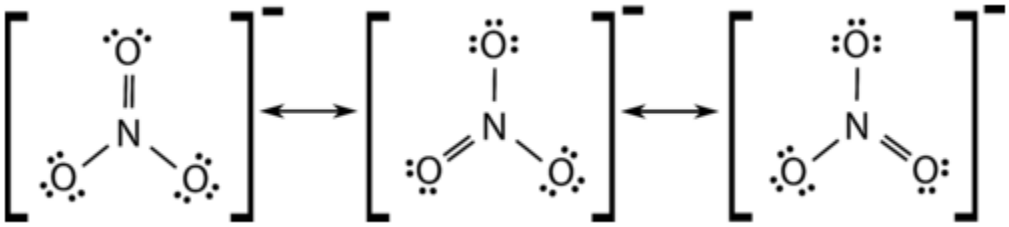


Figure 7.14.4: Resonance structure of nitrate anion. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

YouTube Video

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

View this YouTube Video online

Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

Review

1. How many electrons total are in the ozone structure?

2. What is changed in the two resonance structures of ozone?
3. How can we think of the covalent bonds in ozone?

This page titled 7.14: Resonance is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.10: Resonance by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.15: VSEPR Theory

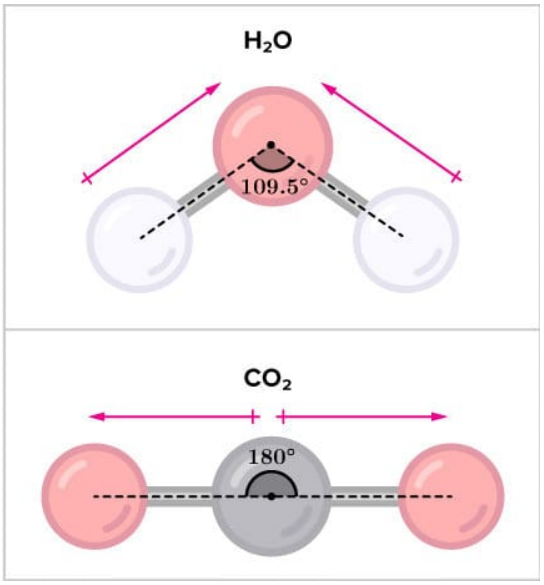


Figure 7.15.1: Copy and Paste Caption here. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Why is the water molecule bent like that?

The characteristic bent shape of the water molecule shown above was a puzzling discovery for scientists at first. The shape allows the molecule to be polar, increasing its boiling point and making it possible for life on earth to exist as we know it. But what makes it bend? The structure is almost the same as carbon dioxide which is known to be a gas at room temperature, why not water too?

Putting atoms together to form compounds can be done on paper or in the lab. However, when the shape of the molecule made in the lab is different from the shape of the molecule drawn on paper, then we need to rethink our ideas and find better explanations.

VSEPR Theory

In 1956, British scientists R.J. Gillespie and R.S. Nyholm recognized that the current model for explaining bond angles did not work well. The theory at that time relied on hybrid orbitals to explain all aspects of bonding. The problem was that this theory gave an incorrect prediction of bond angles for many compounds. They developed a new approach based on earlier work by other scientists that incorporated a consideration of electron pairs in predicting three-dimensional structure.

The **valence shell** is the outermost electron-occupied shell of an atom. The valence shell holds the electrons that are involved in bonding and are the electrons shown in a Lewis structure. The acronym VSEPR stands for the **valence-shell electron pair repulsion** model. The model states that electron pairs will repel each other such that the shape of the molecule will adjust, so that the valence electron-pairs stay as far apart from one another as possible. Molecules can be systematically classified according to the number of bonding pairs of electrons; as well as the number of nonbonding, or lone pairs, around the central atom. For the purposes of the VSEPR model, a double or triple bond is no different in terms of repulsion than a single bond.

YouTube Video

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

View this YouTube Video online

Summary

- VSEPR theory allows more accurate predictions of molecular shape.

Review

1. Who did the major work in developing the VSEPR theory?
2. Where are the electrons that are involved in bonding?
3. What is the basic idea behind the VSEPR theory?

This page titled 7.15: VSEPR Theory is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.13: VSEPR Theory by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.16: Molecular Shapes- No Lone Pairs on Central Atoms

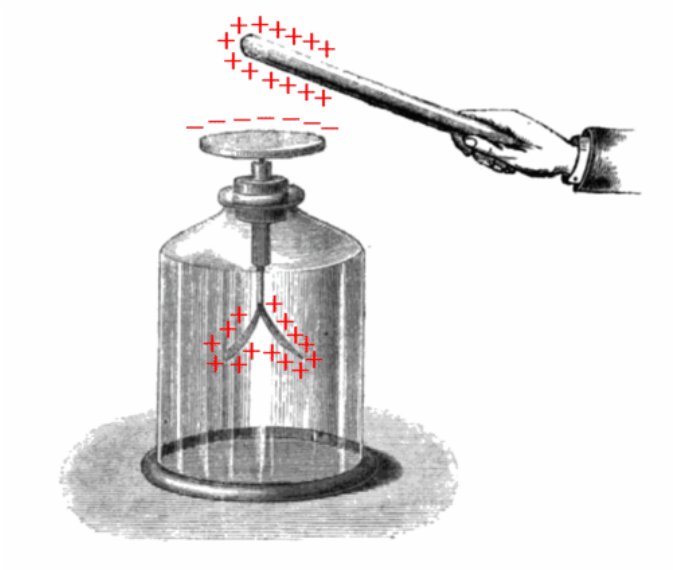


Figure 7.16.1 (Public Domain; Sylvanus P. Thompson via Wikipedia)

How does an electroscope work?

An electroscope is a device used to study charge. When a positively charged object (the rod) nears the upper post, electrons flow to the top of the jar, leaving the two gold leaves positively charged. The leaves repel each other since both hold positive, like charges. The VSEPR theory says that electron pairs, also a set of like charges, will repel each other such that the shape of the molecule will adjust, so that the valence electron pairs stay as far apart from each other as possible.

Central Atom with No Lone Pairs

In order to easily understand the types of molecules possible, we will use a simple system to identify the parts of any molecule:

- A = **central atom** in a molecule
- B = atoms surrounding the central atom

Subscripts after the B will denote the number of B atoms that are bonded to the central A atom. For example, AB_4 is a molecule with a central atom surrounded by four covalently bonded atoms. Again, it does not matter if those bonds are single, double, or triple bonds.

AB_2: Beryllium Hydride (BeH_2)

Beryllium hydride consists of a central beryllium atom with two single bonds to hydrogen atoms. Recall that it violates the octet rule.

$$H{-}Be{-}H$$

According to the requirement that electron pairs maximize their distance from one another, the two bonding pairs in the BeH_2 molecules will arrange themselves on directly opposite sides of the central Be atom. The resulting geometry is a linear molecule, shown in the figure below in a "ball and stick" model.

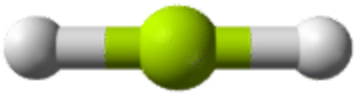


Figure 7.16.2: Beryllium hydride model. (Public Domain; Ben Mills (User:Benjah-bmm27/Wikimedia Commons) via Wikipedia)

The bond angle from $\mathrm{H{-}Be{-}H}$ is 180° because of its linear geometry.

Carbon dioxide is another example of a molecule which falls under the AB_2 category. Its Lewis structure consists of double bonds between the central carbon and the oxygen atoms (see figure below).

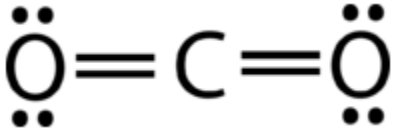


Figure 7.16.3: Carbon dioxide bonding. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The repulsion between the two groups of four electrons (two pairs) is no different than the repulsion of the two groups of two electrons (one pair) in the $\mathrm{BeH_2}$ molecule. Carbon dioxide is also linear (see figure below).

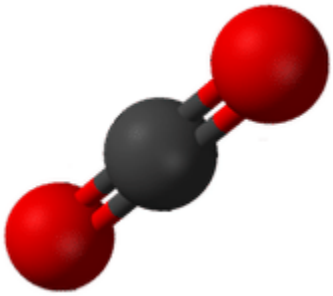


Figure 7.16.4: Carbon dioxide. (Public Domain; User:Benji9072/Wikimedia Commons via Wikipedia)

AB_3: Boron Trifluoride $(\mathrm{BF_3})$

Boron trifluoride consists of a central boron atom with three single bonds to fluorine atoms (see figure below). The boron atom also has an incomplete octet.

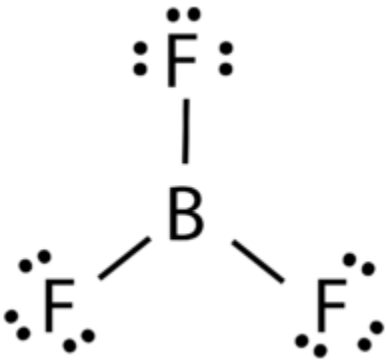


Figure 7.16.5: Boron trifluoride bonding. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The geometry of the $\mathrm{BF_3}$ molecule is called trigonal planar (figure 7.16.3). The fluorine atoms are positioned at the vertices of an equilateral triangle. The $\mathrm{F{-}B{-}F}$ angle is 120° and all four atoms lie in the same plane.

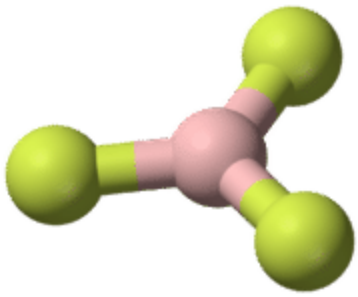


Figure 7.16.6: Boron trifluoride model. (Public Domain; User:Benji90072/Wikimedia Commons via Wikipedia)

AB_4: Methane (CH_4)

Methane is an organic compound that is the primary component of natural gas. Its structure consists of a central carbon atom with four single bonds to hydrogen atoms (see figure below). In order to maximize their distance from one another, the four groups of bonding electrons do not lie in the same plane. Instead, each of the hydrogen atoms lie at the corners of a geometrical shape called a tetrahedron. The carbon atom is at the center of the tetrahedron. Each face of a tetrahedron is an equilateral triangle.

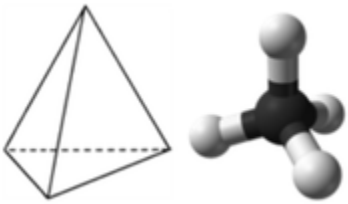


Figure 7.16.7: Tetrahedral structure of methane. (Public Domain; (Left) Pearson Scott Foresman; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Tetrahedron_(PSF).png; (Right) http://commons.wikimedia.org/wiki/File:Methane-CRC-MW-3D-balls.png)

The molecular geometry of the methane molecule is tetrahedral (see figure 7.16.5). The $H-C-H$ bond angles are 109.5^{o}, which is larger than the 90° that they would be if the molecule was planar. When drawing a structural formula for a molecule such as methane, it is advantageous to be able to indicate the three-dimensional character of its shape. The structural formula below is called a perspective drawing. The dotted line bond is to be visualized as receding into the page, while the solid triangle bond is to be visualized as coming out of the page.

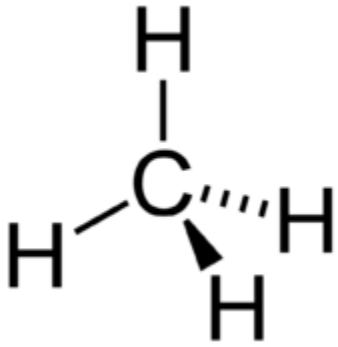


Figure 7.16.8: Methane perspective model. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- Electron pairs repel each other and influence bond angles and molecular shape.

Review

1. What are the bond angles in carbon dioxide?
2. What molecule has bond angles of 109.5°?
3. What is the geometry of the BF_3 molecule?

This page titled 7.16: Molecular Shapes- No Lone Pairs on Central Atoms is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.14: Molecular Shapes- No Lone Pairs on Central Atoms by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.17: Molecular Shapes - Lone Pair(s) on Central Atom



Figure 7.17.1 (Used under license from Shutterstock.com; Image copyright Jandrie Lombard, 2014 via http://www.shutterstock.com)

How can all these clothes fit into such a small space?

When we travel, we often take a lot more stuff than we need. Trying to fit it all into a suitcase can be a real challenge. We may have to repack or just squeeze it all in. Atoms often have to rearrange where the electrons are in order to create a more stable structure.

Central Atom with One or More Lone Pairs

The molecular geometries of molecules change when the central atom has one or more lone pairs of electrons. The total number of electron pairs, both bonding pairs and lone pairs, leads to what is called the **electron domain geometry**. When one or more of the bonding pairs of electrons is replaced with a lone pair, the molecular geometry (actual shape) of the molecule is altered. In keeping with the A and B symbols established in the previous section, we will use E to represent a lone pair on the central atom (A). A subscript will be used when there is more than one lone pair. Lone pairs on the surrounding atoms (B) do not affect the geometry.

AB_3E: Ammonia, NH_3

The ammonia molecule contains three single bonds and one lone pair on the central nitrogen atom (see figure below).

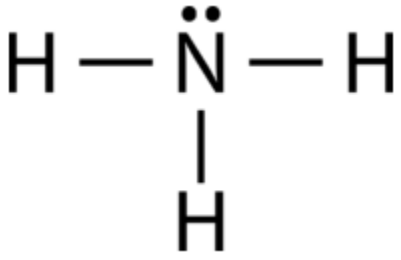


Figure 7.17.2: Lone pair electrons in ammonia. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The domain geometry for a molecule with four electron pairs is tetrahedral, as was seen with CH_4. In the ammonia molecule, one of the electron pairs is a lone pair rather than a bonding pair. The molecular geometry of NH_3 is called trigonal pyramidal (see figure below).

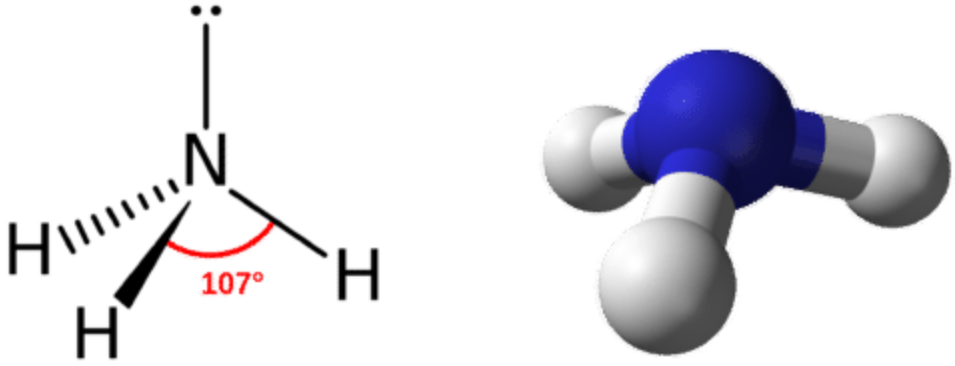


Figure 7.17.3: Ammonia molecule. (Public Domain; (Left) User: Booyabazooka/ Wikimedia Commons; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Ammonia_lone_electron_pair.svg; (Right) http://commons.wikimedia.org/wiki/File:Ammonia-3D-balls-A.png)

Recall that the bond angle in the tetrahedral $\mathrm{CH_4}$ molecule is 109.5°. Again, the replacement of one of the bonded electron pairs with a lone pair compresses the angle slightly. The $\mathrm{H{-}N{-}H}$ angle is approximately 107°.

$\mathrm{AB_2E_2}$: Water, $\mathrm{H_2O}$

A water molecule consists of two bonding pairs and two lone pairs (see figure below).



Figure 7.17.4: Lone pair electrons on water. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

As for methane and ammonia, the domain geometry for a molecule with four electron pairs is tetrahedral. In the water molecule, two of the electron pairs are lone pairs rather than bonding pairs. The molecular geometry of the water molecule is bent. The $\mathrm{H{-}O{-}H}$ bond angle is 104.5°, which is smaller than the bond angle in $\mathrm{NH_3}$ (see figure below).

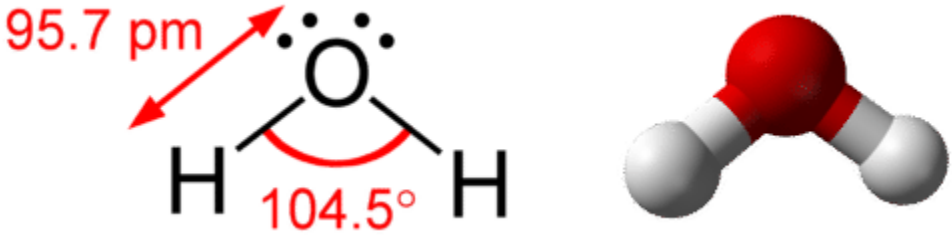


Figure 7.17.5: Water molecule. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Water-dimensions-from-Greenwood%2526Earnshaw-2D.png; (Right) http://commons.wikimedia.org/wiki/File:Water-3D-balls-A.png; http://commons.wikimedia.org/wiki/File:Zeepzieden.jpg)

$\mathrm{AB_4E}$: Sulfur Tetrafluoride, $\mathrm{SF_4}$

The Lewis structure for $\mathrm{SF_4}$ contains four single bonds and a lone pair on the sulfur atom (see figure below).

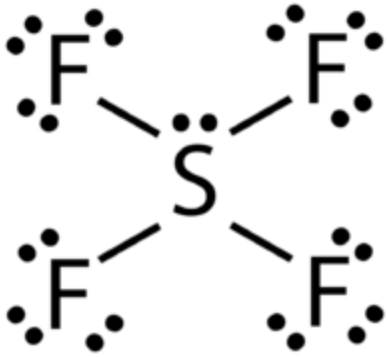


Figure 7.17.6: Lone pair electrons in SF_4. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The sulfur atom has five electron groups around it, which corresponds to the trigonal bipyramidal domain geometry, as in PCl_5 (see figure below). Recall that the trigonal bipyramidal geometry has three equatorial atoms and two axial atoms attached to the central atom. Because of the greater repulsion of a lone pair, it is one of the equatorial atoms that are replaced by a lone pair. The geometry of the molecule is called a distorted tetrahedron, or seesaw.

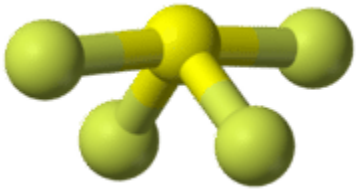


Figure 7.17.7: Ball and stick model for SF_4. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Table 7.17.1: Geometries in Which the Central Atom Has One or More Lone Pairs

Total Number of Electron Pairs	Number of Bonding Pairs	Number of Lone Pairs	Electron Domain Geometry	Molecular Geometry	Examples
3	2	1	Trigonal Planar	Bent	O_3
4	3	1	Tetrahedral	Trigonal Pyramidal	NH_3
4	2	2	Tetrahedral	Bent	H_2O
5	4	1	Trigonal Bipyramidal	Distorted Tetrahedron (Seesaw)	SF_4
5	3	2	Trigonal Bipyramidal	T-shaped	ClF_3
5	2	3	Trigonal Bipyramidal	Linear	I_3^-
6	5	1	Octahedral	Square Pyramidal	BrF_5
6	4	2	Octahedral	Square Planar	XeF_4

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- The presence of lone pair electrons influences the three-dimensional shape of the molecule.

Review

1. Why does water have a bent geometry?
2. Why is ammonia not a planar molecule?
3. How would we write the configuration for xenon tetrafluoride using the ABE system?

This page titled 7.17: Molecular Shapes - Lone Pair(s) on Central Atom is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.15: Molecular Shapes - Lone Pair(s) on Central Atom by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.18: Bond Polarity



Figure 7.18.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What makes people share?

Have you ever spent time with someone you really didn't like? You had nothing in common with them and did not want to have anything to do with them. On the other hand, there are people that you enjoy being with. You have a lot in common and like to share with them. Atoms work the same way. If there are strong differences in their attraction of electrons, one atom gets the electrons and the other atom loses them. If they are similar, they share electrons to form a covalent bond.

Bond Polarity

Electronegativity is defined as the ability of an atom to attract electrons when the atoms are in a compound. Electronegativities of elements are shown in the periodic table below.

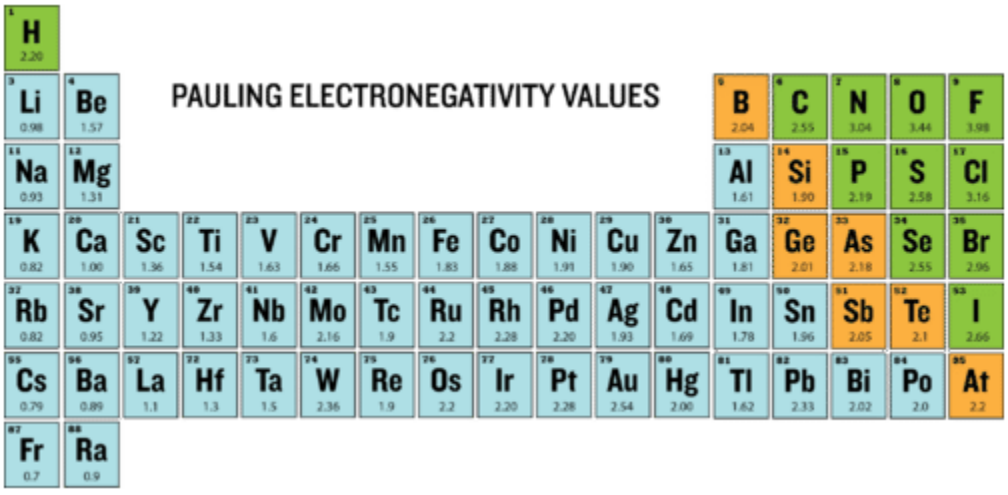


Figure 7.18.2: Electronegativities of elements. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The degree to which a given bond is ionic or covalent is determined by calculating the difference in electronegativity between the two atoms involved in the bond.

As an example, consider the bond that occurs between an atom of potassium and an atom of fluorine. Using the table, the difference in electronegativity is equal to $4.0 - 0.8 = 3.2$. Since the difference in electronegativity is relatively large, the bond between the two atoms is ionic. Since the fluorine atom has a much larger attraction for electrons than the potassium atom does, the valence electron from the potassium atom is completely transferred to the fluorine atom. The diagram below shows how difference in electronegativity relates to the ionic or covalent character of a chemical bond.

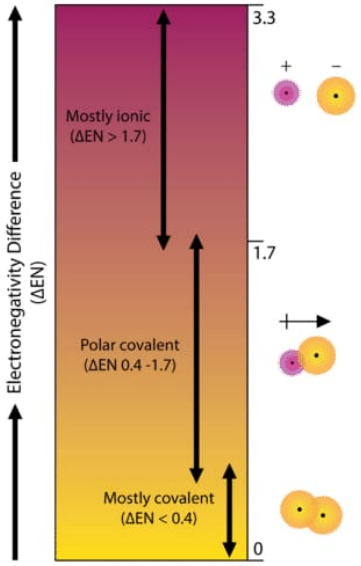


Figure 7.18.3: Bond type is predicted on the difference in electronegativity of the two elements involved in the bond. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.7 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

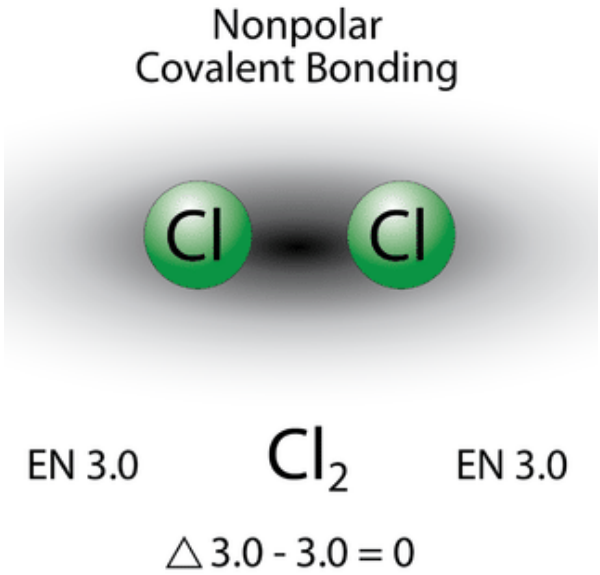


Figure 7.18.4: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the

Cl_2 molecule is symmetrical. Note that molecules in which the electronegativity difference is very small (<0.4) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta EN = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.4 and 1.7 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons, and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

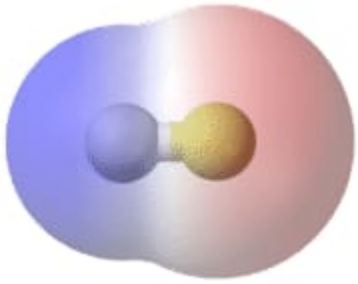


Figure 7.18.5: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Commons Wikimedia, Hydrogen fluoride elpot transparent 3D balls [commons.wikimedia.org])

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

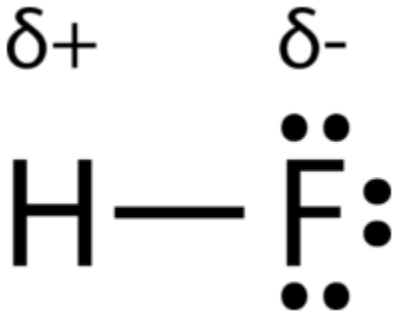


Figure 7.18.6: Use of δ to indicate partial charge. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

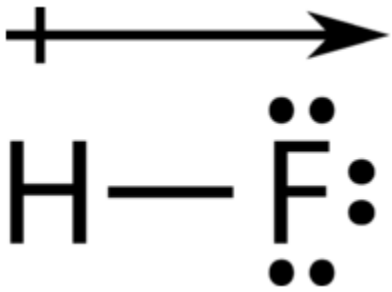


Figure 7.18.7: Use of crossed arrow to indicate polarity. (Credit: Joy Sheng Source: CK-12 Foundation; License: CC BY-NC 3.0)

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Simulation

If it wasn't for our understanding of bond polarity, we'd have a really tough time cleaning our clothing. Use this simulation to understand how the unique polarity of soap molecules helps to clean our clothing!

Summary

- The electronegativity of an atom determines how strongly it attracts electrons to itself.
- The polarity of a bond is affected by the electronegativity values of the two atoms involved in that bond.

Review

1. If two atoms bonded together have an electronegativity difference of 1.9, what is the bond type?
2. What would be the bond type for BH_2?
3. Your friend tells you that the LiF bond is covalent. Are they correct? Why or why not.

This page titled 7.18: Bond Polarity is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.16: Bond Polarity by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

7.19: Polar Molecules

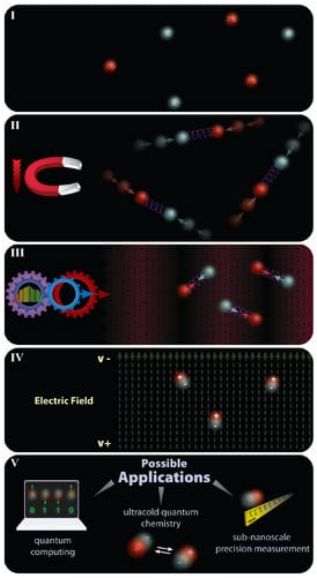


Figure 7.19.1 (Public Domain; G. Kuebler/JILA and the NIST via Patapsco Image Gallery [patapsco.nist.gov])

How cold are ultracold polar molecules?

Ultracold polar molecules have been created by research scientists. These "molecules" consist of Rb and K atoms excited by lasers to form a type of $\mathrm{Rb}{-}\mathrm{K}$ compound where the Rb has a positive charge and the K has a negative charge. The material is formed at temperatures extremely close to absolute zero. The researchers believe these techniques will help them make new reactions and new materials.

Polar Molecules

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 7.19.2: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Polar molecules orient themselves in the presence of an electric field with the positive ends of the molecule being attracted to the negative plate, while the negative ends of the molecules are attracted to the positive plate (see figure below).

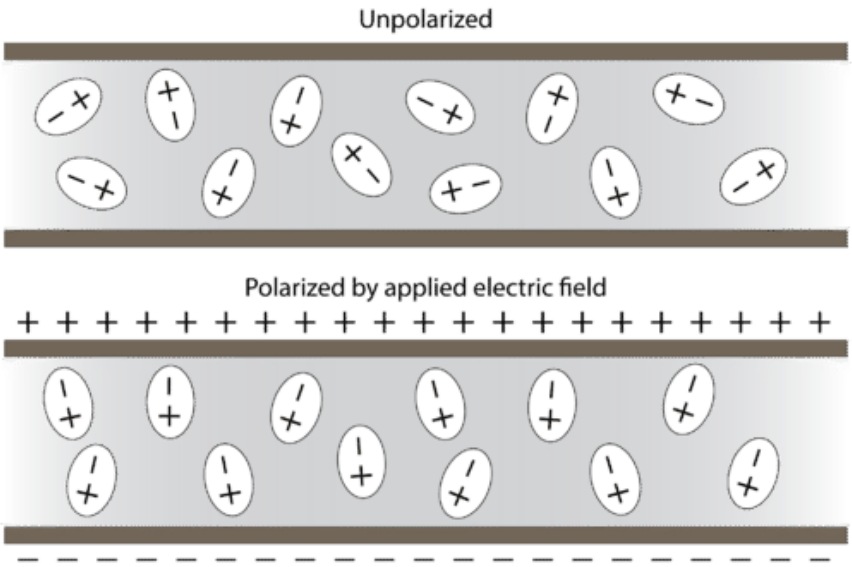


Figure 7.19.3: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves because of the attraction of opposite charges (bottom). (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $(\mathrm{CO_2})$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out, and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out, and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

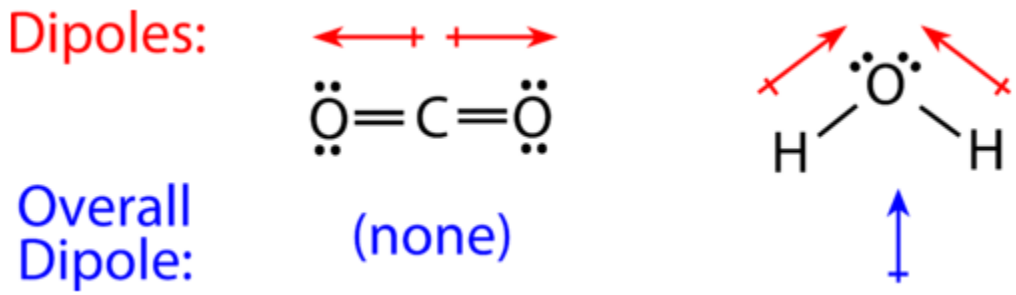


Figure 7.19.4: The molecular geometry of a molecule affects its polarity. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $(\mathrm{BF_3})$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $(\mathrm{NH_3})$ is polar.

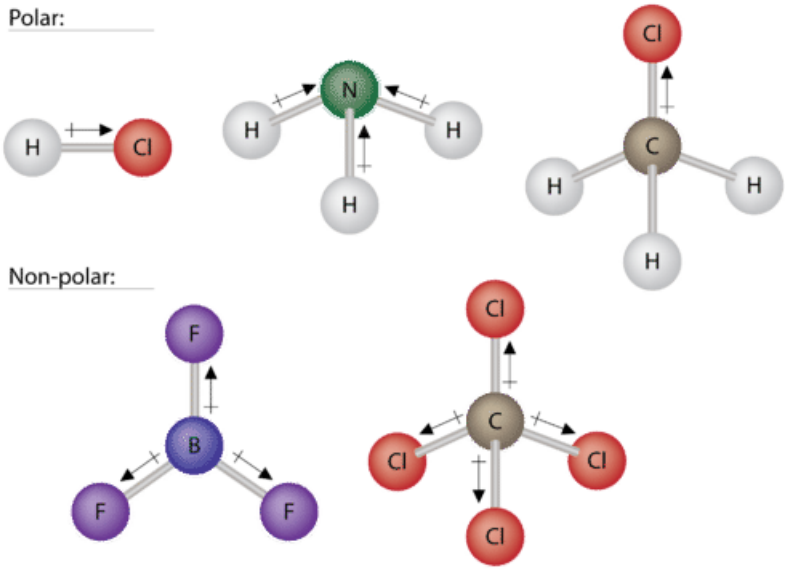


Figure 7.19.5: Some examples of polar and nonpolar molecules based on molecular geometry. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Simulations

Why does water form droplets? How does soap get our clothes clean? Hint: both actions have to do with polarity. Try these simulations to learn more:

How are stains removed from your clothing?

Why do liquids form droplets?

Summary

- Polar molecules result from differences in the electronegativity of the atoms in the molecule.
- Dipoles that are directly opposite one another cancel each other out.

Review

1. Why is a polar molecule called a dipole?
2. The molecular geometry of NH_3 is trigonal pyramidal. What about the molecular geometry of NH_3 makes it a polar molecule?
3. What happens to polar molecules when an electric field is applied?

This page titled 7.19: Polar Molecules is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.17: Polar Molecules by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8: CHEMICAL NOMENCLATURE

The International Union of Pure and Applied Chemistry (IUPAC) is an international federation of organizations that represents chemists in individual countries. IUPAC is best known for its works standardizing nomenclature in chemistry and other fields of science.

8.1: Molecular Formula

8.2: Empirical Formula

8.3: Cations

8.4: Anions

8.5: Transition Metal Ions

8.6: The Stock System of Nomenclature

8.7: Naming Binary Ionic Compounds

8.8: Formulas for Binary Ionic Compounds

8.9: Polyatomic Ions

8.10: Ternary Ionic Compounds- Naming and Formulas

8.11: Binary Molecular Compounds- Naming and Formulas

8.12: Acids - Naming and Formulas

8.13: Bases- Naming and Formulas

Template:HideTOC

This page titled 8: Chemical Nomenclature is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7: Chemical Nomenclature by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.1: Molecular Formula
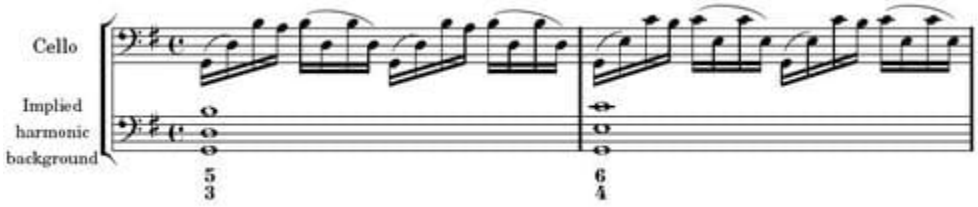


Figure 8.1.1 (Public Domain; J. S. Bach via Commons Wikimedia, Bach cello harmony [commons.wikimedia.org])

Why are music notes a unique "language"?

There are many "universal languages" in the world. Musicians of every culture recognize music embodied in a series of notes on a staff.

This passage from a Bach cello suite could be played by any trained musician from any country, because there is an agreement as to what the symbols on the page mean. In the same way, molecules are represented using symbols that all chemists agree upon.

Molecular Formula

A **molecule** is comprised of two or more atoms that have been chemically combined. 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:

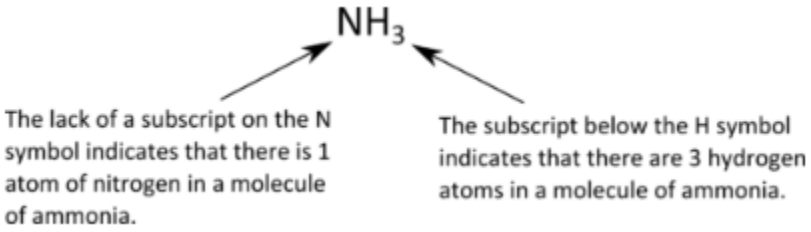


Figure 8.1.2: The molecular formula for ammonia. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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 $\mathrm{N_3H}$ for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

The molecular formula does not tell us anything about the shape of the molecule or where the different atoms are. The molecular formula for sucrose (table sugar) is $\mathrm{C_{12}H_{22}O_{11}}$. This simply tells us the number of carbon, hydrogen, and oxygen atoms in the molecule. There is nothing said about where the individual atoms are located. We need a much more complicated formula (shown below) to communicate that information.

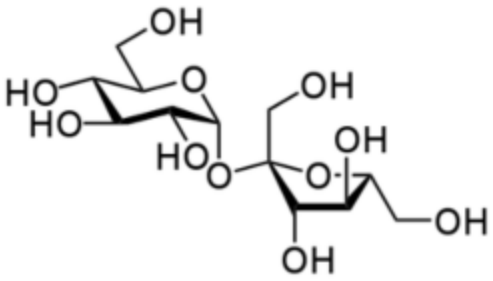


Figure 8.1.3: The molecular structure of sucrose, $C_{12}H_{22}O_{11}$. (Public Domain; User:Calvero/Wikimedia Commons via Commons Wikimedia, Sucrose [commons.wikimedia.org])

YouTube Video

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

View this YouTube Video online

Summary

- A molecular formula tells us which atoms and how many of each type of atom are present in a molecule.
- 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.
- Molecular formulas do not indicate how the atoms are arranged in a molecule.

Review

1. What does a molecular formula tell us?
2. What does a molecular formula not tell us?
3. What do the subscripts mean in a molecular formula?
4. If you wrote $C_6H_{11}O_5C_6H_{11}O_6$ as the molecular formula for sucrose, would that be correct? Explain your answer.
5. Sometimes the formula for acetic acid is written CH_3COOH. Is this a true molecular formula?

This page titled 8.1: Molecular Formula is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.1: Molecular Formula by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.2: Empirical Formula

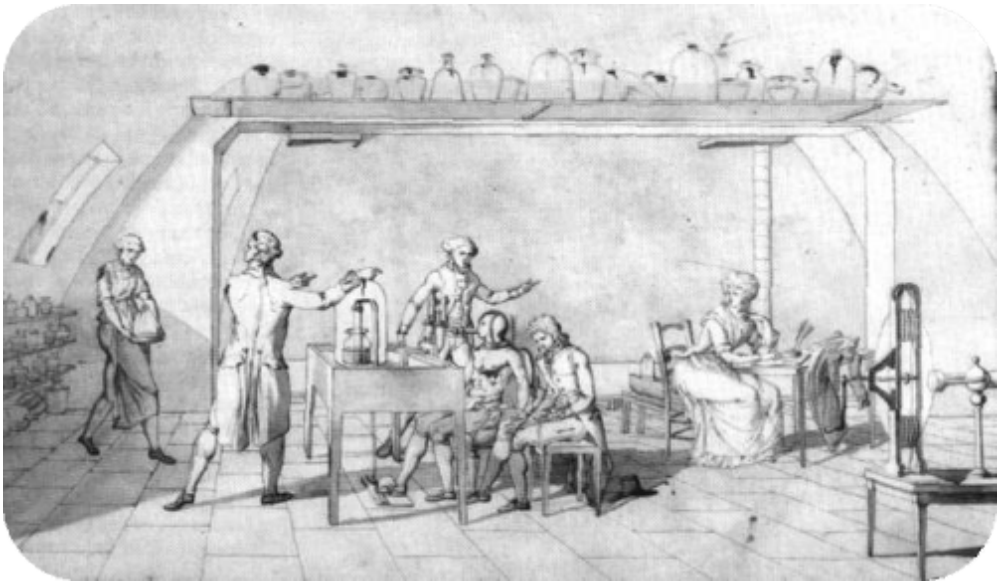


Figure 8.2.1 (Public Domain; Madame Lavoisier via Wikipedia)

What might the scientists in the picture be discussing?

When the French scientist Antoine Lavoisier conducted his experiments, he did not know what the products of each reaction would be. He had to isolate the material (whether he was heating mercury or measuring gases from breathing) and then study its elemental composition before he could understand the processes that were occurring.

Discovering that a new compound exists is the start of a long research project. In order to make this new compound in the lab, we need to know a lot about its structure. Often, the place to start is to determine the elements in the material. Then, we can find out the relative amounts of each element to continue our evaluation of this new material.

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 $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula. 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 H_2O or P_2O_5, then the empirical formula is also the molecular formula.

How do we determine an empirical formula for a compound? Consider a compound composed of carbon, hydrogen, and oxygen. We can analyze the relative amounts of each element in the compound. When we get a percent figure for each element, we now know how many grams of each are in 100 grams of the original material. This allows us to determine the number of moles for each element. The ratios can then be reduced to small whole numbers to give the empirical formula. If we wanted a molecular formula, we would need to determine the molecular weight of the compound.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- The empirical formula gives the lowest whole-number ratio of elements in a compound.
- The empirical formula does not show the actual number of atoms.

Review

1. Define “empirical formula.”
2. Why is $C_6H_{12}O_6$ not considered to be an empirical formula for glucose?
3. Can the empirical formula for a compound be the same as the molecular formula?
4. What do we need to know in order to determine a molecular formula from an empirical formula?
5. Give three examples of compounds whose empirical formulas are the same as their molecular formulas.

This page titled 8.2: Empirical Formula is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.2: Empirical Formula by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.3: Cations



Figure 8.3.1 (Public Domain; User:Gold Guru/Wikipedia via Wikipedia)

Have you ever gone digging for gold?

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

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

Cations

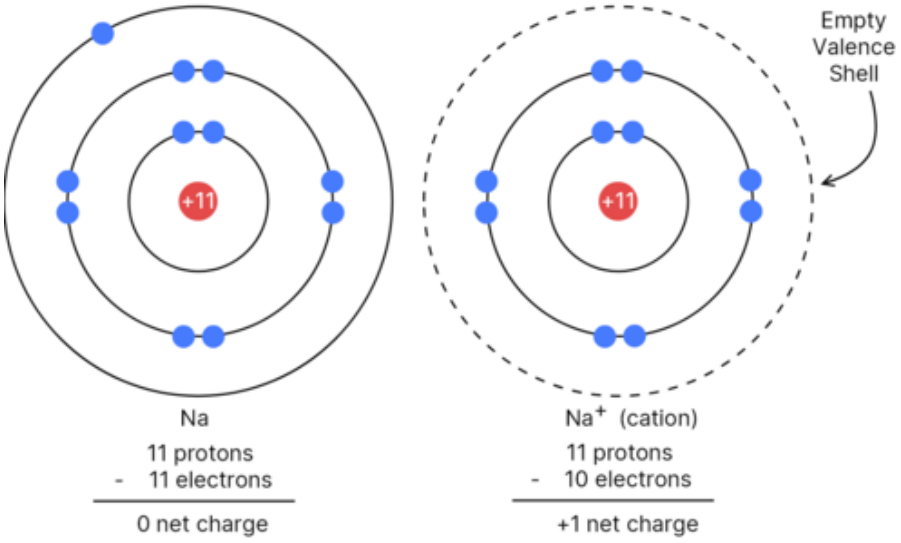


Figure 8.3.2: Sodium loses an electron to become a cation. (CK-12 Curriculum Materials License; CK-12 Foundation via CK-12 Foundation)

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

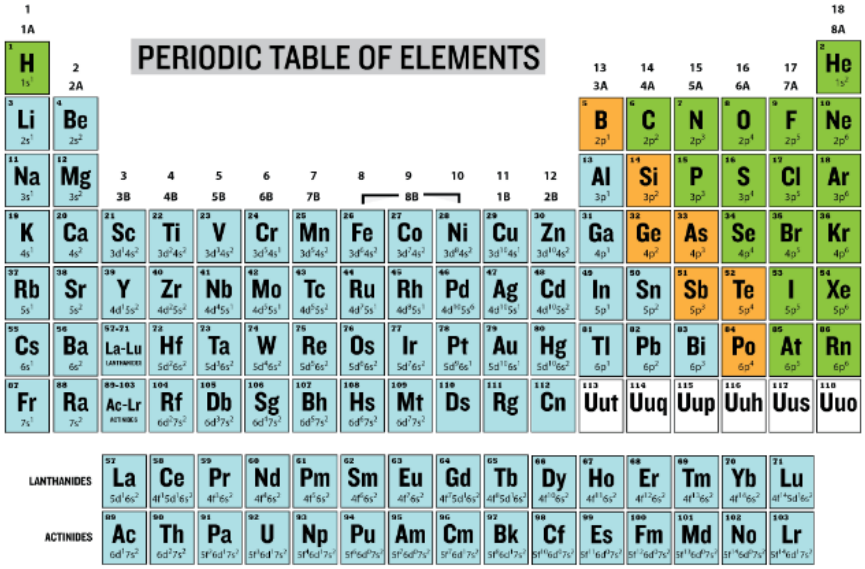


Figure 8.3.3: Periodic Table of Elements, notated with group numbers. (CC BY-NC 3.0; Christopher Auyeung via CK-12)

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

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Applications of Cations

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

Summary

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

Review

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

This page titled 8.3: Cations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.3: Cations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.4: Anions



Figure 8.4.1 (Public Domain; Shane Anderson/NOAA via Commons Wikimedia, Kelp 300 [commons.wikimedia.org])

What does the amount of salt in seaweed tell us?

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

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

Anions

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

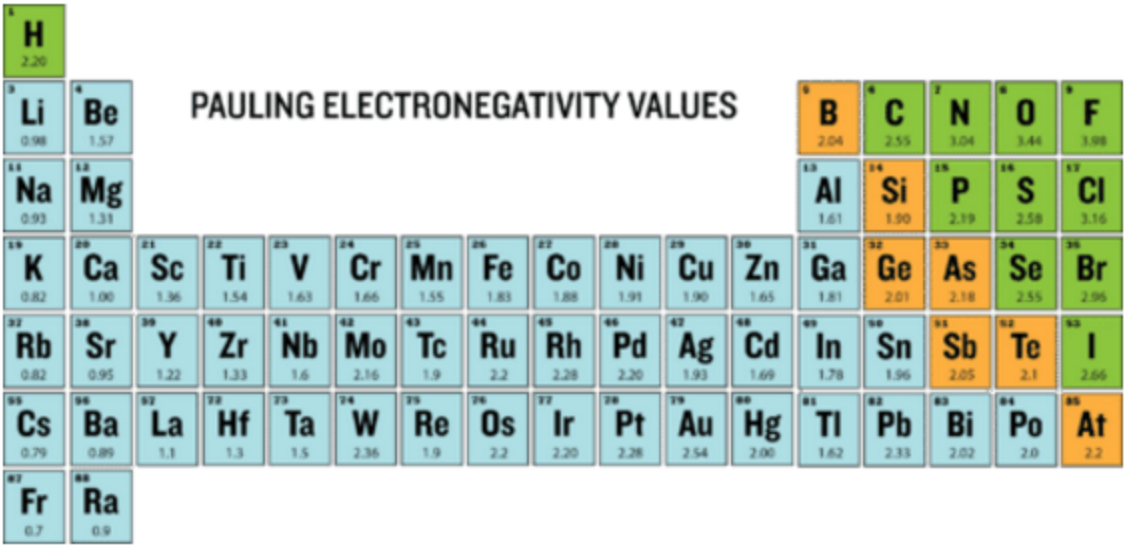


Figure 8.4.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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

Table 8.4.1

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Uses for Anions

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

Summary

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

Review

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

This page titled 8.4: Anions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.4: Anions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.5: Transition Metal Ions



Figure 8.5.1 (Public Domain; the US Mint via Commons Wikimedia, Platinium eagle 101 [commons.wikimedia.org])

What kind of coin is this?

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

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

Transition Metal Ions

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

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org



Figure 8.5.2: (Public Domain; Cepheus, modified by CK-12 Foundation via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

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

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

Common Transition Metal Ions

Table 8.5.1: Common Transition Metal Ions

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

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Uses for Transition Metals

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

Summary

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

Review

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

This page titled 8.5: Transition Metal Ions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.5: Transition Metal Ions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.6: The Stock System of Nomenclature



Figure 8.6.1 (Public Domain; XAVeRY via Wikipedia)

How is an mp3 player designed?

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

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

Naming Compounds Using the Stock System

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

Table 8.6.1

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

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

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

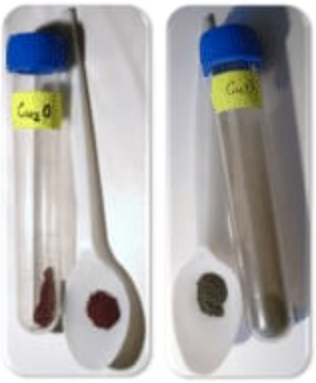


Figure 8.6.2: Copper (I) oxide, a red solid, and copper (II) oxide, a black solid, are different compounds because of the charge of the copper ion. (Public Domain; Ondřej Mangl via Copper(I) oxide: http://commons.wikimedia.org/wiki/File:Cu2O.png; Copper(II) oxide: http://commons.wikimedia.org/wiki/File:Oxid_m%25C4%259B%25C4%258Fnat%25C3%25BD.PNG)

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 8.6: The Stock System of Nomenclature is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.6: The Stock System of Nomenclature by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.7: Naming Binary Ionic Compounds



Figure 8.7.1 (Public Domain; Carl Gustav Calwer, Gustav Jäger, Emil Hochdanz via Wikipedia)

Do you know the proper name for this insect?

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

Binary Ionic Compounds

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

Naming Binary Ionic Compounds

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

Table 8.7.1: Naming Binary Ionic Compounds

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

Notice that in each of the formulas above, the overall charge of the compound is zero. Potassium ion is K^+, while fluoride ion is F^-. Since the magnitude of the charges is equal, the formula contains one of each ion. This would also be the case for a compound such as MgS, in which the ions are Mg^{2+} and S^{2-}. For sodium nitride, the sodium ion is Na^+, while the nitride ion is N^{3-}. In order to make a neutral compound, three of the $1+$ sodium ions are required in order to balance out the single $3-$

nitride ion. So the Na is given a subscript of 3. For calcium phosphide, the calcium ion is Ca^{2+}, while the phosphide ion is P^{3-}. The least common multiple of 2 and 3 is 6. To make the compound neutral, three calcium ions have a total charge of $6+$, while two phosphide ions have a total charge of $6-$. The Ca is given a subscript of 3, while the P is given a subscript of 2.

YouTube Video

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

View this YouTube Video online

Summary

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

Review

1. What is a binary compound?
2. Which of the following is not a binary compound?
   a. NaCl
   b. KH2PO4
   c. KBr
3. Name the following compounds:
   a. NaBr
   b. MgCl2
   c. LiI
   d. CaO
   e. CuBr2
   f. FeO

This page titled 8.7: Naming Binary Ionic Compounds is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.7: Naming Binary Ionic Compounds by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.8: Formulas for Binary Ionic Compounds

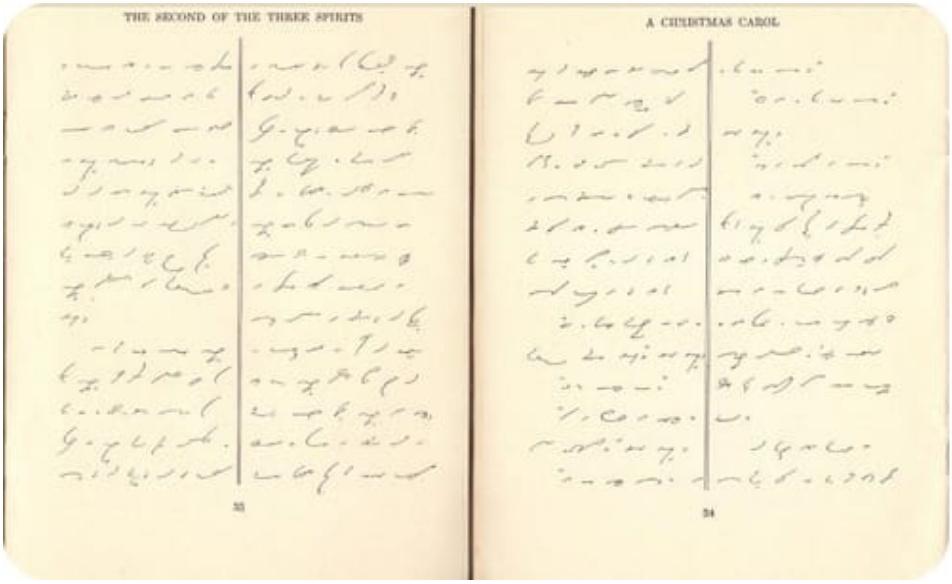


Figure 8.8.1 (Public Domain; Winifred Kenna via Commons Wikimedia, Gregg shorthand A Christmas Carol [commons.wikimedia.org])

How does shorthand work?

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

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

Writing Formulas for Binary Ionic Compounds

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

[redacted]

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

The ions for the compound lithium oxide are:

Li^+ [redacted]

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

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

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 8.8.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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

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

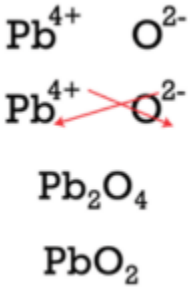


Figure 8.8.3 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

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

Review

1. Write formulas for the binary ionic compounds formed between the following pairs of elements:
   a. cesium and fluorine
   b. calcium and sulfur
   c. aluminum and chlorine
   d. zinc and nitrogen
2. Write the formula and give the name for the compound formed by the following ions:
   a. Fe^{3+} and O^{2-}
   b. Ni^{2+} and S^{2-}
   c. Au^{+} and Cl^{-}
   d. Sn^{4+} and I^{-}
3. Give names for the following compounds:
   a. Ag_2S
   b. PdO
   c. $PtCl_4$
   d. V_2O_5

This page titled 8.8: Formulas for Binary Ionic Compounds is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.8: Formulas for Binary Ionic Compounds by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.9: Polyatomic Ions



Figure 8.9.1 (Public Domain; Huhu via Wikipedia)

Have you ever read the story of Romeo and Juliet?

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

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

Polyatomic Ions

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

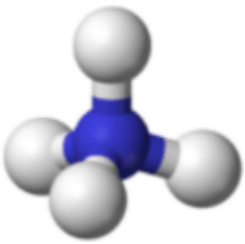


Figure 8.9.2: The ammonium ion $(\mathrm{NH_4}^+)$ is a nitrogen atom (blue) bonded to four hydrogen atoms (white). (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)



Figure 8.9.3: The hydroxide ion (OH^-) is an oxygen atom (red) bonded to a hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

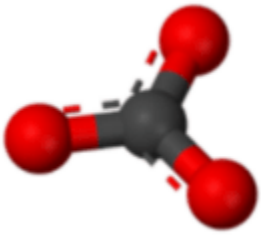


Figure 8.9.4: The carbonate ion (CO_3^{2-}) is a carbon atom (black) bonded to three oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

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

Table 8.9.1: Common Polyatomic Ions

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

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

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

CK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

- ClO_4^-, perchlorate

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

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 8.9: Polyatomic Ions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.9: Polyatomic Ions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.10: Ternary Ionic Compounds- Naming and Formulas



Figure 8.10.1 (Public Domain; Jon Sullivan via Commons Wikimedia, Pepperoni Pizza [commons.wikimedia.org])

How many varieties of pizza can you think of?

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

Ternary Ionic Compounds

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

Naming Ternary Ionic Compounds

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

Table 8.10.1: Examples of Ternary Ionic Compounds

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

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

Writing Formulas for Ternary Ionic Compounds

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

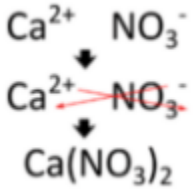


Figure 8.10.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The charge is balanced by the presence of two nitrate ions and one calcium ion. Parentheses are used around the nitrate ion because more than one of the polyatomic ion is needed. If only one polyatomic ion is in a formula, parentheses are not used. As an example, the formula for calcium carbonate is $\mathrm{CaCO_3}$. The carbonate ion carries a $2-$ charge and so exactly balances the $2+$ charge of the calcium ion.

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

Summary

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

Review

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

This page titled 8.10: Ternary Ionic Compounds- Naming and Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.10: Ternary Ionic Compounds: Naming and Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.11: Binary Molecular Compounds- Naming and Formulas



Figure 8.11.1 (Public Domain; Anonymous via Wikipedia)

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

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

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

Molecular Compounds

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

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

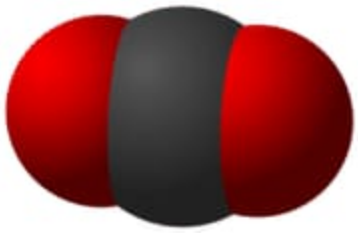


Figure 8.11.2: Carbon dioxide molecules consist of a central carbon atom bonded to two oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

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

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is $\mathrm{H_2O}$. 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 $\mathrm{C_8H_{18}}$.



Figure 8.11.3: Nitrogen dioxide $(\mathrm{NO_2})$ is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines. (Public Domain; User:Greenhorn1/Wikimedia Commons via Wikipedia)

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

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

Table 8.11.1: Numerical Prefixes

Number of Atoms	Prefix
1	*mono-*
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*

Number of Atoms	Prefix
6	*hexa-*
7	*hepta-*
8	*octa-*
9	*nona-*
10	*deca-*

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

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

Some examples of molecular compounds are listed in Table 8.11.2.

Table 8.11.2

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

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

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

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

Review

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

This page titled 8.11: Binary Molecular Compounds- Naming and Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.11: Binary Molecular Compounds: Naming and Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.12: Acids - Naming and Formulas



Figure 8.12.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How is gold tested?

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

Acids

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



Figure 8.12.2: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids. (CC BY 2.0; (A) Becky Cortino; (B) Flickr: isox4 via (A) http://www.flickr.com/photos/mediasavvy/8239231530; (B) http://www.flickr.com/photos/isox4/5167980026)

This is a different type of compound than the others we have seen so far. Acids are molecular, which means that in their pure

state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$\mathrm{HCl} \rightarrow \mathrm{H^+ + Cl^-}$$

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

Naming Acids

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

Table $8.12.1$*: Naming System for Acids*

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

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

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

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

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

Writing Formulas for Acids

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

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

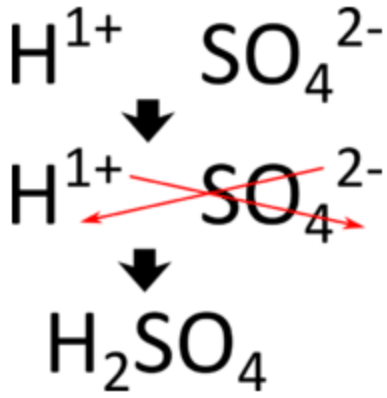


Figure 8.12.3: Crisscross approach to writing the formula for sulfuric acid. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Interactive Content

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

View this Interactive Content online

Summary

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

Review

1. Define “acid.”
2. What is a binary acid?
3. What is an oxoacid?
4. Name the following acids:
   a. H_2SO_4
   b. HCN
   c. HCl
   d. H_3PO_4
5. Write formulas for the following acids:
   a. hydrobromic acid
   b. perchloric acid
   c. nitrous acid

This page titled 8.12: Acids - Naming and Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.12: Acids - Naming and Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

8.13: Bases- Naming and Formulas



Figure 8.13.1 (Public Domain; Unknown Author via Wikipedia)

What different things are these workers doing to make soap?

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

Bases

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

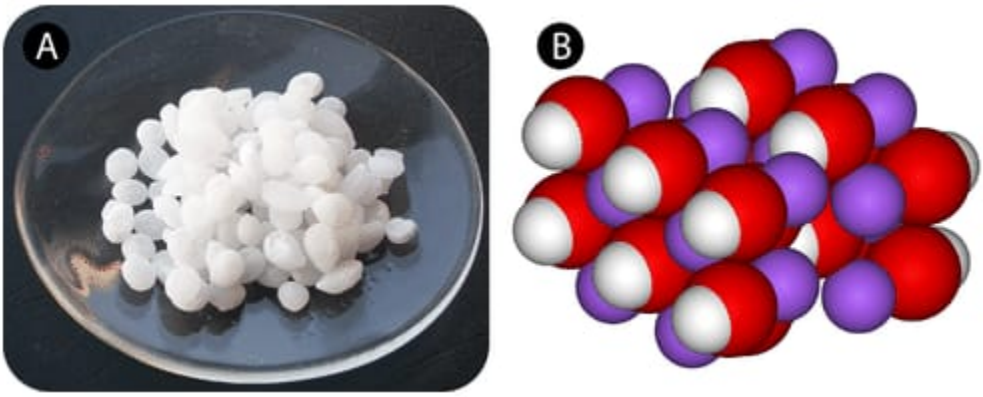


Figure 8.13.2: (A) Sodium hydroxide, a base, is a solid that is typically produced as small white pellets. (B) The structure of sodium hydroxide is an extended three-dimensional network. The purple spheres are the sodium ions $(\mathrm{Na^+})$. The red and white spheres are oxygen and hydrogen atoms, respectively, which are bonded together to form hydroxide ions $(\mathrm{OH^-})$. (Public Domain; (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:SodiumHydroxide.jpg; (B) http://commons.wikimedia.org/wiki/File:Sodium-hydroxide-crystal-3D-vdW.png)

Names and Formulas of Bases

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

Names and Formulas of Bases

Table 8.13.1

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

Notice that because bases are ionic compounds, the number of hydroxides in the formula does not affect the name. The compound must be neutral, so the charges of the ions are balanced just as for other ionic compounds. The sodium ion (Na^+) requires one OH^- ion to balance the charge, so the formula is $NaOH$. Calcium ([redacted]) requires two OH^- ions to balance the charge, so the formula is $Ca(OH)_2$. Hydroxide ion is a polyatomic ion, and must be in parentheses when there is more than one in a formula.

Summary

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

Review

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

This page titled 8.13: Bases- Naming and Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 7.13: Bases: Naming and Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9: THE MOLE

The mole is the unit of measurement in the International System of Units (SI) for amount of substance. It is defined as the amount of a chemical substance that contains as many elementary entities (e.g., atoms, molecules, ions, electrons, or photons). This number is expressed by the Avogadro constant, which has a value of $6.022140857 \times 10^{23} \ mol^{-1}$. The mole is one of the base units of the SI, and has the unit symbol mol.

9.1: Law of Definite Proportions

9.2: Avogadro's Number

9.3: Conversions Between Moles and Atoms

9.4: Molar Mass

9.5: Conversions Between Moles and Mass

9.6: Conversions Between Mass and Number of Particles

9.7: Avogadro's Hypothesis and Molar Volume

9.8: Conversions Between Moles and Gas Volume

9.9: Gas Density

9.10: Mole Road Map

9.11: Percent Composition

9.12: Percent of Water in a Hydrate

9.13: Determining Empirical Formulas

9.14: Determining Molecular Formulas

Template:HideTOC

This page titled 9: The Mole is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10: The Mole by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.1: Law of Definite Proportions



Figure 9.1.1 (Public Domain; Staff Sgt. Christopher Flurry, U.S. Marine Corps via Commons Wikimedia, File:USMC-110609-M-FK922-004.jpg [commons.wikimedia.org])

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

Law of Definite Proportions

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



Figure 9.1.2: Water. (CC BY 2.0; Jenn Durfey via Flickr Photos, Water Faucet [www.flickr.com])

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

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



Figure 9.1.3: Carbon dioxide is produced during the burning process. (Public Domain; Patsy Lynch, FEMA via Commons Wikimedia, FEMA Photograph by Patsy Lynch taken on 04-17-2006 in Missouri [commons.wikimedia.org])

Summary

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

Review

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

Explore More

Use the resource below to answer the questions that follow.

YouTube Video

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

View this YouTube Video online

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

This page titled 9.1: Law of Definite Proportions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 4.4: Law of Definite Proportions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.2: Avogadro's Number



Figure $9.2.1$ (Public Domain; User:OSX/Wikimedia Commons via Wikipedia)

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

It certainly is easy to count objects such as bananas, or something as large as elephants (as long as you stay out of their way). However, counting grains of sugar from a sugar canister would take a long, long time. Atoms and molecules are extremely small—far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years, without stopping, you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists of the past needed a name that could stand for a very large number of items. Amadeo Avogadro (1776-1856), an Italian scientist, provided such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.



Figure 9.2.2: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry. (Public Domain; C. Sentier via Wikipedia)

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as $NaCl$ and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.



Figure 9.2.3: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year. ((left; Left: Michael David Hill, 2005 (Mikiwikipikidikipedia); Right: chrisbb@prodigy.net;(left) Michael David Hill, 2005 (Mikiwikipikidikipedia); (right) chrisbb@prodigy.net via Left: http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; Right: http://www.flickr.com/photos/chrisbrenschmidt/436990097;(left) http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; (right) http://www.flickr.com/photos/chrisbrenschmidt/436990097) CC-BY 2.5; (right) CC-BY 2.0)

YouTube Video

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

View this YouTube Video online

Summary

- A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

This page titled 9.2: Avogadro's Number is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.1: Avogadro's Number by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.3: Conversions Between Moles and Atoms



Figure 9.3.1 (Public Domain; Diane A. Reid/National Cancer Institute via Wikipedia)

Big numbers or little numbers?

Do you hate to type subscripts and superscripts? Even with a good word-processing program, having to click on an icon to get a superscript and then remembering to click off after you type the number can be a real hassle. If we did not know about moles and just knew about numbers of atoms or molecules (those big numbers that require lots of superscripts), life would be much more complicated and we would make many more typing errors.

Conversions Between Moles and Atoms

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Example 9.3.1: Converting Number of Particles to Moles

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

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

Known

- number of C atoms $= 4.72 \times 10^{24}$
- 1 mole $= 6.02 \times 10^{23}$ atoms

Unknown

- 4.72 x 10^{24} = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures.

Suppose that you want to know how many hydrogen atoms are in a mole of water molecules. First, you need to know the chemical formula for water, which is H_2O. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen are in two water molecules? There are $2 \times 2 = 4$ hydrogen atoms. How about in a dozen? In that case, a dozen is 12; so $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers (4 and 24), you multiply the given number of molecules by two atoms of hydrogen per molecule. So, to find the number of hydrogen atoms in a mole of water molecules, the problem can be solved using conversion factors:

$$1 \text{ mol } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2O} = 1.20 \times 10^{24} \text{ atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.

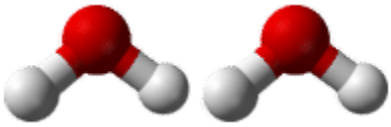


Figure 9.3.2: Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Example 9.3.2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is the sample?

Solution:

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

Known

- $4.89 \times 10^{25} = \text{O}$ atoms
- 1 mole $= 6.02 \times 10^{23}$ molecules H_2SO_4

Unknown

- mol of H_2SO_4 molecules

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ molecule } H_2SO_4}{4 \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{6.02 \times 10^{23} \text{ molecules } H_2SO_4} = 20.3 \text{ mol } H_2SO_4$$

Step 3: Think about your result.

The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.

YouTube Video

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

View this YouTube Video online

Summary

- Methods are described for conversions between moles, atoms, and molecules.

Review

1. What conversion factor would we need to convert moles of helium to atoms of helium?
2. I want to convert atoms to moles. My friend tells me to multiply the number of atoms by 6.02×10^{23} atoms/mole. Is my friend correct?
3. Why do you need to know the formula for a molecule in order to calculate the number of moles of one of the atoms?
4. How many atoms of fluorine are in 5.6×10^{22} molecules of MgF_2?

This page titled 9.3: Conversions Between Moles and Atoms is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.2: Conversions Between Moles and Atoms by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.4: Molar Mass



Figure 9.4.1 (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

When creating a solution, how do I know how much of each substance to put in?

I want to make a solution that contains 1.8 moles of potassium dichromate. You do not have a balance calibrated in moles, but you do have one calibrated in grams. If you know the relationship between moles and the number of grams in a mole, you can use your balance to measure out the needed amount of material.

Molar Mass

Molar mass is defined as the mass of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of lithium is $6.94\ \text{g}$, the molar mass of zinc is $65.38\ \text{g}$, and the molar mass of gold is $196.97\ \text{g}$. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole, or g/mol.

Molar Masses of Compounds

The molecular formula of the compound carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen:

$$12.01\ \text{amu} + 2\,(16.00\ \text{amu}) = 44.01\ \text{amu}$$

The **molecular mass** of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is $44.01\ \text{amu}$.

The molar mass of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of $44.01\ \text{g}$, while one mole of sodium sulfide formula units has a mass of $78.04\ \text{g}$. The molar masses are $44.01\ \text{g/mol}$ and $78.04\ \text{g/mol}$ respectively. In both cases, that is the mass of 6.02×10^{23} representative particles. The representative particle of CO_2 is the molecule, while for Na_2S it is the formula unit.

Example 9.4.1: Molar Mass of a Compound

Calcium nitrate, $\text{Ca}(\text{NO}_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Solution

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

Known

- Formula $= \mathrm{Ca(NO_3)_2}$
- Molar mass $\mathrm{Ca} = 40.08\ \mathrm{g/mol}$
- Molar mass $\mathrm{N} = 14.01\ \mathrm{g/mol}$
- Molar mass $\mathrm{O} = 16.00\ \mathrm{g/mol}$

Unknown

- molar mass $Ca(NO_3)_2$

First we need to analyze the formula. Since the Ca lacks a subscript, there is one Ca atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit and each nitrate ion consists of one nitrogen atom and three oxygen atoms per formula unit. Thus, $1\ \mathrm{mol}$ of calcium nitrate contains $1\ \mathrm{mol}$ of Ca atoms, $2\ \mathrm{mol}$ of N atoms, and $6\ \mathrm{mol}$ of O atoms.

Step 2: Calculate

Use the molar masses of each atom together with the number of atoms in the formula and add together.

$$1\ \mathrm{mol\ Ca} \times \frac{40.08\ \mathrm{g\ Ca}}{1\ \mathrm{mol\ Ca}} = 40.08\ \mathrm{g\ Ca}$$

$$2\ \mathrm{mol\ N} \times \frac{14.01\ \mathrm{g\ N}}{1\ \mathrm{mol\ N}} = 28.02\ \mathrm{g\ N}$$

$$6\ \mathrm{mol\ O} \times \frac{16.00\ \mathrm{g\ O}}{1\ \mathrm{mol\ O}} = 96.00\ \mathrm{g\ O}$$

$$\text{Molar mass of } \mathrm{Ca(NO_3)_2} = 40.08\ \mathrm{g} + 28.02\ \mathrm{g} + 96.00\ \mathrm{g} = 164.10\ \mathrm{g/mol}$$

YouTube Video

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

View this YouTube Video online

Summary

- Calculations are described for the determination of molar mass of an atom or a compound.

Review

1. What is the molar mass of Pb?
2. Where do you find the molar mass of an element?
3. How many moles of Cl are in one mole of the $CaCl_2$?
4. How many moles of H are in one mole of the compound $(NH_4)_3PO_4$?
5. Calculate the molar mass of $CaCl_2$.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

This page titled 9.4: Molar Mass is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.3: Molar Mass by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.5: Conversions Between Moles and Mass



Figure 9.5.1 (Public Domain; User:Estormiz/Wikimedia Commons via Wikipedia)

How can we get more product?

Chemical manufacturing plants are always seeking to improve their processes. One way that improvement comes about is through measuring the amount of material produced in a reaction. By knowing how much is made, the scientists and engineers can try different ways of getting more product at less cost.

Conversions Between Moles and Mass

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment, you need 3.00 moles of calcium chloride $(\mathrm{CaCl_2})$. Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. The molar mass of $\mathrm{CaCl_2}$ is $110.98\ \mathrm{g/mol}$. The conversion factor that can be used is then based on the equality that $1\ \mathrm{mol} = 110.98\ \mathrm{g\ CaCl_2}$. Dimensional analysis will allow you to calculate the mass of $\mathrm{CaCl_2}$ that you should measure.

$$3.00\ \mathrm{mol\ CaCl_2} \times \frac{110.98\ \mathrm{g\ CaCl_2}}{1\ \mathrm{mol\ CaCl_2}} = 333\ \mathrm{g\ CaCl_2}$$

When you measure the mass of $333\ \mathrm{g}$ of $\mathrm{CaCl_2}$, you are measuring 3.00 moles of $\mathrm{CaCl_2}$.



Figure 9.5.2: Calcium chloride is used as a drying agent and as a road deicer. (Public Domain; Martin Walker (User:Walkerma/Wikimedia Commons) via Wikipedia)

Example 9.5.1: Converting Moles to Mass

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

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

Known

- Molar mass of $\mathrm{Cr} = 52.00\ \mathrm{g/mol}$
- $0.560\ \mathrm{mol\ Cr}$

Unknown

- $0.560\ \mathrm{mol\ Cr}$ = ? g

One conversion factor will allow us to convert from the moles of Cr to mass.

Step 2: Calculate.

$$0.560\ \mathrm{mol\ Cr} \times \frac{52.00\ \mathrm{g\ Cr}}{1\ \mathrm{mol\ Cr}} = 29.1\ \mathrm{g\ Cr}$$

Step 3: Think about your result.

Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because of the $0.560\ \mathrm{mol}$.

A similar conversion factor utilizing molar mass can be used to convert from the mass of a substance to moles. In a laboratory situation, you may perform a reaction and produce a certain amount of a product which can be massed. It will often then be necessary to determine the number of moles of the product that was formed. The next problem illustrates this situation.

Example 9.5.2: Converting Mass to Moles

A certain reaction produces $2.81\ \mathrm{g}$ of copper (II) hydroxide, $\mathrm{Cu(OH)_2}$. Determine the number of moles produced in the reaction.

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

Known

- mass = 2.81g

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Unknown

- mol $Cu(OH)_2$

One conversion factor will allow us to convert from mass to moles.

Step 2: Calculate.

First, it is necessary to calculate the molar mass of $\mathrm{Cu(OH)_2}$ from the molar masses of Cu, O, and H. The molar mass is $97.57\ \mathrm{g/mol}$.

$$2.81\ \mathrm{g\ Cu(OH)_2} \times \frac{1\ \mathrm{mol\ Cu(OH)_2}}{97.57\ \mathrm{g\ Cu(OH)_2}} = 0.0288\ \mathrm{mol\ Cu(OH)_2}$$

Step 3: Think about your result.

The relatively small mass of product formed results in a small number of moles.

YouTube Video

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

View this YouTube Video online

Simulation

Practice converting mass to moles in everyday objects in this simulation:

How many atoms are in a gold ring, a grain of sand, or the human body?

Summary

- Calculations involving conversions between moles of a material and the mass of that material are described.

Review

1. You have 19.7 grams of a material and wonder how many moles were formed. Your friend tells you to multiply the mass by grams/mole. Is your friend correct?
2. How many grams of MgO are in 3.500 moles?
3. How many moles of H_2O are in 15.2 grams of pure ice?

This page titled 9.5: Conversions Between Moles and Mass is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.4: Conversions Between Moles and Mass by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.6: Conversions Between Mass and Number of Particles

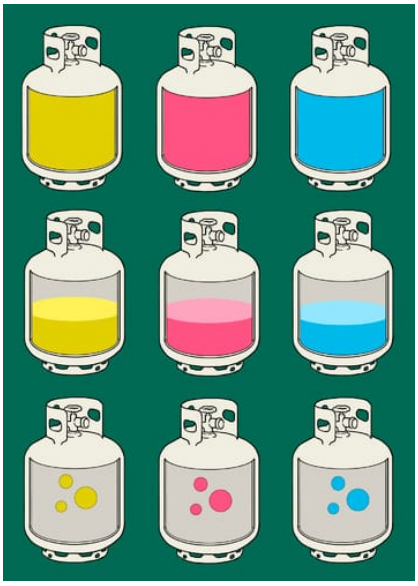


Figure 9.6.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

How much gas is there?

Avogadro was interested in studying gases. He theorized that equal volumes of gases under the same conditions contained the same number of particles. Other researchers studied how many gas particles were in specific volumes of gases. Eventually, scientists were able to develop the relationship between number of particles and mass, using the idea of moles.

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and the mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to grams. In order to convert from mass to number of particles or vice-versa, a conversion to moles is required.



Figure 9.6.2: Conversion from number of particles to mass, or from mass to number of particles requires two steps. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 9.6.1: Converting Mass to Particles

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

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

Known

- Molar mass $\mathrm{Cl_2} = 70.90\ \mathrm{g/mol}$
- $20.0\ \mathrm{g}\ \mathrm{Cl_2}$

Unknown

- number of molecules of Cl_2

Use two conversion factors. The first converts grams of $\mathrm{Cl_2}$ to moles. The second converts moles of $\mathrm{Cl_2}$ to the number of molecules.

Step 2: Calculate.

$$20.0\ \mathrm{g\ Cl_2} \times \frac{1\ \mathrm{mol\ Cl_2}}{70.90\ \mathrm{g\ Cl_2}} \times \frac{6.02 \times 10^{23}\ \mathrm{molecules\ Cl_2}}{1\ \mathrm{mol\ Cl_2}} = 1.70 \times 10^{23}\ \mathrm{molecules\ Cl_2}$$

The problem is done using two consecutive conversion factors. There is no need to explicitly calculate the moles of $\mathrm{Cl_2}$.

Step 3: Think about your result.

Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

YouTube Video

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

View this YouTube Video online

Summary

- Calculations are illustrated for conversions between mass and number of particles.

Review

1. Why can't we convert directly from number of particles to grams?
2. How many atoms of chlorine are present in 1.70×10^{23} molecules Cl_2?
3. How many molecules of BH_3 are in 14.32 grams BH_3?

This page titled 9.6: Conversions Between Mass and Number of Particles is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.5: Conversions Between Mass and Number of Particles by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.7: Avogadro's Hypothesis and Molar Volume



Figure 9.7.1 (Public Domain; User:Mark.murphy/Wikimedia Commons via Wikipedia)

How do scuba divers know if they will run out of gas?

Knowing how much gas is available for a dive is crucial to a diver's survival. The tank on the diver's back is equipped with gauges to indicate how much gas is present and what the pressure is. A basic knowledge of gas behavior allows the diver to assess how long they can stay underwater without developing problems.

Avogadro's Hypothesis and Molar Volume

Volume is a third way to measure the amount of matter, after item count and mass. With liquids and solids, volume varies greatly depending on the density of the substance. This is because solid and liquid particles are packed close together with very little space in between the particles. However, gases are largely composed of empty spaces between the actual gas particles (see figure below).

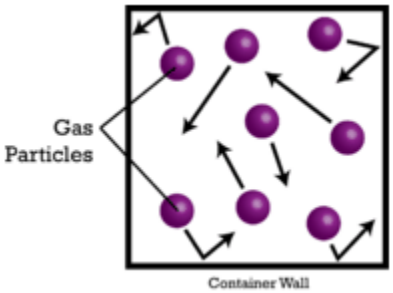


Figure 9.7.2: Gas particles are very small compared to the large amounts of empty space between them. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In 1811, Amadeo Avogadro explained that the volumes of all gases can be easily determined. **Avogadro's hypothesis** states that equal volumes of all gases at the same temperature and pressure contain equal numbers of particles. Since the total volume that a gas occupies is made up primarily of the empty space between the particles, the actual size of the particles themselves is nearly negligible. A given volume of a gas with small light particles, such as hydrogen $(\mathrm{H_2})$, contains the same number of particles as the same volume of a heavy gas with large particles, such as sulfur hexafluoride, $\mathrm{SF_6}$.

Gases are compressible, meaning that when put under high pressure, the particles are forced closer to one another. This decreases the amount of empty space and reduces the volume of the gas. Gas volume is also affected by temperature. When a gas is heated, its molecules move faster and the gas expands. Because of the variation in gas volume due to pressure and temperature changes, the comparison of gas volumes must be done at standard temperature and pressure. **Standard temperature and pressure (STP)** is defined as $0^{\circ}\mathrm{C}\ (273.15\ \mathrm{K})$ and $1\ \mathrm{atm}$ pressure. The **molar volume** of a gas is the volume of one mole of a gas at STP. At STP, one mole (6.02×10^{23} representative particles) of any gas occupies a volume of $22.4\ \mathrm{L}$ (figure below).

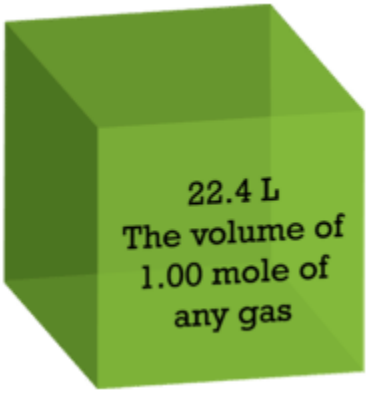
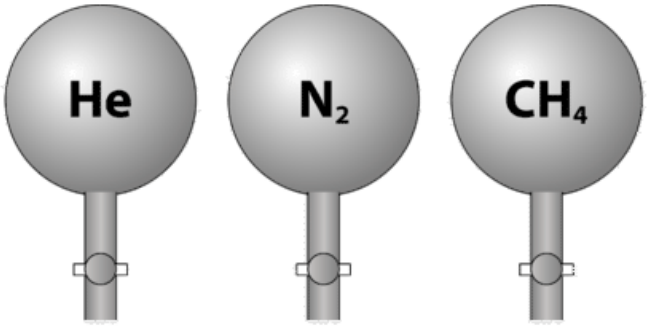


Figure 9.7.3: A mole of any gas occupies $22.4\ \mathrm{L}$ at standard temperature and pressure ($0^{\circ}\mathrm{C}$ and $1\ \mathrm{atm}$). (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The figure below illustrates how molar volume can be seen when comparing different gases. Samples of helium (He), nitrogen $(\mathrm{N_2})$, and methane $(\mathrm{CH_4})$ are at STP. Each contains 1 mole or 6.02×10^{23} particles. However, the mass of each gas is different and corresponds to the molar mass of that gas: $4.00\ \mathrm{g/mol}$ for He, $28.0\ \mathrm{g/mol}$ for $\mathrm{N_2}$, and $16.0\ \mathrm{g/mol}$ for $\mathrm{CH_4}$.



Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Figure 9.7.4: Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of particles. At standard temperature and pressure, 1 mole of any gas occupies $22.4\ \mathrm{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Summary

- Equal volumes of gases at the same conditions contain the same number of particles.
- Standard temperature and pressure is abbreviated (STP).
- Standard temperature is 0°C (273.15 K) and standard pressure is 1 atm.
- At STP, one mole of any gas occupies a volume of 22.4 L

Review

- A container is filled with gas, what do we know about the space actually taken up by a gas?
- Why do we need to do all our comparisons at the same temperature and pressure?
- At standard temperature and pressure, 1 mole of gas is always equal to how many liters?

This page titled 9.7: Avogadro's Hypothesis and Molar Volume is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.6: Avogadro's Hypothesis and Molar Volume by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.8: Conversions Between Moles and Gas Volume

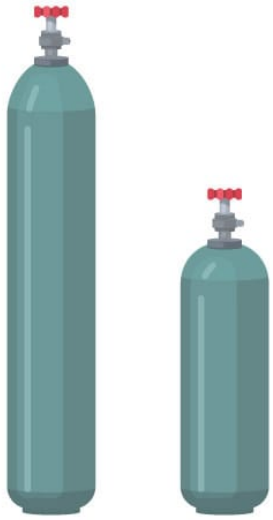


Figure 9.8.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How can you tell how much gas is in these containers?

Small gas tanks are often used to supply gases for chemistry reactions. A gas gauge will give some information about how much is in the tank, but quantitative estimates are needed so that the reaction will be able to proceed to completion. Knowing how to calculate the necessary parameters for gases is very helpful to avoid running out earlier than desired.

Conversions Between Moles and Gas Volume

Molar volume at STP can be used to convert from moles to gas volume and from gas volume to moles. The equality of $1\ \text{mol} = 22.4\ \text{L}$ is the basis for the conversion factor.

Example 9.8.1: Converting Gas Volume to Moles

Many metals react with acids to produce hydrogen gas. A certain reaction produces $86.5\ \text{L}$ of hydrogen gas at STP. How many moles of hydrogen were produced?

Solution

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

Known

- $86.5\ \text{L}\ \text{H}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- moles of H_2

Apply a conversion factor to convert from liters to moles.

Step 2: Calculate.

$$86.5\ \text{L}\ \text{H}_2 \times \frac{1\ \text{mol}\ \text{H}_2}{22.4\ \text{L}\ \text{H}_2} = 3.86\ \text{mol}\ \text{H}_2$$

Step 3: Think about your result.

The volume of gas produced is nearly four times larger than the molar volume. The fact that the gas is hydrogen plays no role in the calculation.

Example 9.8.2: Convertig Moles to Gas Volume

What volume does $4.96\ \text{mol}$ of O_2 occupy at STP?

Solution

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

Known

- $4.96\ \text{mol}\ \text{O}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- volume of O_2

Step 2: Calculate.

$$4.96\ \text{mol} \times 22.4\ \text{L/mol} = 111.1\ \text{L}$$

Step 3: Think about your result.

The volume seems correct given the number of moles.

Example 9.8.3: Converting Volume to Mass

If we know the volume of a gas sample at STP, we can determine how much mass is present. Assume we have $867\ \text{L}$ of N_2 at STP. What is the mass of the nitrogen gas?

Solution

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

Known

- $867\ \text{L}\ \text{N}_2$
- $1\ \text{mol} = 22.4\ \text{L}$
- Molar mass $\text{N}_2 = 28.02\ \text{g/mol}$

Unknown

- mass of N_2

Step 2: Calculate.

We start by determining the number of moles of gas present. We know that 22.4 liters of a gas at STP equals one mole, so:

$$867\ \text{L} \times \frac{1\ \text{mol}}{22.4\ \text{L}} = 38.7\ \text{mol}$$

We also know the molecular weight of N_2 $(28.0\ \text{g/mol})$, so we can then calculate the weight of nitrogen gas in 867 liters:

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

$$38.7 \text{ mol} \times \frac{28 \text{ g}}{1 \text{ mol}} = 1083.6 \text{ g } N_2$$

Step 3: Think about your result.

In a multi-step problem, be sure that the units check out.

YouTube Video

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

View this YouTube Video online

Summary

- Conversions between moles and volume of a gas are shown.

Review

1. In the problems above, why was the gas always at standard temperature and pressure?
2. A container contains 45.2 L of N_2 gas at STP. How many moles of N_2 gas are in the container?
3. If the gas in the previous problem was CH_4 gas at STP instead of N_2 gas, then how many moles of CH_4 gas would there be?

This page titled 9.8: Conversions Between Moles and Gas Volume is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.7: Conversions Between Moles and Gas Volume by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.9: Gas Density



Figure $9.9.1$ (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Why does carbon dioxide sink in air?

When we run a reaction to produce a gas, we expect it to rise into the air. Many students have done experiments where gases such as hydrogen are formed. The gas can be trapped in a test tube held upside-down over the reaction. Carbon dioxide, on the other hand, sinks when it is released. Carbon dioxide has a density greater than air, so it will not rise like the hydrogen gas.

Gas Density

As you know, density is defined as the mass per unit volume of a substance. Since gases all occupy the same volume on a per mole basis, the density of a particular gas is dependent on its molar mass. A gas with a small molar mass will have a lower density than a gas with a large molar mass. Gas densities are typically reported in $\mathrm{g/L}$. Gas density can be calculated from molar mass and molar volume.



Figure 9.9.2: Balloons filled with helium gas float in air because the density of helium is less than the density of air. (Public Domain; Photographer: Warren Denning, courtesy of the Pioneer Balloon Company via Wikipedia)

Example 9.9.1: Gas Density

What is the density of nitrogen gas at STP?

Solution

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

Known

- $\mathrm{N_2} = 28.02\ \mathrm{g/mol}$
- $1\ \mathrm{mol} = 22.4\ \mathrm{L}$

Unknown

- density = ? g/L

Molar mass divided by molar volume yields the gas density at STP.

Step 2: Calculate.

$$\frac{28.02\ \mathrm{g}}{1\ \mathrm{mol}} \times \frac{1\ \mathrm{mol}}{22.4\ \mathrm{L}} = 1.25\ \mathrm{g/L}$$

When set up with a conversion factor, the mol unit cancels, leaving $\mathrm{g/L}$ as the unit in the result.

Step 3: Think about your result.

The molar mass of nitrogen is slightly larger than molar volume, so the density is slightly greater than $1\ \mathrm{g/L}$.

Alternatively, the molar mass of a gas can be determined if the density of the gas at STP is known.

Example 9.9.2: Molar Mass from Gas Density

What is the molar mass of a gas whose density is $0.761\ \mathrm{g/L}$ at STP?

Solution

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

Known

- $\mathrm{N_2} = 28.02\ \mathrm{g/mol}$
- $1\ \mathrm{mol} = 22.4\ \mathrm{L}$

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Unknown

- molar mass = ? g/mol

Molar mass is equal to density multiplied by molar volume.

Step 2: Calculate.

$$\frac{0.761\ \text{g}}{1\ \text{L}} \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 17.0\ \text{g/mol}$$

Step 3: Think about your result.

Because the density of the gas is less than $1\ \text{g/L}$, the molar mass is less than 22.4.

Summary

- Calculations are described showing conversions between molar mass and density for gases.

Review

1. How is density calculated?
2. How is molar mass calculated?
3. What would be the volume of 3.5 moles of a gas?

This page titled 9.9: Gas Density is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.8: Gas Density by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.10: Mole Road Map



Figure 9.10.1\ (Public Domain; Doug Smith (Wikimedia: Xsmith) via Commons Wikimedia, Currituck Sound Map [commons.wikimedia.org])

How do I get from here to there?

If I want to visit the town of Manteo, North Carolina, out on the coast, I will need a map of how to get there. I may have a printed map or I may download directions from the internet, but I need something to get me going in the right direction. Chemistry road maps serve the same purpose. How do I handle a certain type of calculation? There is a process and a set of directions to help.

Mole Road Map

Previously, we saw how the conversions between mass and number of particles required two steps, with moles as the intermediate. This concept can now be extended to also include gas volume at STP. The resulting diagram is referred to as a mole road map (see figure below).

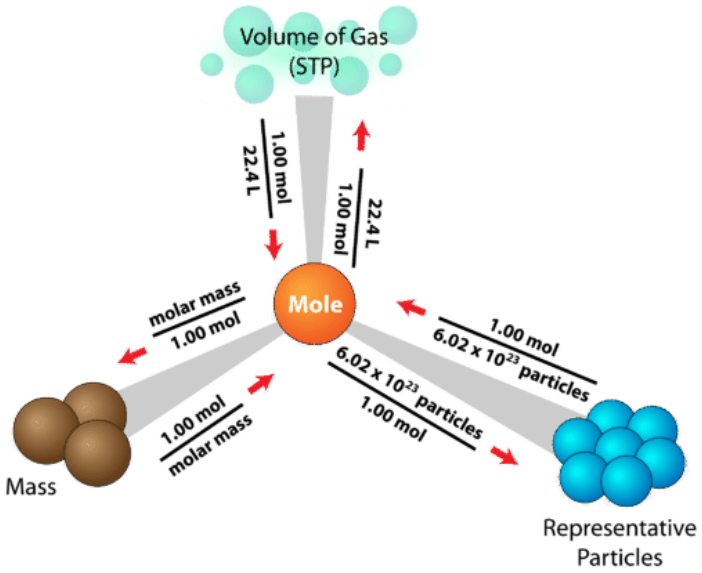


Figure 9.10.2: The mole road map shows the conversion factors needed to interconvert mass, number of particles, and volume of a gas. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The mole is at the center of any calculation involving amount of a substance. The sample problem below is one of many different problems that can be solved using the mole road map.

Example 9.10.1: Mole Road Map

What is the volume of $79.3\ \text{g}$ of neon gas at STP?

Solution

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

Known

- $\text{Ne} = 20.18\ \text{g/mol}$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- volume = ? L

The conversion factors will be grams $\rightarrow$ moles $\rightarrow$ gas volume.

Step 2: Calculate.

$$79.3\ \text{g Ne} \times \frac{1\ \text{mol Ne}}{20.18\ \text{g Ne}} \times \frac{22.4\ \text{L Ne}}{1\ \text{mol Ne}} = 88.0\ \text{L Ne}$$

Step 3: Think about your result.

The given mass of neon is equal to about 4 moles, resulting in a volume that is about 4 times larger than molar volume.

YouTube Video

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

View this YouTube Video online

Summary

- An overall process is given for calculations involving moles, grams, and gas volume.

Review

1. In the problem above, what is the formula weight of neon?
2. What value is at the center of all the calculations?
3. If we had 79.3 grams of Xe, would we expect a volume that is greater than or less than that obtained with neon?

This page titled 9.10: Mole Road Map is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.8: Mole Road Map by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.11: Percent Composition



Figure 9.11.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Is there anything healthy in this jar?

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

Percent Composition

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

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

Percent Composition from Mass Data

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

Example 9.11.1: Percent Composition from Mass

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

Solution

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

Known

- Mass of compound $= 20.00\ \text{g}$
- Mass of $\text{Zn} = 16.07\ \text{g}$

Unknown

- Percent $\mathrm{Zn} = ?\%$
- Percent $\mathrm{O} = ?\%$

Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass.

Step 2: Calculate.

$$\text{Mass of oxygen} = 20.00\ \text{g} - 16.07\ \text{g} = 3.93\ \text{g O}$$

$$\%\ \mathrm{Zn} = \frac{16.07\ \text{g Zn}}{20.00\ \text{g}} \times 100\% = 80.35\%\ \mathrm{Zn}$$

$$\%\ \mathrm{O} = \frac{3.93\ \text{g O}}{20.00\ \text{g}} \times 100\% = 19.65\%\ \mathrm{O}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\ \text{by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example $9.11.2$**: Percent Composition from Chemical Formula**

Dichlorineheptoxide $(\mathrm{Cl_2O_7})$ is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorineheptoxide.

Solution

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

Known

- Mass of Cl in $1\ \text{mol}\ \mathrm{Cl_2O_7} = 70.90\ \text{g}$
- Mass of O in $1\ \text{mol}\ \mathrm{Cl_2O_7} = 112.00\ \text{g}$
- Molar mass of $\mathrm{Cl_2O_7} = 182.90\ \text{g/mol}$

Unknown

- Percent $\mathrm{Cl} = ?\%$
- Percent $\mathrm{O} = ?\%$

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\%\text{Cl} = \frac{70.90\text{ g Cl}}{182.90\text{ g}} \times 100\% = 38.76\%\text{ Cl}$$

$$\%\text{ O} = \frac{112.00\text{ g O}}{182.90\text{ g}} \times 100\% = 61.24\%\text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \text{Cl}$ and $61.24\%\ \text{O}$. Suppose that you need to know the masses of chlorine and oxygen present in a $12.50\ \text{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element:

$$12.50\text{ g } \text{Cl}_2\text{O}_7 \times \frac{38.76\text{ g Cl}}{100\text{ g } \text{Cl}_2\text{O}_7} = 4.845\text{ g Cl}$$

$$12.50\text{ g } \text{Cl}_2\text{O}_7 \times \frac{61.24\text{ g O}}{100\text{ g } \text{Cl}_2\text{O}_7} = 7.655\text{ g O}$$

The sum of the two masses is $12.50\ \text{g}$, the mass of the sample size.

YouTube Video

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

View this YouTube Video online

Science Friday: Stained Glass Conservation

Stained glass from the Middle Ages is often hundreds of years old. Unfortunately, many of these relics are in need of cleaning and maintenance. In this video by Science Friday, conservator Mary Higgins discusses the methods used to protect the stained glass.

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

View this YouTube Video online

Summary

- Processes are described for calculating the percent composition of a material based on mass or on chemical composition.

Review

1. What is the formula for calculating percent composition?
2. What information do you need to calculate percent composition by mass?
3. What do subscripts in a chemical formula tell you?

This page titled 9.11: Percent Composition is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.10: Percent Composition by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.12: Percent of Water in a Hydrate



Figure 9.12.1 (Public Domain; Doug Smith (Wikimedia: Xsmith) via Wikipedia)

Why does the color change?

If you look at a typical bottle of copper sulfate, it will be a bluish-green. If someone were to tell you that copper sulfate is white, you likely would not believe them. However, it turns out that you are both right; it just depends on the copper sulfate. Your blue-green copper sulfate has several water molecules attached to it, while your friend's copper sulfate is anhydrous (no water attached). Why the difference? The water molecules interact with some of the d electrons in the copper ion and produce the color. When the water is removed, the electron configuration changes and the color disappears.

Percent of Water in a Hydrate

Many ionic compounds naturally contain water as part of the crystal lattice structure. A **hydrate** is a compound that has one or more water molecules bound to each formula unit. Ionic compounds that contain a transition metal are often highly colored. Interestingly, it is common for the hydrated form of a compound to be of a different color than the **anhydrous** form, which has no water in its structure. A hydrate can usually be converted to the anhydrous compound by heating. For example, the anhydrous compound cobalt (II) chloride is blue, while the hydrate is a distinctive magenta color.

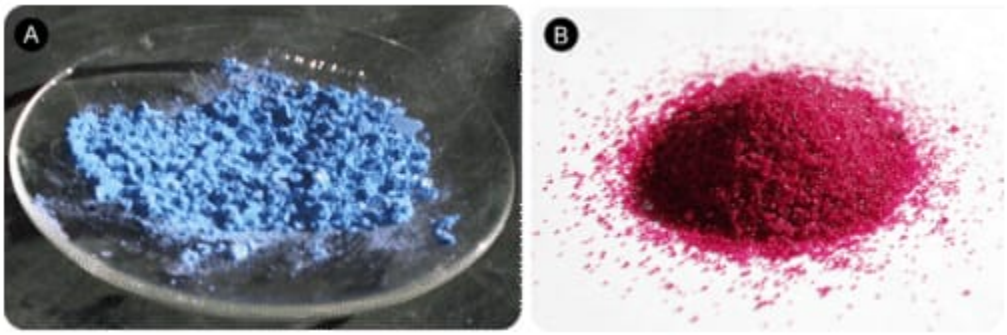


Figure 9.12.2: On the left is anhydrous cobalt (II) chloride, $\mathrm{CoCl_2}$. On the right is the hydrated form of the compound called cobalt (II) chloride hexahydrate, $\mathrm{CoCl_2 \cdot 6H_2O}$. (Public Domain; (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Cobalt%2528II%2529_chloride.jpg; (B) http://commons.wikimedia.org/wiki/File:Cobalt%2528II%2529-chloride-hexahydrate-sample.jpg)

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

The hydrated form of cobalt (II) chloride contains six water molecules in each formula unit. The name of the compound is cobalt (II) chloride hexahydrate and its formula is $\mathrm{CoCl_2 \cdot 6H_2O}$. The formula for water is set apart at the end of the formula with a dot, followed by a coefficient that represents the number of water molecules per formula unit.

It is useful to know the percent of water contained within a hydrate. The sample problem below demonstrates the procedure.

Example $9.12.1$**: Percent of Water in a Hydrate**

Find the percent water in cobalt (II) chloride hexahydrate, $\mathrm{CoCl_2 \cdot 6H_2O}$.

Solution

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

The mass of water in the hydrate is the coefficient (6) multiplied by the molar mass of $\mathrm{H_2O}$. The molar mass of the hydrate is the molar mass of the $\mathrm{CoCl_2}$ plus the mass of water.

Known

- Mass of $\mathrm{H_2O}$ in $1\ \mathrm{mol}$ hydrate $= 108.12\ \mathrm{g}$
- Molar mass of hydrate $= 237.95\ \mathrm{g/mol}$

Unknown

- percent H_2O = ?%

Calculate the percent by mass of water by dividing the mass of $\mathrm{H_2O}$ in 1 mole of the hydrate by the molar mass of the hydrate and multiplying by 100%.

Step 2: Calculate.

$$\%\ \mathrm{H_2O} = \frac{108.12\ \mathrm{g\ H_2O}}{237.95\ \mathrm{g}} \times 100\% = 45.44\%\ \mathrm{H_2O}$$

Step 3: Think about your result.

Nearly half of the mass of the hydrate is composed of water molecules within the crystal.

YouTube Video

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

View this YouTube Video online

Summary

- The process of calculating the percent water in a hydrate is described.

Review

1. What is a hydrate?
2. How can you convert a hydrate to an anhydrous compound?
3. What does hexahydrate mean?

This page titled 9.12: Percent of Water in a Hydrate is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.11: Percent of Water in a Hydrate by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.13: Determining Empirical Formulas



Figure 9.13.1 (Public Domain; Harriet Moore via Wikipedia)

What is occuring in this picture?

In the early days of chemistry, there were few tools available 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 many of these materials had carbon, hydrogen, oxygen, and possibly nitrogen in simple ratios. Chemists had no way to determine the exact amounts of these atoms that were contained in specific molecules.

Determining Empirical Formulas

An **empirical formula** is one that shows the lowest whole-number ratio of the elements in a compound. Because the structure of ionic compounds is an extended three-dimensional network of positive and negative ions, all formulas of ionic compounds are empirical. However, we can also consider the empirical formula of a molecular compound. Ethene is a small hydrocarbon compound with the formula $\mathrm{C_2H_4}$ (see figure below). While $\mathrm{C_2H_4}$ is its molecular formula and represents its true molecular structure, it has an empirical formula of $\mathrm{CH_2}$. The simplest ratio of carbon to hydrogen in ethene is 1:2. There are two ways to view that ratio. Considering one molecule of ethene, the ratio is 1 carbon atom for every 2 atoms of hydrogen. Considering one mole of ethene, the ratio is 1 mole of carbon for every 2 moles of hydrogen. So, the subscripts in a formula represent the mole ratio of the elements in that formula.

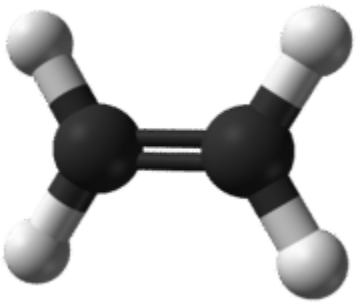


Figure 9.13.2: Ball-and-stick model of ethene, $\mathrm{C_2H_4}$. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

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. The steps to be taken are outlined below.

1. Assume a $100\ \text{g}$ sample of the compound, so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.
3. In order to find a whole-number ratio, divide the moles of each element by whichever of the moles from step 2 is the smallest.
4. If all the moles at this point are whole numbers (or very close), the empirical formula can be written with the moles as the subscript of each element.
5. In some cases, one or more of the moles calculated in step 3 will not be whole numbers. Multiply each of the moles by the smallest whole number that will convert each into a whole number. Write the empirical formula.

Example $9.13.1$**: Determining the Empirical Formula of a Compound**

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

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

Known

- $\%$ of $\text{Fe} = 69.94\%$
- $\%$ of $\text{O} = 30.06\%$

Unknown

- Empirial formula = Fe?O?

Steps to follow are outlined in the text.

Step 2: Calculate.

1. Assume a $100\ \text{g}$ sample.

$$69.94\ \text{g Fe}$$

$$30.06\ \text{g O}$$

2. Convert to moles.

$$69.94\ \text{g Fe} \times \frac{1\ \text{mol Fe}}{55.85\ \text{g Fe}} = 1.252\ \text{mol Fe}$$

$$30.06\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 1.879\ \text{mol O}$$

3. Divide both moles by the smallest of the results.

$$\frac{1.252\ \text{mol Fe}}{1.252} = 1\ \text{mol Fe} \quad \frac{1.879\ \text{mol O}}{1.252} = 1.501\ \text{molO}$$

4/5. Since the moles of O is still not a whole number, both moles can be multiplied by 2, while rounding to a whole number.

$$1\ \text{mol Fe} \times 2 = 2\ \text{mol Fe} \quad 1.501\ \text{mol O} \times 2 = 3\ \text{mol O}$$

The empirical formula of the compound is Fe_2O_3.

Step 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.

YouTube Video

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

View this YouTube Video online

Summary

- A process is described for the calculation of the empirical formula for a compound based on the percent composition of that compound.

Review

1. What is an empirical formula?
2. What does an empirical formula tell you?
3. What does it not tell you?

This page titled 9.13: Determining Empirical Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.12: Determining Empirical Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

9.14: Determining Molecular Formulas

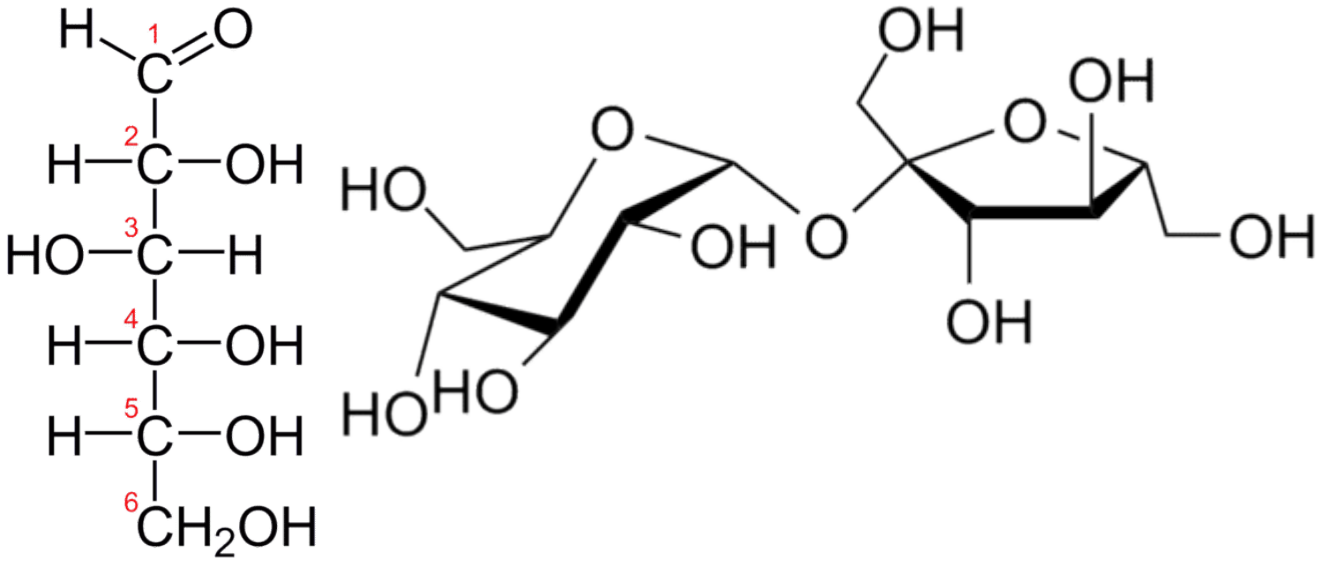


Figure 9.14.1 (Left: Credit: Ben Mills (Wikimedia: Benjah-bmm27) Source: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png License: Public Domain) (Right: Credit: User:glycoform/ Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27); Source: http://commons.wikimedia.org/wiki/File:Sucrose_3Dprojection.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png; License: Public Domain)

How can you determine the differences between these two molecules?

Above 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 molecular weights—this approach allows you to easily tell which compound is which.

Molecular Formulas

Molecular formulas give the kind and number of atoms of each element present in a molecular compound. In many cases, the molecular formula is the same as the empirical formula. The molecular formula of methane is $\mathrm{CH_4}$ and because it contains only one carbon atom, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole-number multiple of the empirical formula. Acetic acid is an organic acid that is the main component of vinegar. Its molecular formula is $\mathrm{C_2H_4O_2}$. Glucose is a simple sugar that cells use as a primary source of energy. Its molecular formula is $\mathrm{C_6H_{12}O_6}$. The structures of both molecules are shown in the figure below. They are very different compounds, yet both have the same empirical formula of $\mathrm{CH_2O}$.

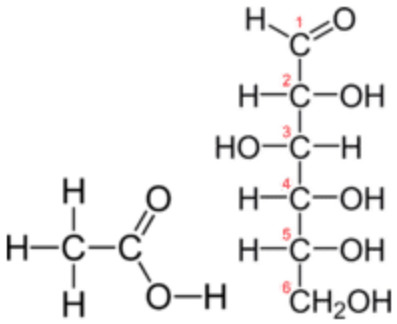


Figure 9.14.2: Acetic acid (left) has a molecular formula of $\mathrm{C_2H_4O_2}$, while glucose (right) has a molecular formula of $\mathrm{C_6H_{12}O_6}$. Both have the empirical formula $\mathrm{CH_2O}$. (Public Domain; (left) Ben Mills (Wikimedia: Benjah-bmm27); (right) Ben Mills (Wikimedia: Benjah-bmm27), User:Yikrazuul/Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27) via Acetic acid: http://commons.wikimedia.org/wiki/File:Acetic-acid-2D-flat.png; Glucose: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png)

Empirical formulas can be determined from the percent composition of a compound. 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)**, which is simply the molar mass represented by the empirical formula.
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 9.14.1: Determining the Molecular Formula of a Compound

The empirical formula of a compound of boron and hydrogen is $\mathrm{BH_3}$. Its molar mass is $27.7\ \mathrm{g/mol}$. Determine the molecular formula of the compound.

Solution

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

Known

- Empirical formula $= \mathrm{BH_3}$
- Molar mass $= 27.7\ \mathrm{g/mol}$

Unknown

- molecular formula = ?

Steps to follow are outlined in the text.

Step 2: Calculate.

$$\text{Empirical formula mass (EFM)} = 13.84\ \mathrm{g/mol}$$

$$\frac{\text{molar mass}}{\text{EFM}} = \frac{17.7}{13.84} = 2$$

$$\mathrm{BH_3} \times 2 = \mathrm{B_2H_6}$$

The molecular formula of the compound is $\mathrm{B_2H_6}$.

Step 3: Think about your result.

The molar mass of the molecular formula matches the molar mass of the compound.

YouTube Video

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

View this YouTube Video online

Summary

- A procedure is described for the calculation of the exact molecular formula of a compound.

Review

1. What is the difference between an empirical formula and a molecular formula?
2. In addition to the elemental analysis, what do you need to know to calculate the molecular formula?
3. What does the empirical formula mass tell you?

This page titled 9.14: Determining Molecular Formulas is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 10.13: Determining Molecular Formulas by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10: CHEMICAL CALCULATIONS INVOLVING CHEMICAL EQUATIONS

10.1: Law of Conservation of Mass
10.2: Word Equations
10.3: Chemical Equations
10.4: Balancing Equations
10.5: Combination Reactions
10.6: Decomposition Reactions
10.7: Combustion Reactions
10.8: Single Replacement Reactions
10.9: Activity Series
10.10: Double Replacement Reactions
10.11: Everyday Stoichiometry
10.12: Mole Ratios
10.13: Mass-Mole Stoichiometry
10.14: Mass-Mass Stoichiometry
10.15: Volume-Volume Stoichiometry
10.16: Mass-Volume Stoichiometry
10.17: Limiting Reactant
10.18: Determining the Limiting Reactant
10.19: Theoretical Yield and Percent Yield

This page titled 10: Chemical Calculations Involving Chemical Equations is shared under a license and was authored, remixed, and/or curated by LibreTexts.

10.1: Law of Conservation of Mass



Figure 10.1.1 (CC BY 2.0; Douglas Heriot via Flickr)

Have you ever lost a screw?

The following situation happens all too often. You have taken apart a piece of equipment to clean it up. When you put the equipment back together, somehow you have an extra screw or two. Or you find out that a screw is missing that was a part of the original equipment. In either case, you know something is wrong. You expect to end up with the same amount of material that you started with, not with more or less than what you had originally.

Law of Conservation of Mass

By the late 1700s, chemists accepted the definition of an element as a substance that cannot be broken down into a simpler substance by ordinary chemical means. It was also clear that elements combine with one another to form more complex substances called compounds. The chemical and physical properties of these compounds are different than the properties of the elements from which they were formed. There were questions about the details of these processes.

In the 1790s, a greater emphasis began to be placed on the quantitative analysis of chemical reactions. Accurate and reproducible measurements of the masses of reacting elements and the compounds they form led to the formulation of several basic **laws**. One of these is called the law of **conservation of mass**, which states that during a chemical reaction, the total mass of the **products** must be equal to the total mass of the **reactants**. In other words, mass cannot be created or destroyed during a chemical reaction, but is always conserved.

As an example, consider the reaction between silver nitrate and sodium chloride. These two compounds will dissolve in water to form silver chloride and sodium nitrate. The silver chloride does not dissolve in water, so it forms a solid that we can filter off. When we evaporate the water, we can recover the sodium nitrate formed. If we react 58.5 grams of sodium chloride with 169.9 grams of silver nitrate, we start with 228.4 grams of materials. After the reaction is complete and the materials separated, we find that we have formed 143.4 grams of silver chloride and 85.0 grams of sodium nitrate, giving us a total mass of 228.4 grams for the products. So, the total mass of reactants equals the total mass of products, a proof of the law of conservation of mass.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Summary

- The law of conservation of mass states that, during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants.

Review

1. The law of conservation of mass states that, during a chemical reaction, the total _____ of the products must be equal to the total _____ of the reactants.
2. Describe an example of the law of conservation of mass.

This page titled 10.1: Law of Conservation of Mass is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

10.2: Word Equations

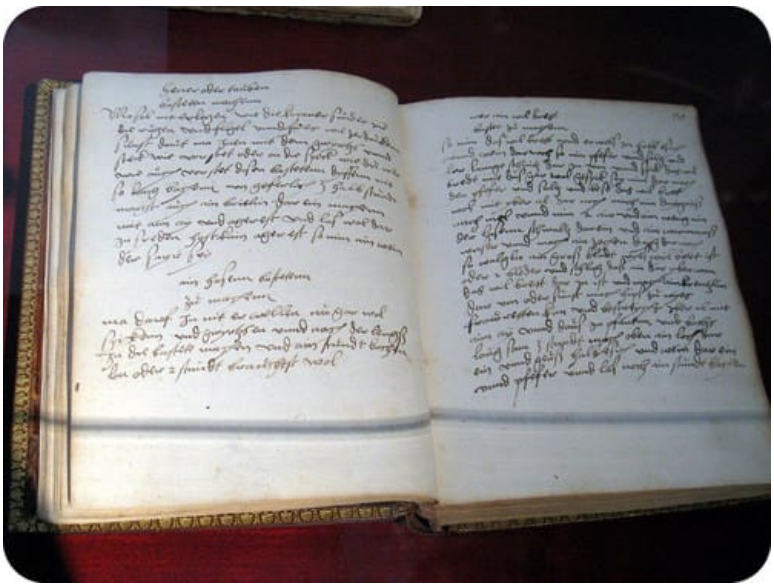


Figure 10.2.1 (Public Domain; User:Daderot/Wikimedia Commons via Wikipedia)

What's for Dinner?

Various ways of recording recipes have developed over the centuries. The cookbook shown above was written by a woman who probably collected all her own recipes. Later, printed cookbooks became available (even guys had no excuse for not being able to cook). Today we can find recipes on a number of internet sites and can quickly search for information on how to cook anything we want. Reading a recipe sometimes requires we understand a few codes and symbols (what's the difference between a tsp and a Tsp?), but the information on what we start with and what we end up with is there.

Writing Chemical Equations

Chemical reactions are occurring all around us. Plants use sunlight to drive their photosynthetic process and produce energy. Cars and other vehicles burn gasoline in order to power their engines. Batteries use electrochemical reactions to produce energy and power many everyday devices. Many chemical reactions are going on within the human body as well, particularly during the digestion of food.

In math class, you have written and solved many mathematical equations. Chemists keep track of chemical reactions by writing equations as well. In any chemical reaction, one or more substances—called **reactants**—are converted into one or more new substances—called **products**. The general form of the equation for such a process looks like this:

$$\text{Reactants} \rightarrow \text{Products}$$

Unlike a math equation, a chemical equation does not use an equal sign. Instead, the arrow is called a yield sign and so the equation is described as "reactants yield products".

Word Equations

You can describe a chemical reaction by writing a **word equation**. When silver metal is exposed to sulfur, it reacts to form silver sulfide. Silver sulfide is commonly known as tarnish and turns the surface of silver objects dark and streaky black (see figure below). The sulfur that contributes to tarnish can come from traces of sulfur in the air, or from foods such as eggs. The word equation for the process is:

$$\text{Silver} + \text{sulfur} \rightarrow \text{Silver sulfide}$$

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

The silver and the sulfur are the reactants in the equation, while the silver sulfide is the product.



Figure 10.2.2**:** You can see dark spots of tarnished (silver sulfide) forming on this ring as it reacts with sulfur compounds in the air. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Another common chemical reaction is the burning of methane gas. Methane is the major component of natural gas and is commonly burned on a gas stove or in a Bunsen burner (figure below). Burning is a chemical reaction in which some type of fuel is reacted with oxygen gas. The products of the reaction in the burning of methane, as well as other fuels, are carbon dioxide and water. The word equation for this reaction is:

$$\text{Methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$



Figure 10.2.3: A Bunsen burner is commonly used to heat substances in a chemistry lab. Methane is reacted with oxygen to form carbon dioxide and water. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Word equations can be very useful, but do have one major drawback—they cannot be used for any quantitative work. A word equation does not tell how many moles of each material are needed, or how many moles of product are formed.

YouTube Video

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

View this YouTube Video online

Summary

- Word equations are used to describe the conversion of reactants to products.

Review

1. Write the generic form of a chemical reaction.
2. What are reactants?
3. What are products?

This page titled 10.2: Word Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.1: Word Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.3: Chemical Equations



Figure 10.3.1 (Public Domain; jons2 at pdphoto.org via Wikipedia)

How do you make shrimp gumbo?

Shrimp gumbo is one of many dishes that are part of the Cajun culture in Louisiana. It's a spicy dish that needs careful control of all the ingredients so that it has a "kick", but is not overwhelming. Recipes convey not only what the preparation entails, but also describe how much of each ingredient is needed, and the details of how to cook the meal. Similarly, we need this type of information in order to carry out chemical reactions successfully and safely.

Chemical Equations

Word equations are time-consuming to write and do not prove to be convenient for many of the things that chemists need to do with equations. A **chemical equation** is a representation of a chemical reaction that displays the reactants and products with chemical formulas. The chemical equation for the reaction of methane with oxygen is shown:

$$\mathrm{CH_4 + O_2 \rightarrow CO_2 + H_2O}$$

The equation above, called a **skeleton equation**, is an equation that shows only the formulas of the reactants and products with nothing to indicate the relative amounts. The first step in writing an accurate chemical equation is to write the skeleton equation, making sure that the formulas of all substances involved are written correctly. All reactants are written to the left of the yield arrow, separated from one another by a plus sign. Likewise, products are written to the right of the yield arrow, also separated with a plus sign.

It is often important to know the physical states of the reactants and products taking part in a reaction. To do this, put the appropriate symbol in parentheses after each formula: (s) for solid, (l) for liquid, (g) for gas, and (aq) for an aqueous (water-based) solution. The previous reaction becomes:

$$\mathrm{CH_4\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)}$$

The table below shows a listing of symbols used in chemical equations. Some, such as the double arrow which represents equilibrium, and the use of a catalyst in a reaction, will be detailed in other concepts.

Table 10.3.1: Symbols Used in Chemical Equations

Symbols	Description
$+$	Used to separate multiple reactants or products.
$\rightarrow$	Yield sign; separates reactants from products.
$\rightleftharpoons$	Replaces the yield sign for reversible reactions that reach equilibrium.
(s)	Reactant or product in the solid state.
(l)	Reactant or product in the liquid state.
(g)	Reactant or product in the gaseous state.
(aq)	Reactant or product in an aqueous solution (dissolved in water).
$\overset{\text{Pt}}{\rightarrow}$	Formula written above the arrow is used as a catalyst in the reaction.
$\overset{\Delta}{\rightarrow}$	Triangle indicates that the reaction is being heated.

YouTube Video

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

View this YouTube Video online

Summary

- A chemical equation is representation a chemical reaction using chemical formulas reactants and products, as well as symbols.
- Symbols used in chemical equations are described and explained.

Review

1. What does a skeleton equation tell you?
2. Why would you want to know the physical state of materials in a chemical reaction?
3. In a chemical reaction what does symbol → mean?
4. If I see Δ over the arrow, what does this indicate?

This page titled 10.3: Chemical Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.2: Chemical Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.4: Balancing Equations



Figure 10.4.1 (Public Domain; Joseph Allen via Wikipedia)

Any leftovers?

When you cook a meal, quite often there are leftovers because you prepared more than people would eat at one sitting. Sometimes when you repair a piece of equipment, you end up with what are called "pocket parts"—small pieces you put in your pocket because you're not sure where they belong. Chemistry tries to avoid leftovers and pocket parts. In normal chemical processes, we cannot create or destroy matter (law of conservation of mass). If we start out with ten carbon atoms, we need to end up with ten carbon atoms. John Dalton's atomic theory said that chemical reactions basically involve the rearrangement of atoms. Chemical equations need to follow these principles in order to be correct.

Balancing Chemical Equations

A **balanced equation** is a chemical equation in which mass is conserved and there are equal numbers of atoms of each element on both sides of the equation. We can write a chemical equation for the reaction of carbon with hydrogen gas to form methane $(\mathrm{CH_4})$:

$$\underset{2\text{ C atoms}}{\mathrm{C}\,(s)} + \underset{2\text{ H atoms}}{\mathrm{H_2}\,(g)} \rightarrow \underset{1\text{ C atom},\,4\text{ H atoms}}{\mathrm{CH_4}\,(g)}$$

In order to write a correct equation, you must first write the correct skeleton equation with the correct chemical formulas. Recall that hydrogen is a diatomic molecule and so is written as $\mathrm{H_2}$.

When we count the number of atoms of both elements, shown under the equation, we see that the equation is not balanced. There are only 2 atoms of hydrogen on the reactant side of the equation, while there are 4 atoms of hydrogen on the product side. We can balance the above equation by adding a coefficient of 2 in front of the formula for hydrogen.

$$\mathrm{C}\,(s) + 2\mathrm{H_2}\,(g) \rightarrow \mathrm{CH_4}\,(g)$$

A **coefficient** is a small whole number placed in front of a formula in an equation in order to balance it. The 2 in front of the $\mathrm{H_2}$ means that there are a total of $2 \times 2 = 4$ atoms of hydrogen as reactants. Visually, the reaction looks like the figure below.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

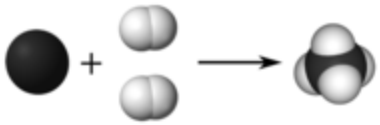


Figure 10.4.2: Reaction between carbon and hydrogen to form methane. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27 via Wikipedia and http://commons.wikimedia.org/wiki/File:Methane-3D-space-filling.png)

In the balanced equation, there is one atom of carbon and four atoms of hydrogen on both sides of the arrow. Below are guidelines for writing and balancing chemical equations.

1. Determine the correct chemical formulas for each reactant and product.
2. Write the skeleton equation.
3. Count the number of atoms of each element that appears as a reactant and as a product. If a polyatomic ion is unchanged on both sides of the equation, count it as a unit.
4. Balance each element one at a time by placing coefficients in front of the formulas.
   a. It is best to begin by balancing elements that only appear in one chemical formula on each side of the equation.
   b. No coefficient is written for a 1.
   c. NEVER change the subscripts in a chemical formula—you can only balance equations by using coefficients.
5. Check each atom or polyatomic ion to be sure that they are equal on both sides of the equation.
6. Make sure that all coefficients are in the lowest possible ratio. If necessary, reduce to the lowest ratio.

Example 10.4.1: Balancing Chemical Equations

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Step 1: Plan the problem.

Follow the steps for writing and balancing a chemical equation listed in the text.

Step 2: Solve.

Write the skeleton equation with the correct formulas.

$$\mathrm{Pb(NO_3)_2}\,(aq) + \mathrm{NaCl}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s)$$

Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3}^-$ ions	1 $\mathrm{NO_3}^-$ ions
1 Na atom	1 Na atom
1 Cl atom	2 Cl atoms

The nitrate ions and the chlorine atoms are unbalanced. Start by placing a 2 in front of the NaCl. This increases the reactant counts to 2 Na atoms and 2 Cl atoms. Then place a 2 in front of the $\mathrm{NaNO_3}$. The result is:

$$\mathrm{Pb(NO_3)_2}\,(aq) + 2\mathrm{NaCl}\,(aq) \rightarrow 2\mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s)$$

The new count for each atom and polyatomic ion becomes:

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3}^-$ ions	2 $\mathrm{NO_3}^-$ ions
2 Na atom	2 Na atom
2 Cl atom	2 Cl atoms

cK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

Step 3: Think about your result.

The equation is now balanced since there are equal numbers of atoms of each element on both sides of the equation.

YouTube Video

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

View this YouTube Video online

Simulation

Practice balancing chemical equations with this simulation:

Can you balance a chemical equation?

Summary

- The process of balancing chemical equations is described.

Review

1. What is the law of conservation of mass?
2. How did Dalton describe the process of a chemical reaction?
3. Why don't we change the subscripts in order to balance an equation?

This page titled 10.4: Balancing Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.3: Balancing Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.5: Combination Reactions



Figure 10.5.1 (Public Domain; Rim: User:Relaxatiallc/Wikimedia Commons; Wheel: Christopher Ziemnowicz via Rim: http://commons.wikimedia.org/wiki/File:Ats_amgpenta.jpg; Wheel: http://commons.wikimedia.org/wiki/File:1975_AACA_AMC_Pacer_X_red-white_wheel.jpg)

How useful is a wheel rim?

A wheel rim is not very useful by itself. Driving on the rim can damage it and make for a very rough ride. When the rim is combined with a tire, the product can be put on a car and used for a safe and comfortable ride. The two separate items combine to make something that improves the car ride.

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:

$$\mathrm{A} + \mathrm{B} \rightarrow \mathrm{AB}$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

$$2\mathrm{Na}\,(s) + \mathrm{Cl_2}\,(g) \rightarrow 2\mathrm{NaCl}\,(s)$$

In order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules: $\mathrm{H_2}$, $\mathrm{N_2}$, $\mathrm{O_2}$, $\mathrm{F_2}$, $\mathrm{Cl_2}$, $\mathrm{Br_2}$, and $\mathrm{I_2}$.

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.

$$2\mathrm{Mg}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{MgO}\,(s)$$

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Sulfur reacts with oxygen to form sulfur dioxide.

$$\mathrm{S}\,(s) + \mathrm{O_2}\,(g) \rightarrow \mathrm{SO_2}\,(g)$$

When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide.

$$2\mathrm{S}\,(s) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{SO_3}\,(g)$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide:

$$2\mathrm{Fe}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{FeO}\,(s)$$

$$4\mathrm{Fe}\,(s) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{Fe_2O_3}\,(s)$$

YouTube Video

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

View this YouTube Video online

Example $10.5.1$**: Combination Reactions**

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

Step 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 $\mathrm{K^+}$, while the oxide ion is $\mathrm{O^{2-}}$.

Step 2: Solve.

The skeleton (unbalanced) equation:

$$\mathrm{K}\,(s) + \mathrm{O_2}\,(g) \rightarrow \mathrm{K_2O}\,(s)$$

The equation is then easily balanced with coefficients.

$$4\mathrm{K}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{K_2O}\,(s)$$

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Step 3: Think about your result.

Formulas are correct and the resulting combination reaction is balanced.

Exercise 10.5.1

Can you build a balanced, combination reaction for the formation of aluminum oxide (Al_2O_3)?

Answer available at the end of this section.

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:

$$2\mathrm{CO}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{CO_2}\,(g)$$

Two compounds may also react to form a more complex compound. A very common example is the reaction of an oxide with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide.

$$\mathrm{CaO}\,(s) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{Ca(OH)_2}\,(aq)$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$\mathrm{SO_3}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2SO_4}\,(aq)$$

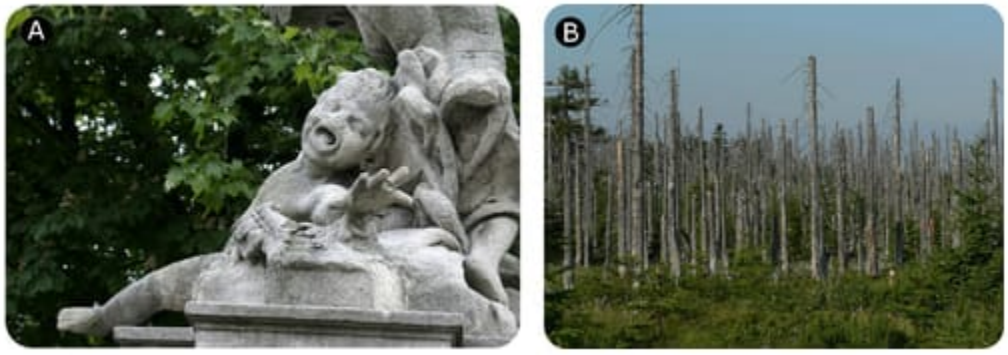


Figure 10.5.2: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain. ((A; (A) Nipik; (B) Nino Barbieri via (A) http://commons.wikimedia.org/wiki/File:Acid_rain_woods1.JPG; (B) http://commons.wikimedia.org/wiki/File:Pollution_-_Damaged_by_acid_rain.jpg) Public Domain; (B) CC by 2.5)

Summary

- Combination reactions occur when two or more substances combine to form a new substance.

Review

1. What are combination reactions?
2. Write the product of the following reaction: $Mg + H_2O \rightarrow$
3. Is $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ a combination reaction? Explain your answer.

Answers

Exercise 10.5.1

$4Al + 3O_2 \rightarrow 2Al_2O_3$

This page titled 10.5: Combination Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.4: Combination Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.6: Decomposition Reactions

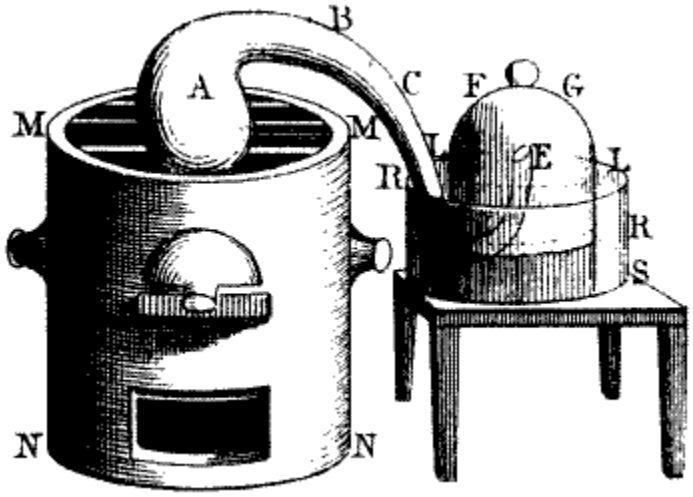


Figure 10.6.1 (Public Domain; Madame Lavoisier, modified by User: Cdang/Wikimedia Commons via Wikipedia)

How does a decomposition reaction work?

Antoine Lavoisier is widely known as the “father of modern chemistry.” He was one of the first to study chemical reactions in detail. Lavoisier reacted mercury with oxygen to form mercuric oxide as part of his studies on the composition of the atmosphere. He was then able to show that the decomposition of mercuric oxide produced mercury and oxygen. The diagram above shows the apparatus used by Lavoisier to study the formation and decomposition of mercuric oxide.

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$\mathrm{AB} \rightarrow \mathrm{A} + \mathrm{B}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas.

$$2\mathrm{HgO}\,(s) \rightarrow 2\mathrm{Hg}\,(l) + \mathrm{O_2}\,(g)$$

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 10.6.2: *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.* (Public Domain; Ben Mills (User:Benjah-bmm27/Wikimedia Commons) via Wikipedia)

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide:

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g)$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water:

$$2\mathrm{NaOH}\,(s) \rightarrow \mathrm{Na_2O}\,(s) + \mathrm{H_2O}\,(g)$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water:

$$\mathrm{H_2CO_3}\,(aq) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

YouTube Video

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

View this YouTube Video online

Exercise 10.6.1

Can you build a balanced Chemical reaction for the decomposition of Potassium chlorate ($KClO_3$)?

Answer available at the end of this section.

Example 10.6.1: Decomposition Reactions

When an electric current is passed through pure water, it decomposes into its elements. Write a balanced equation for the decomposition of water.

Solution

Step 1: 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.

Step 2: Solve.

The skeleton (unbalanced) equation:

$$\mathrm{H_2O}\,(l) \overset{\text{elec}}{\rightarrow} \mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$

Note the abbreviation "elec" above the arrow to indicate the passage of an electric current to initiate the reaction. Balance the equation.

$$2\mathrm{H_2O}\,(l) \overset{\text{elec}}{\rightarrow} 2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$

Step 3: Think about your result.

The products are elements and the equation is balanced.

YouTube Video

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

View this YouTube Video online

Simulations

How does an airbag explode at just the right time in the event of a car accident? Learn more in this simulation:

How does a chemical reaction inflate and airbag

Summary

- A definition of decomposition reaction and example reactions are given.

Review

1. What is a decomposition reaction?
2. What is usually needed for a decomposition reaction to take place?
3. Are elements always the product of a decomposition reaction?

Answers

Exercise 10.6.1

$2KClO_3 \rightarrow 2KCl + 3O_2$

This page titled 10.6: Decomposition Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.5: Decomposition Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.7: Combustion Reactions



Figure 10.7.1 (CC BY 2.0; Nina Hale via Wikipedia)

How do you cook the perfect marshmallow?

Roasting marshmallows over an open fire is a favorite past-time for campers, outdoor cook-outs, and just gathering around a fire in the back yard. 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 $\mathrm{O_2}$ as one reactant. The combustion of hydrogen gas produces water vapor:

$$\mathrm{2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g)}$$

Notice that this reaction also qualifies as a combination reaction.

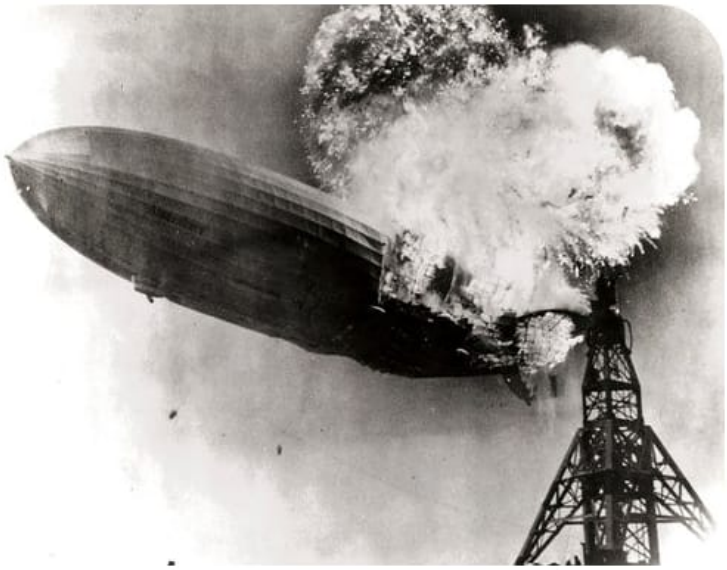


Figure 10.7.2: Explosion of the Hindenberg. (Public Domain; Gus Pasquerella/US Navy via Wikipedia)

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 $(\mathrm{C_3H_8})$ is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$\mathrm{C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)}$$

Example 10.7.1: Combustion Reactions

Ethanol can be used as a fuel source in an alcohol lamp. The formula for ethanol is $\mathrm{C_2H_5OH}$. 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:

$$\mathrm{C_2H_5OH\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)}$$

Balance the equation.

$$\mathrm{C_2H_5OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 3H_2O\,(g)}$$

Step 3: Think about your result.

Combustion reactions must have oxygen as a reactant. Note that the water produced is in the gas state, rather than the liquid state, because of the high temperatures that accompany a combustion reaction.

YouTube Video

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

View this YouTube Video online

Summary

- Combustion reaction is defined and examples are given.

Review

1. What is needed for a combustion reaction to take place?
2. What is formed in any combustion reaction?
3. Mercury reacts with oxygen to form mercuric oxide. Is this a combustion reaction?
4. What are the products of any combustion reaction involving a hydrocarbon?

This page titled 10.7: Combustion Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.6: Combustion Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.8: Single Replacement Reactions



Figure 10.8.1 (Pixabay License; wisteriax via Pixabay)

Why is the silver dark?

The platter and pitcher shown above provides an example of tarnish, a chemical reaction caued when silver metal reacts with hydrogen sulfide gas produced by some industrial processes or as a result of decaying animal or plant materials:

$$2\mathrm{Ag} + \mathrm{H_2S} \rightarrow \mathrm{Ag_2S} + \mathrm{H_2}$$

The tarnish can be removed using a number of polishes, but the process also removes a small amount of silver along with the tarnish.

Single-Replacement Reactions

A **single-replacement reaction** is a reaction in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$\mathrm{A} + \mathrm{BC} \rightarrow \mathrm{AC} + \mathrm{B}$$

In this general reaction, element A is a metal and replaces element B (also a metal) in the compound. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$\mathrm{Y} + \mathrm{XZ} \rightarrow \mathrm{XY} + \mathrm{Z}$$

(Where Y is a nonmetal and replaces the nonmetal Z in the compound with X.)

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\mathrm{Mg}\,(s) + \mathrm{Cu(NO_3)_2}\,(aq) \rightarrow \mathrm{Mg(NO_3)_2}\,(aq) + \mathrm{Cu}\,(s)$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Hydrogen Replacement

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$\mathrm{Zn}\,(s) + 2\mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\,(aq) + \mathrm{H_2}\,(g)$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal.



Figure 10.8.2: Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Public Domain; User: Chemicalinterest/Wikimedia Commons via Wikipedia)

Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2\mathrm{Na}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{NaOH}\,(aq) + \mathrm{H_2}\,(g)$$



Figure 10.8.3: Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite. (Public Domain; User:Ajhalls/Wikimedia Commons via Wikipedia)

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine:

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l)$$

The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

YouTube Video

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

View this YouTube Video online

Summary

- The activity series describes the relative reactivities of metals and halogens.

Review

1. What is a metal replacement reaction?
2. Will a non-metal replace a metal?
3. What is the most reactive halogen?
4. What products will I get if I add potassium metal to water?

This page titled 10.8: Single Replacement Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.7: Single Replacement Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.9: Activity Series



Figure 10.9.1: Copy and Paste Caption here. (Public Domain; Sodium: User: Ajhalls/Wikimedia Commons; Silver: User:Daderot/Wikimedia Commons via Sodium: http://commons.wikimedia.org/wiki/File:Large_Sodium_Explosion.jpg; Silver: http://commons.wikimedia.org/wiki/File:Camp_cup_and_tumbler,_1795-1800,_Paul_Revere_silver_collection,_Worcester_Art_Museum_-_IMG_7624.JPG)

What's the difference between the two pictures above?

We see above 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. Both metals have a single s electron in their outer shell, so you would predict similar reactivities. However, we have a better tool that allows us to make better prediction about what will react with what.

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 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 10.9.1: Activity Series

Activity of Metals		Activity of Halogens
Elements	**Reaction Occurring**	**Elements**
Li K Ba Sr Ca Na	React with cold water, replacing hydrogen.	F_2 Cl_2 Br_2 I_2

	Activity of Metals	Activity of Halogens
Mg Al Zn Cr Fe Cd	React with steam, but not cold water, replacing hydrogen.	
Co Ni Sn Pb	Do not react with water. React with acids, replacing hydrogen.	
H_2		
Cu Hg Ag Pt Au	Unreactive with water or acids.	

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.

$$\mathrm{Ni}\,(s) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow \mathrm{Ni(NO_3)_2}\,(aq) + \mathrm{Pb}\,(s)$$

$$\mathrm{Ni}\,(s) + \mathrm{Fe(NO_3)_3}\,(aq) \rightarrow \mathrm{NR\ (no\ reaction)}$$

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.

CK-12 LICENSED UNDER CK-12 Science
© CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Example $10.9.1$**: Single-Replacement Reactions**

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. $\mathrm{Al}\,(s) + \mathrm{Zn(NO_3)_2}\,(aq) \rightarrow$
b. $\mathrm{Ag}\,(s) + \mathrm{HCl}\,(aq) \rightarrow$

Solution

Step 1: Plan the problem.

For A, compare the placements of aluminum and zinc on the activity series. For B, compare the placements of silver and hydrogen.

Step 2: 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:

$$2\mathrm{Al}\,(s) + 3\mathrm{Zn(NO_3)_2}\,(aq) \rightarrow 2\mathrm{Al(NO_3)_3}\,(aq) + 3\mathrm{Zn}\,(s)$$

Since silver is below hydrogen, it is not capable of replacing hydrogen in a reaction with an acid.

$$\mathrm{Ag}\,(s) + \mathrm{HCl}\,(aq) \rightarrow \mathrm{NR}$$

YouTube Video

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

View this YouTube Video online

Summary

- Metals and halogens are ranked according to their ability to displace other metals or halogens below them in the series.

Review

1. What does the activity series tell us?
2. Can a metal undergo any of the reactions listed below it in the series?

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

3. List two metals that cobalt will displace and two that will displace it.

This page titled 10.9: Activity Series is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.8: Activity Series by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.10: Double Replacement Reactions



Figure 10.10.1 (Public Domain; F. S. Church, published in Harper's Weekly, January 17, 1874, p. 61 via Wikipedia)

Want to trade?

The practice of barter (trading one thing for another) has been in existence for a very long time. In the illustration above, items like chickens were bartered for newspapers. You have something I want, and I have something you want. So we trade and we each have something new. Some chemical reactions are like that. Compounds swap parts and you have new materials.

Double-Replacement Reactions

A **double-replacement reaction** is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement (also called double-displacement) reaction is:

$$\mathrm{AB} + \mathrm{CD} \rightarrow \mathrm{AD} + \mathrm{CB}$$

In this reaction, A and C are positively-charged cations, while B and D are negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

Formation of a Precipitate

A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs:

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

There are very strong attractive forces that occur between $\mathrm{Pb^{2+}}$ and $\mathrm{I^-}$ ions and the result is a brilliant yellow precipitate (see figure below). The other product of the reaction, potassium nitrate, remains soluble.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org



Figure 10.10.2: Formation of lead iodide precipitate. (CC-BY 2.0; Paige Powers - "Lead Iodide" via Flickr)

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas.

$$\mathrm{Na_2S}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{H_2S}\,(g)$$

Formation of a Molecular Compound

Another kind of double-replacement reaction is one that produces a molecular compound as one of its products. Many examples in this category are reactions that produce water. When aqueous hydrochloric acid is reacted with aqueous sodium hydroxide, the products are aqueous sodium chloride and water:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Example 10.10.1: Double-Replacement Reactions

Write a complete and balanced chemical equation for the following double-replacement reactions. One product is indicated as a guide.

a. $\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow$ (hydrogen cyanide gas is formed)
b. $(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow$ (a precipitate of barium sulfate forms)

Solution

Step 1: Plan the problem.

In A, the production of a gas drives the reaction. In B, the production of a precipitate drives the reaction. In both cases, use the ionic charges of both reactants to construct the correct formulas of the products.

Step 2: Solve.

A. The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partners, the balanced equation is:

$$\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow \mathrm{NaBr}\,(aq) + \mathrm{HCN}\,(g)$$

B. Ammonium ion and nitrate ion are $1+$ and $1-$ respectively, while barium and sulfate are $2+$ and $2-$. This must be taken into account when exchanging partners and writing the new formulas. Then, the equation is balanced.

$$(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow 2\mathrm{NH_4NO_3}\,(aq) + \mathrm{BaSO_4}\,(s)$$

Step 3: Think about your result.

Both are double replacement reactions. All formulas are correct and the equations are balanced.

Occasionally, a reaction will produce both a gas and a molecular compound. The reaction of a sodium carbonate solution with

hydrochloric acid produces aqueous sodium chloride, carbon dioxide gas, and water.

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\ \text{DEL}\ \mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

YouTube Video

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

View this YouTube Video online

Summary

- The double-replacement reaction generally takes the form of AB + CD → AD + CB where A and C are positively-charged cations, while B and D are negatively-charged anions.
- In a double replacement reactions, typically one of the products is a precipitate, a gas, or a molecular compound.

Review

1. In a double-replacement reaction, what type of compounds are usually the reactants? A double-replacement reaction occurs between sodium sulfide and hydrogen chloride. Write the rest of the chemical equation and balance it.
2. Na_2S + HCl →
3. In double-replacement reaction, one of three possible types of products usually form. What are the three types?

This page titled 10.10: Double Replacement Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 11.9: Double Replacement Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.11: Everyday Stoichiometry



Figure 10.11.1 (Public Domain; User:Skatebiker/Wikipedia via Wikipedia)

How much equipment do you need for an experiment?

You are in charge of setting out the lab equipment for a chemistry experiment. If you have twenty students in the lab (and they will be working in teams of two) and the experiment calls for three beakers and two test tubes, how much glassware do you need to set out? Figuring this out involves a type of balanced equation and the sort of calculations that you would do for a chemical reaction.

Everyday Stoichiometry

You have learned about chemical equations and the techniques used in order to balance them. Chemists use balanced equations to allow them to manipulate chemical reactions in a quantitative manner. Before we look at a chemical reaction, let's consider an equation for the ideal ham sandwich.



Figure 10.11.2: The ideal ham sandwich. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Our ham sandwich is composed of 2 slices of ham (H), a slice of cheese (C), a slice of tomato (T), 5 pickles (P), and 2 slices of bread (B). The equation for our sandwich is:

$$2\mathrm{H} + \mathrm{C} + \mathrm{T} + 5\mathrm{P} + 2\mathrm{B} \rightarrow \mathrm{H_2CTP_5B_2}$$

Now let us suppose that you are having some friends over and need to make five ham sandwiches. How much of each sandwich ingredient do you need? You take the number of each ingredient required for one sandwich (its coefficient in the above equation) and multiply by five. Using ham and cheese as examples, and using a conversion factor, you can calculate:

$$5\mathrm{H_2CTP_5B_2} \times \frac{2\ \mathrm{H}}{1\mathrm{H_2CTP_5B_2}} = 10\ \mathrm{H}$$

$$5\mathrm{H_2CTP_5B_2} \times \frac{1\mathrm{C}}{1\mathrm{H_2CTP_5B_2}} = 5\ \mathrm{C}$$

The conversion factors contain the coefficient of each specific ingredient as the numerator and the formula of one sandwich as the denominator. The result is what you would expect. In order to make five ham sandwiches, you need 10 slices of ham and 5 slices of cheese.

This type of calculation demonstrates the use of stoichiometry. **Stoichiometry** is the calculation of the amount of substances in a chemical reaction from the balanced equation. The sample problem below is another stoichiometry problem involving ingredients of the ideal ham sandwich.

Example $10.11.1$**: Ham Sandwich Stoichiometry**

Kim looks in the refrigerator and finds that she has 8 slices of ham. In order to make as many sandwiches as possible, how many pickles does she need? Use the equation above.

Solution

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

- Have 8 ham slices (H)
- $2\ \mathrm{H} = 5\ \mathrm{P}$ (conversion factor)

Unknown

- How many pickles (P) needed?

The coefficients for the two reactants (ingredients) are used to make a conversion factor between ham slices and pickles.

Step 2: Solve.

$$8\ \mathrm{H} \times \frac{5\ \mathrm{P}}{2\ \mathrm{H}} = 20\ \mathrm{P}$$

Since 5 pickles combine with 2 ham slices in each sandwich, 20 pickles are needed to fully combine with 8 ham slices.

Step 3: Think about your result.

The 8 ham slices will make 4 ham sandwiches. With 5 pickles per sandwich, the 20 pickles are used in the 4 sandwiches.

Summary

- An example of everyday stoichiometry is given.

Review

1. A smoothie contains 1 banana (B), 4 strawberries (St), 1 container of yogurt (Y), and 3 ice cubes (Ic). Write a balanced

equation to describe the relationship.
2. Write a conversion factor to show the relationship between the number of ice cubes and the number of smoothies produced.
3. How many strawberries would you need to make 12 smoothies?

This page titled 10.11: Everyday Stoichiometry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.1: Everyday Stoichiometry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.12: Mole Ratios



Figure 10.12.1 (Public Domain; User:Ilemme/Wikipedia via Commons Wikimedia, File:Col 03 scup porch chairs.JPG [commons.wikimedia.org])

What does this porch need?

You want to add some sections to the porch seen above. Before you go to the hardware store to buy lumber, you need to determine the unit composition (the material between two large uprights). You count how many posts, how many boards, how many rails – then you decide how many sections you want to add before you calculate the amount of building material needed for your porch expansion.

Mole Ratios

Stoichiometry problems can be characterized by two things: (1) the information *given* in the problem, and (2) the information that is to be solved for, referred to as the *unknown*. The given and the unknown may both be reactants, both be products, or one may be a reactant while the other is a product. The amounts of the substances can be expressed in moles. However, in a laboratory situation, it is common to determine the amount of a substance by finding its mass in grams. The amount of a gaseous substance may be expressed by its volume. In this concept, we will focus on the type of problem where both the given and the unknown quantities are expressed in moles.

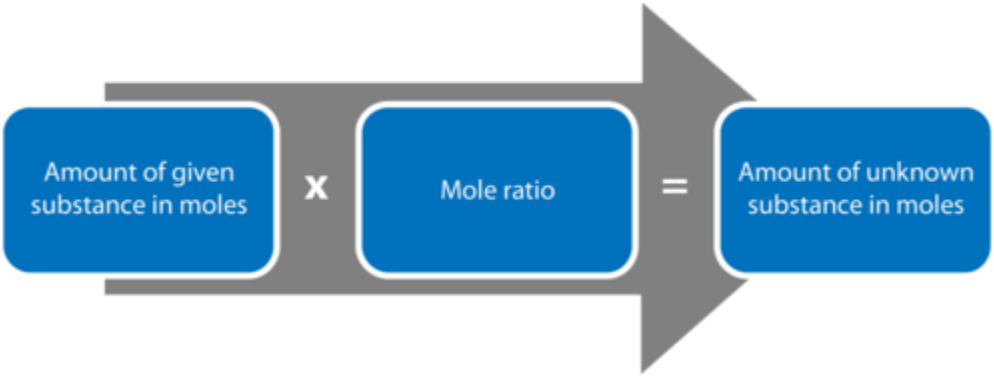


Figure 10.12.2: Mole ratio relationship. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Chemical equations express the amounts of reactants and products in a reaction. The coefficients of a balanced equation can

represent either the number of molecules or the number of moles of each substance. The production of ammonia $(\mathrm{NH_3})$ from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightarrow 2\mathrm{NH_3}\,(g)$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

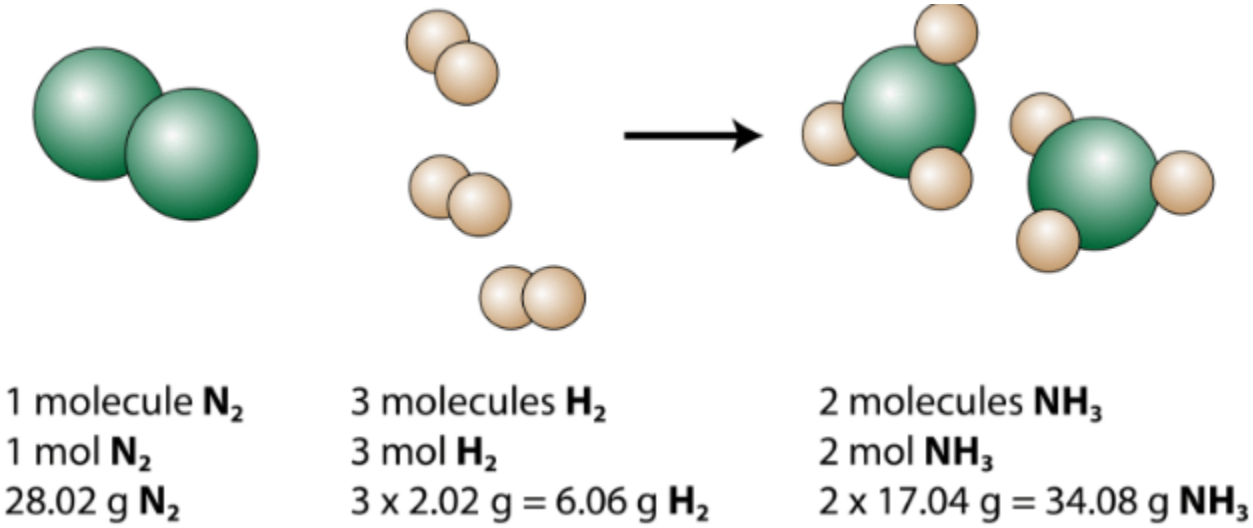


Figure 10.12.3: This representation of the production of ammonia from nitrogen and hydrogen shows several ways to interpret the quantitative information of a chemical reaction. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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 amount 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. $1\ \mathrm{mol}$ of nitrogen has a mass of $28.02\ \mathrm{g}$, while $3\ \mathrm{mol}$ of hydrogen has a mass of $6.06\ \mathrm{g}$, and $2\ \mathrm{mol}$ of ammonia has a mass of $34.08\ \mathrm{g}$.

$$28.02\ \mathrm{g}\ \mathrm{N_2} + 6.06\ \mathrm{g}\ \mathrm{H_2} \rightarrow 34.08\ \mathrm{g}\ \mathrm{NH_3}$$

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.



Figure 10.12.4: An apparatus for running the Haber process. (Public Domain; User: JGvBerkel/Wikimedia Commons via Commons Wikimedia, Haber Ammonia [commons.wikimedia.org])

A **mole ratio** is a conversion factor that relates the amounts in moles of any two substances in a chemical reaction. The numbers in a conversion factor come from the coefficients of the balanced chemical equation. The following six mole ratios can be written for the ammonia forming reaction above.

$$\frac{1\text{ mol }N_2}{3\text{ mol }H_2} \quad or \quad \frac{3\text{ mol }H_2}{1\text{ mol }N_2}$$
$$\frac{1\text{ mol }N_2}{2\text{ mol }NH_3} \quad or \quad \frac{2\text{ mol }NH_3}{1\text{ mol }N_2}$$
$$\frac{3\text{ mol }H_2}{2\text{ mol }NH_3} \quad or \quad \frac{2\text{ mol }NH_3}{3\text{ mol }H_2}$$

In a mole ratio problem, the given substance, expressed in moles, is written first. The appropriate conversion factor is chosen in order to convert from moles of the given substance to moles of the unknown.

Example 10.12.1: Mole Ratio

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

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

Known

- given: H_2 = 4.20 mol

Unknown

- mol of NH_3

The conversion is from $\text{mol } H_2$ to $\text{mol } NH_3$. The problem states that there is an excess of nitrogen, so we do not need to be concerned with any mole ratio involving N_2. Choose the conversion factor that has the NH_3 in the numerator and the H_2 in the denominator.

Step 2: Solve.

$$4.20\text{ mol }H_2 \times \frac{2\text{ mol }NH_3}{3\text{ mol }H_2} = 2.80\text{ mol }NH_3$$

CK-12 LICENSED UNDER CC BY NC SA © CK-12 Foundation • Visit us at ck12.org

The reaction of $4.20\ \text{mol}$ of hydrogen with excess nitrogen produces $2.80\ \text{mol}$ of ammonia.

Step 3: Think about your result.

The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

YouTube Video

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

View this YouTube Video online

Summary

- Mole ratios allow comparison of the amounts of any two materials in a balanced equation.
- Calculations can be made to predict how much product can be obtained from a given number of moles of reactant.

Review

- If a reactant is in excess, why do we not worry about the mole ratios involving that reactant?
- What is the mole ratio of H to N in the ammonia molecule?
- The formula for ethanol is CH_3CH_2OH. What is the mole ratio of H to C in this molecule?

This page titled 10.12: Mole Ratios is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.2: Mole Ratios by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.13: Mass-Mole Stoichiometry

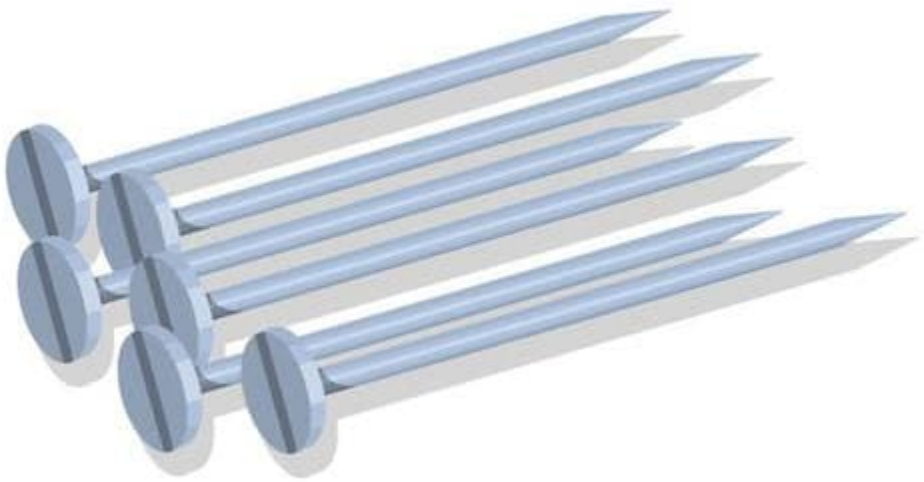


Figure 10.13.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Need nails?

When you are doing a large construction project, you have a good idea of how many nails you will need (lots!). When you go to the hardware store, you don't want to sit there and count out several hundred nails. You can buy nails by weight, so you determine how many nails are in a pound, calculate how many pounds you need, and you're on your way to begin building.

While the mole ratio is ever-present in all stoichiometry calculations, amounts of substances in the laboratory are most often measured by mass. Therefore, we need to use mole-mass calculations in combination with mole ratios to solve several different types of mass-based stoichiometry problems.

Mass to Moles Problems

In this type of problem, the mass of one substance is given, usually in grams. From this, you are to determine the amount in moles of another substance that will either react with or be produced from the given substance.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown}$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation.

Example 10.13.1: Mass-Mole Stoichiometry

Tin metal reacts with hydrogen fluoride to produce tin (II) fluoride and hydrogen gas, according to the following balanced equation.

$$\text{Sn}\,(s) + 2\text{HF}\,(g) \rightarrow \text{SnF}_2\,(s) + \text{H}_2\,(g)$$

How many moles of hydrogen fluoride are required to react completely with $75.0\ \text{g}$ of tin?

Solution:

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

Known

- Given: $75.0\ \text{g Sn}$
- Molar mass of $\text{Sn} = 118.69\ \text{g/mol}$
- $1\ \text{mol Sn} = 2\ \text{mol HF}$ (mole ratio)

Unknown

- mol HF

Use the molar mass of Sn to convert the grams of Sn to moles. Then use the mole ratio to convert from mol Sn to mol HF. This will be done in a single two-step calculation.

$\text{g Sn} \rightarrow \text{mol Sm} \rightarrow \text{mol HF}$

Step 2: Solve.

$$75.0\ \text{g Sn} \times \frac{1\ \text{mol Sn}}{118.69\ \text{g Sn}} \times \frac{2\ \text{mol HF}}{1\ \text{mol Sn}} = 1.26\ \text{mol HF}$$

Step 3: Think about your result.

The mass of tin is less than one mole, but the 1:2 ratio means that more than one mole of HF is required for the reaction. The answer has three significant figures because the given mass has three significant figures.

Moles to Mass Problems

In this type of problem, the amount of one substance is given in moles. From this, you are to determine the mass of another substance that will either react with or be produced from the given substance.

$$\text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{mass of unknown}$$

The moles of the given substance are first converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Then, the moles of the unknown are converted into mass in grams by use of the molar mass of that substance from the periodic table.

Example $10.13.2$**: Mole-Mass Stoichiometry**

Hydrogen sulfide gas burns in oxygen to produce sulfur dioxide and water vapor:

$$2\text{H}_2\text{S}\,(g) + 3\text{O}_2\,(g) \rightarrow 2\text{SO}_2\,(g) + 2\text{H}_2\text{O}\,(g)$$

What mass of oxygen gas is consumed in a reaction that produces $4.60\ \text{mol SO}_2$?

Solution:

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

Known

- Given: $4.60\ \text{mol SO}_2$
- $2\ \text{mol SO}_2 = 3\ \text{mol O}_2$ (mole ratio)
- Molar mass of $\text{O}_2 = 32.00\ \text{g/mol}$

Unknown

- mass O_2 = ? g

Use the mole ratio to convert from $\text{mol } SO_2$ to $\text{mol } O_2$. Then convert $\text{mol } O_2$ to grams. This will be done in a single two-step calculation.

$$\text{mol } SO_2 \rightarrow \text{mol } O_2 \rightarrow \text{g } O_2$$

Step 2: Solve.

$$4.60 \text{ mol } SO_2 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } SO_2} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 221 \text{ g } O_2$$

Step 3: Think about your result.

According to the mole ratio, $6.90 \text{ mol } O_2$ is produced with a mass of 221 g. The answer has three significant figures because the given number of moles has three significant figures.

YouTube Video

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

View this YouTube Video online

Summary

- Calculations involving conversions of mass to moles and moles to mass are described.

Review

1. In the first problem, what would happen if you multiply grams Sn by 118.69 grams/mole Sn?
2. Why is a balanced equation needed?
3. Does the physical form of the material matter for these calculations?

This page titled 10.13: Mass-Mole Stoichiometry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.3: Mass-Mole Stoichiometry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.14: Mass-Mass Stoichiometry



Figure 10.14.1 (CC by 2.0; Pauline Mak via Flickr)

How many walnuts are needed to equal 250 grams?

I want to send 250 grams of shelled walnuts to a friend (don't ask why- just go with the question). How many walnuts in shells do I need to buy? To figure this out, I need to know how much the shell of a walnut weighs (about 40% of the total weight of the unshelled walnut). I can then calculate the mass of walnuts that will yield 250 grams of shelled walnuts, and then determine how many walnuts I need to buy.

Mass to Mass Problems

Mass-mass calculations are the most practical of all mass-based stoichiometry problems. Moles cannot be measured directly, while the mass of any substance can generally be easily measured in the lab. This type of problem is three steps, and is a combination of the two previous types.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{mass of unknown}$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Finally, the moles of the unknown are converted to mass by use of its molar mass.

Example 10.14.1**: Mass-Mass Stoichiometry**

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.

Solution

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

Known

- Given: $45.7\ \text{g}\ \text{NH}_4\text{NO}_3$
- $1\ \text{mol}\ \text{NH}_4\text{NO}_3 = 1\ \text{mol}\ \text{N}_2\text{O} = 2\ \text{mol}\ \text{H}_2\text{O}$
- Molar mass of $\text{NH}_4\text{NO}_3 = 80.06\ \text{g/mol}$
- Molar mass of $\text{N}_2\text{O} = 44.02\ \text{g/mol}$
- Molar mass of $\text{H}_2\text{O} = 18.02\ \text{g/mol}$

Unknown

- Mass $\text{N}_2\text{O} = ?\ \text{g}$
- Mass $\text{H}_2\text{O} = ?\ \text{g}$

Perform two separate three-step mass-mass calculations, as shown below.

$$\text{g}\ \text{NH}_4\text{NO}_3 \rightarrow \text{mol}\ \text{NH}_4\text{NO}_3 \rightarrow \text{mol}\ \text{N}_2\text{O} \rightarrow \text{g}\ \text{N}_2\text{O}$$

$$\text{g}\ \text{NH}_4\text{NO}_3 \rightarrow \text{mol}\ \text{NH}_4\text{NO}_3 \rightarrow \text{mol}\ \text{H}_2\text{O} \rightarrow \text{g}\ \text{H}_2\text{O}$$

Step 2: Solve.

$$45.7\ \text{g}\ \text{NH}_4\text{NO}_3 \times \frac{1\ \text{mol}\ \text{NH}_4\text{NO}_3}{80.06\ \text{g}\ \text{NH}_4\text{NO}_3} \times \frac{1\ \text{mol}\ \text{N}_2\text{O}}{1\ \text{mol}\ \text{NH}_4\text{NO}_3} \times \frac{44.02\ \text{g}\ \text{N}_2\text{O}}{1\ \text{mol}\ \text{N}_2\text{O}} = 25.1\ \text{g}\ \text{N}_2\text{O}$$

$$45.7\ \text{g}\ \text{NH}_4\text{NO}_3 \times \frac{1\ \text{mol}\ \text{NH}_4\text{NO}_3}{80.06\ \text{g}\ \text{NH}_4\text{NO}_3} \times \frac{2\ \text{H}_2\text{O}}{1\ \text{mol}\ \text{NH}_4\text{NO}_3} \times \frac{18.02\ \text{g}\ \text{H}_2\text{O}}{1\ \text{mol}\ \text{H}_2\text{O}} = 20.6\ \text{g}\ \text{H}_2\text{O}$$

Step 3: 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.

YouTube Video

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

View this YouTube Video online

Summary

- Mass-mass calculations involve converting the mass of a reactant to moles of reactant, then using mole ratios to determine moles of product which can then be converted to mass of product.

Review

1. If matter is neither created nor destroyed, why can't we just go directly from grams of reactant to grams of product?
2. Why is it important to get the subscripts correct in the formulas?
3. Why do the coefficients need to be correct?

This page titled 10.14: Mass-Mass Stoichiometry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.4: Mass-Mass Stoichiometry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.15: Volume-Volume Stoichiometry



Figure 10.15.1 (Public Domain; User:Fluffball70/Wikipedia via Wikipedia)

How much propane is left in the tank?

As the weather gets warmer, more and more people want to cook out on the back deck or back yard. Many folks still use charcoal for grilling because of the added flavor. But increasing numbers of back yard cooks like to use a propane grill. The gas burns clean, the grill is ready to go as soon as the flame is lit – but how do you know how much propane is left in the tank? You can buy gauges at hardware stores that measure gas pressure and tell you how much is left in the tank.

Volume-Volume Stoichiometry

Avogadro's hypothesis states that equal volumes of all gases at the same temperature and pressure (0°C and $1\ \text{atm}$) occupies a volume of $22.4\ \text{L}$. These characteristics make stoichiometry problems involving gases at STP very straightforward. Consider the reaction of nitrogen and oxygen gases to form nitrogen dioxide:

$$\begin{array}{ccccc} \text{N}_2\,(g) & + & 2\text{O}_2\,(g) & \rightarrow & 2\text{NO}_2\,(g) \\ 1\ \text{molecule} & & 2\ \text{molecules} & & 2\ \text{molecules} \\ 1\ \text{mol} & & 2\ \text{mol} & & 2\ \text{mol} \\ 1\ \text{volume} & & 2\ \text{volumes} & & 2\ \text{volumes} \end{array}$$

Because of Avogadro's work, we know that the mole ratios between substances in a gas-phase reaction are also volume ratios. The six possible volume ratios for the above equation are:

$$\begin{array}{c} \dfrac{1\ \text{volume}\ \text{N}_2}{2\ \text{volumes}\ \text{O}_2}\ \text{or}\ \dfrac{2\ \text{volumes}\ \text{O}_2}{1\ \text{volume}\ \text{N}_2} \\ \dfrac{1\ \text{volume}\ \text{N}_2}{2\ \text{volumes}\ \text{NO}_2}\ \text{or}\ \dfrac{2\ \text{volumes}\ \text{NO}_2}{1\ \text{volume}\ \text{N}_2} \\ \dfrac{2\ \text{volumes}\ \text{O}_2}{2\ \text{volumes}\ \text{NO}_2}\ \text{or}\ \dfrac{2\ \text{volumes}\ \text{NO}_2}{2\ \text{volumes}\ \text{O}_2} \end{array}$$

The volume ratios above can easily be used when the volume of one gas in a reaction is known, and you need to determine the volume of another gas that will either react with or be produced from the first gas. The pressure and temperature conditions of both gases need to be the same.

Example 10.15.1: Volume-Volume Stoichiometry

The combustion of propane gas produces carbon dioxide and water vapor.

$$\mathrm{C_3H_8}\,(g) + 5\mathrm{O_2}\,(g) \rightarrow 3\mathrm{CO_2}\,(g) + 4\mathrm{H_2O}\,(g)$$

What volume of oxygen is required to completely combust $0.650\ \mathrm{L}$ of propane? What volume of carbon dioxide is produced in the reaction?

Solution:

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

Known

- Given: $0.650\ \mathrm{L}\ \mathrm{C_3H_8}$
- 1 volume $\mathrm{C_3H_8} = 5$ volumes $\mathrm{O_2}$
- 1 volume $\mathrm{C_3H_8} = 3$ volumes $\mathrm{CO_2}$

Unknown

- Volume $\mathrm{O_2} =?\ \mathrm{L}$
- Volume $\mathrm{CO_2} =?\ \mathrm{L}$

Two separate calculations can be done using the volume ratios.

Step 2: Solve.

$$0.650\ \mathrm{L\ C_3H_8} \times \frac{5\ \mathrm{L\ O_2}}{1\ \mathrm{L\ C_3H_8}} = 3.25\ \mathrm{L\ O_2}$$

$$0.650\ \mathrm{L\ C_3H_8} \times \frac{3\ \mathrm{L\ CO_2}}{1\ \mathrm{L\ C_3H_8}} = 1.95\ \mathrm{L\ CO_2}$$

Step 3: Think about your result.

Because the coefficients of the $\mathrm{O_2}$ and the $\mathrm{CO_2}$ are larger than that of the $\mathrm{C_3H_8}$, the volumes for those two gases are greater. Note that total volume is not necessarily conserved in a reaction because moles are not necessarily conserved. In this reaction, 6 total volumes of reactants become 7 total volumes of products.

Summary

- Calculations of volume-volume ratios are based on Avogadro's hypothesis.
- Pressures and temperatures of the gases involved need to be the same.

Review

1. What is Avogadro's hypothesis?
2. How much volume is occupied by one mole of a gas at STP?
3. In the sample problem above, assume we combust 1.3 L of propane. How much CO_2 will be produced?

This page titled 10.15: Volume-Volume Stoichiometry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.5: Volume-Volume Stoichiometry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.16: Mass-Volume Stoichiometry



Figure 10.16.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How much azide is needed to fill an air bag?

Cars and many other vehicles have air bags in them. In case of a collision, a reaction is triggered so that the rapid decomposition of sodium azide produces nitrogen gas, filling the air bag. If too little sodium azide is used, the air bag will not fill completely and will not protect the person in the vehicle. Too much sodium azide could cause the formation of more gas that the bag can safely handle. If the bag breaks from the excess gas pressure, all protection is lost.

Mass to Volume and Volume to Mass Problems

Chemical reactions frequently involve both solid substances whose masses can be measured, as well as gases, for which volume measurements are more appropriate. Stoichiometry problems of this type are called either mass-volume or volume-mass problems.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{volume of unknown}$$

$$\text{volume of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{mass of unknown}$$

Because both types of problems involve a conversion from either moles of gas to volume or vice-versa, we can use the molar volume of $22.4\ \text{L/mol}$, provided that the conditions for the reaction are at STP.

Example 10.16.1: Mass-Volume Stoichiometry

Aluminum metal reacts rapidly with aqueous sulfuric acid to produce aqueous aluminum sulfate and hydrogen gas:

$$2\,\text{Al}\,(s) + 3\text{H}_2\text{SO}_4\,(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3\,(aq) + 3\text{H}_2\,(g)$$

Determine the volume of hydrogen gas produced at STP when a $2.00\ \text{g}$ piece of aluminum completely reacts.

Solution:

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

Known

- Given: $2.00\ \text{g Al}$
- Molar mass $\text{Al} = 26.98\ \text{g/mol}$
- $2\ \text{mol Al} = 3\ \text{mol H}_2$

Unknown

- volume H_2 = ?

The grams of aluminum will first be converted to moles. Then the mole ratio will be applied to convert to moles of hydrogen gas. Finally, the molar volume of a gas will be used to convert to liters of hydrogen.

$$\text{g Al} \rightarrow \text{mol Al} \rightarrow \text{mol H}_2 \rightarrow \text{L H}_2$$

Step 2: Solve.

$$2.00\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} \times \frac{3\ \text{mol H}_2}{2\ \text{mol Al}} \times \frac{22.4\ \text{L H}_2}{1\ \text{mol H}_2} = 2.49\ \text{L H}_2$$

Step 3: Think about your result.

The volume result is in liters. For much smaller amounts, it may be convenient to convert to milliliters. The answer here has three significant figures. Because the molar volume is a measured quantity of $22.4\ \text{L/mol}$, three is the maximum number of significant figures for this type of problem.

Example $10.16.2$**: Volume-Mass Stoichiometry**

Calcium oxide is used to remove sulfur dioxide generated in coal-burning power plants, according to the following reaction.

$$2\text{CaO}\,(s) + 2\text{SO}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CaSO}_4\,(s)$$

What mass of calcium oxide is required to react completely with $1.4 \times 10^3\ \text{L}$ of sulfur dioxide?

Solution:

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

Known

- Given: $1.4 \times 10^3\ \text{L} = \text{SO}_2$
- $2\ \text{mol SO}_2 = 2\ \text{mol CaO}$
- Molar mass $\text{CaO} = 56.08\ \text{g/mol}$

Unknown

- mass CaO = ? g

The volume of SO_2 will be converted to moles, followed by the mole ratio, and finally a conversion of moles of CaO to grams.

$$\text{L SO}_2 \rightarrow \text{mol SO}_2 \rightarrow \text{mol CaO} \rightarrow \text{g CaO}$$

Step 2: Solve.

$$1.4 \times 10^3\ \text{L SO}_2 \times \frac{1\ \text{mol SO}_2}{22.4\ \text{L SO}_2} \times \frac{2\ \text{mol CaO}}{2\ \text{mol SO}_2} \times \frac{56.08\ \text{g CaO}}{1\ \text{mol CaO}} = 3.5 \times 10^3\ \text{g CaO}$$

Step 3: Think about your result.

The resultant mass could be reported as $3.5\ \text{kg}$, with two significant figures. Even though the 2:2 mole ratio does not mathematically affect the problem, it is still necessary for unit conversion.

CK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

YouTube Video

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

View this YouTube Video online

Summary

- Calculations are described for determining the amount of gas formed in a reaction.
- Calculations are described for determining amounts of a material needed to react with a gas.

Review

1. What are the conditions for all gases in these calculations?
2. How can you tell if all the ratios were set up correctly?
3. Why was 2 mol CaO/2mol SO_2 included in the second example if it did not affect the final number?

This page titled 10.16: Mass-Volume Stoichiometry is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.6: Mass-Volume Stoichiometry by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.17: Limiting Reactant

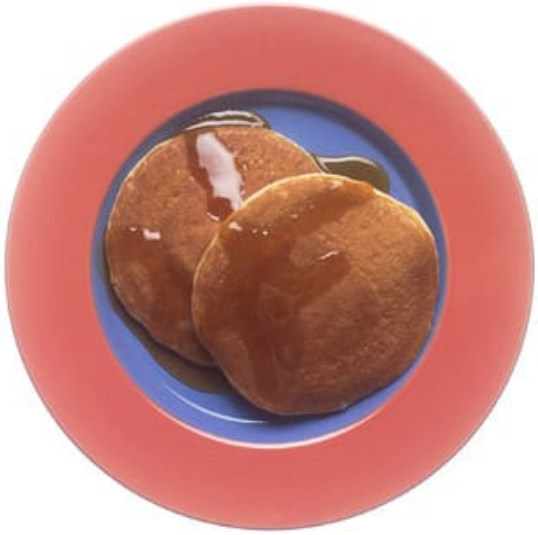


Figure 10.17.1 (Public Domain; Renee Comet and the National Cancer Institute; CK-12 Foundation - Christopher Auyeung via Commons Wikimedia, Pancakes [commons.wikimedia.org])

Don't you hate running out of cooking ingredients?

Cooking is a great example of everyday chemistry. In order to correctly follow a recipe, a cook needs to make sure that he has plenty of all the necessary ingredients in order to make his dish. Let us suppose that you are deciding to make some pancakes for a large group of people. The recipe on the box indicates that the following ingredients are needed for each batch of pancakes:

1 cup of pancake mix

$\frac{3}{4}$ cup milk

1 egg

1 tablespoon vegetable oil

Now you check the pantry and the refrigerator and see that you have the following ingredients available:

2 boxes of pancake mix (8 cups)

Half a gallon of milk (4 cups)

2 eggs

Full bottle of vegetable oil (about 3 cups)

The question that you must ask is: how many batches of pancakes can I make? The answer is two. Even though you have enough pancake mix, milk, and oil to make many more batches of pancakes, you are *limited* by the fact that you only have two eggs. As soon as you have made two batches of pancakes, you will be out of eggs and your "reaction" will be complete.

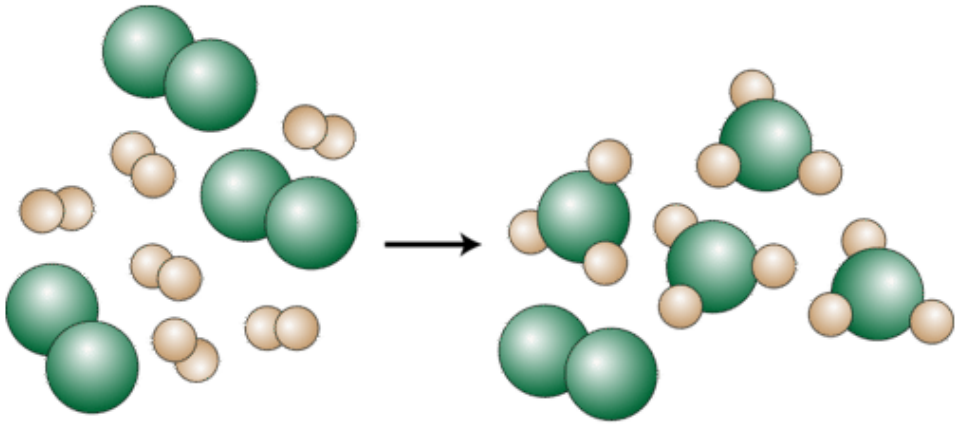
Limiting Reactant

For a chemist, the balanced chemical equation is the recipe that must be followed. As you have seen earlier, the Haber process is a reaction in which nitrogen gas is combined with hydrogen gas to form ammonia. The balanced equation is shown below.

$$N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$$

We know that the coefficients of the balanced equation indicate the mole ratio that is required for this reaction to occur. One mole of N_2 will react with three moles of H_2 to form two moles of NH_3.

Now let us suppose that a chemist were to react three moles of N_2 with six moles of H_2 (see figure below).



	Reactants		Products
Before reaction	3 mol N_2	6 mol H_2	0 mol NH_3
After reaction	1 mol N_2	0 mol H_2	4 mol NH_3

Figure 10.17.2: Reaction in presence of limiting reagent. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

So what happened in this reaction? The chemist started with 3 moles of N_2. You may think of this as being 3 times as much as the "recipe" (the balanced equation) requires since the coefficient for the N_2 is a 1. However, the 6 moles of H_2 that the chemist started with is only two times as much as the "recipe" requires, since the coefficient for the H_2 is a 3 and $3 \times 2 = 6$. So, after the reaction is complete, the hydrogen gas will be completely used up; while there will be 1 mole of nitrogen gas left over. Finally, the reaction will produce 4 moles of NH_3 because that is also two times as much as shown in the balanced equation. The overall reaction that occurred in words:

$$2\text{ mol } N_2 + 6\text{ mol } H_2 \rightarrow 4\text{ mol } NH_3$$

All the amounts are doubled from the original balanced equation.

The **limiting reactant** (or **limiting reagent**) is the reactant that determines the amount of product that can be formed in a chemical reaction. The reaction proceeds until the limiting reactant is completely used up. In our example above, the H_2 is the limiting reactant. The **excess reactant** (or **excess reagent**) is the reactant that is initially present in a greater amount than will eventually be reacted. In other words, there is always excess reactant left over after the reaction is complete. In the above example, the N_2 is the excess reactant.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- The amount of limiting reactant determines how much product will be formed in a chemical reaction.

Review

1. In the Haber reaction illustrated above, how do we know that hydrogen is the limiting reactant?
2. What if hydrogen were left over?
3. Which material would be limiting if no hydrogen or nitrogen were left over?

This page titled 10.17: Limiting Reactant is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.7: Limiting Reactant by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.18: Determining the Limiting Reactant



Figure $10.18.1$ (Public Domain; Renee Comet and the National Cancer Institute via Wikipedia)

Who's coming for dinner?

You have ten people that show up for a dinner party. One of the guest brings twenty brownies for dessert. The decision about serving dessert is easy: two brownies are placed on every plate. If someone wants more brownies, they will have to wait until they go to the store. There are only enough brownies for everyone to have two.

Determining the Limiting Reactant

In the real world, amounts of reactants and products are typically measured by mass or by volume. It is first necessary to convert the given quantities of each reactant to moles in order to identify the limiting reactant.

Example $10.18.1$: Determining the Limiting Reactant

Silver metal reacts with sulfur to form silver sulfide according to the following balanced equation:

$$2\mathrm{Ag}\,(s) + \mathrm{S}\,(s) \rightarrow \mathrm{Ag_2S}\,(s)$$

What is the limiting reactant when $50.0\ \mathrm{g\ Ag}$ is reacted with $10.0\ \mathrm{g\ S}$?

Solution:

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

Known

- Given: $50.0\ \mathrm{g\ Ag}$
- Given: $10.0\ \mathrm{g\ S}$

Unknown

- limiting reactant

Use the atomic masses of Ag and S to determine the number of moles of each present. Then, use the balanced equation to calculate the number of moles of sulfur that would be needed to react with the number of moles of silver present. Compare this result to the actual number of moles of sulfur present.

Step 2: Solve.

First, calculate the number of moles of Ag and S present:

$$50.0\ \mathrm{g\ Ag} \times \frac{1\ \mathrm{mol\ Ag}}{107.87\ \mathrm{g\ Ag}} = 0.464\ \mathrm{mol\ Ag}$$

$$10.0\ \mathrm{g\ S} \times \frac{1\ \mathrm{mol\ S}}{32.07\ \mathrm{g\ S}} = 0.312\ \mathrm{mol\ S}$$

Second, find the moles of S that would be required to react with all of the given Ag:

$$0.464\ \mathrm{mol\ Ag} \times \frac{1\ \mathrm{mol\ S}}{2\ \mathrm{mol\ Ag}} = 0.232\ \mathrm{mol\ S\ (required)}$$

The amount of S actually present is 0.312 moles. The amount of S that is required to fully react with all of the Ag is 0.232 moles. Since there is more sulfur present than what is required to react, the sulfur is the excess reactant. Therefore, silver is the limiting reactant.

Step 3: Think about your result.

The balanced equation indicates that the necessary mole ratio of Ag to S is 2:1. Since there were not twice as many moles of Ag present in the original amounts, that makes silver the limiting reactant.

There is a very important point to consider about the preceding problem. Even though the mass of silver present in the reaction $(50.0\ \mathrm{g})$ was greater than the mass of sulfur $(10.0\ \mathrm{g})$, silver was the limiting reactant. This is because chemists must always convert to molar quantities and consider the mole ratio from the balanced chemical equation.

There is one other component to be determined in a limiting reactant problem—the quantity of the excess reactant that will be left over after the reaction is complete. We will return to the sample problem, above, to answer this question below.

Example $10.18.2$: Determining the Amount of Excess Reactant Left Over

What is the mass of excess reactant remaining when $50.0\ \mathrm{g\ Ag}$ reacts with $10.0\ \mathrm{g\ S}$?

$$2\mathrm{Ag}\,(s) + \mathrm{S}\,(s) \rightarrow \mathrm{Ag_2S}\,(s)$$

Solution:

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

Known

- Excess reactant $= 0.312\ \mathrm{mol\ S}$ (from Example $10.18.1$)
- Amount of excess reactant needed $= 0.232\ \mathrm{mol\ S}$ (from Example $10.18.1$)

Unknown

- Mass of excess reactant remaining after the reaction = ? g

Subtract the amount (in moles) of the excess reactant that will react from the amount that is originally present. Convert moles to grams.

Step 2: Solve.

$$0.312\ \mathrm{mol\ S} - 0.232\ \mathrm{mol\ S} = 0.080\ \mathrm{mol\ S\ (remaining\ after\ reaction)}$$

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

$$0.080 \text{ mol S} \times \frac{32.07 \text{ g S}}{1 \text{ mol S}} = 2.57 \text{ g S}$$

There are 2.57 g of sulfur remaining when the reaction is complete.

Step 3: Think about your result.

There were 10.0 g of sulfur present before the reaction began. If 2.57 g of sulfur remain after the reaction, then 7.43 g S reacted.

$$7.43 \text{ g S} \times \frac{1 \text{ mol S}}{32.07 \text{ g S}} = 0.232 \text{ mol S}$$

This is the amount of sulfur that reacted. The problem is internally consistent.

Interactive Content

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

View this Interactive Content online

Interactive Content

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

View this Interactive Content online

Summary

- Determining the limiting reactant requires that all mass quantities first be converted to moles to evaluate the equation.

Review

1. Why do all mass values need to be converted to moles before determining the limiting reactant?
2. Silver metal reacts with sulfur to form silver sulfide according to the following balanced equation:

$$2\text{Ag}\,(s) + \text{S}\,(s) \rightarrow \text{Ag}_2\text{S}\,(s)$$

   a. If 0.700 moles Ag is reacted with 10.0 g S, is sulfur or aluminum the limiting reactant?
   b. How many grams of Ag_2S will be produced?

This page titled 10.18: Determining the Limiting Reactant is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.8: Determining the Limiting Reactant by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

10.19: Theoretical Yield and Percent Yield



Figure 10.19.1 (Public Domain; Linda Bartlett and the National Cancer Institute via Wikipedia)

Can we save some money?

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 do not 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 often 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:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

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 previously indicated. 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 10.19.1: Calculating the Theoretical Yield and the Percent Yield

Potassium chlorate decomposes upon slight heating in the presence of a catalyst, according to the reaction below.

$$2\mathrm{KClO_3}\,(s) \rightarrow 2\mathrm{KCl}\,(s) + 3\mathrm{O_2}\,(g)$$

In a certain experiment, $40.0\ \mathrm{g\ KClO_3}$ is heated until it completely decomposes. What is the theoretical yield of oxygen gas? The experiment is performed, the oxygen gas is collected, and its mass is found to be $14.9\ \mathrm{g}$. What is the percent yield for the reaction?

Solution

First, we will calculate the theoretical yield based on the stoichiometry.

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

Known

- Given: Mass of $\mathrm{KClO_3} = 40.0\ \mathrm{g}$
- Molar mass $\mathrm{KClO_3} = 122.55\ \mathrm{g/mol}$
- Molar mass $\mathrm{O_2} = 32.00\ \mathrm{g/mol}$

Unknown

- theoretical yield O_2 = ? g

Apply stoichiometry to convert from the mass of a reactant to the mass of a product:

$$\mathrm{g\ KClO_3} \rightarrow \mathrm{mol\ KClO_3} \rightarrow \mathrm{mol\ O_2} \rightarrow \mathrm{g\ O_2}$$

Step 2: Solve.

$$40.0\ \mathrm{g\ KClO_3} \times \frac{1\ \mathrm{mol\ KClO_3}}{122.55\ \mathrm{g\ KClO_3}} \times \frac{3\ \mathrm{mol\ O_2}}{2\ \mathrm{mol\ KClO_3}} \times \frac{32.00\ \mathrm{g\ O_2}}{1\ \mathrm{mol\ O_2}} = 15.7\ \mathrm{g\ O_2}$$

The theoretical yield of $\mathrm{O_2}$ is $15.7\ \mathrm{g}$.

Step 3: Think about your result.

The mass of oxygen gas must be less than the $40.0\ \mathrm{g}$ of potassium chlorate that was decomposed.

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

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

Known

- Actual yield $= 14.9\ \mathrm{g}$
- Theoretical yield $= 15.7\ \mathrm{g}$

Unknown

- Percent yield = ? %

$$\mathrm{Percent\ Yield} = \frac{\mathrm{Actual\ Yield}}{\mathrm{Theoretical\ Yield}} \times 100\%$$

Use the percent yield equation above.

Step 2: Solve.

CK-12 LICENSED UNDER CC BY NC SA
© CK12 Foundation • Visit us at ck12.org

$$\text{Percent Yield} = \frac{14.9\text{ g}}{15.7\text{ g}} \times 100\% = 94.9\%$$

Step 3: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

YouTube Video

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

View this YouTube Video online

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.

Review

1. What do we need in order to calculate theoretical yield?
2. If I spill some of the product before I weigh it, how will that affect the actual yield?
3. How will spilling some of the product affect the percent yield?
4. I make a product and weigh it before it is dry. How will that affect the actual yield?

This page titled 10.19: Theoretical Yield and Percent Yield is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 12.9: Theoretical Yield and Percent Yield by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11: STATES OF MATTER

11.1: Physical Properties and Intermolecular Forces

11.2: Hydrogen Bonding

11.3: Kinetic Molecular Theory

11.4: Gas Pressure

11.5: Atmospheric Pressure

11.6: Pressure Units and Conversions

11.7: Average Kinetic Energy and Temperature

11.8: Physical Properties and Intermolecular Forces

11.9: Surface Tension

11.10: Evaporation

11.11: Vapor Pressure

11.12: Boiling

11.13: Vapor Pressure Curves

11.14: Melting

11.15: Sublimation

11.16: Heating and Cooling Curves

11.17: General Phase Diagram

11.18: Phase Diagram for Water

Template:HideTOC

This page titled 11: States of Matter is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13: States of Matter by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.1: Physical Properties and Intermolecular Forces

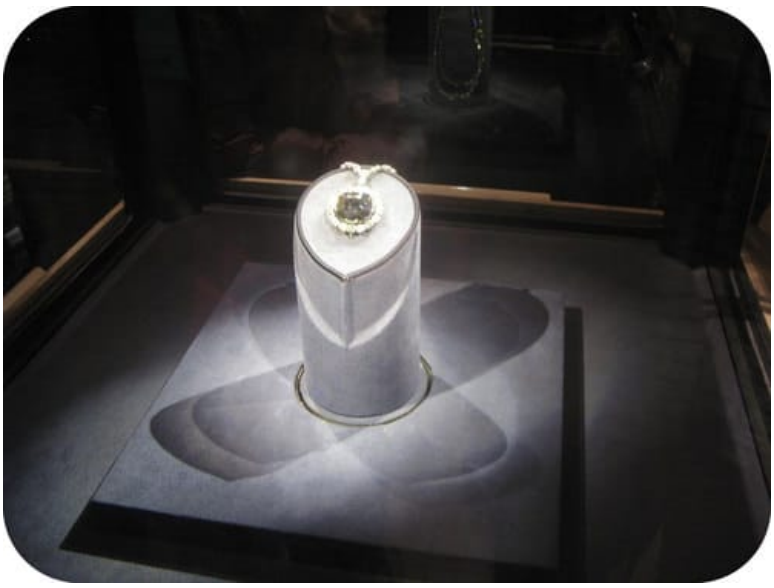


Figure 11.1.1 (Public Domain; Moskey71 via https://commons.wikimedia.org/wiki/F...opeDiamond.JPG)

Can you guess how much this is worth?

Carbon is an interesting and versatile element. There are over twenty million known compounds containing carbon, encompassed in the growing field of organic chemistry. The element itself can exist in two major forms. Diamond is a form of carbon that is extremely hard and is one of the few materials that can scratch glass. The other form of carbon is graphite, a very soft material that we find in “lead” pencils. The two forms differ mainly in how the carbon atoms are connected to one another. The differences in the arrangement of atoms affect the properties of the material.

Physical Properties and Intermolecular Forces

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. **Molecular compounds**, sometimes called covalent compounds, display a wide range of physical properties due to the different types of intermolecular attractions such as different kinds of polar interactions. The melting and boiling points of molecular compounds are generally quite low compared to those of **ionic compounds**. This is because the energy required to disrupt the **intermolecular forces** between molecules is far less than the energy required to break the ionic bonds in a crystalline ionic compound. Since molecular compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. Ionic compounds do not conduct electricity in the solid state because of their rigid structure, but conduct well when either molten or dissolved into a solution. The water solubility of molecular compounds is variable and depends primarily on the type of intermolecular forces involved. Substances that exhibit hydrogen bonding or dipole-dipole forces are generally water soluble, whereas those that exhibit only London dispersion forces are generally insoluble. Most, but not all, ionic compounds are quite soluble in water. The table below summarizes some of the differences between ionic and molecular compounds.

Table 11.1.1*: Comparison of Ionic and Molecular Compounds*

Property	Ionic Compounds	Molecular Compounds
Type of elements	metal and nonmetal	nonmetals only
Bonding	ionic – transfer of electron(s) between atoms	covalent – sharing of pair(s) of electrons between atoms
Representative unit	formula unit	molecule

cK-12 LICENSED UNDER CK-12 License © CK-12 Foundation • Visit us at ck12.org

Property	Ionic Compounds	Molecular Compounds
Physical state at room temp.	solid	gas, liquid, or solid
Water solubility	usually high	variable
Melting and boiling temps	generally high	generally low
Electrical conductivity	good when molten or in solution	poor

One type of molecular compound behaves quite differently than that described so far. A covalent network solid is a compound in which all of the atoms are connected to one another by covalent bonds. Diamond is composed entirely of carbon atoms, each bonded to four other carbon atoms in a tetrahedral geometry. Melting a covalent network solid is not accomplished by overcoming the relatively weak intermolecular forces. Rather, all of the covalent bonds must be broken, a process that requires extremely high temperatures. Diamond, in fact, does not melt at all. Instead, it vaporizes to a gas at temperatures above 3500°C.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Summary

- The physical properties of a material are affected by the intermolecular forces holding the molecules together.

Review

1. Are melting points of molecular compounds generally higher or lower than those of ionic compounds?
2. Do ionic compounds conduct electricity in the solid state?
3. What types of substances are generally water-soluble?

This page titled 11.1: Physical Properties and Intermolecular Forces is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

11.2: Hydrogen Bonding

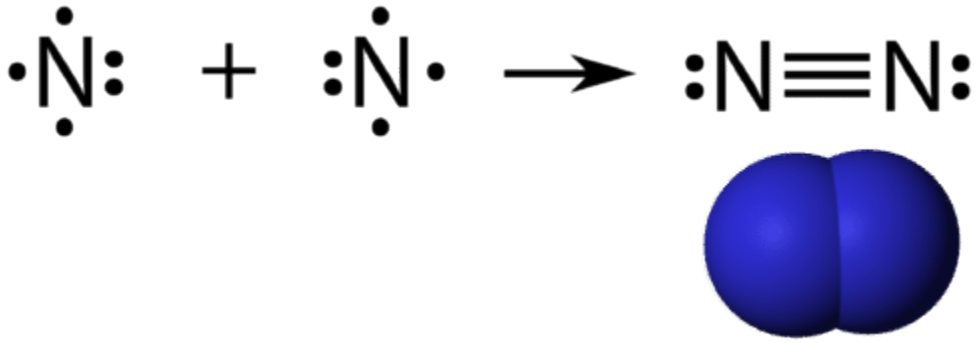


Figure 11.2.1 (CC BY-NC-SA 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Nitrogen-3D-vdW.png)

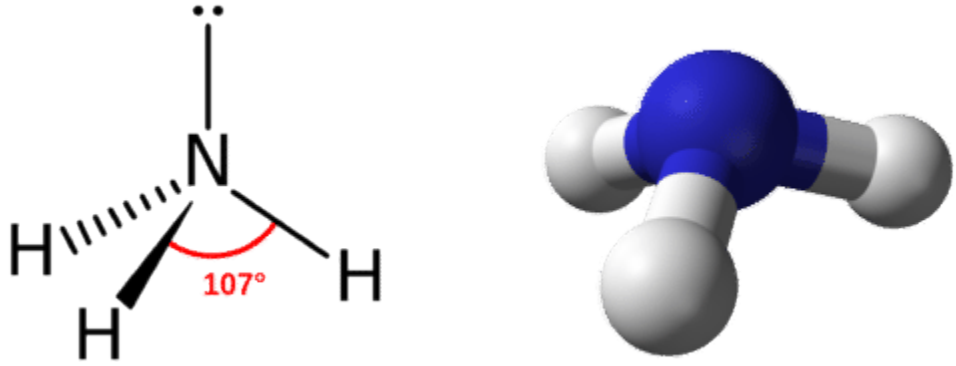


Figure 11.2.2 (Public Domain; (Left) User:Booyabazooka/Wikimedia Commons; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Ammonia_lone_electron_pair.svg; (Right) http://commons.wikimedia.org/wiki/File:Ammonia-3D-balls-A.png)

What's the difference between these two molecules?

A rough rule of thumb is that higher molecular-weight materials have higher boiling points than their lower molecular-weight counterparts. More energy is needed to move the larger molecule from the liquid state to the vapor state. However, ammonia has a boiling point of $-33.34^{\circ}\mathrm{C}$ and a molecular weight of 17, while nitrogen (molecular weight 28) has a boiling point of $-195.8^{\circ}\mathrm{C}$. The lighter ammonia molecule must have other factors that influence its physical properties.

Hydrogen Bonding

The attractive force between water molecules is a dipole interaction. The hydrogen atoms are bound to the highly electronegative oxygen atom (which also possesses two lone pair sets of electrons, making for a very polar bond). The partially positive hydrogen atom of one molecule is then attracted to the oxygen atom of a nearby water molecule (see figure below).

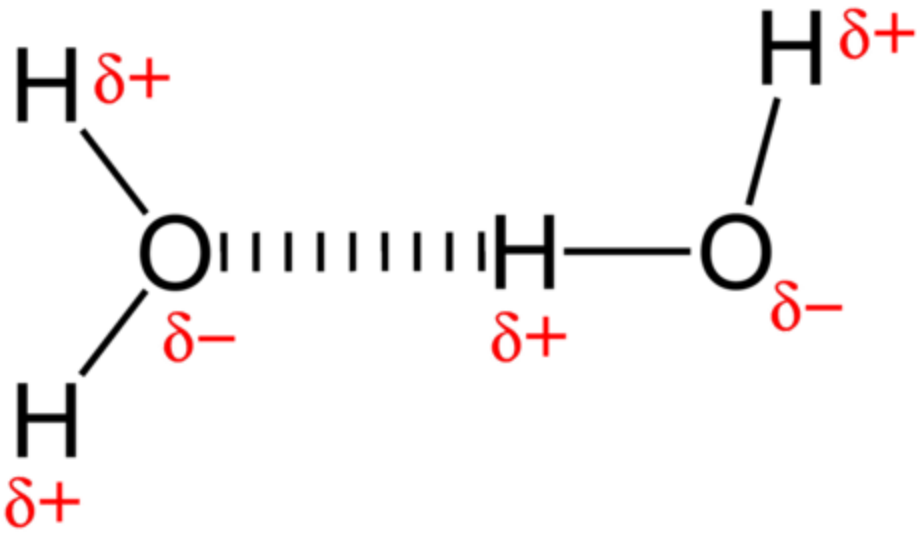


Figure 11.2.3: A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on an oxygen atom of a neighboring water molecule. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

A **hydrogen bond** is an intermolecular attractive force in which a hydrogen atom that is covalently bonded to a small, highly electronegative atom is attracted to a lone pair of electrons on an atom in a neighboring molecule. Hydrogen bonds are very strong compared to other dipole interactions. The strength of a typical hydrogen bond is about 5% of that of a covalent bond.

Hydrogen bonding occurs only in molecules where hydrogen is covalently bonded to one of three elements: fluorine, oxygen, or nitrogen. These three elements are so electronegative that they withdraw the majority of the electron density in the covalent bond with hydrogen, leaving the H atom very electron-deficient. The H atom nearly acts as a bare proton, leaving it very attracted to lone pair electrons on a nearby atom.

The hydrogen bonding that occurs in water leads to some unusual, but very important, properties. Most molecular compounds that have a mass similar to water are gases at room temperature. Because of the strong hydrogen bonds, water molecules are able to stay condensed in the liquid state. The figure below shows how the bent shape, and two hydrogen atoms per molecule, allows each water molecule to be able to hydrogen bond to two other molecules.

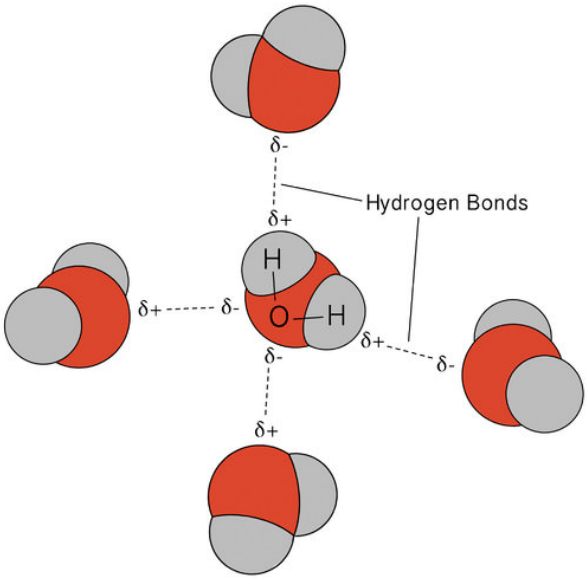


Figure 11.2.4: Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two hydrogen atoms per molecule. (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When

water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become permanent and form a very specific network (see figure below).

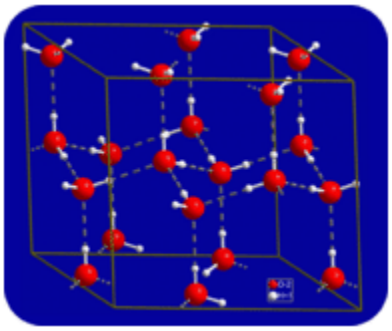


Figure 11.2.5: When water freezes to ice, the hydrogen bonding network becomes permanent. Each oxygen atom has an approximately tetrahedral geometry—two covalent bonds and two hydrogen bonds. (Public Domain; User:Materialscientist/Wikimedia Commons via Wikipedia)

The bent shape of the molecules leads to gaps in the hydrogen bonding network of ice. Ice has a very unusual property—its solid state is less dense than its liquid state. Ice floats in water. Virtually all other substances are denser in the solid state than in the liquid state. Hydrogen bonds play a very important biological role in the physical structures of proteins and nucleic acids.

Interactive Content

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

View this Interactive Content online

Simulation

Why does water form droplets? Hint: it has to do with the attraction between water molecules. Learn more (and see how many drops of water you can get a penny can hold) in this simulation.

Summary

- Hydrogen bonds form when a H attached to a N, O, or F atom interacts with another N, O, or F atom.

Review

1. How strong is a hydrogen bond?
2. What happens when H is covalently bonded to N, O, or F?
3. How does the shape of the water molecule affect the properties of ice?

This page titled 11.2: Hydrogen Bonding is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 9.19: Hydrogen Bonding by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.3: Kinetic Molecular Theory



Figure 11.3.1 (Public Domain; Airman Maebel Y. Tinoko/U.S. Navy via Wikipedia)

How much oxygen is in this container?

Approximately 20% of the atmosphere is oxygen. This gas is essential for life. In environments where oxygen is in low supply, it can be provided from a tank. Since gases are very compressible, a large amount of oxygen can be stored in a relatively small container. When it is released, the volume expands and the pressure decreases. The gas is then available for ventilation under normal pressure.

Kinetic-Molecular Theory

The kinetic-molecular theory explains the states of matter, and is based on the idea that matter is composed of tiny particles that are always in motion. This theory helps explain observable properties and behaviors of solids, liquids, and **gases**. However, kinetic-molecular theory is most easily understood as it applies to gases, and it is with gases that we will begin our detailed study. The theory applies specifically to a model of gas called an **ideal gas**. An ideal gas is an imaginary gas whose behavior perfectly fits all the assumptions of the kinetic-molecular theory. In reality, gases are not ideal, but are very close to being so under most everyday conditions.

The kinetic-molecular theory, as it applies to gases, has five basic assumptions:

1. **Gases consist of very large numbers of tiny spherical particles that are far apart from one another compared to their size**. The particles of a gas may be either atoms or molecules. The distance between the particles of a gas is much, much greater than the distance between the particles of a liquid or a solid. Most of the volume of a gas, therefore, is composed of the empty space between the particles. In fact, the volume of the particles themselves is considered to be insignificant compared to the volume of the empty space.
2. **Gas particles are in constant rapid motion in random directions**. The fast motion of gas particles gives them a relatively large amount of kinetic energy. Recall that kinetic energy is the energy that an object possesses because of its motion. The particles of a gas move in straight-line motion until they collide with another particle, or with one of the walls of the gas container.
3. **Collisions between gas particles and between particles and the container walls are elastic collisions**. An elastic collision is one in which there is no overall loss of kinetic energy. Kinetic energy may be transferred from one particle to another during an elastic collision, but there is no change in the total energy of the colliding particles.
4. **There are no forces of attraction or repulsion between gas particles.** Attractive forces are responsible for particles of a real gas condensing together to form a liquid. It is assumed that the particles of an ideal gas have no such attractive forces.

The motion of each particle is completely independent of the motion of all other particles.

5. **The average kinetic energy of gas particles is dependent upon the temperature of the gas**. As the temperature of a sample of gas is increased, the speeds of the particles are increased. This results in an increase in the kinetic energy of the particles. Not all particles of gas in a sample have the same speed, and so they do not have the same kinetic energy. The temperature of a gas is proportional to the average kinetic energy of the gas particles.

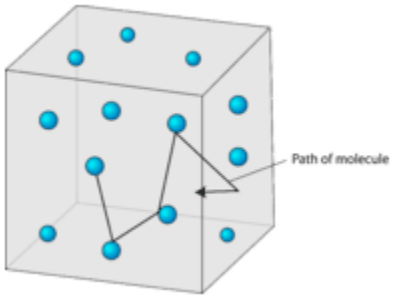


Figure 11.3.2: Gas particles are in random straight-line motion according to the kinetic-molecular theory. The space between particles is very large compared to the particle size. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Summary

- Assumptions of the kinetic-molecular theory:
  - Gases consist of very large numbers of tiny spherical particles that are far apart from one another compared to their size.
  - Gas particles are in constant rapid motion in random directions.
  - Collisions between gas particles and between particles and the container walls are elastic collisions.
  - There are no forces of attraction or repulsion between gas particles.
  - The average kinetic energy of gas particles is dependent upon the temperature of the gas.

Review

1. Describe the motion of gas particles.
2. What kind of collisions occurs?
3. What is the relationship between the kinetic energy of gas particles and the temperature of the gas?

This page titled 11.3: Kinetic Molecular Theory is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.1: Kinetic Molecular Theory by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.4: Gas Pressure



Figure 11.4.1 (Public Domain; Meeta via Wikipedia)

How are hot air balloons able to move smoothly in the air?

Many people enjoy riding in hot air balloons. Some use them for romantic picnics and marriage proposals. Others race in competitions. Being above the earth gives a whole new perspective on the world around us. As the beginning of a hot air balloon ride, the balloon is flat because the pressure inside the balloon equals the pressure outside. When the air inside the balloon is heated, the speed of movement of those air molecules increases and the pressure goes up. After a while the balloon is completely expanded and the flight is ready to take off.

Gas Pressure

Pressure is defined as the force per unit area on a surface.

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

When a person stands on the floor, his feet exert pressure on the surface. That pressure is related to both the mass of the person and the surface area of his feet. If the person were holding a heavy object, the pressure would increase because of a greater force. Alternatively, if the person stands on his toes, the pressure also increases because of a decrease in the surface area.

Gas molecules also exert pressure. Earth's atmosphere exerts pressure because gravity acts on the huge number of gas particles contained in the atmosphere, holding it in place. Pressure is also exerted by a small sample of gas, such as that which is contained in a balloon. Gas pressure is the pressure that results from **collisions** of gas particles with an object. Inside the balloon, the gas particles collide with the balloon's inner walls. It is those collisions which keep the balloon inflated. If the gas particles were to suddenly stop moving, the balloon would instantly deflate. The figure below is an illustration of gas particles exerting pressure inside a container.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

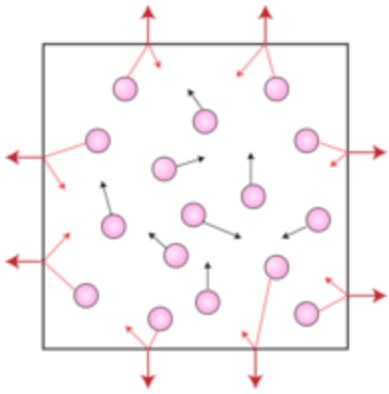


Figure 11.4.2: Collision of gas particles with container wall. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The pressure inside the hot air balloon is affected by temperature. As the molecules heat up, they move faster and strike the inside wall of the balloon harder. This increased motion of the gas particles increases the force on an area of the balloon, producing a rise in the pressure.

YouTube Video

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

View this YouTube Video online

Summary

- Pressure is defined as $\frac{\text{force}}{\text{volume}}$.
- Gas pressure is the result of collisions between gas particles and an object.
- An increase in temperature will produce an increase in pressure of a gas.

Review

1. What is pressure?
2. How does a gas create pressure on an object or container?
3. What would happen to the pressure if gas particles suddenly stopped moving?
4. How does temperature affect pressure?

This page titled 11.4: Gas Pressure is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.2: Gas Pressure by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.5: Atmospheric Pressure

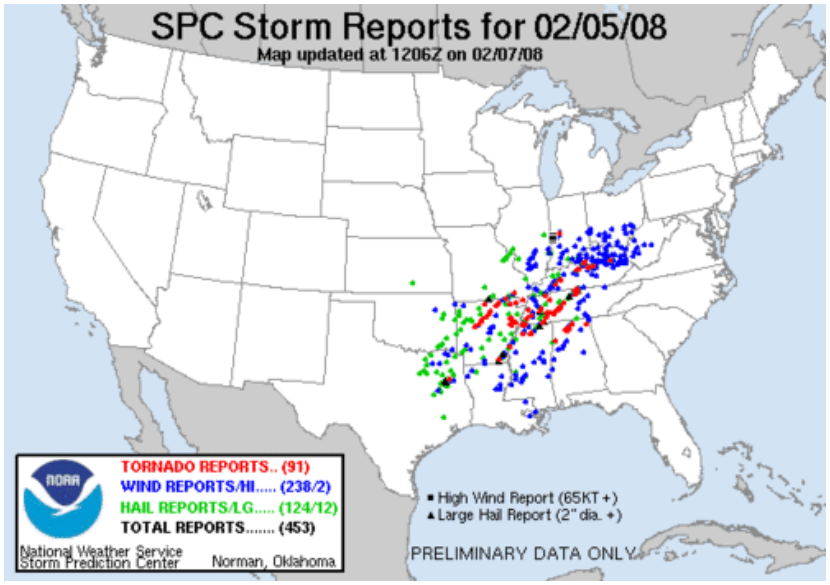


Figure 11.5.1 (Public Domain; the Storm Prediction Center/NOAA via Wikipedia)
A large difference in pressure between two areas is an important criterion for generating storm reports

What do storm reports reveal?

The pressure in the atmosphere is an important factor in determining what the weather will be like. If the barometric pressure is high in an area, this will cause air to move to a region of lower pressure. The greater the difference in pressure between the two areas, the stronger the winds will develop. Under certain conditions, the winds can produce a tornado (a violent rotating column of air that reaches from a thunderstorm down to the ground).

Atmospheric Pressure

Atmospheric pressure is the pressure exerted by gas particles in Earth's atmosphere as those particles collide with objects. A **barometer** is an instrument used to measure atmospheric pressure. A traditional mercury barometer consists of an evacuated tube immersed in a container of mercury. Air molecules push down on the surface of the mercury. Because the inside of the tube is a vacuum, the mercury rises inside the tube. The height to which the mercury rises is dependent on the external air pressure.

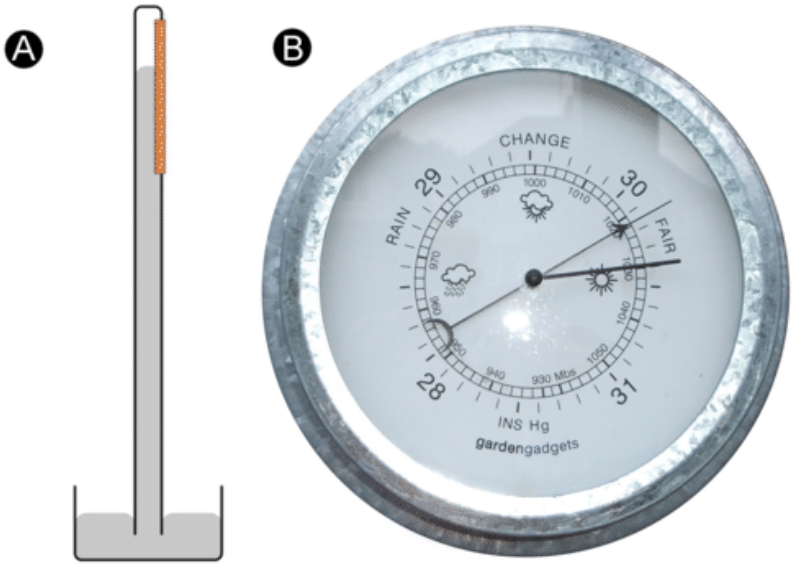


Figure 11.5.2: (A) A barometer measures atmospheric pressure as the height of a column of mercury. (B) A modern aneroid barometer in the form of a dial is used by meteorologists to help them predict upcoming weather. (Public Domain; (A) User: Danomagnum/Wikimedia Commons; (B) User:Agnellous/Wikimedia Commons via (A) http://commons.wikimedia.org/wiki/File:MercuryBarometer.svg; (B) http://commons.wikimedia.org/wiki/File:Barometer.JPG)

A more convenient barometer, called an **aneroid barometer**, measures pressure by the expansion and contraction of a small spring within an evacuated metal capsule.

Atmospheric Pressure and Altitude

At sea level, a mercury column will rise a distance of $760\ \text{mm}$. This **atmospheric pressure** is reported as $760\ \text{mm Hg}$ (millimeters of mercury). At higher altitudes, the atmospheric pressure is decreased and so the column of mercury will not rise as high. On the summit of Mt. Everest (elevation of $8848\ \text{m}$), the air pressure is $253\ \text{mm Hg}$. Atmospheric pressure is slightly dependent on weather conditions. On the graph below, we can see the decrease in atmospheric pressure as the altitude increases. At sea level, the atmospheric pressure would be a little over $100\ \text{kPa}$ (one atmosphere or $760\ \text{mm Hg}$). If we climb to the top of Mount Everest (the highest mountain in the world at 29,029 feet or 8848 meters), the atmospheric pressure will drop to slightly over $30\ \text{kPa}$ (about 0.30 atmospheres or $228\ \text{mm Hg}$). This marked decrease in atmospheric pressure results in much lower levels of oxygen. Teams that climb this mountain must bring supplies of oxygen with them in order to breathe at these high altitudes.

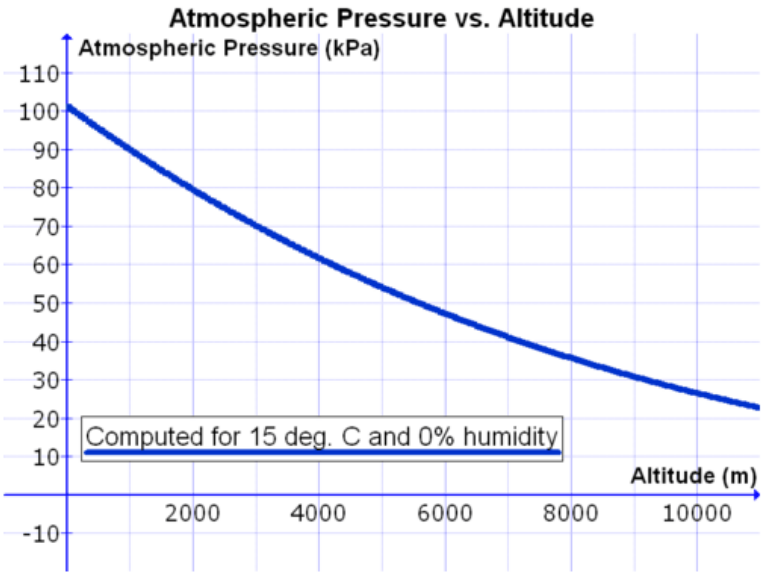


Figure 11.5.3: Effect of altitude on atmospheric pressure. (Public Domain; User: Geek.not.nerd/Wikimedia Commons via Wikipedia)

YouTube Video

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

View this YouTube Video online

Simulation

Ever try cooking at the top of a mountain or at sea level? Compare how atmospheric pressure, vapor pressure and boiling point relate in this simulation:

Does water boil at 100°C everywhere?

Summary

- Atmospheric pressure is the pressure exerted by gas particles in Earth's atmosphere as those particles collide with objects.
- A barometer measures atmospheric pressure.
- Atmospheric pressure decreases as the altitude increases.

Review

1. Define atmospheric pressure.
2. What is an aneroid barometer?
3. How does the atmospheric pressure change as the altitude increases?

This page titled 11.5: Atmospheric Pressure is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.3: Atmospheric Pressure by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.6: Pressure Units and Conversions



Figure 11.6.1 (Public Domain; User:IFCAR/Wikimedia Commons via Wikipedia)

Can you guess how old this car is?

There are several benefits to maintaining the proper air pressure in a car tire. The ride is smoother and safer than it is with too low of pressure. The car gets better gas mileage and the tires do not wear as fast. The recommended pressure for that model of car (usually somewhere between 32-$35\ \text{psi}$) is generally listed in the owner's manual or stamped somewhere inside the door. The pressure on the tire is the maximum pressure for that tire, not the recommended one. Tire pressure is best measured when the tire is cold, since driving the car will heat up the air in the tire and increase the pressure.

Pressure Units and Conversions

A barometer measures gas pressure by the height of the column of mercury. One unit of gas pressure is the millimeter of mercury (mm Hg). An equivalent unit to the mm Hg is called the torr, in honor of the inventor of the barometer, Evangelista Torricelli. The **pascal** (Pa) is the standard unit of pressure. A pascal is a very small amount of pressure, so the most useful unit for everyday gas pressures is the kilopascal (kPa). A kilopascal is equal to 1000 pascals. Another commonly used unit of pressure is the **atmosphere** (atm). Standard atmospheric pressure is defined as $1\ \text{atm}$ of pressure and is equal to $760\ \text{mm Hg}$ and $101.3\ \text{kPa}$. Atmospheric pressure is also often stated as pounds per square inch (psi). The atmospheric pressure at sea level is $14.7\ \text{psi}$.

$$1\ \text{atm} = 760\ \text{mm Hg} = 760\ \text{torr} = 101.3\ \text{kPa} = 14.7\ \text{psi}$$

It is important to be able to convert between different units of pressure. To do so, we will use the equivalent standard pressures shown above.

Example 11.6.1: Pressure Unit Conversions

The atmospheric pressure in a mountainous location is measured to be $613\ \text{mm Hg}$. What is this pressure in atm and in kPa ?

Solution

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

Known

- Given: $613\ \text{mm Hg}$

- $1\ \text{atm} = 760\ \text{mm Hg}$
- $101.3\ \text{kPa} = 760\ \text{mm Hg}$

Unknown

- Pressure $=?\ \text{atm}$
- Pressure $=?\ \text{kPa}$

Use conversion factors from the equivalent pressure units to convert from mm Hg to atm and from mm Hg to kPa.

Step 2: Solve.

$$613\ \text{mm Hg} \times \frac{1\ \text{atm}}{760\ \text{mm Hg}} = 0.807\ \text{atm}$$

$$613\ \text{mm Hg} \times \frac{101.3\ \text{kPa}}{760\ \text{mm Hg}} = 81.7\ \text{kPa}$$

Step 3: Think about your result.

The air pressure is about 80% of standard atmospheric pressure at sea level. For significant figure purposes, the standard pressure of $760\ \text{mm Hg}$ has three significant figures.

YouTube Video

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

View this YouTube Video online

Summary

- Calculations are described for converting between different pressure units.

Review

1. What instrument measures gas pressure by the height of the column of mercury?
2. 1 atm = ___ torr
3. 32.02 atm = ___ kPa
4. 542 mmHg = ___ psi
5. The pressure in a car tire is 35 psi. How many atmospheres is that?

This page titled 11.6: Pressure Units and Conversions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.4: Pressure Units and Conversions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.7: Average Kinetic Energy and Temperature



Figure 11.7.1 (Public Domain; Danny Meyer/U.S. Air Force via Wikipedia)

How much energy does it take to hit a baseball?

Kinetic energy is the energy of motion. Any object that is moving possesses kinetic energy. Baseball involves a great deal of kinetic energy. The pitcher throws a ball, imparting kinetic energy to the ball. When the batter swings, the motion of swinging creates kinetic energy in the bat. The collision of the bat with the ball changes the direction and speed of the ball, with the idea of kinetic energy being involved again.

Kinetic Energy and Temperature

As stated in the kinetic-molecular theory, the temperature of a substance is related to the average **kinetic energy** of the particles of that substance. When a substance is heated, some of the absorbed energy is stored within the particles, while some of the energy increases the motion of the particles. This is registered as an increase in the temperature of the substance.

Average Kinetic Energy

At any given temperature, not all of the particles of a sample of matter have the same kinetic energy. Instead, the particles display a wide range of kinetic energies. Most of the particles have a kinetic energy near the middle of the range. However, a small number of particles have kinetic energies a great deal lower or a great deal higher than the average (see figure below).

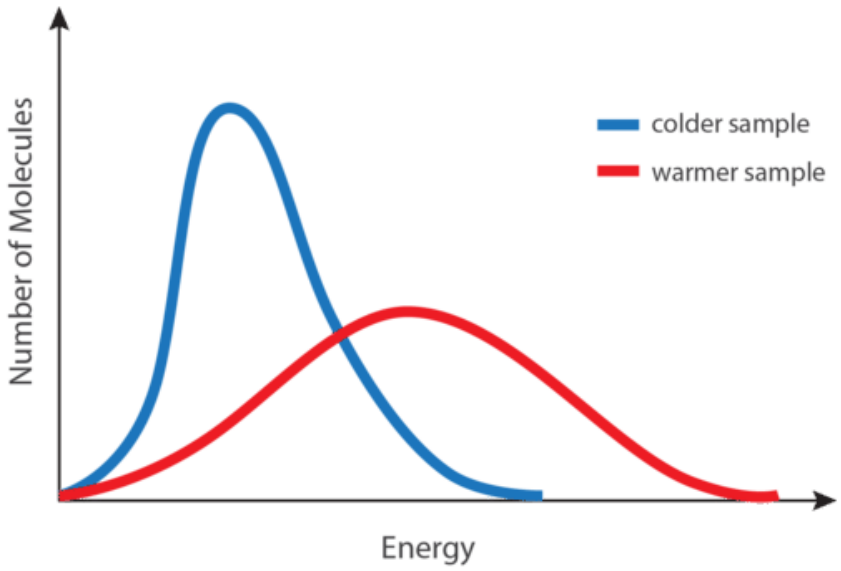


Figure 11.7.2: A distribution of molecular kinetic energies as a function of temperature. The blue curve is for a low temperature, while the red curve is for a high temperature. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The blue curve in the figure above is for a sample of matter at a relatively low temperature, while the red curve is for a sample at a relatively high temperature. In both cases, most of the particles have intermediate kinetic energies, close to the average. Notice that as the temperature increases, the range of kinetic energies increases and the distribution curve "flattens out". At a given temperature, the particles of any substance have the same average kinetic energy.

YouTube Video

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

View this YouTube Video online

Absolute Zero

As a sample of matter is continually cooled, the average kinetic energy of its particles decreases. Eventually, one would expect the particles to stop moving completely. **Absolute zero** is the temperature at which the motion of particles theoretically ceases. Absolute zero has never been attained in the laboratory, but temperatures on the order of 1×10^{-10} K have been achieved. The **Kelvin** temperature scale is the scale that is based on molecular motion, and so absolute zero is also called 0 K. The Kelvin temperature of a substance is directly proportional to the average kinetic energy of the particles of the substance. For example, the particles in a sample of hydrogen gas at 200 K have twice the average kinetic energy as the particles in a hydrogen sample at 100 K.



Figure 11.7.3: Helium gas liquefies at $4\ \mathrm{K}$, or four degrees above absolute zero. Liquid helium is used as a coolant for large superconducting magnets, and must be stored in insulated metal canisters. (CC BY 2.0; Michael Pereckas (Flickr: Beige Alert) via Flickr)

Summary

- Kinetic energy is the energy of motion.
- At a given temperature, individual particles of a substance have a range of kinetic energies.
- The motion of particles theoretically ceases at absolute zero.

Review

1. What is kinetic energy?
2. If the temperature increases, will particles move faster or slower than they would at a lower temperature?
3. What is absolute zero?

This page titled 11.7: Average Kinetic Energy and Temperature is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.5: Average Kinetic Energy and Temperature by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.8: Physical Properties and Intermolecular Forces

Carbon is an interesting and versatile element. There are over twenty million known compounds containing carbon, encompassed in the growing field of organic chemistry. The element itself can exist in two major forms. Diamond is a form of carbon that is extremely hard, and is one of the few materials that can scratch glass. The other form of carbon is graphite, a very soft material that we find in "lead" pencils. The two forms differ mainly in how the carbon atoms are connected to one another. The differences in the arrangement of atoms affect the properties of the material.

Physical Properties and Intermolecular Forces

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. **Molecular compounds**, sometimes called covalent compounds, display a wide range of physical properties due to the different types of intermolecular attractions, such as different kinds of polar interactions. The melting and boiling points of molecular compounds are generally quite low compared to those of **ionic compounds**. This is because the energy required to disrupt the **intermolecular forces** between molecules is far less than the energy required to break the ionic bonds in a crystalline ionic compound. Since molecular compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. Ionic compounds do not conduct electricity in the solid state because of their rigid structure, but conduct well when either molten or dissolved into a solution. The water solubility of molecular compounds is variable and depends primarily on the type of intermolecular forces involved. Substances that exhibit hydrogen bonding or dipole-dipole forces are generally water soluble, whereas those that exhibit only London dispersion forces are generally insoluble. Most, but not all, ionic compounds are quite soluble in water. The table below summarizes some of the differences between ionic and molecular compounds.

Comparison of Ionic and Molecular Compounds

Table 11.8.1: Comparison of Ionic and Molecular Compounds

Property	Ionic Compounds	Molecular Compounds
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - transfer of electron(s) between atoms	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor

One type of molecular compound behaves quite differently than those described thus far. A covalent network solid is a compound in which all of the atoms are connected to one another by covalent bonds. Diamond is composed entirely of carbon atoms, each bonded to four other carbon atoms in a tetrahedral geometry. Melting a covalent network solid is not accomplished by overcoming the relatively weak intermolecular forces. Rather, all of the covalent bonds must be broken, a process that requires extremely high temperatures. Diamond, in fact, does not melt at all. Instead, it vaporizes to a gas at temperatures above $3500^{\circ}\mathrm{C}$.

Summary

- The physical properties of a material are affected by the intermolecular forces holding the molecules together.
- The melting and boiling points of molecular compounds are generally quite low compared to those of ionic compounds.

This page titled 11.8: Physical Properties and Intermolecular Forces is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.6: Physical Properties and Intermolecular Forces by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.9: Surface Tension



Figure 11.9.1 (CC by 2.0; Gordon Wrigley via Flickr)

How is this insect able to stand on water?

The next time you are by a still body of water, take a close look at the creatures scooting along on the surface. You may see insects seemingly floating on top of the water. These creatures are known by a variety of names including water skaters, water striders, and pond skaters. They take advantage of a property called surface tension to stay above the water and not sink. The force they exert downward is less than the forces exerted among the water molecules on the surface of the pond, so the insect does not penetrate the surface of the water.

Surface Tension

Molecules within a liquid are pulled equally in all directions by intermolecular forces. However, molecules at the surface are pulled downwards and sideways by other liquid molecules, but not upwards away from the surface. The overall effect is that the surface molecules are pulled into the liquid, creating a surface that is tightened like a film (see A in the figure below). The **surface tension** of a liquid is a measure of the elastic force within the liquid's surface. Liquids that have strong intermolecular forces, like the hydrogen bonding in water, exhibit the greatest surface tension. Surface tension allows objects that are denser than water, such as the paper clip shown in B in the figure below, to nonetheless float on its surface. It is also responsible for the beading up of water droplets on a freshly waxed car because there are no attractions between the polar water molecules and the nonpolar wax.

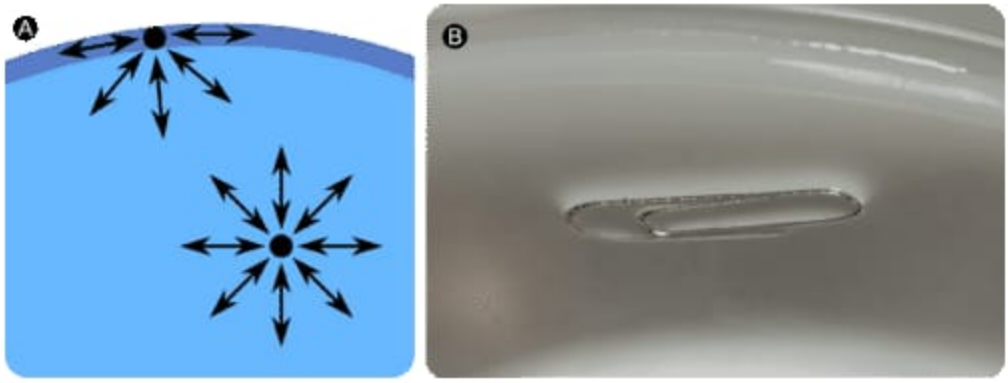


Figure 11.9.2: (A) Molecules at the surface of a liquid are pulled downwards into the liquid, creating a tightened surface. (B) Surface tension allows a paper clip to float on water's surface. (Public Domain; (A) User: Füsiahh/Wikimedia Commons; (B) User: Kaldari/Wikimedia Commons via (A) http://commons.wikimedia.org/wiki/File:Wassermolek%25C3%25BCleInTr%25C3%25B6pfchen-2.svg; (B) http://commons.wikimedia.org/wiki/File:Water_surface_tension_2.jpg)

Other liquids, such as diethyl ether, do not demonstrate strong surface tension interactions. The intermolecular forces for the ether are relatively weak dipole-dipole interactions that do not draw the molecules together as tightly as hydrogen bonds would.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Science Friday: Candy Corn in Space

Candy corn is a very tasty treat. In this video by Science Friday, astronaut Don Pettit uses Candy Corn to demonstrate the effects of hydrophobic and hydrophilic interactions.

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

View this YouTube Video online

Summary

- The surface tension of a liquid is a measure of the elastic force in the liquid's surface.
- Liquids with strong intermolecular forces have higher surface tensions than liquids with weaker forces.

Review

1. Define surface tension.
2. What is responsible for the strong surface tension in water?
3. Does diethyl ether have a stronger or weaker surface tension than water?

This page titled 11.9: Surface Tension is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.6: Surface Tension by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.10: Evaporation



Figure 11.10.1 (Public Domain; User:Mcheath/Wikipedia via Wikipedia)

What is the box on the house's roof?

On the roof of the house pictured below is a device known as a "swamp cooler". This piece of equipment traces its origin back to the ancient Egyptians, who hung wet blankets across the doors of their homes. As the warm air passed through the blankets, water would evaporate and cool the air. The royalty went one step further, and had servants fan wet cloths over jugs of water to get more evaporation and cooling.

The origin of the term "swamp cooler" is not known; they certainly don't work in a swamp. The best conditions for cooling include a high temperature (over $80^{\circ}\mathrm{F}$) and a low humidity (preferably less than 30%). These coolers work well in desert areas, but don't provide any cooling in the humid areas of the country.

Evaporation

A puddle of water left undisturbed eventually disappears. The liquid molecules escape into the gas phase, becoming water vapor. **Vaporization** is the process in which a liquid is converted to a gas. **Evaporation** is the conversion of a liquid to its vapor below the boiling temperature of the liquid. If the water is instead kept in a closed container, the water vapor molecules do not have a chance to escape into the surroundings, and so the water level does not change. As some water molecules become vapor, an equal number of water vapor molecules condense back into the liquid state. **Condensation** is the change of state from a gas to a liquid.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

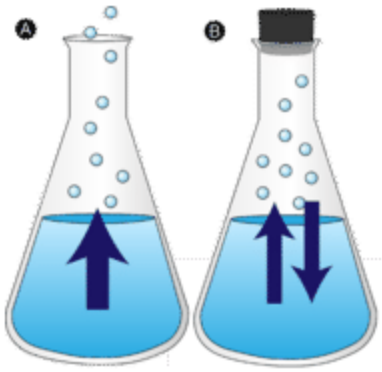


Figure 11.10.2: Evaporation (A) and condensation (B). (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In order for a liquid molecule to escape into the gas state, the molecule must have enough kinetic energy to overcome the intermolecular attractive forces in the liquid. Recall that a given liquid sample will have molecules with a wide range of kinetic energies. Liquid molecules that have this certain threshold kinetic energy escape the surface and become vapor. As a result, the liquid molecules that remain now have lower kinetic energy. As evaporation occurs, the temperature of the remaining liquid decreases. You have observed the effects of evaporative cooling. On a hot day, the water molecules in your perspiration absorb body heat and evaporate from the surface of your skin. The evaporation process leaves the remaining perspiration cooler, which in turn absorbs more heat from your body.

A given liquid will evaporate more quickly when it is heated. This is because the heating process results in a greater fraction of the liquid's molecules having the necessary kinetic energy to escape the surface of the liquid. The figure below shows the kinetic energy distribution of liquid molecules at two temperatures. The numbers of molecules that have the required kinetic energy to evaporate are shown in the shaded area under the curve at the right. The higher temperature liquid (T_2) has more molecules that are capable of escaping into the vapor phase than the lower temperature liquid (T_1).

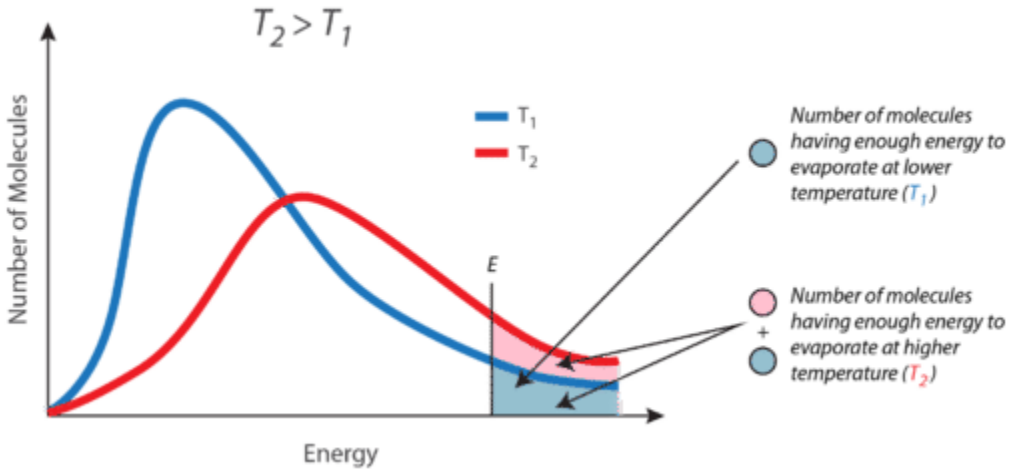


Figure 11.10.3: Kinetic energy distribution curves for a liquid at two temperatures T_1 and T_2. The shaded area represents the molecules with enough kinetic energy to escape the liquid and become vapor. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

YouTube Video

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

View this YouTube Video online

Interactive Element

Explore how heat and temperature relate to phase change in this simulation:

Does the temperature of water rise while it is boiling?

Summary

- Evaporation is the conversion of a liquid to its vapor below the boiling temperature of the liquid.
- Condensation is the change of state from a gas to a liquid.
- As the temperature increases, the rate of evaporation increases.

Review

1. Define vaporization.
2. Define evaporation.
3. Define condensation.
4. How does temperature affect the rate of evaporation?

This page titled 11.10: Evaporation is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.7: Evaporation by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.11: Vapor Pressure

What causes this toy to move?

The drinking duck is a toy that many kids (and adults) enjoy playing with. You can see the drinking duck in action in the video below:

YouTube Video

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

View this YouTube Video online

The motion of the duck illustrates a physical principle called vapor pressure. As the vapor pressure changes, the liquid in the duck moves up and down, causing the duck to move.

Vapor Pressure

When a partially filled container of liquid is sealed with a stopper, some liquid molecules at the surface evaporate into the **vapor phase**. However, the vapor molecules cannot escape from the container. So, after a certain amount of time, the space above the liquid reaches a point where it cannot hold any more vapor molecules. Now, some of the vapor molecules condense back into a liquid. The system reaches the point where the rate of evaporation is equal to the rate of condensation (see figure below). This is considered a dynamic equilibrium between the liquid and vapor phase.

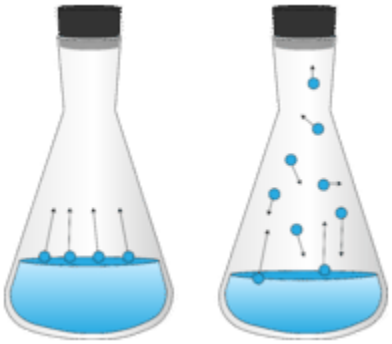


Figure 11.11.1: Equilibrium between liquid phase and vapor phase. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

A dynamic equilibrium can be illustrated by an equation with a double arrow, meaning that the reaction is occurring in both directions and at the same rate.

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_2O}\,(g)$$

The forward direction represents the evaporation process, while the reverse direction represents the condensation process.

Because they cannot escape the container, the vapor molecules above the surface of the liquid exert a pressure on the walls of the container. The **vapor pressure** is a measure of the pressure (force per unit area) exerted by a gas above a liquid in a sealed container. Vapor pressure is a property of a liquid based on the strength of its intermolecular forces. A liquid with weak intermolecular forces evaporates more easily and has a higher vapor pressure. A liquid with stronger intermolecular forces does not evaporate easily, and thus has a lower vapor pressure. For example, diethyl ether is a nonpolar liquid with weak dispersion

forces. Its vapor pressure at $20^{\circ}\mathrm{C}$ is $58.96\ \mathrm{kPa}$. Water is a polar liquid whose molecules are attracted to one another by relatively strong hydrogen bonding. The vapor pressure of water at $20^{\circ}\mathrm{C}$ is only $2.33\ \mathrm{kPa}$, far less than that of diethyl ether.

Vapor Pressure and Temperature

Vapor pressure is dependent upon temperature. When the liquid in a closed container is heated, more molecules escape the **liquid phase** and evaporate. The greater number of vapor molecules strike the container walls more frequently, resulting in an increase in pressure. The table below shows the temperature dependence of the vapor pressure of three liquids.

Vapor Pressure (in kPa of Three Liquids at Different Temperatures

Table 11.11.1: Vapor Pressure (in kPa of Three Liquids at Different Temperatures

	$0^{\circ}\mathrm{C}$	$20^{\circ}\mathrm{C}$	$40^{\circ}\mathrm{C}$	$60^{\circ}\mathrm{C}$	$80^{\circ}\mathrm{C}$	$100^{\circ}\mathrm{C}$
Water	0.61	2.33	7.37	19.92	47.34	101.33
Ethanol	1.63	5.85	18.04	47.02	108.34	225.75
Diethyl Ether	24.70	58.96	122.80	230.65	399.11	647.87

Notice that the temperature dependence of the vapor pressure is not linear. From $0^{\circ}\mathrm{C}$ to $80^{\circ}\mathrm{C}$, the vapor pressure of water increases by $46.73\ \mathrm{kPa}$, while it increases by $53.99\ \mathrm{kPa}$ in only a span of twenty degrees from $80^{\circ}\mathrm{C}$ to $100^{\circ}\mathrm{C}$.

YouTube Video

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

View this YouTube Video online

YouTube Video

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

View this YouTube Video online

Simulation

Ever try to boil water on the top of a mountain? Explore the relationship between altitude, vapor pressure and boiling point in this simulation:

Does water boil at 100°C everywhere?

Summary

- Vapor pressure is a measure of the pressure exerted by a gas above a liquid in a sealed container.
- Strong intermolecular forces produce a lower rate of evaporation and a lower vapor pressure.
- Weak intermolecular forces produce a higher rate of evaporation and a higher vapor pressure.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

- As the temperature increases, the vapor pressure increases.

Review

1. Define vapor pressure.
2. How do intermolecular forces affect vapor pressure?
3. How does temperature affect vapor pressure?

This page titled 11.11: Vapor Pressure is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.8: Vapor Pressure by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.12: Boiling



Figure 11.12.1 (Public Domain; the Earth Observatory/NASA via Commons Wikimedia, Mount Everest [commons.wikimedia.org])

Why is it harder to breathe at the top of Mount Everest than at the bottom?

At 29,029 feet $(8848\ \mathrm{m})$, Mount Everest in the Himalayan range, on the border between China and Nepal, is the highest point on the earth. Its altitude presents many practical problems to climbers. The oxygen content of the air is much lower than at sea level, making it necessary to bring oxygen tanks along (although a few climbers have reached the peak without oxygen). One other problem is that of boiling water for cooking food. Although water boils at $100^{\circ}\mathrm{C}$ at sea level, the boiling point on top of Mount Everest is only about $70^{\circ}\mathrm{C}$. This difference makes it very difficult to get a decent cup of tea (which definitely frustrates some British climbers).

Boiling Point

As a liquid is heated, the average kinetic energy of its particles increases. The rate of evaporation increases as more and more molecules are able to escape the liquid's surface into the vapor phase. Eventually a point is reached when the molecules all throughout the liquid have enough kinetic energy to **vaporize**. At this point, the liquid begins to boil. The **boiling point** is the temperature at which the vapor pressure of a liquid is equal to the external pressure. The figure below illustrates the boiling of liquid.

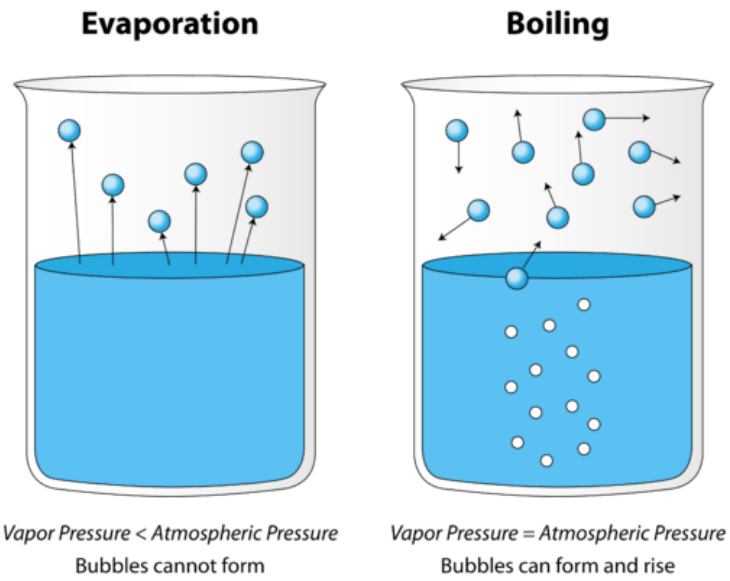


Figure 11.12.2: Comparison between evaporation and boiling. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In the picture on the left, the liquid is below its boiling point, yet some of the liquid evaporates. On the right, the temperature has been increased until bubbles begin to form in the body of the liquid. When the vapor pressure inside the bubbles is equal to the external atmospheric pressure, the bubbles rise to the surface of the liquid and burst. The temperature at which this process occurs is the boiling point of the liquid.

The normal boiling point is the temperature at which the vapor pressure of the liquid is equal to standard pressure. Because atmospheric pressure can change based on location, the boiling point of a liquid changes with the external pressure. The normal boiling point is a constant because it is defined relative to the standard atmospheric pressure of $760\ \text{mm Hg}$ (or $1\ \text{atm}$ or $101.3\ \text{kPa}$).

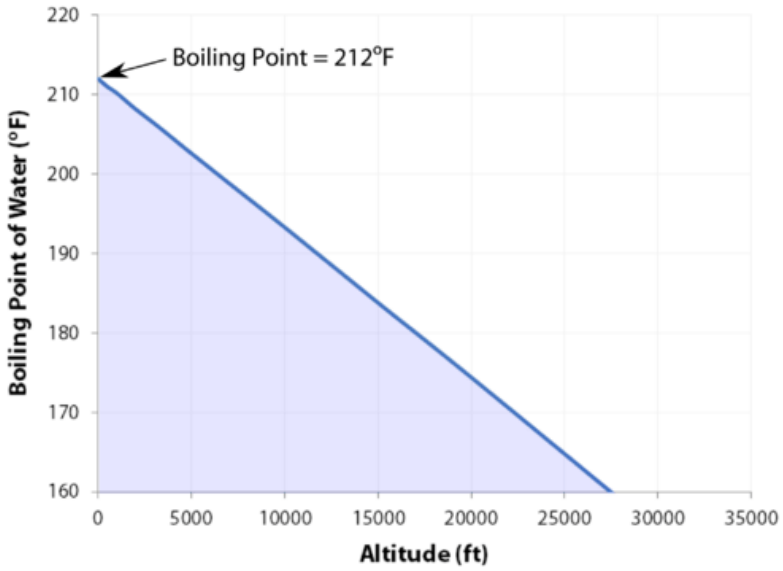


Figure 11.12.3: Influence of altitude on the boiling point of water. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Summary

- The boiling point is the temperature at which the vapor pressure of a liquid is equal to the external pressure.
- As the altitude increases, the boiling point decreases.

Review

1. What happens when a liquid is heated?
2. What is the boiling point?
3. What is the pressure at which the normal boiling point is determined?
4. Are the boiling point and normal boiling point always the same value, sometimes the same value, or never the same value? Explain your answer.

Explore More

Use the video below to answer the following questions:

YouTube Video

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

View this YouTube Video online

1. The boiling point of water is 100°C. What is an example of liquid that boils a temperature higher than the boiling point of water?
2. What is an example of liquid that boiling point is below room temperature?
3. What phase change is the opposite of boiling?

This page titled 11.12: Boiling is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.9: Boiling by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.13: Vapor Pressure Curves



Figure 11.13.1 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

What is the fastest way to boil water?

One of the first lessons in cooking is how to boil water. Yes, it sounds simple, but there are a couple of hints that speed things up. One hint is to put a lid on the pot. The picture above has water boiling uncovered with the steam escaping to the atmosphere. If the lid is on the pot, less water will be boiled off and the water will boil faster. The buildup of pressure inside the pot helps speed up the boiling process.

Vapor Pressure Curves

The boiling points of various liquids can be illustrated in a **vapor pressure curve** (figure below). A vapor pressure curve is a graph of vapor pressure as a function of temperature. To find the normal boiling point of liquid, a horizontal line is drawn from the y-axis at a pressure equal to standard pressure. A vertical line starting at the x-axis can be drawn connected to the point at which the standard pressure and the vapor pressure curve of a liquid intersect—the corresponding temperature is the boiling point of that liquid.

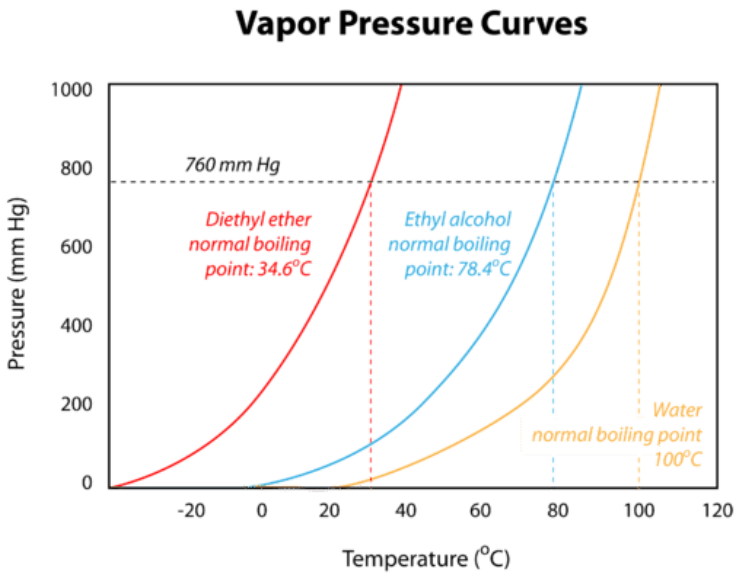


Figure 11.13.2: Vapor pressure curves. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The boiling point of a liquid also correlates to the strength of its intermolecular forces. Recall that diethyl ether has weak dispersion forces, which means that the liquid has a high vapor pressure. The weak forces also mean that it does not require a large input of energy to make diethyl ether boil, and so it has a relatively low normal boiling point of $34.6^{\circ}\mathrm{C}$. Water, with its much stronger hydrogen bonding, has a low vapor pressure and a higher normal boiling point of $100^{\circ}\mathrm{C}$.

As stated earlier, boiling points are affected by external pressure. At higher altitudes, the atmospheric pressure is lower. With less pressure pushing down on the surface of the liquid, it boils at a lower temperature. This can also be seen from the vapor pressure curves. If one draws a horizontal line at a lower vapor pressure, it intersects each curve at a lower temperature. The boiling point of water is $100^{\circ}\mathrm{C}$ at sea level, where the atmospheric pressure is standard. In Denver, Colorado at $1600\ \mathrm{m}$ above sea level, the atmospheric pressure is about $640\ \mathrm{mm\ Hg}$ and water boils at about $95^{\circ}\mathrm{C}$. On the summit of Mt. Everest, the atmospheric pressure is about $255\ \mathrm{mm\ Hg}$, and water boils at only $70^{\circ}\mathrm{C}$. On the other hand, water boils at greater than $100^{\circ}\mathrm{C}$ if the external pressure is higher than standard. Pressure cookers do not allow the vapor to escape and the vapor pressure increases. Since water in a pressure cooker boils at a temperature above $100^{\circ}\mathrm{C}$, the food cooks more quickly.



Figure 11.13.3: Pressure cooker. (CC BY 2.0; Dinner Series via Flickr, Dinner Series [www.flickr.com])

The effect of decreased air pressure can be demonstrated by placing a beaker of water in a vacuum chamber. At a low enough pressure, about $20\ \mathrm{mm\ Hg}$, water will boil at room temperature.

YouTube Video

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

View this YouTube Video online

Summary

- A vapor pressure curve is a graph of vapor pressure as a function of temperature.
- Boiling points are affected by external pressure.

Review

1. What does a vapor pressure curve show?
2. Why does diethyl ether have a low boiling point?
3. What intermolecular forces hold water molecules together?
4. Why does water boil at a lower temperature when at a high altitude?

Explore More

Use the resource below to answer the questions that follow.

YouTube Video

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

View this YouTube Video online

1. What does a vacuum pump do?
2. What is the role of the O-ring?
3. What happens when the vacuum pump is turned on?
4. What happened when the vacuum pump was turned on?
5. What was the temperature of this boiling water?

This page titled 11.13: Vapor Pressure Curves is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.10: Vapor Pressure Curves by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.14: Melting



Figure 11.14.1 (Public Domain; Jonas Bergsten (Wikimedia: Bergsten) via Wikipedia)

Have you ever gone ice skating?

In the winter, many people find the snow and ice beautiful. They enjoy getting out to ski or ice-skate. Others don't find that time of year to be so much fun. When the snow melts, the roads get very sloppy and messy. Those people look forward to spring when all the ice and snow are gone and the weather is warmer.

Melting Point

Solids are similar to liquids in that both are condensed states, with particles that are far closer together than those of a gas. However, while liquids are fluid, solids are not. The particles of most solids are packed tightly together in an orderly arrangement. The motion of individual atoms, ions, or molecules in a solid is restricted to **vibrational motion** about a fixed point. Solids are almost completely **incompressible** and are the densest of the three states of matter.

As a solid is heated, its particles vibrate more rapidly as it absorbs kinetic energy. Eventually, the organization of the particles within the solid structure begins to break down, and the solid starts to melt. The **melting point** is the temperature at which a solid changes into a liquid. At its melting point, the disruptive vibrations of the particles of the solid overcome the attractive forces operating within the solid. As with boiling points, the melting point of a solid is dependent on the strength of those attractive forces. Sodium chloride (NaCl) is an ionic compound that consists of a multitude of strong ionic bonds. Sodium chloride melts at $801^{\circ}\mathrm{C}$. Ice (solid $\mathrm{H_2O}$) is a molecular compound whose molecules are held together by hydrogen bonds. Though hydrogen bonds are the strongest of the intermolecular forces, the strength of hydrogen bonds is much less than that of ionic bonds. The melting point of ice is 0 °C.

The melting point of a solid is the same as the freezing point of the liquid. At that temperature, the solid and liquid states of the substance are in equilibrium. For water, this equilibrium occurs at $0^{\circ}\mathrm{C}$.

$$\mathrm{H_2O}\,(s) \rightleftharpoons \mathrm{H_2O}\,(l)$$

We tend to think of solids as those materials that are solid at room temperature. However, all materials have melting points of some sort. Gases become solids at extremely low temperatures, and liquids will also become solid if the temperature is low enough.

YouTube Video

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

View this YouTube Video online

The table below gives the melting points of some common materials.

Table 11.14.1*: Melting Points of Common Materials*

Materials	Melting Point (°C)
Hydrogen	-259
Oxygen	-219
Diethyl ether	-116
Ethanol	-114
Water	0
Pure silver	961
Pure gold	1063
Iron	1538

YouTube Video

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

View this YouTube Video online

Simulation

Explore how heat and temperature relate to phase change in this simulation:

Does the temperature of water rise while it is boiling?

Summary

- The melting point is the temperature at which a solid changes into a liquid.
- Intermolecular forces have a strong influence on melting point.

Review

1. Define melting point.
2. What happens when a material melts?
3. Would you expect ethane (C_2H_6) to have a higher or lower melting point than water? Explain your answer.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

This page titled 11.14: Melting is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.11: Melting by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.15: Sublimation



Figure 11.15.1 (Public Domain; Conrad Poirier via Wikipedia)

How did early settlers do their laundry in the winter?

Of course, they could wash the clothes inside, but where to dry them? There were no dryers available, so the clothes had to be hung up to dry. In the winter, you would expect that ice would form on the clothes, but that didn't happen. Even in freezing weather, the water might freeze initially, but would eventually go off as a vapor – from solid ice directly to the gas stage. There are still families today who have to rely on this cumbersome process to get their laundry done.

Sublimation

Vapor Pressure of a Solid

"Sublimation" surveys the vapor pressure of a liquid and its dependence upon temperature. Solids also have a vapor pressure, though it is generally much less than that of a liquid. A snow bank will gradually disappear even if the temperature stays below $0^{\circ}\mathrm{C}$. The snow does not melt, but instead passes directly from the solid state to the vapor state. **Sublimation** is the change of state from a solid to a gas, without passing through the liquid state.

Examples of Sublimation

Sublimation of Iodine

When iodine is heated sublimation can be readily observed. Iodine's vapor is a distinctive purple color and has a very strong scent, making it easy to detect. The video below shows the sublimation and deposition of iodine. The beaker containing iodine is covered with a round-bottom flask that contains ice. When the iodine is heated it sublimes. When iodine vapors cool, deposition occurs. Deposition is the change of state from a gas to a solid.

YouTube Video

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

View this YouTube Video online

Dry Ice

"Dry ice", or solid carbon dioxide, is a substance that sublimes at atmospheric pressures. Dry ice is very cold $(-78^{\circ}\mathrm{C})$ and so is used as a coolant for goods, such as ice cream, that must remain frozen during shipment. Because the dry ice sublimes rather than melts, there is no liquid mess associated with its change of state as it warms. As you may have seen in demonstrations, dry ice merely transforms from a solid to a cold, steamy looking gas.

YouTube Video

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

View this YouTube Video online

Ferrocene

Ferrocene (an iron-containing compound) is usually purified by sublimation. A sample of crude ferrocene is gently heated causing it to sublime. When the ferrocene is cooled, reddish ferrocene crystals deposited on the outside of that tube as shown in the figure below.



Figure 11.15.2: Ferrocene purified by sublimation. (Public Domain; User: Bucki/De.Wikipedia via Wikipedia)

Summary

- Sublimation is the change of state from a solid to a gas, without passing through the liquid state.
- Deposition is the change of state from a gas to a solid.
- Carbon dioxide is an example of a material that easily undergoes sublimation.

Review

1. Define sublimation.
2. What is another name for solid carbon dioxide?
3. How is ferrocene purified?
4. Name a substance which undergoes sublimation.

This page titled 11.15: Sublimation is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.12: Sublimation by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.16: Heating and Cooling Curves

During the lifetime of Mark Twain (real name: Samuel Langhorne Clemens, 1835-1910), the steamboat was a major means of transportation on the rivers and lakes of the United States. Twain himself was a steamboat pilot on the Mississippi River for a period of time, and took his pen name from the measurement of water depth (twelve feet, which was a safe depth for the boats). The boats got their power from steam—liquid water converted to a gas at high temperatures. The steam would push the pistons of the engine, causing the paddle wheels to turn and propel the boat.

Heating Curves

Imagine that you have a block of ice that is at a temperature of $-30^{\circ}\mathrm{C}$, well below its **melting** point. The ice is in a closed container. As heat is steadily added to the ice block, the water molecules will begin to vibrate faster and faster as they absorb kinetic energy. Eventually, when the ice has warmed to $0^{\circ}\mathrm{C}$, the added energy will start to break apart the hydrogen bonding that keeps the water molecules in place when it is in the **solid** form. As the ice melts, its temperature does not rise. All of the energy that is being put into the ice goes into the melting process and not into any increase in temperature. During the melting process, the two states—solid and **liquid**—are in equilibrium with one another. If the system was isolated at that point and no energy was allowed to enter or leave, the ice-water mixture at $0^{\circ}\mathrm{C}$ would remain. Temperature is always constant during a change of state.

Continued heating of the water after the ice has completely melted will now increase the kinetic energy of the liquid molecules and the temperature will rise. Assuming that the atmospheric pressure is standard, the temperature will rise steadily until it reaches $100^{\circ}\mathrm{C}$. At this point, the added energy from the heat will cause the liquid to begin to vaporize. As with the previous state change, the temperature will remain at $100^{\circ}\mathrm{C}$ while the water molecules are going from the liquid to the **gas** or vapor state. Once all the liquid has completely boiled away, continued heating of the steam (since the container is closed) will increase its temperature above $100^{\circ}\mathrm{C}$.

The experiment described above can be summarized in a graph called a heating curve (figure below).

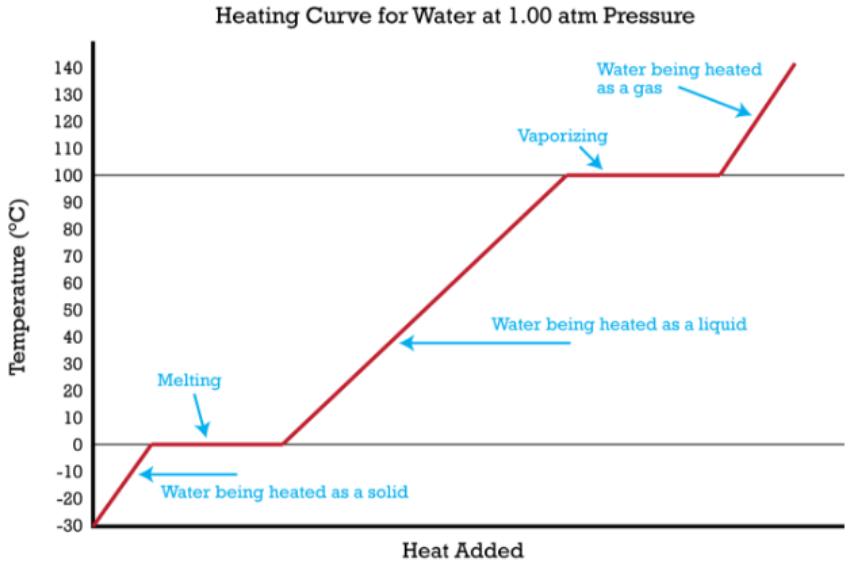


Figure 11.16.1: In the heating curve of water, the temperature is shown as heat is continually added. Changes of state occur during plateaus, because the temperature is constant.

The change of state behavior of all substances can be represented with a heating curve of this type. The melting and boiling points of the substance can be determined by the horizontal lines or plateaus on the curve. Other substances have melting and boiling points that are different from those of water. An exception to this blueprint heating curve is for a substance such as carbon dioxide, which sublimes rather than melts at standard pressure. The heating curve for carbon dioxide would have only one plateau, at the **sublimation** temperature of $\mathrm{CO_2}$.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

The entire experiment could be run in reverse. Steam above $100^{\circ}\mathrm{C}$ could be steadily cooled down to $100^{\circ}\mathrm{C}$, at which point it would condense to liquid water. The water could then be cooled to $0^{\circ}\mathrm{C}$, at which point continued cooling would freeze the water to ice. The ice could then be cooled to a point below $0^{\circ}\mathrm{C}$. This could be diagrammed in a cooling curve that would be the reverse of the heating curve.

Summary of State Changes

All of the changes of state that occur between solid, liquid, and gas are summarized in the diagram in the figure below. **Freezing** is the opposite of melting, and both represent the equilibrium between the solid and liquid states. **Evaporation** occurs when a liquid turns to a gas. **Condensation** is the opposite of vaporization, and both represent the equilibrium between the liquid and gas states. **Deposition** is the opposite of sublimation, and both represent the equilibrium between the solid and gas states.

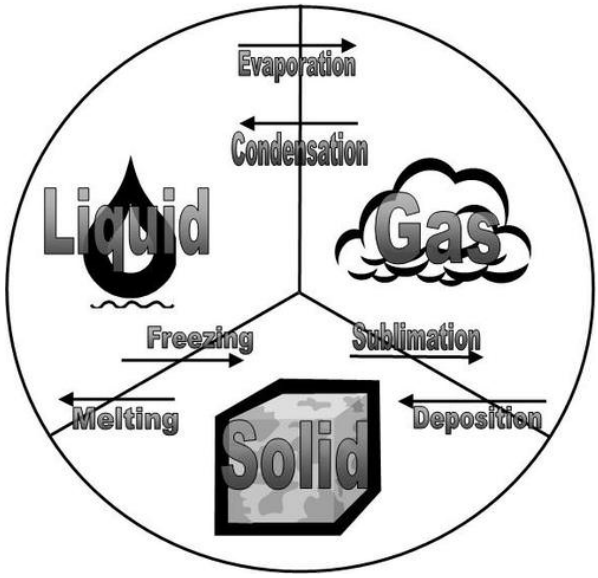


Figure 11.16.2: Solid, liquid, and gas states with the terms for each change of state that occurs between them.

Summary

- A change of state can be brought about by putting heat into a system or removing it from the system.
- The temperature of a system will not change as long as the substance is undergoing a change from solid to liquid to gas, as well as the reverse.
- Freezing is the opposite of melting.
- Evaporation occurs when a liquid turns to a gas.
- Condensation is the opposite of vaporization.
- Deposition is the opposite of sublimation.

This page titled 11.16: Heating and Cooling Curves is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.18: Heating and Cooling Curves by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.17: General Phase Diagram

Many rockets use a combination of kerosene and liquid oxygen for their fuel. Oxygen can be reduced to the liquid state either by cooling or by using high pressure. In the case of a rocket, since the oxygen is in a container essentially out in the open, maintaining a temperature of $-183^{\circ}\mathrm{C}$ (the boiling point of oxygen) is not very practical. However, high pressure can be used to force the oxygen into tanks, causing it to liquefy so that it can mix with the kerosene, and provide a powerful ignition to move the rocket.

Phase Diagrams

The relationship among the solid, liquid, and vapor (gas) states of a substance can be shown as a function of temperature and pressure in a single diagram. A **phase diagram** is graph showing the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas phases. Examine the general phase diagram shown in the figure below. In each of the three colored regions of the diagram, the substance is in a single state (or phase). The dark lines that act as the boundary between those regions represent the conditions under which the two phases are in equilibrium.

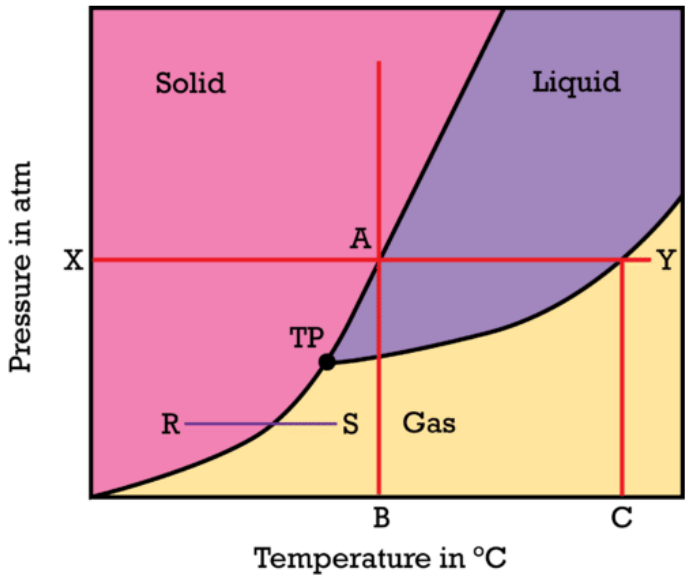


Figure 11.17.1: General phase diagram, which shows the state (phase) of a substance as a function of its temperature and pressure.

Find the X on the pressure axis and presume that the value of X is standard pressure of $1\ \mathrm{atm}$. As one moves left to right across the red line, the temperature of the solid substance is being increased while the pressure remains constant. When point A is reached, the substance melts and the temperature B on the horizontal axis represents the normal melting point of the substance. Moving further to the right, the substance boils at point Y and so point C on the horizontal axis represents the normal boiling point of the substance. As the temperature increases at a constant pressure, the substance changes from solid to liquid to gas.

Start right above point B on the temperature axis and follow the red line vertically. At very low pressure, the particles of the substance are far apart from one another and the substance is in the gas state. As the pressure is increased, the particles of the substance are forced closer and closer together. Eventually the particles are pushed so close together that attractive forces cause the substance to condense into the liquid state. Continually increasing the pressure on the liquid will eventually cause the substance to solidify. For the majority of substances, the solid state is denser than the liquid state and so putting a liquid under great pressure will cause it to turn into a solid. The line segment R-S represents the process of sublimation, where the substance changes directly from a solid to a gas. At a sufficiently low pressure, the liquid phase does not exist. The point labeled TP is called the **triple point**. The triple point is the one condition of temperature and pressure where the solid, liquid, and vapor states of a substance can all coexist at equilibrium.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Summary

- A phase diagram graphs the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas states.
- The triple point is the one condition of temperature and pressure where the solid, liquid, and vapor states of a substance can all coexist at equilibrium.

This page titled 11.17: General Phase Diagram is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.19: General Phase Diagram by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

11.18: Phase Diagram for Water

A specific consistency of snow is required to make the best snowballs. Dry snow can be tightly pressed, and will form snowballs because the higher pressure causes the snowflakes to melt somewhat. However, when you release the pressure, the snow goes back to a more solid form and the flakes no longer stick together. Ideally, instead, the snow needs to be a little bit wet so that the particles will stick together.

Phase Diagram for Water

Water is a unique substance in many ways. One of these special properties is the fact that solid water (ice) is less dense than liquid water just above the freezing point. The phase diagram for water is shown in the figure below.

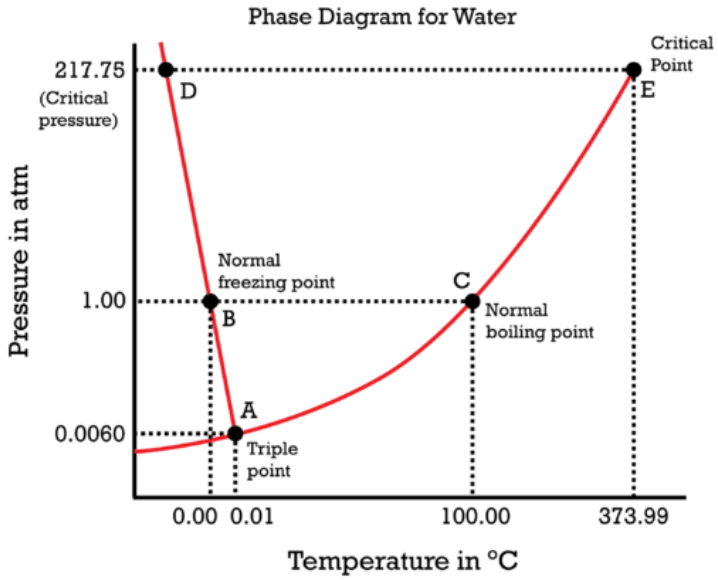


Figure 11.18.1: Phase diagram for water.

Notice one key difference between last section's general phase diagram, and the above phase diagram for water: in water's diagram, the slope of the line between the solid and liquid states is negative rather than positive. The reason is that water is an unusual substance, in that its solid state is less dense than the liquid state. Ice floats in liquid water. Therefore, a pressure change has the opposite effect on those two phases. If ice is relatively near its melting point, it can be changed into liquid water by the application of pressure. The water molecules are actually closer together in the liquid phase than they are in the solid phase.

Refer again to water's phase diagram (figure above). Notice point E, labeled the **critical point**. What does that mean? At $373.99^{\circ}\mathrm{C}$, particles of water in the gas phase are moving very, very rapidly. At any temperature higher than that, the gas phase cannot be made to liquefy, no matter how much pressure is applied to the gas. The **critical pressure** (P_{C}) is the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid. For water, the critical pressure is very high, 217.75 atm. The critical point is the intersection point of the **critical temperature** and the critical pressure.

Summary

- Solid water is less dense than liquid water just above the freezing point.
- The critical temperature (T_{C}) of a substance is the highest temperature at which the substance can possibly exist as a liquid.
- The critical pressure (P_{C}) is the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid.
- The critical point is the intersection point of the critical temperature and the critical pressure.

This page titled 11.18: Phase Diagram for Water is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 13.20: Phase Diagram for Water by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

12: Gas Laws

This page titled 12: Gas Laws is shared under a license and was authored, remixed, and/or curated by LibreTexts.

13: SOLUTIONS

13.1: Solute-Solvent Combinations

13.2: Rate of Dissolution

13.3: Saturated and Unsaturated Solutions

13.4: How Temperature Influences Solubility

13.5: Supersaturated Solutions

13.6: Predicting Precipitates Using Solubility Rules

13.7: Percent Solutions

13.8: Molarity

13.9: Preparing Solutions

13.10: Dilution

13.11: Molality

13.12: Calculating Molar Mass

13.13: Electrolytes and Nonelectrolytes

13.14: Strong and Weak Electrolytes

13.15: Molecular and Ionic Equations

13.16: Net Ionic Equations

Template:HideTOC

This page titled 13: Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16: Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.1: Solute-Solvent Combinations

Solute-Solvent Combinations

The focus of "Chapter 15: Water" was water's role in the formation of aqueous solutions. We examined the primary characteristics of a solution and how water is able to dissolve solid solutes; we differentiated between a solution, a suspension, and a colloid. There are many examples of solutions that do not involve water at all, or that involve solutes that are not solids. The table below summarizes the possible combinations of solute-solvent states, along with examples of each.

Table 13.1.1*: Solute-Solvent Combinations*

Solute State	Solvent State	Example
liquid	gas	water in air
gas	gas	oxygen in nitrogen (gas mixture)
solid	liquid	salt in water
liquid	liquid	alcohol in water
gas	liquid	carbon dioxide in water
solid	solid	zinc in copper (brass alloy)
liquid	solid	mercury in silver and tin (dental amalgam)

Gas-Gas Solutions

Our air is a homogenous mixture of many different gases and therefore qualifies as a solution. Approximately 78% of the atmosphere is nitrogen, making it the solvent for this solution. The next major constituent is oxygen (about 21%), followed by the inert gas argon (0.9%), carbon dioxide (0.03%), and trace amounts of neon, methane, helium, and other gases.

Solid-Solid Solutions

Solid-solid solutions such as brass, bronze, and sterling silver are called alloys. Bronze (composed mainly of copper with added tin) was widely used in making weapons in times past, dating back to at least 2400 B.C. This metal alloy was hard and tough, but was eventually replaced by iron.

Liquid-Liquid Solutions

Perhaps the most familiar liquid-solid solution is dental amalgam, used to fill teeth when there is a cavity. Approximately 50% of the amalgam material is liquid mercury to which a powdered alloy of silver, tin, and copper is added. Mercury is used because it binds well with the solid metal alloy. However, the use of mercury-based dental amalgam has gone under question in recent years, because of concerns regarding the toxicity of mercury.

Summary

This page titled 13.1: Solute-Solvent Combinations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.1: Solute-Solvent Combinations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.2: Rate of Dissolution

Many people enjoy a cold glass of iced tea on a hot summer day. Some like it unsweetened, while others like to put sugar in it. How sugar dissolves in the tea depends on two factors: how much sugar was put into the tea, and how cold it is. Tea usually has to be stirred for a while to get all the sugar dissolved.

Rate of Dissolution

Dissolution is the process by which a solute dissolves into a solvent and forms a solution. We know that the dissolution of a solid by water depends upon the collisions that occur between the solvent molecules and the particles in the solid crystal. Anything that can be done to increase the frequency of those collisions and/or to give those collisions more energy will increase the rate of dissolution. Imagine that you were trying to dissolve some sugar in a glassful of tea. A packet of granulated sugar would dissolve faster than a cube of sugar. The rate of dissolution would be increased by stirring, or agitating the solution. Finally, the sugar would dissolve faster in hot tea than it would in cold tea.

Surface Area

The rate at which a solute dissolves depends upon the size of the solute particles. Dissolution is a surface phenomenon, since it depends on solvent molecules colliding with the outer surface of the solute. A given quantity of solute dissolves faster when it is ground into small particles, rather than in the form of large pieces, because more **surface area** is exposed. A packet of granulated sugar exposes far more surface area to the solvent and dissolves more quickly than a sugar cube.

Agitation of the Solution

Dissolving sugar in water will occur more quickly if the water is stirred. The stirring allows fresh solvent molecules to continually be in contact with the solute. If it is not stirred, then the water right at the surface of the solute becomes saturated with dissolved sugar molecules, meaning that it is more difficult for additional solute to dissolve. The sugar cube would eventually dissolve because random motions of the water molecules would bring enough fresh solvent into contact with the sugar, but the process would take much longer. It is important to realize that neither stirring nor breaking up a solute affect the overall amount of solute that dissolves—these actions only affect the rate of dissolution.

Temperature

Heating up a solvent gives the molecules more kinetic energy. The increased rapid motion means that the solvent molecules collide with the solute with greater frequency, and that the collisions occur with more force. Both factors increase the rate at which the solute dissolves. As we will see in the next section, a temperature change not only affects the rate of dissolution, but also affects the amount of solute that dissolves.

Summary

This page titled 13.2: Rate of Dissolution is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.2: Rate of Dissolution by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.3: Saturated and Unsaturated Solutions

When compounds are synthesized, contaminating materials are often mixed in with them. The process of recrystallization can be used to remove these impurities. The crystals are dissolved in a hot solvent, forming a solution. When the solvent is cooled, the compound is no longer as soluble and will precipitate out of solution, leaving contaminating materials still dissolved.

Saturated and Unsaturated Solutions

Table salt (NaCl) readily dissolves in water. Suppose that you have a beaker of water to which you add some salt, stirring until it dissolves. You add more and that dissolves. You keep adding more and more salt, eventually reaching a point at which no more of the salt will dissolve—no matter how long or how vigorously you stir it. Why? On the molecular level, we know that action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by water molecules. What also happens is that some of the dissolved ions collide back again with the crystal and remain there. **Recrystallization** is the process of dissolved solute returning to the solid state. At some point, the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged. **Solution equilibrium** is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate. The solution equilibrium for the dissolution of sodium chloride can be represented by one of two equations:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{NaCl}\,(aq)$$

While the above equation shows the change of state back and forth between solid and aqueous solution, the preferred equation also shows the dissociation that occurs as an ionic solid dissolves:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{Na}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

When the solution equilibrium point is reached and no more solute will dissolve, the solution is said to be saturated. A **saturated solution** is a solution that contains the maximum amount of solute that is capable of dissolving. At $20^{\circ}\mathrm{C}$, the maximum amount of NaCl that will dissolve in \100. \: \text{g}\) of water is $36.0\ \mathrm{g}$. If any more NaCl is added past this point, it will not dissolve, because the solution is saturated. What if more water is added to the solution instead? In that case, more NaCl would be capable of dissolving in the additional solvent. An **unsaturated solution** is a solution that contains less than the maximum amount of solute that is capable of being dissolved. The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

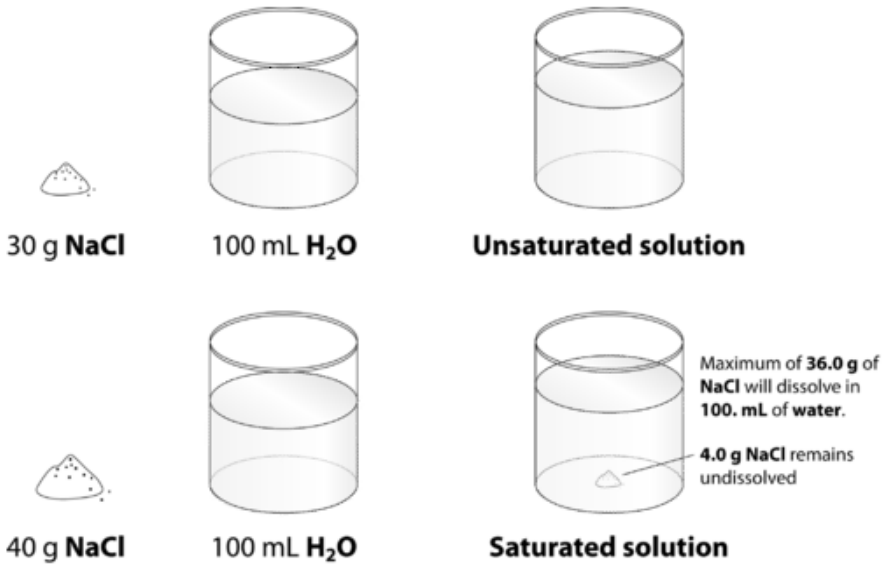


Figure 13.3.1: When $30.0\ \mathrm{g}$ of NaCl is added to $100\ \mathrm{mL}$, it all dissolves, forming an unsaturated solution. When $40.0\ \mathrm{g}$ is added, $36.0\ \mathrm{g}$ dissolves and $4.0\ \mathrm{g}$ remains undissolved, forming a saturated solution.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

Summary

- Recrystallization is the process of dissolved solute returning to the solid state.
- A saturated solution is a solution that contains the maximum amount of solute that is capable of dissolving.
- An unsaturated solution is a solution that contains less than the maximum amount of solute that is capable of being dissolved.
- Solution equilibrium exists when the rate of dissolution equals the rate of recrystallization.

This page titled 13.3: Saturated and Unsaturated Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.3: Saturated and Unsaturated Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.4: How Temperature Influences Solubility

Nuclear power plants require large amounts of water to generate steam for turbines and to cool equipment. They are usually situated near bodies of water to use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per $100\ \text{g}$ of solvent. The solubility of sodium chloride in water is $36.0\ \text{g}$ per $100\ \text{g}$ water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. In this section, we will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (see figure below).

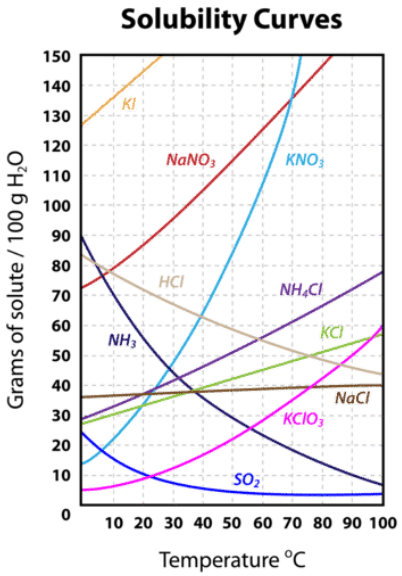


Figure 13.4.1: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep, and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, NH_3, and SO_2—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ \text{g}$ of KNO_3 is added to $100\ \text{g}$ of water at 30°C. According to the solubility curve, approximately $48\ \text{g}$ of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since $48\ \text{g}$ is less than $80\ \text{g}$. We can also determine that there will be $80 - 48 = 32\ \text{g}$ of undissolved KNO_3 remaining at the bottom of the container. In a second scenario, suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about $107\ \text{g}$. The solution, in this case, is unsaturated since it contains only the original $80\ \text{g}$ of dissolved solute. Suppose in a third case, that the solution is cooled all the way down to 0°C. The solubility at 0°C is about $14\ \text{g}$, meaning that $80 - 14 = 66\ \text{g}$ of the KNO_3 will recrystallize.

Summary

- The solubility of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
- A solubility curve is a graph of the solubility vs. temperature
- The solubility of a solid in water increases with an increase in temperature.
- Gas solubility decreases as the temperature increases.

This page titled 13.4: How Temperature Influences Solubility is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.4: How Temperature Influences Solubility by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.5: Supersaturated Solutions

In the picture below, a thermal pack is applied to the back. Small packs can be used either for heating or cooling, depending on the material used. A heat pack contains a supersaturated solution of material such as sodium acetate. The solution is clear until a small metal trigger is activated. The sodium acetate then crystallizes out of solution and generates heat in the process.



Figure 13.5.1: A thermal pack is applied to heat or cool muscles.

Supersaturated Solutions

Some solutes, such as sodium acetate, do not recrystallize easily. Suppose an exactly saturated solution of sodium acetate is prepared at $50^{\circ}\mathrm{C}$. As it cools back to room temperature, no crystals appear in the solution, even though the solubility of sodium acetate is lower at room temperature. A **supersaturated solution** is a solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature. The recrystallization of the excess dissolved solute in a supersaturated solution can be initiated by the addition of a tiny crystal of solute, called a seed crystal. The seed crystal provides a nucleation site on which the excess dissolved crystals can begin to grow. Recrystallization from a supersaturated solution is typically very fast.

Summary

- A supersaturated solution is a solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature.
- When a seed crystal is added to the solution, a supersaturated solution can recrystallize.

This page titled 13.5: Supersaturated Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.5: Supersaturated Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.6: Predicting Precipitates Using Solubility Rules

Predicting the weather is tricky business. A thorough examination of a large amount of data is needed to make the daily forecast. Wind patterns, historical data, barometric pressure—these and many other data are fed into computers that then use a set of rules to predict what will happen based on past history.

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all.

It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (shown in Table 13.6.1).

Table 13.6.1: Solubility Rules for Ionic Compounds in Water

Soluble	Compounds containing the alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+) and ammonium ion (NH_4^+).
Soluble	Compounds containing the nitrate ion (NO_3^-), acetate ion (CH_3COO^-), chlorate ion (ClO_3^-), and bicarbonate ion (HCO_3^-).
Mostly Soluble	Compounds containing the chloride ion (Cl^-), bromide ion (Br^-), and iodide ion (I^-). Exceptions are those of silver (Ag^+), mercury (I) (Hg_2^{2+}), and lead (II) ([redacted]).
Mostly Soluble	Compounds containing the sulfate ion (SO_4^{2-}). Exceptions are those of silver (Ag^+), calcium ([redacted]), strontium ([redacted]), barium ([redacted]), mercury (I) (Hg_2^{2+}), and lead (II) ([redacted]).
Mostly Insoluble	Compounds containing the carbonate ion (CO_3^{2-}), phosphate ion (PO_4^{3-}), chromate ion (CrO_4^{2-}), sulfide ion ([redacted]), and silicate ion (SiO_3^{2-}). Exceptions are those of the alkali metals and ammonium.
Mostly Insoluble	Compounds containing the hydroxide ion (OH^-). Exceptions are those of the alkali metals and the barium ion ([redacted]).

For practice using the solubility rules, predict if a precipitate will form when solutions of cesium bromide and lead (II) nitrate are mixed.

$$Cs^+(aq) + Br^-(aq) + [redacted](aq) + 2NO_3^-(aq) \rightarrow ?$$

The potential precipitates from a double-replacement reaction are cesium nitrate and lead (II) bromide. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, the cesium and nitrate ions are spectator ions and the lead (II) bromide is a precipitate. The balanced net ionic reaction is:

$$[redacted](aq) + 2Br^-(aq) \rightarrow PbBr_2(s)$$

Summary

This page titled 13.6: Predicting Precipitates Using Solubility Rules is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.19: Predicting Precipitates Using Solubility Rules by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.7: Percent Solutions

There are human cultures that do not recognize numbers above three. Anything greater than that is simply referred to as "much" or "many". Although this form of calculation may seem very limited, American culture does the same thing to a certain degree. For example, there are several ways to express the amount of solute in a solution in a quantitative manner. The **concentration** of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution. A **concentrated** solution is one that has a relatively large amount of dissolved solute. A **dilute** solution is one that has a relatively small amount of dissolved solute. However, those terms are vague, and it is often necessary to express concentration with numbers.

Percent Solutions

One way to describe the concentration of a solution is by the percent of a solute in the solvent. The percent can further be determined in one of two ways: (1) the ratio of the mass of the solute divided by the mass of the solution or (2) the ratio of the volume of the solute divided by the volume of the solution.

Mass Percent

When the solute in a solution is a solid, a convenient way to express the concentration is by mass percent $\left(\frac{\text{mass}}{\text{mass}}\right)$, which is the grams of solute per $100\ \text{g}$ of solution.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Suppose that a solution was prepared by dissolving $25.0\ \text{g}$ of sugar into $100\ \text{g}$ of water. The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25\ \text{g sugar}}{125\ \text{g solution}} \times 100\% = 20\%\ \text{sugar}$$

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 solvent to use. For example, you need to make $3000\ \text{g}$ of a 5% solution of sodium chloride. You can rearrange and solve for the mass of solute:

$$\text{mass of solute} = \frac{\text{percent by mass}}{100\%} \times \text{mass of solution} = \frac{5\%}{100\%} \times 3000\ \text{g} = 150\ \text{g NaCl}$$

You would need to weigh out $150\ \text{g}$ of NaCl and add it to $2850\ \text{g}$ of water. Notice that it was necessary to subtract the mass of the $\text{NaCl}\ (150\ \text{g})$ from the mass of solution $(3000\ \text{g})$ to calculate the mass of the water that would need to be added.

Volume Percent

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution, expressed as a percent, yields the percent by volume $\left(\frac{\text{volume}}{\text{volume}}\right)$ of the solution. If a solution is made by adding $40\ \text{mL}$ of ethanol to $200\ \text{mL}$ of water, the percent by volume is:

$$\begin{aligned}\text{Percent by volume} &= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \\ &= \frac{40\ \text{mL ethanol}}{240\ \text{mL solution}} \times 100\% \\ &= 16.7\%\ \text{ethanol}\end{aligned}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

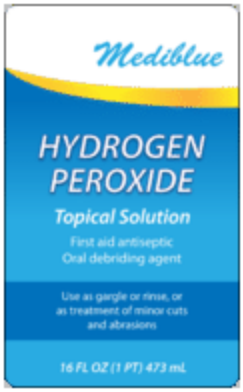


Figure 13.7.1: Hydrogen peroxide is commonly labeled as a 3% by volume solution for use as a disinfectant.

Summary

- The concentration of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution.
- A concentrated solution has a relatively large amount of dissolved solute.
- A dilute solution has a relatively small amount of dissolved solute.
- Techniques for calculation of percent mass and percent volume solution concentrations are described.

This page titled 13.7: Percent Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.7: Percent Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.8: Molarity

Chemists deal with amounts of molecules every day. Chemical reactions are described as so many molecules of compound A reacting with so many molecules of compound B to form so many molecules of compound C. When we determine how much reagent to use, we need to know the number of molecules in a given volume of the reagent. Percent solutions only tell us the number of grams, not molecules. A $100\ \text{mL}$ solution of $2\%\ \text{NaCl}$ will have a very different number of molecules than a 2% solution of CsCl. So, we need another way to talk about numbers of molecules.

Molarity

Chemists primarily need the concentration of solutions to be expressed in a way that accounts for the number of particles that react according to a particular chemical equation. Since percentage measurements are based on either mass or volume, they are generally not useful for chemical reactions. A concentration unit based on moles is preferable. The **molarity** (M) of a solution is the number of moles of solute dissolved in one liter of solution. To calculate the molarity of a solution, divide the moles of solute by the volume of the solution expressed in liters:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

Note that the volume is in liters of solution and not liters of solvent. When a molarity is reported, the unit is the symbol M and is read as "molar". For example a solution labeled as $1.5\ \text{M}\ \text{NH}_3$ is read as "1.5 molar ammonia solution".

Example 13.8.1

A solution is prepared by dissolving $42.23\ \text{g}$ of NH_4Cl into enough water to make $500.0\ \text{mL}$ of solution. Calculate its molarity.

Solution

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

Known

- Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$
- Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$
- Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$

Unknown

The mass of the ammonium chloride is first converted to moles. Then the molarity is calculated by dividing by liters. Note that the given volume has been converted to liters.

Step 2: Solve.

$$42.23\ \text{g}\ \text{NH}_4\text{Cl} \times \frac{1\ \text{mol}\ \text{NH}_4\text{Cl}}{53.50\ \text{g}\ \text{NH}_4\text{Cl}} = 0.7893\ \text{mol}\ \text{NH}_4\text{Cl}$$

$$\frac{0.7893\ \text{mol}\ \text{NH}_4\text{Cl}}{0.5000\ \text{L}} = 1.579\ \text{M}$$

Step 3: Think about your result.

The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol}\ \text{NH}_4\text{Cl}$. Four significant figures are appropriate.

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.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Example 13.8.2

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

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

Known

- Molarity $= 0.250\ \text{M}$
- Volume $= 3.00\ \text{L}$
- Molar mass $\text{KMnO}_4 = 158.04\ \text{g/mol}$

Unknown

Moles of solute is calculated by multiplying molarity by liters. Then, moles is converted to grams.

Step 2: Solve.

$$\text{mol KMnO}_4 = 0.250\ \text{M KMnO}_4 \times 3.00\ \text{L} = 0.750\ \text{mol KMnO}_4$$

$$0.750\ \text{mol KMnO}_4 \times \frac{158.04\ \text{g KMnO}_4}{1\ \text{mol KMnO}_4} = 119\ \text{g KMnO}_4$$

Step 3: Think about your result.

When $119\ \text{g}$ of potassium permanganate is dissolved into water to make $3.00\ \text{L}$ of solution, the molarity is $0.250\ \text{M}$.

Summary

- The molarity (M) of a solution is the number of moles of solute dissolved in one liter of solution.
- Calculations using the concept of molarity are described.

This page titled 13.8: Molarity is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.8: Molarity by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.9: Preparing Solutions

Back in the "good ol' days" (a variable time frame, dependent on who you ask), many cooks did not bother with careful measurements. They simply "knew" how much flour to use, or how much water to add. Most of us need somewhat more precise ways to measure when we cook. Chemists are very careful when they prepare solutions, because the results of their experiments need to be quantitative. Simply "knowing" is not accurate enough for scientific purposes.

Preparing Solutions

If you are attempting to prepare $1.00\ \mathrm{L}$ of a $1.00\ \mathrm{M}$ solution of NaCl, you would obtain $58.44\ \mathrm{g}$ of sodium chloride. However, you cannot simply add the sodium chloride to $1.00\ \mathrm{L}$ of water. After the solute dissolves, the volume of the solution will be slightly greater than one liter, because the hydrated sodium and chloride ions take up space in the solution. Instead, a **volumetric flask** needs to be used. Volumetric flasks come in a variety of sizes (see image below) and are designed to allow a chemist to prepare a solution of only one specific volume.



Figure $13.9.1$: Volumetric flasks.

In other words, you cannot use a 1-liter volumetric flask to make $500\ \mathrm{mL}$ of a solution. It can only be used to prepare 1 liter of a solution. The steps to follow when preparing a solution with a 1-liter volumetric flask are outlined below, and shown in the figure below.

1. The appropriate mass of solute is weighed out and added to a volumetric flask that has been about half-filled with distilled water.

2. The solution is swirled until all of the solute dissolves.

3. More distilled water is carefully added up to the line etched on the neck of the flask.

4. The flask is capped and inverted several times to completely mix.

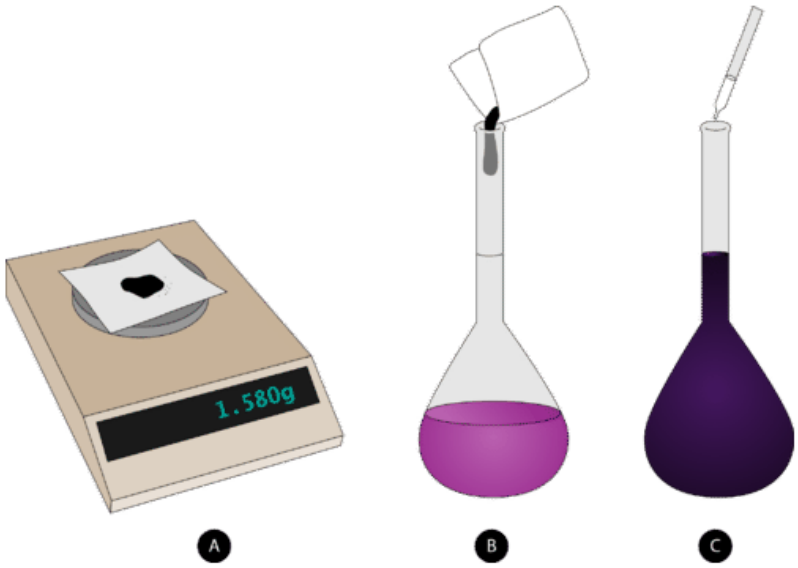


Figure 13.9.2: Steps to follow in preparing a solution of known molarity: (A) weigh out correct mass of solute, (B) dissolve into solvent in a volumetric flask, and (C) add solvent to the fill line on the flask and mix.

Summary

This page titled 13.9: Preparing Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.9: Preparing Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.10: Dilution

Muriatic acid (another name for HCl) is widely used for cleaning concrete and masonry surfaces. The acid must be diluted before use to get it down to a safer strength. Commercially available at concentrations of about 18%, muriatic acid can be used to remove scales and deposits (usually composed of basic materials).

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the number of moles of the solute does not change, while the volume of the solution increases. We can set up an equality between the moles of the solute before the dilution (1) and the moles of the solute after the dilution (2).

$$\mathrm{mol}_1 = \mathrm{mol}_2$$

Since the moles of solute in a solution is equal to the molarity multiplied by the liters, we can set those equal.

$$M_1 \times L_1 = M_2 \times L_2$$

Finally, because the two sides of the equation are set equal to one another, the volume can be described in any unit that we choose, as long as that unit is the same on both sides. Our equation for calculating the molarity of a diluted solution becomes:

$$M_1 \times V_1 = M_2 \times V_2 \tag{13.10.1}$$

Suppose that you have $100.\ \mathrm{mL}$ of a $2.0\ \mathrm{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \mathrm{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2:

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \mathrm{M} \times 100.\ \mathrm{mL}}{500.\ \mathrm{mL}} = 0.40\ \mathrm{M\ HCl}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves 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 13.10.1

Nitric acid $(\mathrm{HNO_3})$ is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \mathrm{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \mathrm{L}$ of a $0.50\ \mathrm{M}$ solution?

Solution

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

Known

- Stock $\mathrm{HNO_3} = 16\ \mathrm{M}$
- $V_2 = 8.00\ \mathrm{L}$
- $M_2 = 0.50\ \mathrm{M}$

Unknown

The unknown in the equation is V_1, the volume of the concentrated stock solution.

Step 2: Solve

Using Equation 13.10.1 we can solve for V_1:

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.50\ \text{M} \times 8.00\ \text{L}}{16\ \text{M}} = 0.25\ \text{L} = 250\ \text{mL}$$

Step 3: 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 $0.5\ \text{M}$.

Dilutions can be performed in the laboratory with various tools, depending on the volumes required and the desired accuracy. The images below illustrate the use of two different types of pipettes. In the first figure, a glass **pipette** is being used to transfer a portion of a solution to a graduated cylinder. Use of a pipette rather than a graduated cylinder for the transfer improves accuracy. The second figure shows a **micropipette**, which is designed to quickly and accurately dispense small volumes. Micropipettes are adjustable and come in a variety of sizes.



Figure 13.10.1: Volumetric pipette and micropipette.

Summary

- A process is described to calculate dilutions.
- Pipettes and micropipettes are used to transfer and dispense solution.

This page titled 13.10: Dilution is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.10: Dilution by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.11: Molality

For many purposes, the use of molarity is very convenient. However, when we want to know the concentration of solute present in situations where there are temperature changes, molarity will not work. The volume of the solution will change somewhat with temperature, enough to affect the accuracy of data observations and calculations. Another parameter is needed, one not affected by the temperature of the material we are studying.

Molality

A final way to express the concentration of a solution is by its molality. The **molality** (m) of a solution is the moles of solute divided by the kilograms of solvent. A solution that contains $1.0\ \text{mol}$ of NaCl dissolved into $1.0\ \text{kg}$ of water is a "one-molal" solution of sodium chloride. The symbol for molality is a lower-case m written in italics.

$$\text{Molality}\,(m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{\text{mol}}{\text{kg}}$$

Molality differs from molarity only in the denominator. While molarity is based on the liters of solution, molality is based on the kilograms of solvent. Concentrations expressed in molality are used when studying properties of solutions related to vapor pressure and temperature changes. Molality is used because its value does not change with changes in temperature. The volume of a solution, on the other hand, is slightly dependent upon temperature.

Example 13.11.1

Determine the molality of a solution prepared by dissolving $28.60\ \text{g}$ of glucose $(\text{C}_6\text{H}_{12}\text{O}_6)$ into $250\ \text{g}$ of water.

Solution

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

Known

- Mass solute $= 28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6$
- Mass solvent $= 250\ \text{g} = 0.250\ \text{kg}$
- Molar mass $\text{C}_6\text{H}_{12}\text{O}_6 = 180.18\ \text{g/mol}$

Unknown

Convert grams of glucose to moles and divide by the mass of the water in kilograms.

Step 2: Solve.

$$28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{180.18\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6} = 0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6$$

$$\frac{0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{0.250\ \text{kg}\ \text{H}_2\text{O}} = 0.635\ m\ \text{C}_6\text{H}_{12}\text{O}_6$$

Step 3: Think about your result.

The answer represents the moles of glucose per kilogram of water and has three significant figures.

Molality and molarity are closely related in value for dilute aqueous solutions because the density of those solutions is relatively close to $1.0\ \text{g/mL}$. This means that $1.0\ \text{L}$ of solution has nearly a mass of $1.0\ \text{kg}$. As the solution becomes more concentrated, its density will not be as close to $1.0\ \text{g/mL}$ and the molality value will be different than the molarity. For solutions with solvents other than water, the molality will be very different than the molarity. Make sure that you are paying attention to which quantity is being used in a given problem.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

Summary

- The molality (m) of a solution is the moles of solute divided by the kilograms of solvent.
- The value of molality does not change with temperature.

This page titled 13.11: Molality is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.11: Molality by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.12: Calculating Molar Mass

Putting antifreeze into a radiator will keep an engine from freezing. By knowing how cold the weather will get and how much water is in the radiator, one can determine how much antifreeze to add to achieve a specific desired freezing point depression. This is made possible by knowing what antifreeze is. Can things be switched around, so as to get some information about the properties of the antifreeze (such as its molecular weight) from the freezing point decrease? As it turns out, this can be done fairly easily and accurately.

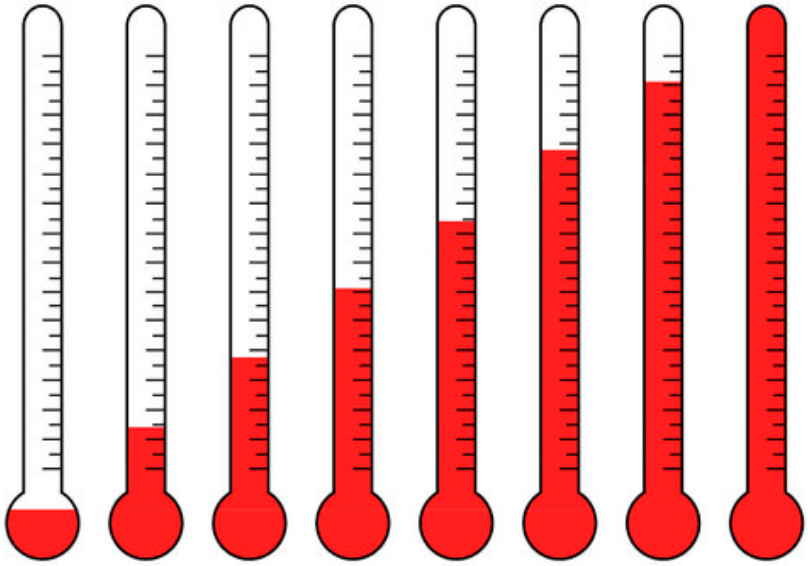


Figure 13.12.1: Changes in temperature.

Calculating Molar Mass

In the laboratory, freezing point or boiling point data can be used to determine the molar mass of an unknown solute. Since we know the relationship between a decrease in freezing point and the concentration of solute, if we dissolve a known mass of our unknown solute into a known amount of solvent, we can calculate the molar mass of the solute. The K_f or K_b of the solvent must be known. We also need to know if the solute is an electrolyte or a nonelectrolyte. If the solvent is an electrolyte, we would need to know the number of ions produced when it dissociates.

Example 13.12.1

$38.7\ \text{g}$ of a nonelectrolyte is dissolved into $218\ \text{g}$ of water. The freezing point of the solution is measured to be -5.53°C. Calculate the molar mass of the solute.

Solution

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

Known

- $\Delta T_f = -5.53^{\circ}\text{C}$
- Mass $\text{H}_2\text{O} = 218\ \text{g} = 0.218\ \text{kg}$
- Mass solute $= 38.7\ \text{g}$
- $K_f\,(\text{H}_2\text{O}) = -1.86^{\circ}\text{C}/m$

Unknown

Use the freezing point depression (ΔT_f) to calculate the molality of the solution. Then use the molality equation to calculate the moles of solute. Then divide the grams of solute by the moles to determine the molar mass.

Step 2: Solve.

$$m = \frac{\Delta T_f}{K_f} = \frac{-5.53^\circ\text{C}}{-1.86^\circ\text{C}/m} = 2.97\ m$$
$$\text{mol solute} = m \times \text{kg}\ \text{H}_2\text{O} = 2.97\ m \times 0.218\ \text{kg} = 0.648\ \text{mol}$$
$$\frac{38.7\ \text{g}}{0.648\ \text{mol}} = 59.7\ \text{g/mol}$$

Step 3: Think about your result.

The molar mass of the unknown solute is $59.7\ \text{g/mol}$. Knowing the molar mass is an important step in determining the identity of an unknown. A similar problem could be done with the change in boiling point.

Summary

This page titled 13.12: Calculating Molar Mass is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.16: Calculating Molar Mass by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.13: Electrolytes and Nonelectrolytes

People around the world jog for exercise. For the most part, jogging can be a healthy way to stay fit. However, problems can develop for those who jog in the heat. Excessive sweating can lead to electrolyte loss, which can be life-threatening. Early symptoms of electrolyte deficiency can include nausea, fatigue, and dizziness. If not treated, individuals can experience muscle weakness and increased heart rate (which could lead to a heart attack). Sports drinks can be consumed to restore electrolytes quickly in the body.

Electrolytes and Nonelectrolytes

An **electrolyte** is a compound that conducts an electric current when it is in an aqueous solution or melted. In order to conduct a current, a substance must contain mobile ions that can move from one electrode to the other. All ionic compounds are electrolytes. When ionic compounds dissolve, they break apart into ions which are then able to conduct a current (**conductivity**). Even insoluble ionic compounds such as $\mathrm{CaCO_3}$ are electrolytes because they can conduct a current in the molten (melted) state.

A **nonelectrolyte** is a compound that does not conduct an electric current in either aqueous solution or in the molten state. Many molecular compounds, such as sugar or ethanol, are nonelectrolytes. When these compounds dissolve in water, they do not produce ions. The figure below illustrates the difference between an electrolyte and a nonelectrolyte.

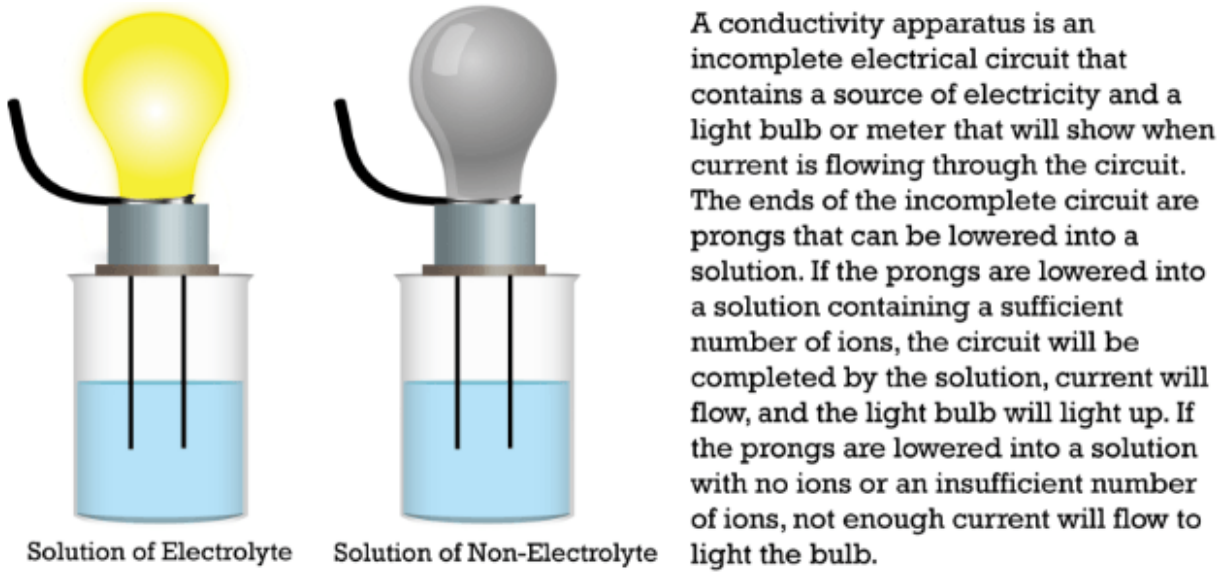


Figure 13.13.1: Conductivity apparatus.

Roles of Electrolytes in the Body

Several electrolytes play important roles in the body. Here are a few significant electrolytes:

1. Calcium - in bones and teeth. Also important for muscle contraction, blood clotting, and nerve function.

2. Sodium - found outside the cell. Mainly involved in water balance and nerve signaling.

3. Potassium - major cation inside the cell. Important for proper functioning of heart, muscles, kidneys, and nerves.

4. Magnesium - in bone and cells. Involved in muscle, bone, nervous system, and takes part in many biochemical reactions.

Summary

- Electrolytes conduct electric current when in solution or melted.
- Nonelectrolytes do not conduct electric current when in solution or melted.
- Some electrolytes play important roles in the body.

This page titled 13.13: Electrolytes and Nonelectrolytes is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 15.7: Electrolytes and Nonelectrolytes by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.14: Strong and Weak Electrolytes

Car batteries are used around the world to provide the power to start car engines. One essential component of car batteries is the strong electrolyte sulfuric acid. In the battery, this material ionizes into hydrogen ions and sulfate ions. As the battery is used, the concentrations of these ions decreases. Older batteries had openings in the top where new sulfuric acid could be added to replenish the supply. Today, batteries are sealed to prevent leakage of hazardous sulfuric acid.

Strong and Weak Electrolytes

Some polar molecular compounds are nonelectrolytes when they are in their pure state, but become electrolytes when they are dissolved in water. Hydrogen chloride (HCl) is a gas in its pure molecular state and is a nonelectrolyte. However, when HCl is dissolved in water, it conducts a current well because the HCl molecule ionizes into hydrogen and chloride ions.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

HCl dissolved into water is called hydrochloric acid. A **strong electrolyte** is a solution in which a large fraction of the dissolved solute exists as ions. Ionic compounds, and some polar compounds, are completely broken apart into ions and thus conduct a current very well—this makes them strong electrolytes.

Some other polar molecular compounds become electrolytes upon being dissolved into water, but do not ionize to very great extent. Gaseous nitrous acid ionizes in solution to hydrogen ions and nitrite ions, but does so very weakly. Aqueous nitrous acid is composed of only about 5% ions and 95% intact nitrous acid molecules. A **weak electrolyte** is a solution in which only a small fraction of the dissolved solute exists as ions. The equation showing the ionization of a weak electrolyte utilizes a double arrow indicating an equilibrium between the reactants and products.

$$\mathrm{HNO_2}\,(g) \rightleftharpoons \mathrm{H}^{+}\,(aq) + \mathrm{NO_2}^{-}\,(aq)$$

Summary

- A strong electrolyte exists mainly as ions in solution.
- A weak electrolyte has only a small amount of ionization in solution.

This page titled 13.14: Strong and Weak Electrolytes is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 15.9: Strong and Weak Electrolytes by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.15: Molecular and Ionic Equations

One of the unfortunate byproducts of industrialized society is acid rain. Sulfur dioxide from burning coal and nitrogen oxides from vehicle emissions both form acids. When these acids react with limestone (calcium carbonate), reactions occur that dissolve the limestone and release water and carbon dioxide. Over a period of time, serious damage is caused to the limestone structure.

Molecular and Ionic Equations

When ionic compounds are dissolved into water, the polar water molecules break apart the solid crystal lattice, resulting in the hydrated ions being evenly distributed through the water. This process is called dissociation and is the reason that all ionic compounds are strong electrolytes. When two different ionic compounds that have been dissolved in water are mixed, a chemical reaction may occur between certain pairs of the hydrated ions. Consider the double-replacement reaction that occurs when a solution of sodium chloride is mixed with a solution of silver nitrate:

$$\mathrm{NaCl}\,(aq) + \mathrm{AgNO_3}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{AgCl}\,(s)$$

The driving force behind this reaction is the formation of the silver chloride precipitate (see figure below).

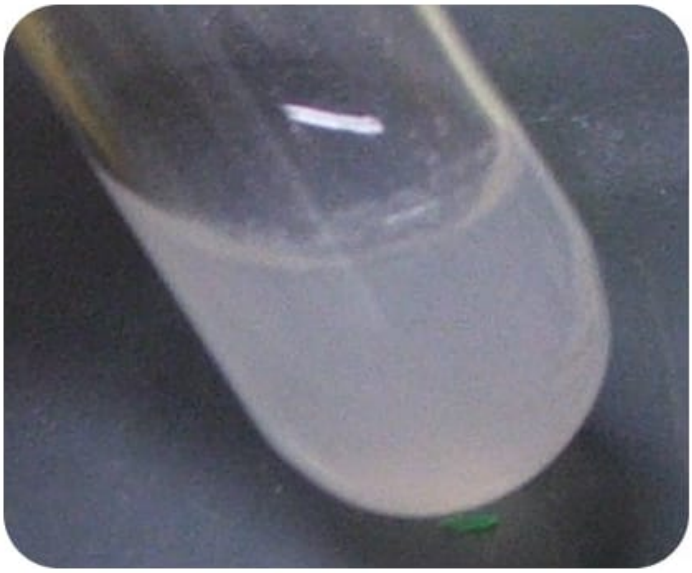


Figure 13.15.1: The white silver chloride precipitate instantly forms when a solution of silver nitrate is added to a solution of sodium chloride.

This is called a molecular equation. A **molecular equation** is an equation in which the formulas of the compounds are written as though all substances exist as molecules. However, there is a better way to show what is happening in this reaction. All of the aqueous compounds should be written as ions, because they are present in the water as separated ions—a result of their dissociation.

$$\mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) + \mathrm{NO_3{}^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{NO_3{}^-}\,(aq) + \mathrm{AgCl}\,(s)$$

This equation is called an **ionic equation**, an equation in which dissolved ionic compounds are shown as free ions.

Some other double-replacement reactions do not produce a precipitate as one of the products. The production of a gas and/or a molecular compound such as water may also drive the reaction. For example, consider the reaction of a solution of sodium carbonate with a solution of hydrochloric acid (HCl). The products of the reaction are aqueous sodium chloride, carbon dioxide, and water. The balanced molecular equation is:

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

The ionic equation is:

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

$$2\mathrm{Na}^{+}\left(aq\right)+{\mathrm{CO}_3}^{2-}\left(aq\right)+2\mathrm{H}^{+}\left(aq\right)+2\mathrm{Cl}^{-}\left(aq\right)\rightarrow 2\mathrm{Na}^{+}\left(aq\right)+2\mathrm{Cl}^{-}\left(aq\right)+\mathrm{CO}_2\left(g\right)+\mathrm{H}_2\mathrm{O}\left(l\right)$$

A single-replacement reaction is one in which an element replaces another element in a compound. An element is in either the solid, liquid, or gas state and is not an ion. The example below shows the reaction of solid magnesium metal with aqueous silver nitrate to form aqueous magnesium nitrate and silver metal.

Balanced molecular equation:

$$\mathrm{Mg}\left(s\right)+2\mathrm{AgNO}_3\left(aq\right)\rightarrow \mathrm{Mg}(\mathrm{NO}_3)_2\left(aq\right)+2\mathrm{Ag}\left(s\right)$$

Ionic equation:

$$\mathrm{Mg}\left(s\right)+2\mathrm{Ag}^{+}\left(aq\right)+2{\mathrm{NO}_3}^{-}\left(aq\right)\rightarrow \mathrm{Mg}^{2+}\left(aq\right)+2{\mathrm{NO}_3}^{-}\left(aq\right)+\mathrm{Ag}\left(s\right)$$

This type of single-replacement reaction is called a metal replacement. Other common categories of single-replacement reactions are: hydrogen replacement and halogen replacement.

Summary

- A molecular equation is an equation in which the formulas of the compounds are written as though all substances exist as molecules.
- An ionic equation is one in which dissolved ionic compounds are shown as free ions.
- Examples of molecular and ionic equations are shown.

This page titled 13.15: Molecular and Ionic Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.17: Molecular and Ionic Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

13.16: Net Ionic Equations

At sports events around the world, a small number of athletes fiercely compete on fields and in stadiums. They get tired, dirty, and sometimes hurt as they try to win the game. Surrounding them are thousands of spectators watching and cheering. Would the game be different without the spectators? Definitely! Spectators provide encouragement to the team and generate enthusiasm. Although the spectators are not playing the game, they are certainly a part of the process.

Net Ionic Equations

We can write a molecular equation for the formation of silver chloride precipitate:

$$\mathrm{NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl}$$

The corresponding ionic equation is:

$$\mathrm{Na^+\,(aq) + Cl^-\,(aq) + Ag^+\,(aq) + NO_3{}^-\,(aq) \rightarrow Na^+\,(aq) + NO_3{}^-\,(aq) + AgCl\,(s)}$$

If you look carefully at the ionic equation, you will notice that the sodium ion and the nitrate ion appear unchanged on both sides of the equation. When the two solutions are mixed, neither the $\mathrm{Na^+}$ nor the $\mathrm{NO_3{}^-}$ ions participate in the reaction. They can be eliminated from the reaction.

$$\mathrm{\cancel{Na^+\,(aq)} + Cl^-\,(aq) + Ag^+\,(aq) + \cancel{NO_3{}^-\,(aq)} \rightarrow \cancel{Na^+\,(aq)} + \cancel{NO_3{}^-\,(aq)} + AgCl\,(s)}$$

A **spectator ion** is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction. In the above reaction, the sodium ion and the nitrate ion are both spectator ions. The equation can now be written without the spectator ions:

$$\mathrm{Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)}$$

The **net ionic equation** is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction. Notice that in writing the net ionic equation, the positively-charged silver cation was written first on the reactant side, followed by the negatively-charged chloride anion. This is somewhat customary because that is the order in which the ions must be written in the silver chloride product. However, it is not absolutely necessary to order the reactants in this way.

Net ionic equations must be balanced by both mass and charge. Balancing by mass means ensuring that there are equal masses of each element on the product and reactant sides. Balancing by charge means making sure that the overall charge is the same on both sides of the equation. In the above equation, the overall charge is zero, or neutral, on both sides of the equation. As a general rule, if you balance the molecular equation properly, the net ionic equation will end up being balanced by both mass and charge.

Example 13.16.1

When aqueous solutions of copper (II) chloride and potassium phosphate are mixed, a precipitate of copper (II) phosphate is formed. Write a balanced net ionic equation for this reaction.

Solution

Step 1: Plan the problem.

Write and balance the molecular equation first, making sure that all formulas are correct. Then write the ionic equation, showing all aqueous substances as ions. Carry through any coefficients. Finally, eliminate spectator ions and write the net ionic equation.

Step 2: Solve.

Molecular equation:

$$3\mathrm{CuCl_2}\,(aq) + 2\mathrm{K_3PO_4}\,(aq) \rightarrow 6\mathrm{KCl}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Ionic equation:

$$3[\blacksquare]\,(aq) + 6\mathrm{Cl^-}\,(aq) + 6\mathrm{K^+}\,(aq) + 2{\mathrm{PO_4}}^{3-}\,(aq) \rightarrow 6\mathrm{K^+}\,(aq) + 6\mathrm{Cl^-}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Notice that the balance of the equation is carried through when writing the dissociated ions. For example, there are six chloride ions on the reactant side because the coefficient of 3 is multiplied by the subscript of 2 on the copper (II) chloride formula. The spectator ions, $\mathrm{K^+}$ and $\mathrm{Cl^-}$, can be eliminated.

Net ionic equation:

$$3[\blacksquare]\,(aq) + 2{\mathrm{PO_4}}^{3-}\,(aq) \rightarrow \mathrm{Cu_3(PO_4)_2}\,(s)$$

Step 3: Think about your result.

For a precipitation reaction, the net ionic equation always shows the two ions that come together to form the precipitate. The equation is balanced by mass and charge.

Summary

- A spectator ion is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction.
- The net ionic equation is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction.
- An example of writing a net ionic equation is outlined.

This page titled 13.16: Net Ionic Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 16.18: Net Ionic Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14: ACIDS AND BASES

14.1: Properties of Acids

14.2: Properties of Bases

14.3: Arrhenius Acids

14.4: Arrhenius Bases

14.5: Brønsted-Lowry Acids and Bases

14.6: Brønsted-Lowry Acid-Base Reactions

14.7: Ion-Product of Water

14.8: The pH Scale

14.9: Calculating pH of Acids and Bases

14.10: The pOH Concept

14.11: Strong and Weak Acids and Acid Ionization Constant (K_a)

14.12: Strong and Weak Bases and Base Ionization Constant

14.13: Calculating Acid and Base Dissociation Constants

14.14: Calculating pH of Weak Acid and Base Solutions

14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

14.16: Titration Experiment

14.17: Titration Calculations

14.18: Titration Curves

14.19: Indicators

14.20: Hydrolysis of Salts - Equations

14.21: Calculating pH of Salt Solutions

Template:HideTOC

This page titled 14: Acids and Bases is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21: Acids and Bases by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.1: Properties of Acids

A cup of coffee first thing in the morning can help start the day. However, 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. These properties are:

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:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g)$$

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.

Summary

This page titled 14.1: Properties of Acids is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.1: Properties of Acids by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.2: Properties of Bases

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.



Figure $14.2.1$***:*** *Phenolphthalein indicator in presence of base.*

Please note that tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words, don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate.

Summary

This page titled 14.2: Properties of Bases is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.2: Properties of Bases by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.3: Arrhenius Acids

Venus is the planet nearest to Earth, but has a very different and hostile environment. It has a surface temperature that averages around $450^{\circ}\mathrm{C}$. The atmosphere is composed of carbon dioxide, but clouds of sulfuric acid move through the upper atmosphere, helping to create the extremely unfriendly conditions. Because of these conditions, Venus is not a place you want to visit on vacation.

Arrhenius Acids

Swedish chemist Svante Arrhenius (1859-1927) was the first to propose a theory to explain the observed behavior of acids and bases. Because of their ability to conduct a current, he knew that both acids and bases contained ions in solution. An **Arrhenius acid** is a compound that ionizes to yield hydrogen ions (H^{+}) in aqueous solution.

Acids are molecular compounds with ionizable hydrogen atoms. Only hydrogen atoms that are part of a highly polar covalent bond are ionizable. Hydrogen chloride (HCl) is a gas at room temperature and under normal pressure. The $\mathrm{H}-\mathrm{Cl}$ bond in hydrogen chloride is a polar bond. The hydrogen atom is electron deficient because of the higher electronegativity of the chlorine atom. Consequently, the hydrogen atom is attracted to the lone pair of electrons in a water molecule when HCl is dissolved in water. The result is that the $\mathrm{H}-\mathrm{Cl}$ bond breaks, with both bonding electrons remaining with the Cl, forming a chloride ion. The H^{+} ion attaches to the water molecule, forming a polyatomic ion called the hydronium ion. The **hydronium ion** $(\mathrm{H_3O^{+}})$ can be thought of as a water molecule with an extra attached hydrogen ion.

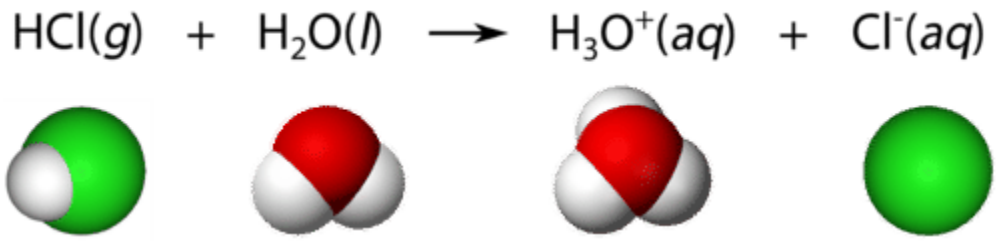


Figure $14.3.1$: Formation of a hydronium ion. (CC BY-NC; CK-12)

Equations showing the ionization of an acid in water are frequently simplified by omitting the water molecule:

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

This is merely a simplification of the previous equation, but it is commonly used. Any hydrogen ions in an aqueous solution will be attached to water molecules as hydronium ions.

Not all hydrogen atoms in molecular compounds are ionizable. In methane $(\mathrm{CH_4})$, the hydrogen atoms are covalently bonded to carbon in bonds that are only slightly polar. The hydrogen atoms are not capable of ionizing, and methane has no acidic properties. Acetic acid $(\mathrm{CH_3COOH})$ belongs to a class of acids called organic acids. There are four hydrogen atoms in the molecule, but only the one hydrogen that is bonded to an oxygen atom is ionizable.

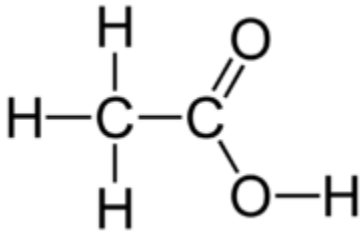


Figure $14.3.2$: The $\mathrm{O}-\mathrm{H}$ bond can be ionized to yield the H^{+} ion and the acetate ion. The other hydrogen atoms in this molecule are not acidic. (CC BY-NC; CK-12)

The table below lists some of the more common acids:

Common Acids

Table 14.3.1: Common Acids	
Acid Name	**Formula**
Hydrochloric acid	HCl
Nitric acid	$\mathrm{HNO_3}$
Sulfuric acid	$\mathrm{H_2SO_4}$
Phosphoric acid	$\mathrm{H_3PO_4}$
Acetic acid	$\mathrm{CH_3COOH}$
Hypochlorous acid	HClO

A **monoprotic acid** is an acid that contains only one ionizable hydrogen. Hydrochloric acid and acetic acid are monoprotic acids. A **polyprotic acid** is an acid that contains multiple ionizable hydrogens. Most common polyprotic acids are either diprotic (such as $\mathrm{H_2SO_4}$) or triprotic (such as $\mathrm{H_3PO_4}$).

Summary

- An Arrhenius acid is a compound that ionizes to yield hydrogen ions $(\mathrm{H^+})$ in aqueous solution.
- Examples of Arrhenius acids are given.

This page titled 14.3: Arrhenius Acids is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.3: Arrhenius Acids by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.4: Arrhenius Bases

Sodium hydroxide is a versatile chemical. It can be used for such mundane purposes as cleaning clogged drains. Several commercial preparations contain sodium hydroxide for this purpose. It also has a number of applications in the food processing field. Ice cream is thickened using NaOH. If olives are soaked in a solution containing sodium hydroxide and other chemicals, the olives will turn black. Soft pretzels are made chewy by the application of sodium hydroxide to the food. This compound has been widely used in the synthesis of plastics, for etching aluminum, for paint removal, and is employed in the dehorning of cattle.

Arrhenius Bases

An **Arrhenius base** is a compound that ionizes to yield hydroxide ions (OH^-) in aqueous solution. The table below lists several of the more common bases.

Table 14.4.1*: Common Bases*

Base Name	Formula
Sodium hydroxide	NaOH
Potassium hydroxide	KOH
Magnesium hydroxide	$\mathrm{Mg(OH)_2}$
Calcium hydroxide	$\mathrm{Ca(OH)_2}$

All of the bases listed in the table are solids at room temperature. Upon dissolving in water, each dissociates into a metal cation and the hydroxide ion.

$$\mathrm{NaOH}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

Sodium hydroxide is a very caustic substance also known as lye. Lye is used as a rigorous cleaner and is an ingredient in the manufacture of soaps. Care must be taken with strong bases like sodium hydroxide, as exposure can lead to severe burns (see figure below).



Figure 14.4.1: This foot has severe burns due to prolonged contact with a solution of sodium hydroxide, also known as lye. (CC BY-NC; CK-12)

Sodium belongs to the group of elements called the alkali metals. An **alkaline solution** is another name for a solution that is basic. All alkali metals react readily with water to produce the metal hydroxide and hydrogen gas. The resulting solutions are basic.

$$2\mathrm{K}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{KOH}\,(aq) + \mathrm{H_2}\,(g)$$

Bases that consist of an alkali metal cation and the hydroxide anion are all very soluble in water. Compounds of the Group 2 metals (the alkaline earth metals) are also basic. However, these compounds are generally not as soluble in water. Therefore, the dissociation reactions for these compounds are shown as equilibrium reactions:

$$\mathrm{Mg(OH)_2}\,(s) \overset{\mathrm{H_2O}}{\rightleftharpoons} \mathrm{Mg^{2+}}\,(aq) + 2\mathrm{OH^-}\,(aq)$$

These relatively insoluble hydroxides were some of the compounds discussed in the context of the solubility product constant (K_{sp}). The solubility of magnesium hydroxide is $0.0084\ \mathrm{g}$ per liter of water at $25^{\circ}\mathrm{C}$. Because of its low solubility, magnesium hydroxide is not as dangerous as sodium hydroxide. In fact, magnesium hydroxide is the active ingredient in a product called milk of magnesia, which is used as an antacid or a mild laxative.

Summary

- Arrhenius base is a compound that ionizes to yield hydroxide ions $(\mathrm{OH^-})$ in aqueous solution.
- Examples of Arrhenius bases are given.

This page titled 14.4: Arrhenius Bases is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.4: Arrhenius Bases by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.5: Brønsted-Lowry Acids and Bases

The Arrhenius concept of acids and bases was a significant contribution to the scientific understanding of acids and bases. It replaced and expanded Lavoisier's original idea that all acids contained oxygen. However, the Arrhenius theory also had its shortcomings. It did not take into account the role of the solvent. In addition, this concept did not deal with acid-base behavior in solvents such as benzene, where there could be no ionization. So, formation of a new theory was imperative, which built on the findings of Arrhenius, but also went beyond them.

Brønsted-Lowry Acids and Bases

The Arrhenius definition of acids and bases is somewhat limited. There are some compounds whose properties suggest that they are either acidic or basic, but which do not qualify according to the Arrhenius definition. An example is ammonia $(\mathrm{NH_3})$. Its aqueous solution turns litmus blue, it reacts with acids, and displays all the other properties of a base. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by Danish chemist Johannes Brønsted (1879-1947) and English chemist Thomas Lowry (1874-1936). A **Brønsted-Lowry acid** is a molecule or ion that donates a hydrogen ion in a reaction. A **Brønsted-Lowry base** is a molecule or ion that accepts a hydrogen ion in a reaction. A hydrogen ion is commonly referred to as a proton, and so acids and bases are proton donors and proton acceptors, respectively, according to the Brønsted-Lowry definition. All substances that are categorized as acids and bases under the Arrhenius definition are also defined as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

What kind of molecule would qualify as a Brønsted-Lowry base? These molecules need to be able to accept a hydrogen ion (or proton). Two possibilities come to mind: an anion that can form a neutral compound with a proton, or a molecule in which one or more atoms has lone-pair electrons. The most obvious anion is the Arrhenius base $\mathrm{OH^-}$. This ion can form a water molecule with a proton by accepting the proton. The acetate anion $\mathrm{CH_3COO^-}$ is another anion that can combine with a proton to form acetic acid. Lone-pair electron groups would include the nitrogen atom (see figure below). The two electrons at the top of the nitrogen atom are not connected in any type of bond, but they do interact readily with a bare proton.

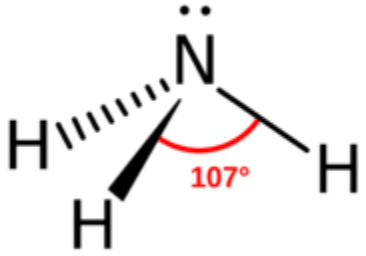


Figure $14.5.1$: *Lone pair electrons on nitrogen.* (CC BY-NC; CK-12)

Oxygen is another atom with lone pair electrons that can function as a Brønsted-Lowry base.



Figure $14.5.2$: Electrons in an oxygen atom. (CC BY-NC; CK-12)

The two single electrons (to the left and bottom of the atom) can form single covalent bonds with other atoms, while the two pairs of double electrons (top and right) are available to interact with a hydrogen ion.

Summary

- A Brønsted-Lowry acid is a molecule or ion that donates a hydrogen ion in a reaction.
- A Brønsted-Lowry base is a molecule or ion that accepts a hydrogen ion in a reaction.

This page titled 14.5: Brønsted-Lowry Acids and Bases is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.5: Brønsted-Lowry Acids and Bases by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.6: Brønsted-Lowry Acid-Base Reactions

The Roman god Janus was considered the god of gates, doors, beginnings, and endings. He is portrayed with two faces, looking in two directions at once. Janus would have been in a good position to look at the acid-base reactions we see in this concept, since they are equilibrium reactions involving two different forms of both acids and bases.

Brønsted-Lowry Acid-Base Reactions

An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another. When ammonia is dissolved in water, it undergoes the following reversible reaction.

$$\underset{\text{base}}{\mathrm{NH_3}\,(aq)} + \underset{\text{acid}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{NH_4^+}\,(aq)} + \underset{\text{base}}{\mathrm{OH^-}\,(aq)}$$

In this reaction, the water molecule is donating a proton to the ammonia molecule. The resulting products are the ammonium ion and the hydroxide ion. The water is acting as a Brønsted-Lowry acid, while the ammonia is acting as a Brønsted-Lowry base. The hydroxide ion that is produced causes the solution to be basic.

We can also consider the reverse reaction in the above equation. In that reaction, the ammonia ion donates a proton to the hydroxide ion. The ammonium ion is a Brønsted-Lowry acid, while the hydroxide ion is a Brønsted-Lowry base. Most Brønsted-Lowry acid-base reactions can be analyzed in this way. One acid and one base are reactants, and one acid and one base are products.

In the above reaction, water acted as an acid, which may seem a bit unexpected. Water can also act as a base in a Brønsted-Lowry acid-base reaction, as long as it reacts with a substance that is a better proton donor. Shown below is the reaction of water with the hydrogen sulfate ion.

$$\underset{\text{acid}}{\mathrm{HSO_4^-}\,(aq)} + \underset{\text{base}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{H_3O^+}\,(aq)} + \underset{\text{base}}{\mathrm{SO_4^{2-}}\,(aq)}$$

So, water is capable of being either an acid or a base, a characteristic called amphoterism. An **amphoteric substance** is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.

Conjugate Acids and Bases

When a substance that is acting as a Brønsted-Lowry acid donates its proton, it becomes a base in the reverse reaction. In the reaction above, the hydrogen sulfate ion $(\mathrm{HSO_4^-})$ donates a proton to water and becomes a sulfate ion $(\mathrm{SO_4^{2-}})$. The $\mathrm{HSO_4^-}$ and the $\mathrm{SO_4^{2-}}$ are linked to one another by the presence or absence of the $\mathrm{H^+}$ ion. A **conjugate acid-base pair** is a pair of substances related by the loss or gain of a single hydrogen ion. A **conjugate acid** is the particle produced when a base accepts a proton. The hydrogen sulfate ion is the conjugate acid of the sulfate ion. A **conjugate base** is the particle produced when an acid donates a proton. The sulfate ion is the conjugate base of the hydrogen sulfate ion.

In the reaction illustrated below, water serves both as acid and base simultaneously. One water molecule serves as an acid and donates a proton. The other water molecule functions as a base by accepting the proton.

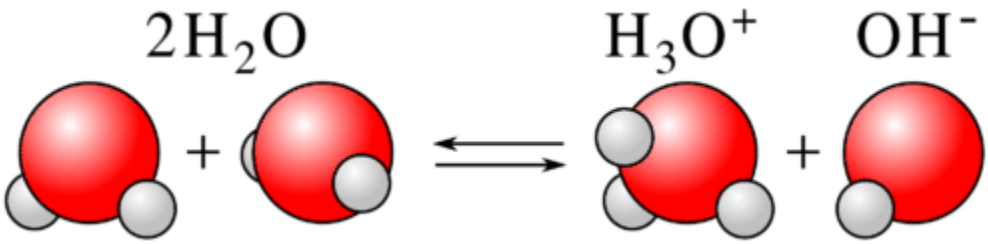


Figure 14.6.1: Water molecules as a conjugate acid-base pair. (CC BY-NC; CK-12)

A typical Brønsted-Lowry acid-base reaction contains two conjugate acid-base pairs, as shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{PO_4}^{3-}\,(aq) \rightleftharpoons \mathrm{NO_2}^{-}\,(aq) + \mathrm{HPO_4}^{2-}\,(aq)$$

One conjugate acid-base pair is $\mathrm{NHO_2/NO_2}^{-}$, while the other pair is $\mathrm{HPO_4}^{2-}/\mathrm{PO_4}^{3-}$.

Summary

- An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another.
- A conjugate acid-base pair is a pair of substances related by the loss or gain of a single hydrogen ion.
- A conjugate acid is the particle produced when a base accepts a proton.
- A conjugate base is the particle produced when an acid donates a proton.
- Examples of conjugate acid-base pairs are given.

This page titled 14.6: Brønsted-Lowry Acid-Base Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.6: Brønsted-Lowry Acid-Base Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.7: Ion-Product of Water

At one time, you could take the little caps off the top of a car battery and check the condition of the sulfuric acid inside. If it got low, you could add more acid. But, sulfuric acid is hazardous stuff, so the batteries are now sealed to protect people. Because of the acid's dangerous nature, it is not a good idea to cut into a battery to see what it looks like—you could get acid burns.

The Ion-Product of Water

The **self-ionization of water** (the process in which water ionizes to hydronium ions and hydroxide ions) occurs to a very limited extent. When two molecules of water collide, there can be a transfer of a hydrogen ion from one molecule to the other. The products are a positively charged hydronium ion and a negatively charged hydroxide ion.

$$\mathrm{H_2O}\,(l) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

We often use the simplified form of the reaction:

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The equilibrium constant for the self-ionization of water is referred to as the ion-product for water and is given the symbol K_w.

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}]$$

The **ion-product of water** (K_w) is the mathematical product of the concentration of hydrogen ions and hydroxide ions. Note that $\mathrm{H_2O}$ is not included in the ion-product expression because it is a pure liquid. The value of K_w is very small, in accordance with a reaction that favors the reactants. At $25^\circ\mathrm{C}$, the experimentally determined value of K_w in pure water is 1.0×10^{-14}.

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

In pure water, the concentrations of hydrogen and hydroxide ions are equal to one another. Pure water or any other aqueous solution in which this ratio holds is said to be neutral. To find the molarity of each ion, the square root of K_w is taken.

$$[\mathrm{H_+}] = [\mathrm{OH^-}] = 1.0 \times 10^{-7}$$

An **acidic solution** is a solution in which the concentration of hydrogen ions is greater than the concentration of hydroxide ions. For example, hydrogen chloride ionizes to produce $\mathrm{H^+}$ and $\mathrm{Cl^-}$ ions upon dissolving in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

This increases the concentration of $\mathrm{H^+}$ ions in the solution. According to Le Chatelier's principle, the equilibrium represented by $\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$ is forced to the left, towards the reactant. As a result, the concentration of the hydroxide ion decreases.

A **basic solution** is a solution in which the concentration of hydroxide ions is greater than the concentration of hydrogen ions. Solid potassium hydroxide dissociates in water to yield potassium ions and hydroxide ions.

$$\mathrm{KOH}\,(s) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The increase in concentration of the $\mathrm{OH^-}$ ions causes a decrease in the concentration of the $\mathrm{H^+}$ ions and the ion-product of $[\mathrm{H^+}]\,[\mathrm{OH^-}]$ remains constant.

Example 14.7.1

Hydrochloric acid (HCl) is a strong acid, meaning it is 100% ionized in solution. What is the $[\mathrm{H^+}]$ and the $[\mathrm{OH^-}]$ in a solution of $2.0 \times 10^{-3}\ \mathrm{M}\ \mathrm{HCl}$?

Solution

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

Known

- $[\mathrm{HCl}] = 2.0 \times 10^{-3}\ \mathrm{M}$
- $K_\mathrm{w} = 1.0 \times 10^{-14}$

Unknown

- $[\mathrm{H^+}] = ?\ \mathrm{M}$
- $[\mathrm{OH^-}] = ?\ \mathrm{M}$

Because HCl is 100% ionized, the concentration of $\mathrm{H^+}$ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one $\mathrm{H^+}$ ion and one $\mathrm{Cl^-}$ ion. The concentration of $\mathrm{OH^-}$ can then be determined from the $[\mathrm{H^+}]$ and K_w.

Step 2: Solve.

$$[\mathrm{H^+}] = 2.0 \times 10^{-3}\,\mathrm{M}$$

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

$$[\mathrm{OH^-}] = \frac{K_\mathrm{w}}{[\mathrm{H^+}]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12}\,\mathrm{M}$$

Step 3: Think about your result.

The $[\mathrm{H^+}]$ is much higher than the $[\mathrm{OH^-}]$ because the solution is acidic. As with other equilibrium constants, the unit for K_w is customarily omitted.

Summary

- The self-ionization of water is described and an ionization constant for the process is stated.
- Acidic and basic solutions are defined.
- Calculations using K_w are illustrated.

This page titled 14.7: Ion-Product of Water is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.8: Ion-Product of Water by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.8: The pH Scale

Grapefruit juice has a pH of somewhere between 2.9 and 3.3, depending on the specific product. Excessive exposure to this juice can cause erosion of tooth enamel and can lead to tooth damage. The acids in grapefruit juice are carbon-based, with citric acid being one of the major constituents. This compound has three ionizable hydrogens on each molecule which contribute to the relatively low pH of the juice. Another component of grapefruit juice is malic acid, containing two ionizable hydrogens per molecule.

The pH Scale

Expressing the acidity of a solution by using the molarity of the hydrogen ion is cumbersome because the quantities are generally very small. Danish scientist Søren Sørensen (1868-1939) proposed an easier system for indicating the concentration of H^+ called the pH scale. The letters pH stand for the power of the hydrogen ion. The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration:

$$\mathrm{pH} = -\log\ [\mathrm{H}^+]$$

In pure water or a neutral solution, the $[\mathrm{H}^+] = 1.0 \times 10^{-7}\ \mathrm{M}$. Substituting into the pH expression:

$$\mathrm{pH} = -\log\left[1.0 \times 10^{-7}\right] = -\left(-7.00\right) = 7.00$$

The pH of pure water or any neutral solution is thus 7.00. For recording purposes, the numbers to the right of the decimal point in the pH value are the significant figures. Since 1.0×10^{-7} has two significant figures, the pH can be reported as 7.00.

A logarithmic scale condenses the range of acidity to numbers that are easy to use. Consider a solution with $[\mathrm{H}^+] = 1.0 \times 10^{-4}\ \mathrm{M}$. That is a hydrogen-ion concentration that is 1000 times higher than the concentration in pure water. The pH of such a solution is 4.00, a difference of just 3 pH units. Notice that when the $[\mathrm{H}^+]$ is written in scientific notation and the coefficient is 1, the pH is simply the exponent with the sign changed. The pH of a solution with $[\mathrm{H}^+] = 1 \times 10^{-2}\ \mathrm{M}$ is 2 and the pH of a solution with $[\mathrm{H}^+] = 1 \times 10^{-10}\ \mathrm{M}$ is 10.

As we saw earlier, a solution with $[\mathrm{H}^+]$ higher than 1.0×10^{-7} is acidic, while a solution with $[\mathrm{H}^+]$ lower than 1.0×10^{-7} is basic. Consequently, solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic. Figure 14.8.1 illustrates this relationship, along with some examples of various solutions.

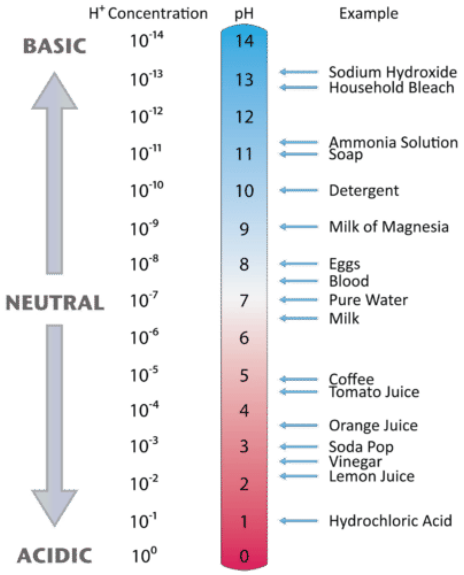


Figure 14.8.1: The pH values for several common materials. (CC BY-NC; CK-12)

Summary

- The pH of a solution is the negative logarithm of the hydrogen-ion concentration.
- Solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic.
- pH values for several common materials are listed.

This page titled 14.8: The pH Scale is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.9: The pH Scale by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.9: Calculating pH of Acids and Bases

It's not uncommon to see a tropical fish tank in homes or businesses. These brightly-colored creatures are relaxing to watch, but do require a certain amount of maintenance in order for them to survive. Tap water is usually too alkaline when it comes out of the faucet, so some adjustments need to be made. The pH of the water will change over time while it is in the tank, which means it needs to be tested every so often. Any fish tank caretaker has a chance to be a chemist for their fish!

Calculating pH of Acids and Bases

Calculation of pH is simple when there is a $1 \times 10^{\text{power}}$ problem. However, in real life, this is rarely the situation. If the coefficient is not equal to 1, a calculator must be used to find the pH. For example, the pH of a solution with $[\text{H}^+] = 2.3 \times 10^{-5}\ \text{M}$ can be found as shown below.

$$\text{pH} = -\log\left[2.3 \times 10^{-5}\right] = 4.64$$

When the pH of a solution is known, the concentration of the hydrogen ion can be calculated. The inverse of the logarithm (or antilog) is the 10^x key on a calculator.

$$[\text{H}^+] = 10^{-\text{pH}}$$

For example, suppose that you have a solution with a pH of 9.14. To find the $[\text{H}^+]$ use the 10^x key.

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-9.14} = 7.24 \times 10^{-10}\ \text{M}$$

Hydroxide Ion Concentration and pH

As we saw earlier, the hydroxide ion concentration of any aqueous solution is related to the hydrogen ion concentration through the value of K_w. We can use that relationship to calculate the pH of a solution of a base.

Example 14.9.1

Sodium hydroxide is a strong base. Find the pH of a solution prepared by dissolving $1.0\ \text{g}$ of NaOH into enough water to make $1.0\ \text{L}$ of solution.

Solution

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

Known

- Mass $\text{NaOH} = 1.0\ \text{g}$
- Molar mass $\text{NaOH} = 40.00\ \text{g/mol}$
- Volume solution $= 1.0\ \text{L}$
- $K_\text{w} = 1.0 \times 10^{-14}$

Unknown

First, convert the mass of NaOH to moles. Second, calculate the molarity of the NaOH solution. Because NaOH is a strong base and is soluble, the $[\text{OH}^-]$ will be equal to the concentration of the NaOH. Third, use K_w to calculate the $[\text{H}^+]$ in the solution. Lastly, calculate the pH.

Step 2: Solve.

$$1.00 \ \cancel{\text{g NaOH}} \times \frac{1 \text{ mol NaOH}}{40.00 \ \cancel{\text{g NaOH}}} = 0.025 \text{ mol NaOH}$$

$$\text{Molarity} = \frac{0.025 \text{ mol NaOH}}{1.00 \text{ L}} = 0.025 \text{ M NaOH} = 0.025 \text{ M OH}^-$$

$$[\text{H}^+] = \frac{K_\text{w}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.025 \text{ M}} = 4.0 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log [\text{H}^+] = -\log \left(4.0 \times 10^{-13}\right) = 12.40$$

Step 3: Think about your result.

The solution is basic and so its pH is greater than 7. The reported pH is rounded to two decimal places because the original mass and volume has two significant figures.

Summary

This page titled 14.9: Calculating pH of Acids and Bases is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.10: Calculating pH of Acids and Bases by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.10: The pOH Concept

Soap Lake has a long history as a healing place. Indian tribes would put aside their rivalries when they came to the lake to enjoy the high mineral content of the water. In the days before good antibiotics, thousands of visitors would come and enjoy the soothing waters of the lake. Soap Lake is alkaline, with water quality though to be similar to that of the moons of the planet Jupiter.

The pOH Concept

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:

$$\text{pOH} = -\log\left[\text{OH}^-\right]$$

The pH of a solution can be related to the pOH. Consider a solution with a pH $= 4.0$. The $[\text{H}^+]$ of the solution would be 1.0×10^{-4} M. Dividing K_w by this yields a $[\text{OH}^-]$ of 1.0×10^{-10} M. Finally the pOH of the solution equals $-\log\left(1.0 \times 10^{-10}\right) = 10$. This example illustrates the following relationship.

$$\text{pH} + \text{pOH} = 14$$

The pOH scale is similar to the pH scale in that a pOH of 7 is indicative of a neutral solution. A basic solution has a pOH less than 7, while an acidic solution has a pOH of greater than 7. The pOH is convenient to use when finding the hydroxide ion concentration from a solution with a known pH.

Example 14.10.1

Find the hydroxide concentration of a solution with a pH of 4.42.

Solution

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

Known

- pH $= 4.42$
- pH + pOH $= 14$

Unknown

First, the pOH is calculated, followed by the $[\text{OH}^-]$.

Step 2: Solve.

$$\text{pOH} = 14 - \text{pH} = 14 - 4.42 = 9.58$$
$$\left[\text{OH}^-\right] = 10^{-\text{pOH}} = 10^{-9.58} = 2.6 \times 10^{-10}\ \text{M}$$

Step 3: Think about your result.

The pH is that of an acidic solution, and the resulting hydroxide-ion concentration is less than 1×10^{-7} M. The answer has two significant figures because the given pH has two decimal places.

The diagram below shows all of the interrelationships between $[\text{H}^+]$, $[\text{OH}^-]$, pH, and pOH.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

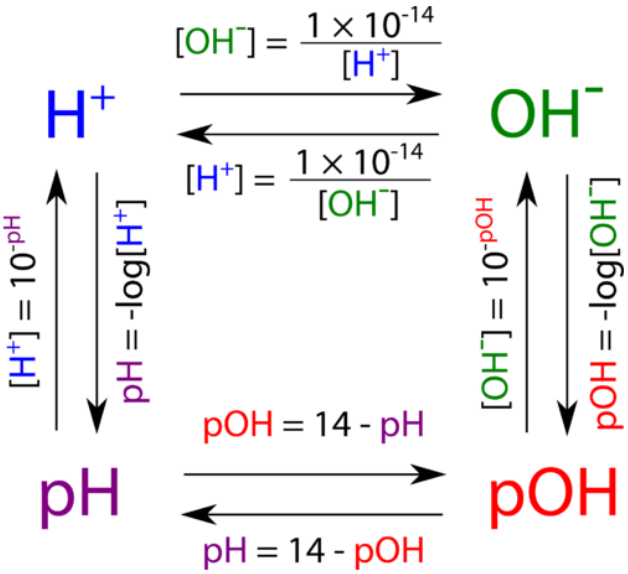


Figure 14.10.1: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH. (CC BY-NC; CK-12)

Summary

- The pOH of a solution is the negative logarithm of the hydroxide-ion concentration.
- Calculations involving pOH are described.

This page titled 14.10: The pOH Concept is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.11: The pOH Concept by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.11: Strong and Weak Acids and Acid Ionization Constant (K_{a})

The etching of glass is a slow process that can produce beautiful artwork. Traditionally, the glass has been treated with dilute hydrofluoric acid which gradually dissolves the glass under it. Parts of the piece that should not be etched are covered with wax or some other non-reactive material. In more recent times, compounds such as ammonium bifluoride have been used. Whichever chemical is employed, the artist must be very careful not to get any on their skin.

Strong and Weak Acids and Acid Ionization Constant

Acids are classified as either strong or weak, based on their ionization in water. A **strong acid** is an acid which is completely ionized in an aqueous solution. Hydrogen chloride (HCl) ionizes completely into hydrogen ions and chloride ions in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

A **weak acid** is an acid that ionizes only slightly in an aqueous solution. Acetic acid (found in vinegar) is a very common weak acid. Its ionization is shown below.

$$\mathrm{CH_3COOH}\,(aq) \rightleftharpoons \mathrm{H}^{+}\,(aq) + \mathrm{CH_3COO}^{-}\,(aq)$$

The ionization of acetic acid is incomplete, and so the equation is shown with a double arrow. The extent of ionization of weak acids varies, but is generally less than 10%. A $0.10\ \mathrm{M}$ solution of acetic acid is only about 1.3% ionized, meaning that the equilibrium strongly favors the reactants.

Weak acids, like strong acids, ionize to yield the H^{+} ion and a conjugate base. Because HCl is a strong acid, its conjugate base (Cl^{-}) is extremely weak. The chloride ion is incapable of accepting the H^{+} ion and becoming HCl again. In general, the stronger the acid, the weaker its conjugate base. Likewise, the weaker the acid, the stronger its conjugate base.

Table $14.11.1$*: Relative Strengths of Acids and their Conjugate Bases*

Acid	Conjugate Base
Strong Acids	
HCl (hydrochloric acid) (strongest)	Cl^{-} (chloride ion) (weakest)
$\mathrm{H_2SO_4}$ (sulfuric acid)	$\mathrm{HSO_4}^{-}$ (hydrogen sulfate ion)
$\mathrm{HNO_3}$ (nitric acid)	$\mathrm{NO_3}^{-}$ (nitrate ion)
Weak Acids	
$\mathrm{H_3PO_4}$ (phosphoric acid)	$\mathrm{H_2PO_4}^{-}$ (dihydrogen phosphate ion)
$\mathrm{CH_3COOH}$ (acetic acid)	$\mathrm{CH_3COO}^{-}$ (acetate ion)
$\mathrm{H_2CO_3}$ (carbonic acid)	$\mathrm{HCO_3}^{-}$ (hydrogen carbonate ion)
HCN (hydrocyanic acid) (weakest)	CN^{-} (cyanide ion) (strongest)

Strong acids are 100% ionized in solution. Weak acids are only slightly ionized. Phosphoric acid is stronger than acetic acid, and so is ionized to a greater extent. Acetic acid is stronger than carbonic acid, and so on.

The Acid Ionization Constant, K_{a}

The ionization for a general weak acid, HA, can be written as follows:

$$\mathrm{HA}\,(aq) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{A}^{-}\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_a = \frac{[H^+]\,[A^-]}{[HA]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution. Therefore, the numerical value of K_a is a reflection of the strength of the acid. Weak acids with relatively higher K_a values are stronger than acids with relatively lower K_a values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the K_a value approaches infinity. For this reason, K_a values are generally reported for weak acids only.

The table below is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table $14.11.2$*: Acid Ionization Constants at* $25°C$

Name of Acid	Ionization Equation	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$ $HSO_4 \rightleftharpoons H^+ + SO_4^{2-}$	very large 1.3×10^{-2}
Oxalic acid	$H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$ $HC_2O_4 \rightleftharpoons H^+ + C_2O_4^{2-}$	6.5×10^{-2} 6.1×10^{-5}
Phosphoric acid	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ $H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2--}$ $HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$	7.5×10^{-3} 6.2×10^{-8} 4.8×10^{-13}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.2×10^{-7} 4.8×10^{-11}
Hydrocyanic acid	$HCN \rightleftharpoons H^+ + CN^-$	4.9×10^{-10}

Summary

- A strong acid is an acid which is completely ionized in an aqueous solution.
- A weak acid is an acid that ionizes only slightly in an aqueous solution.
- The acid ionization constant (K_a) is defined.

This page titled 14.11: Strong and Weak Acids and Acid Ionization Constant \(\left(K_\text{a} \right)\) is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.12: Strong and Weak Acids and Acid Ionization Constant \(\left(K_\text{a} \right)\) by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.12: Strong and Weak Bases and Base Ionization Constant

All the complex electronics and apparatuses in a space shuttle generate heat, as do the astronauts. The shuttles have a complex arrangement of systems to dissipate that heat into outer space. One of the components of this system is a series of coils filled with ammonia that are located on the outside of the shuttle. Ammonia absorbs the heat and then releases it into space as the gas circulates through the coils. This approach is both inexpensive and effective.

Strong and Weak Bases and Base Ionization Constant, K_{b}

As with acids, bases can either be strong or weak, depending on the extent of their ionization. A **strong base** is a base that ionizes completely in an aqueous solution. The most common strong bases are soluble metal hydroxide compounds such as potassium hydroxide. Some metal hydroxides are not as strong, simply because they are not as soluble. Calcium hydroxide is only slightly soluble in water, but the portion that does dissolve also dissociates into ions.

A **weak base** is a base that ionizes only slightly in an aqueous solution. Recall that a base can be defined as a substance that accepts a hydrogen ion from another substance. When a weak base such as ammonia is dissolved in water, it accepts an H^+ ion from water, forming the hydroxide ion and the conjugate acid of the base, the ammonium ion.

$$\mathrm{NH_3}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{NH_4}^+\,(aq) + \mathrm{OH}^-\,(aq)$$

The equilibrium greatly favors the reactants and the extent of ionization of the ammonia molecule is very small.

An equilibrium expression can be written for the reactions of weak bases with water. Because the concentration of water is extremely large and virtually constant, the water is not included in the expression. A **base ionization constant** (K_{b}) is the equilibrium constant for the ionization of a base. For ammonia, the expression is:

$$K_{\mathrm{b}} = \frac{[\mathrm{NH_4}^+]\,[\mathrm{OH}^-]}{[\mathrm{NH_3}]}$$

The numerical value of K_{b} is a reflection of the strength of the base. Weak bases with relatively high K_{b} values are stronger than bases with relatively low K_{b} values. The table below is a listing of base ionization constants for several weak bases.

Table 14.12.1*: Base Ionization Constants at* $25^{\circ}\mathrm{C}$

Name of Base	Ionization Equation	K_{b}
Methylamine	$\mathrm{CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3}^+ + \mathrm{OH}^-$	5.6×10^{-4}
Ammonia	$\mathrm{NH_3 + H_2O \rightleftharpoons NH_4}^+ + \mathrm{OH}^-$	1.8×10^{-5}
Pyridine	$\mathrm{C_5H_5N + H_2O \rightleftharpoons C_5H_5NH}^+ + \mathrm{OH}^-$	1.7×10^{-9}
Acetate ion	$\mathrm{CH_3COO}^- + \mathrm{H_2O \rightleftharpoons CH_3COOH} + \mathrm{OH}^-$	5.6×10^{-10}
Fluoride ion	$\mathrm{F}^- + \mathrm{H_2O \rightleftharpoons HF} + \mathrm{OH}^-$	1.4×10^{-11}
Urea	$\mathrm{H_2NCONH_2 + H_2O \rightleftharpoons H_2NCONH_3}^+ + \mathrm{OH}^-$	1.5×10^{-14}

Notice that the conjugate base of a weak acid is also a strong base. For example, the acetate ion has a small tendency to accept a hydrogen ion from water to form acetic acid and the hydroxide ion.

Summary

- A strong base is a base that ionizes completely in an aqueous solution.
- A weak base is a base that ionizes only slightly in an aqueous solution.
- A base ionization constant (K_{b}) is the equilibrium constant for the ionization of a base.
- The conjugate base of a weak acid is also a strong base.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

This page titled 14.12: Strong and Weak Bases and Base Ionization Constant is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.13: Strong and Weak Bases and Base Ionization Constant by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.13: Calculating Acid and Base Dissociation Constants

The pH meter was invented because Florida orange growers needed a way to test the acidity of their fruit. The first meter was invented by Arnold Beckman, who went on to form Beckman Instruments. Beckman's business was very successful, and he used much of his fortune to fund science education and research. The Beckman family donated $40 million to build the Beckman Institute at the University of Illinois.

Calculating K_{a} and K_{b}

The numerical value of K_{a} and K_{b} can be determined from an experiment. A solution of known concentration is prepared and its pH is measured with an instrument called a **pH meter**.



Figure 14.13.1: A pH meter is a laboratory device that provides quick, accurate measurements of the pH of solutions. (CC BY-NC; CK-12)

Example 14.13.1

A $0.500\ \mathrm{M}$ solution of formic acid is prepared and its pH is measured to be 2.04. Determine the K_{a} for formic acid.

Solution

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

Known

- Initial $[\mathrm{HCOOH}] = 0.500\ \mathrm{M}$
- pH $= 2.04$

Unknown

First, the pH is used to calculate the $[\mathrm{H}^+]$ at equilibrium. An ICE table is set up in order to determine the concentrations of HCOOH and HCOO^- at equilibrium. All concentrations are then substituted into the K_{a} expression and the K_{a} value is calculated.

Step 2: Solve.

$$[\mathrm{H}^+] = 10^{-\mathrm{pH}} = 10^{-2.04} = 9.12 \times 10^{-3}\ \mathrm{M}$$

Since each formic acid molecule that ionizes yields one H^+ ion and one formate ion (HCOO^-), the concentrations of H^+ and HCOO^- are equal at equilibrium. We assume that the initial concentrations of each ion are zero, resulting in the following ICE table.

	HCOOH	$\mathrm{H^+}$	$\mathrm{HCOO^-}$
Initial	0.500	0	0
Change	-9.12×10^{-3}	$+9.12 \times 10^{-3}$	$+9.12 \times 10^{-3}$
Equilibrium	0.491	9.12×10^{-3}	9.12×10^{-3}

Now, substituting into the K_a expression gives:

$$K_\mathrm{a} = \frac{[\mathrm{H^+}]\,[\mathrm{HCOO^-}]}{[\mathrm{HCOOH}]} = \frac{\left(9.12 \times 10^{-3}\right)\left(9.12 \times 10^{-3}\right)}{0.491} = 1.7 \times 10^{-4}$$

Step 3: Think about your result.

The value of K_a is consistent with that of a weak acid. Two significant figures are appropriate for the answer, since there are two digits after the decimal point in the reported pH.

Similar steps can be taken to determine the K_b of a base. For example, a $0.750\ \mathrm{M}$ solution of the weak base ethylamine $(\mathrm{C_2H_5NH_2})$ has a pH of 12.31.

$$\mathrm{C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3{}^+ + OH^-}$$

Since one of the products of the ionization reaction is the hydroxide ion, we need to first find the $[\mathrm{OH^-}]$ at equilibrium. The pOH is $14 - 12.31 = 1.69$. The $[\mathrm{OH^-}]$ is then found from $10^{-1.69} = 2.04 \times 10^{-2}\ \mathrm{M}$. The ICE table is then set up as shown below.

	$\mathrm{C_2H_5NH_2}$	$\mathrm{C_2H_5NH_3{}^+}$	$\mathrm{OH^-}$
Initial	0.750	0	0
Change	-2.04×10^{-2}	$+2.04 \times 10^{-2}$	$+2.04 \times 10^{-2}$
Equilibrium	0.730	2.04×10^{-2}	2.04×10^{-2}

Substituting into the K_b expression yields the K_b for ethylamine.

$$K_\mathrm{b} = \frac{[\mathrm{C_2H_5NH_3{}^+}]\,[\mathrm{OH^-}]}{[\mathrm{C_2H_5NH_2}]} = \frac{\left(2.04 \times 10^{-2}\right)\left(2.04 \times 10^{-2}\right)}{0.730} = 5.7 \times 10^{-4}$$

Summary

This page titled 14.13: Calculating Acid and Base Dissociation Constants is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.14: Calculating Acid and Base Dissociation Constants by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.14: Calculating pH of Weak Acid and Base Solutions

Bees are beautiful creatures that help plants flourish. They carry pollen from one plant to another to facilitate plant growth and development. However, they can be troublesome when they sting! For those who are allergic to bee venom, this can be a serious, life-threatening problem. For all other humans, it can be a painful experience. When stung by a bee, one first-aid treatment is to apply a paste of baking soda (sodium bicarbonate) to the stung area. This weak base helps with the itching and swelling that accompanies the bee sting.

Calculating pH of Weak Acid and Base Solutions

The K_a and K_b values have been determined for a great many acids and bases, as shown in Tables 21.12.2 and 21.13.1. These can be used to calculate the pH of any solution of a weak acid or base whose ionization constant is known.

Example 14.14.1

Calculate the pH of a $2.00\ \text{M}$ solution of nitrous acid (HNO_2). The K_a for nitrous acid is 4.5×10^{-4}.

Solution

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

Known

- Initial $[\text{HNO}_2] = 2.00\ \text{M}$
- $K_a = 4.5 \times 10^{-4}$

Unknown

First, an ICE table is set up with the variable x used to signify the change in concentration of the substance due to ionization of the acid. Then the K_a expression is used to solve for x and calculate the pH.

Step 2: Solve.

	HNO_2	H^+	NO_2^-
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00 - x$	x	x

The K_a expression and value are used to set up an equation to solve for x.

$$K_a = 4.5 \times 10^{-4} = \frac{(x)(x)}{2.00 - x} = \frac{x^2}{2.00 - x}$$

The quadratic equation is required to solve this equation for x. However, a simplification can be made of the fact that the extent of ionization of weak acids is small. The value of x will be significantly less than 2.00, so the "$-x$" in the denominator can be dropped.

$$4.5 \times 10^{-4} = \frac{x^2}{2.00 - x} \approx \frac{x^2}{2.00}$$

$$x = \sqrt{4.5 \times 10^{-4}\,(2.00)} = 2.9 \times 10^{-2}\ \text{M} = [\text{H}^+]$$

Since the variable x represents the hydrogen-ion concentration, the pH of the solution can now be calculated.

$$\text{pH} = -\log[\text{H}^+] = -\log\left[2.9 \times 10^{-2}\right] = 1.54$$

Step 3: Think about your result.

> The pH of a $2.00\ \mathrm{M}$ solution of a strong acid would be equal to $-\log(2.00) = -0.30$. The higher pH of the $2.00\ \mathrm{M}$ nitrous acid is consistent with it being a weak acid and therefore not as acidic as a strong acid would be.

The procedure for calculating the pH of a solution of a weak base is similar to that of the weak acid in the example. However, the variable x will represent the concentration of the hydroxide ion. The pH is found by taking the negative logarithm to get the pOH, followed by subtracting from 14 to get the pH.

Summary

- The procedure for calculating the pH of a weak acid or base is illustrated.

This page titled 14.14: Calculating pH of Weak Acid and Base Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.15: Calculating pH of Weak Acid and Base Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

Pouring concrete and working it are messy jobs. In the process, a lot of wastewater with an alkaline pH is generated. Often, regulations require that this wastewater be cleaned up at the site. One practical way to neutralize the basic pH is to bubble $\mathrm{CO_2}$ into the water. The carbon dioxide forms a weak acid (carbonic acid, $\mathrm{H_2CO_3}$) in solution which serves to bring the alkaline pH down to something closer to neutral.

Neutralization Reactions and Net Ionic Equations for Neutralization Reactions

A **neutralization reaction** is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water. The aqueous sodium chloride that is produced in the reaction is called a salt. A **salt** is an ionic compound composed of a cation from a base and an anion from an acid. A salt is essentially any ionic compound that is neither an acid nor a base.

Strong Acid-Strong Base Reactions

When equal amounts of a strong acid such as hydrochloric acid are mixed with a strong base such as sodium hydroxide, the result is a neutral solution. The products of the reaction do not have the characteristics of either an acid or a base. The balanced molecular equation is:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Chemical reactions occurring in aqueous solution are more accurately represented with a net ionic equation. The full ionic equation for the neutralization of hydrochloric acid by sodium hydroxide is written as follows:

$$\mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{H_2O}\,(l)$$

Since the acid and base are both strong, they are fully ionized and so are written as ions, as is the NaCl formed as a product. The sodium and chloride ions are spectator ions in the reaction, leaving the following as the net ionic reaction.

$$\mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l)$$

All neutralization reactions of a strong acid with a strong base simplify to the net ionic reaction of hydrogen ion combining with hydroxide ion to produce water.

What if the acid is a diprotic acid such as sulfuric acid? The balanced molecular equation now involves a 1:2 ratio between acid and base.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + \mathrm{H_2O}\,(l)$$

In order for the reaction to be a full neutralization, twice as many moles of NaOH must react with the $\mathrm{H_2SO_4}$. The sodium sulfate salt is soluble, and so the net ionic reaction is again the same. Different mole ratios occur for other polyprotic acids or bases with multiple hydroxides such as $\mathrm{Ca(OH)_2}$.

Reactions Involving a Weak Acid or Weak Base

Reactions where at least one of the components is weak do not generally result in a neutral solution. The reaction between weak nitrous acid and strong potassium hydroxide is shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{KOH}\,(aq) \rightarrow \mathrm{KNO_2}\,(aq) + \mathrm{H_2O}\,(l)$$

In order to write the net ionic equation, the weak acid must be written as a molecule since it does not ionize to a great extent in water. The base and the salt are fully dissociated.

CK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org

$$\mathrm{HNO_2}\,(aq) + \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{NO_2{}^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The only spectator ion is the potassium ion, resulting in the net ionic equation:

$$\mathrm{HNO_2}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{NO_2{}^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The strong hydroxide ion essentially "forces" the weak nitrous acid to become ionized. The hydrogen ion from the acid combines with the hydroxide ion to form water, leaving the nitrite ion as the other product. The resulting solution is not neutral (pH $= 7$), but instead is slightly basic.

Reactions can also involve a weak base and strong acid, resulting in a solution that is slightly acidic. The molecular and net ionic equations for the reaction of hydrochloric acid and ammonia are shown below.

$$\begin{aligned}&\mathrm{HCl}\,(aq) + \mathrm{NH_3}\,(aq) \rightarrow \mathrm{NH_4Cl}\,(aq)\\&\mathrm{H^+}\,(aq) + \mathrm{NH_3}\,(aq) \rightarrow \mathrm{NH_4{}^+}\,(aq) \qquad (\mathrm{Cl^-\ is\ a\ spectator\ ion})\end{aligned}$$

Reactions between acids and bases that are both weak may result in solutions that are neutral, acidic, or basic.

Summary

- A neutralization reaction is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water.
- Equations for acid-base neutralizations are given.
- Net ionic equations for neutralization reactions are given.

This page titled 14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.16: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.16: Titration Experiment

A large amount of current research involves the development of biodiesel fuels. Often this material can be made from used vegetable oils. The vegetable oil is treated with lye to create the biofuel. In the oils is a variable amount of acid that needs to be determined, so that the workers will know how much lye to add to make the final fuel. Before the lye is added, the native vegetable oil is titrated to find out how much free acid is present. Then, the amount of lye added can be adjusted to take into account the amount needed to neutralize these free acids.

Titration Experiment

In the neutralization of hydrochloric acid by sodium hydroxide, the mole ratio of acid to base is 1:1.

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

One mole of HCl would be fully neutralized by one mole of NaOH. If instead the hydrochloric acid was reacted with barium hydroxide, the mole ratio would be 2:1.

$$2\mathrm{HCl}\,(aq) + \mathrm{Ba(OH)_2}\,(aq) \rightarrow \mathrm{BaCl_2}\,(aq) + 2\mathrm{H_2O}\,(l)$$

Now two moles of HCl would be required to neutralize one mole of $\mathrm{Ba(OH)_2}$. The mole ratio ensures that the number of moles of $\mathrm{H^+}$ ions supplied by the acid is equal to the number of $\mathrm{OH^-}$ ions supplied by the base. This must be the case for neutralization to occur. The **equivalence point** is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

In the laboratory, it is useful to have an experiment where the unknown concentration of an acid or a base can be determined. This can be accomplished by performing a controlled neutralization reaction. A **titration** is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. Many titrations are acid-base neutralization reactions, though other types of titrations can also be performed.

In order to perform an acid-base titration, the chemist must have a way to visually detect that the neutralization reaction has occurred. An **indicator** is a substance that has a distinctly different color when in an acidic or basic solution. A commonly used indicator for strong acid-strong base titrations is phenolphthalein. Solutions in which a few drops of phenolphthalein have been added turn from colorless to brilliant pink as the solution turns from acidic to basic. The steps in a titration reaction are outlined below.

1. A measured volume of an acid of unknown concentration is added to an Erlenmeyer flask.
2. Several drops of an indicator are added to the acid and mixed by swirling the flask.
3. A buret is filled with a base solution of known molarity.
4. The stopcock of the buret is opened and base is slowly added to the acid, while the flask is constantly swirled to ensure mixing. The stopcock is closed at the exact point at which the indicator just changes color.

The **standard solution** is the solution in a titration whose concentration is known. In the titration described above, the base solution is the standard solution. It is very important in a titration to add the solution from the buret slowly, so that the point at which the indicator changes color can be found accurately. The **end point** of a titration is the point at which the indicator changes color. When phenolphthalein is the indicator, the end point will be signified by a faint pink color.

cK-12 LICENSED UNDER CK-12 License
© CK-12 Foundation • Visit us at ck12.org



Figure 14.16.1: Phenolphthalein in basic solution. (CC BY-NC; CK-12)

Summary

- The equivalence point is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.
- A titration is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. (Many titrations are acid-base neutralization reactions.)
- An indicator is a substance that has a distinctly different color when in an acidic or basic solution.
- The process for carrying out a titration is described.

This page titled 14.16: Titration Experiment is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.17: Titration Experiment by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.17: Titration Calculations

The manufacture of soap requires a number of chemistry techniques. One necessary piece of information is the saponification number. This is the amount of base needed to hydrolyze a certain amount of fat to produce the free fatty acids that are an essential part of the final product. The fat is heated with a known amount of base (usually NaOH or KOH). After hydrolysis is complete, the leftover base is titrated to determine how much was needed to hydrolyze the fat sample.

Titration Calculations

At the equivalence point in a neutralization, the moles of acid are equal to the moles of base.

$$\text{moles acid} = \text{moles base}$$

Recall that the molarity (M) of a solution is defined as the moles of the solute divided by the liters of solution (L). So the moles of solute are therefore equal to the molarity of a solution multiplied by the volume in liters.

$$\text{moles solute} = \mathrm{M} \times \mathrm{L}$$

We can then set the moles of acid equal to the moles of base.

$$\mathrm{M}_A \times \mathrm{V}_A = \mathrm{M}_B \times \mathrm{V}_B$$

M_A is the molarity of the acid, while M_B is the molarity of the base. V_A and V_B are the volumes of the acid and base, respectively.

Suppose that a titration is performed and $20.70\ \mathrm{mL}$ of $0.500\ \mathrm{M}\ \mathrm{NaOH}$ is required to reach the end point when titrated against $15.00\ \mathrm{mL}$ of HCl of unknown concentration. The above equation can be used to solve for the molarity of the acid.

$$\mathrm{M}_A = \frac{\mathrm{M}_B \times \mathrm{V}_B}{\mathrm{V}_A} = \frac{0.500\ \mathrm{M} \times 20.70\ \mathrm{mL}}{15.00\ \mathrm{mL}} = 0.690\ \mathrm{M}$$

The higher molarity of the acid compared to the base in this case means that a smaller volume of the acid is required to reach the equivalence point.

The above equation works only for neutralizations in which there is a 1:1 ratio between the acid and the base. The example below demonstrates the technique to solve a titration problem for a titration of sulfuric acid with sodium hydroxide.

Example 14.17.1

In a titration of sulfuric acid against sodium hydroxide, $32.20\ \mathrm{mL}$ of $0.250\ \mathrm{M}\ \mathrm{NaOH}$ is required to neutralize $26.60\ \mathrm{mL}$ of $\mathrm{H_2SO_4}$. Calculate the molarity of the sulfuric acid.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + 2\mathrm{H_2O}\,(l)$$

Solution

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

- Molarity $\mathrm{NaOH} = 0.250\ \mathrm{M}$
- Volume $\mathrm{NaOH} = 32.20\ \mathrm{mL}$
- Volume $\mathrm{H_2SO_4} = 26.60\ \mathrm{mL}$

Unknown

First determine the moles of NaOH in the reaction. From the mole ratio, calculate the moles of $\mathrm{H_2SO_4}$ that reacted. Finally, divide the moles of $\mathrm{H_2SO_4}$ by its volume to get the molarity.

Step 2: Solve.

CK-12 LICENSED UNDER CC BY NC SA
© CK-12 Foundation • Visit us at ck12.org

$$\text{mol NaOH} = \text{M} \times \text{L} = 0.250\ \text{M} \times 0.03220\ \text{L} = 8.05 \times 10^{-3}\ \text{mol NaOH}$$

$$8.05 \times 10^{-3}\ \text{mol NaOH} \times \frac{1\ \text{mol } H_2SO_4}{2\ \text{mol NaOH}} = 4.03 \times 10^{-3}\ \text{mol } H_2SO_4$$

$$\frac{4.03 \times 10^{-3}\ \text{mol } H_2SO_4}{0.02660\ \text{L}} = 0.151\ \text{M } H_2SO_4$$

Step 3: Think about your result.

The volume of H_2SO_4 required is smaller than the volume of $NaOH$ because of the two hydrogen ions contributed by each molecule.

Summary

- The process of calculating concentration from titration data is described and illustrated.

This page titled 14.17: Titration Calculations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.18: Titration Calculations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.18: Titration Curves

The x-y plot that we know of as a graph was the brainchild of the French mathematician-philosopher Rene Descartes (1596-1650). His studies in mathematics led him to develop what was known as "Cartesian geometry", including the concept of today's graphs. The coordinates are often referred to as Cartesian coordinates.

Titration Curves

As base is added to acid at the beginning of a titration, the pH rises very slowly. Nearer to the equivalence point, the pH begins to rapidly increase. If the titration is a strong acid with a strong base, the pH at the equivalence point is equal to 7. A bit past the equivalence point, the rate of change of the pH again slows down. A **titration curve** is a graphical representation of the pH of a solution during a titration. The figure below shows two different examples of a strong acid-strong base titration curve. On the left is a titration in which the base is added to the acid, and so the pH progresses from low to high. On the right is a titration in which the acid is added to the base. In this case, the pH starts out high and decreases during the titration. In both cases, the equivalence point is reached when the moles of acid and base are equal and the pH is 7. This also corresponds to the color change of the indicator.

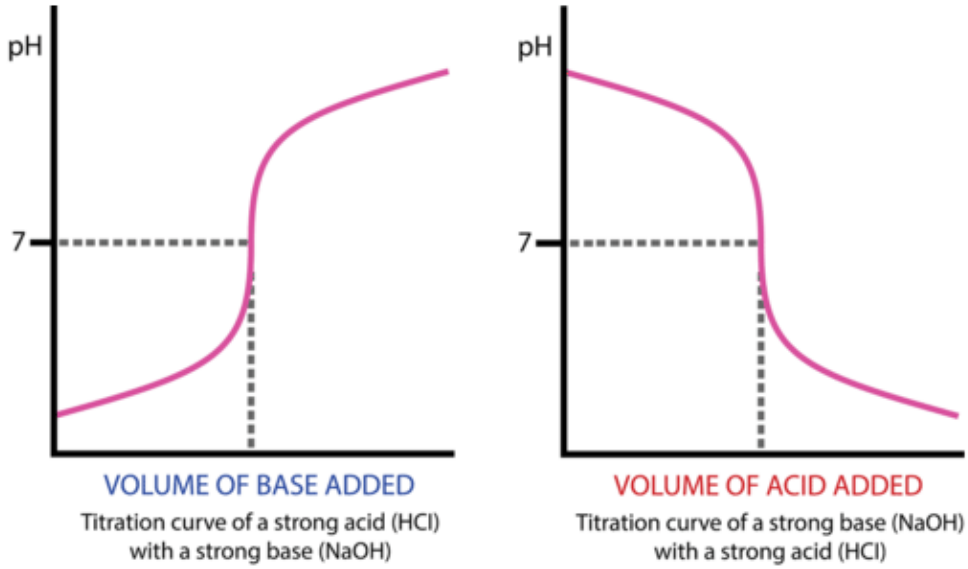


Figure 14.18.1: A titration curve shows the pH changes that occur during the titration of an acid with a base. On the left, base is being added to acid. On the right, acid is being added to base. In both cases, the equivalence point is at pH 7. (CC BY-NC; CK-12)

Titration curves can also be generated in the case of a weak acid-strong base titration or a strong acid-weak base titration. The general shape of the titration curve is the same, but the pH at the equivalence point is different. In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point. In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.

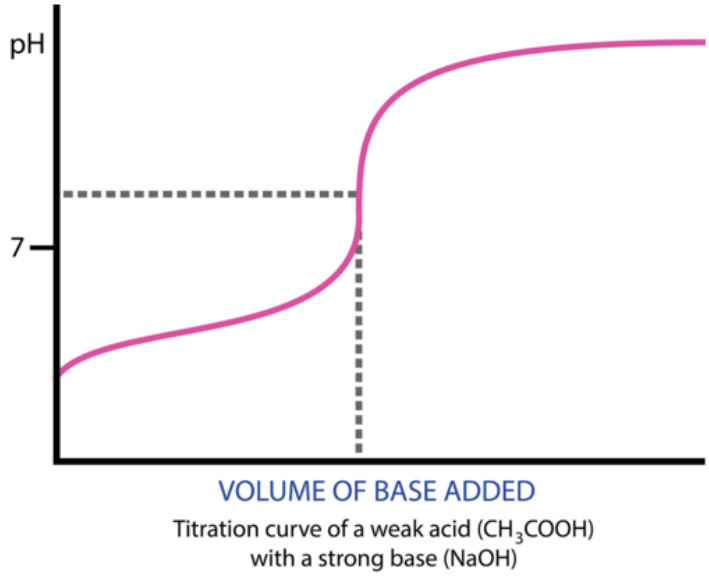


Figure 14.18.2: Titration curve of weak acid and strong base. (CC BY-NC; CK-12)

Summary

- A titration curve is a graphical representation of the pH of a solution during a titration.
- In a strong acid-strong base titration, the equivalence point is reached when the moles of acid and base are equal and the pH is 7.
- In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point.
- In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.

This page titled 14.18: Titration Curves is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.19: Titration Curves by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.19: Indicators

"Boil Them Cabbage Down" is an old bluegrass song. Many people enjoy the music, but chemistry students also enjoy the product of boiled cabbage. Extracting the anthocyanin dye from cabbage leaves with boiling water gives a solution that is red when acidic, purple when neutral, and green to yellow when basic.

Indicators

An **acid-base indicator** is a substance that displays different colors when in the presence of an acid or a base. How does that work? An indicator is a weak acid that ionizes within a known pH range, usually about 2 pH units. We can represent the protonated form of the indicator molecule as HIn and the deprotonated form as $\mathrm{In^-}$. The following equilibrium exists for the indicator:

$$\mathrm{HIn}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{In^-}\,(aq)$$

According to Le Chatelier's principle, the addition of $\mathrm{H^+}$ ions (as in a low pH solution) drives the equilibrium to the left and the protonated HIn predominates. The addition of $\mathrm{OH^-}$ (as in a high pH solution) decreases the $\mathrm{H^+}$ concentration and drives the equilibrium to the right and the deprotonated $\mathrm{In^-}$ predominates. To be useful as an indicator, the two forms must be different colors. In the case of phenolphthalein, the protonated form is colorless, while the deprotonated form is pink. The figure below shows a variety of acid-base indicators that can be used in titration experiments.

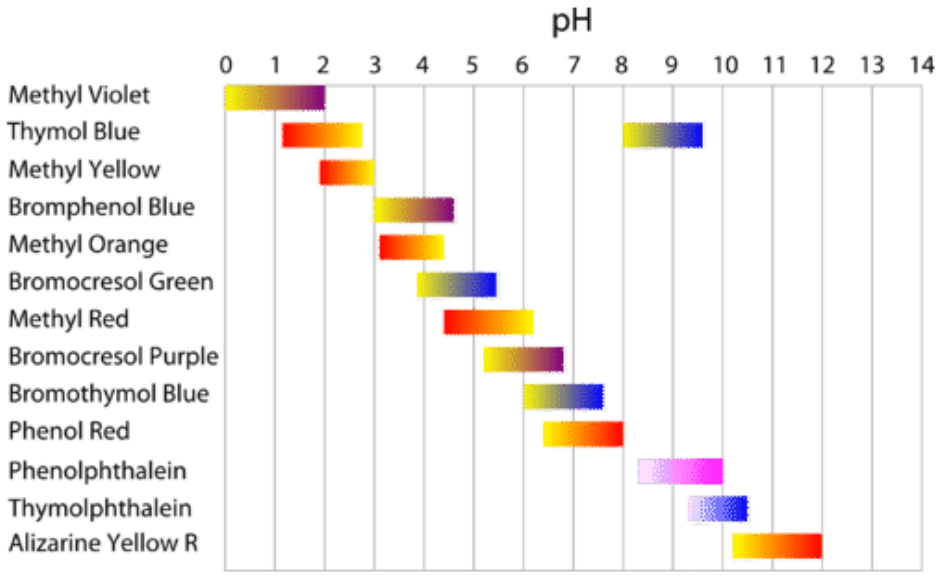


Figure $14.19.1$: pH indicators. (CC BY-NC; CK-12)

Depending on the pH at the equivalence point, the appropriate indicator must be chosen. For example, bromphenol blue has a yellow color below a pH of about 3 and a blue-violet color above a pH of about 4. Bromphenol blue would not be a good choice as the indicator for a strong acid-strong base titration, because the pH is 7 at the equivalence point. Instead, it could be used for a strong acid-weak base titration, where the pH at the equivalence point is lower.

Most indicators have two colored forms. Universal indicator displays the entire rainbow of colors from low pH to high pH (see figure below). Universal indicator is used to make pH paper, which can be used to quickly test solutions for their approximate pH.



Figure 14.19.2: Universal indicator tape. (CC BY-NC; CK-12)

Summary

- An acid-base indicator is a substance that displays different colors when in the presence of an acid or a base.
- Examples of acid-base indicators are given.

This page titled 14.19: Indicators is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.20: Indicators by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.20: Hydrolysis of Salts - Equations

Baking seems easy with all the pre-mixed items available. ("just add water and stir"). However, there is a good amount of chemistry involved in baking with ingredients that are measured out. One important ingredient is baking powder. The fluffiness in the final product of a non-yeast recipe is usually due to the carbon dioxide formed by baking powder. One popular brand uses a mix of sodium bicarbonate and sodium aluminum sulfate to produce the $\mathrm{CO_2}$. The reaction is:

$$3\mathrm{NaHCO_3} + \mathrm{NaAl(SO_4)_2} \rightarrow \mathrm{Al(OH)_3} + 2\mathrm{Na_2SO_4} + 3\mathrm{CO_2}$$

If all goes well, the biscuits rise, the pancakes are fluffy, and everybody is happy.

Hydrolysis of Salts: Equations

A salt is an ionic compound that is formed when an acid and a base neutralize each other. While it may seem that salt solutions are always neutral, they can frequently be either acidic or basic.

Consider the salt formed when the weak acid hydrofluoric acid is neutralized by the strong base sodium hydroxide. The molecular and net ionic equations are shown below.

$$\mathrm{HF}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaF}\,(aq) + \mathrm{H_2O}\,(l)$$
$$\mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l)$$

Since sodium fluoride is soluble, the sodium ion is a spectator ion in the neutralization reaction. The fluoride ion is capable of reacting, to a small extent, with water, accepting a proton.

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

The fluoride ion is acting as a weak Brønsted-Lowry base. The hydroxide ion that is produced as a result of the above reaction makes the solution slightly basic. **Salt hydrolysis** is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.

Salts That Form Basic Solutions

When solid sodium fluoride is dissolved into water, it completely dissociates into sodium ions and fluoride ions. The sodium ions do not have any capability of hydrolyzing, but the fluoride ions hydrolyze to produce a small amount of hydrofluoric acid and hydroxide ion.

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

Salts that are derived from the neutralization of a weak acid (HF) by a strong base (NaOH) will always produce salt solutions that are basic.

Salts That Form Acidic Solutions

Ammonium chloride $(\mathrm{NH_4Cl})$ is a salt that is formed when the strong acid HCl is neutralized by the weak base $\mathrm{NH_3}$. Ammonium chloride is soluble in water. The chloride ion produced is incapable of hydrolyzing because it is the conjugate base of the strong acid HCl. In other words, the $\mathrm{Cl^-}$ ion cannot accept a proton from water to form HCl and $\mathrm{OH^-}$, as the fluoride ion did in the previous section. However, the ammonium ion is capable of reacting slightly with water, donating a proton and so acting as an acid.

$$\mathrm{NH_4{}^+}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{NH_3}\,(aq)$$

Salts That Form Neutral Solutions

A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7. An example is

sodium chloride, formed from the neutralization of HCl by NaOH. A solution of NaCl in water has no acidic or basic properties, since neither ion is capable of hydrolyzing. Other salts that form neutral solutions include potassium nitrate $(\mathrm{KNO_3})$ and lithium bromide (LiBr). The table below summarizes how to determine the acidity or basicity of a salt solution.

Table 14.20.1

Salt formed from:	Salt Solution
Strong acid + Strong base	Neutral
Strong acid + Weak base	Acidic
Weak acid + Strong base	Basic

Salts formed from the reaction of a weak acid and a weak base are more difficult to analyze due to competing hydrolysis reactions between the cation and the anion. These salts are not considered in this chapter's concept.

Summary

- Salt hydrolysis is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.
- Salts that are derived from the neutralization of a weak acid by a strong base will always produce salt solutions that are basic.
- Salts that are derived from the neutralization of a strong acid by a weak base will always produce salt solutions that are acidic.
- A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7.

This page titled 14.20: Hydrolysis of Salts - Equations is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.21: Hydrolysis of Salts - Equations by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

14.21: Calculating pH of Salt Solutions

Many enjoy a cool dip in a swimming pool on a hot day, but may not realize the work needed to keep that water safe and healthy. The ideal pH for a swimming pool is around 7.2. The pH will change as a result of many factors. Adjustment can be accomplished with different chemicals, depending on the tested pH. High pH can be lowered with liquid HCl (unsafe material) or sodium bisulfate. The bisulfate anion is a weak acid and can dissociate partially in solution. To increase pH, use sodium carbonate. The carbonate anion forms an equilibrium with protons that results in some formation of carbon dioxide.

Calculating pH of Salt Solutions

It is often helpful to be able to predict the effect a salt solution will have on the pH of a certain solution. Knowledge of the relevant acidity or basicity constants allows us to carry out the necessary calculations.

Example 14.21.1

If we dissolve NaF in water, we get the following equilibrium:

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

The pH of the resulting solution can be determined if the K_b of the fluoride ion is known. $20.0\ \mathrm{g}$ of sodium fluoride is dissolved in enough water to make $500.0\ \mathrm{mL}$ of solution. Calculate the pH of the solution. The K_b of the fluoride ion is 1.4×10^{-11}.

Solution:

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

Known

- Mass $\mathrm{NaF} = 20.0\ \mathrm{g}$
- Molar mass $\mathrm{NaF} = 41.99\ \mathrm{g/mol}$
- Volume solution $= 0.5000\ \mathrm{L}$
- K_b of $\mathrm{F^-} = 1.4 \times 10^{-11}$

Unknown

The molarity of the $\mathrm{F^-}$ solution can be calculated from the mass, molar mass, and solution volume. Since NaF completely dissociates, the molarity of the NaF is equal to the molarity of the $\mathrm{F^-}$ ion. An ICE table (below) can be used to calculate the concentration of $\mathrm{OH^-}$ produced and then the pH of the solution.

Step 2: Solve.

$$20.0\ \cancel{\mathrm{g\ NaF}} \times \frac{1\ \cancel{\mathrm{mol\ NaF}}}{41.99\ \cancel{\mathrm{g\ NaF}}} \times \frac{1\ \mathrm{mol\ F^-}}{1\ \cancel{\mathrm{mol\ NaF}}} = 0.476\ \mathrm{mol\ F^-}$$

$$\frac{0.476\ \mathrm{mol\ F^-}}{0.5000\ \mathrm{L}} = 0.953\ \mathrm{M\ F^-}$$

$$\text{Hydrolysis equation: } \mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

	$\mathrm{F^-}$	HF	$\mathrm{OH^-}$
Initial	0.953	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.953 - x$	x	x

$$K_b = 1.4 \times 10^{-11} = \frac{(x)(x)}{0.953 - x} = \frac{x^2}{0.953 - x} \approx \frac{x^2}{0.953}$$
$$x = [OH^-] = \sqrt{1.4 \times 10^{-11}(0.953)} = 3.65 \times 10^{-6}\ M$$
$$pOH = -\log\left(3.65 \times 10^{-6}\right) = 5.44$$
$$pH = 14 - 5.44 = 8.56$$

Step 3: Think about your result.

The solution is slightly basic due to the hydrolysis of the fluoride ion.

Salts That Form Acidic Solutions

When the ammonium ion dissolves in water, the following equilibrium exists:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NH_3(aq)$$

The production of hydronium ions causes the resulting solution to be acidic. The pH of a solution of ammonium chloride can be found in a very similar way to the sodium fluoride solution in the previous example. However, since the ammonium chloride is acting as an acid, it is necessary to know the K_a of NH_4^+, which is 5.6×10^{-10}. We will find the pH of a $2.00\ M$ solution of NH_4Cl. Because the NH_4Cl completely ionizes, the concentration of the ammonium ion is $2.00\ M$.

$$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

Again, an ICE table (below) is set up in order to solve for the concentration of the hydronium (or H^+) ion produced.

	NH_4^+	H^+	NH_3
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00 - x$	x	x

Now substituting into the K_a expression gives:

$$K_a = 5.6 \times 10^{-10} = \frac{x^2}{2.00 - x} \approx \frac{x^2}{2.00}$$
$$x = [H^+] = \sqrt{5.6 \times 10^{-10}(2.00)} = 3.3 \times 10^{-5}\ M$$
$$pH = -\log\left(3.3 \times 10^{-5}\right) = 4.48$$

A salt produced from a strong acid and a weak base yields a solution that is acidic.

Summary

This page titled 14.21: Calculating pH of Salt Solutions is shared under a CK-12 license and was authored, remixed, and/or curated by CK12 via source content that was edited to the style and standards of the LibreTexts platform.

- 21.22: Calculating pH of Salt Solutions by CK12 is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

INDEX

A

acid ionization constant
14.11: Strong and Weak Acids and Acid Ionization Constant (K_a)

Aufbau principle
6.9: Aufbau Principle

B

base ionization constant
14.12: Strong and Weak Bases and Base Ionization Constant

C

chemical formula
8.8: Formulas for Binary Ionic Compounds

chromatography
4.13: Separating Mixtures

D

distillation
4.13: Separating Mixtures

E

equivalence point
14.16: Titration Experiment

evaporation
11.10: Evaporation
4.13: Separating Mixtures

G

gas density
9.9: Gas Density

H

hydrolysis
14.20: Hydrolysis of Salts - Equations

I

indicator
14.16: Titration Experiment

ionic bonding
7.6: Ionic Bonding

L

law of definite proportions
9.1: Law of Definite Proportions

length
3.3: Length and Volume

lone pair
7.17: Molecular Shapes - Lone Pair(s) on Central Atom

M

meniscus
11.9: Surface Tension

N

neutralization reaction
14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

noble gas configuration
6.7: Noble Gas Configuration

P

ph meters
14.13: Calculating Acid and Base Dissociation Constants

pOH
14.10: The pOH Concept

polyatomic ions
7.7: Polyatomic Ions
8.9: Polyatomic Ions

R

resonance
7.14: Resonance

S

salt
14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

salt hydrolysis
14.20: Hydrolysis of Salts - Equations

scienfitic notation
2.8: Scientific Notation

standard solution
14.16: Titration Experiment

stock system
8.6: The Stock System of Nomenclature

strong base
14.12: Strong and Weak Bases and Base Ionization Constant

surface tension
11.9: Surface Tension

T

titration
14.16: Titration Experiment

titration curve
14.18: Titration Curves

V

valence electrons
7.12: Valence Electrons

volume
3.3: Length and Volume

W

weak base
14.12: Strong and Weak Bases and Base Ionization Constant

This page titled Index is shared under a license and was authored, remixed, and/or curated by LibreTexts.

GLOSSARY

S

Sample Word 1 | Sample Definition 1

This page titled Glossary is shared under a license and was authored, remixed, and/or curated by LibreTexts.

DETAILED LICENSING

Overview

Title: Introductory Chemistry at Solano College - Updated 2023/04/03

Webpages: 207

Applicable Restrictions: Noncommercial

All licenses found:

- CK-12 License: 90.8% (188 pages)
- Undeclared: 9.2% (19 pages)

By Page

- Introductory Chemistry at Solano College - Updated 2023/04/03 - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Introduction To Chemistry - *Undeclared*
    - 1.1: Scope of Chemistry - *CK-12 License*
    - 1.2: Areas of Chemistry - *CK-12 License*
    - 1.3: Research - *CK-12 License*
    - 1.4: Scientific Problem Solving - *CK-12 License*
  - 2: Numbers From Measurements - *Undeclared*
    - 2.1: Accuracy and Precision - *CK-12 License*
    - 2.2: Percent Error - *CK-12 License*
    - 2.3: Measurement Uncertainty - *CK-12 License*
    - 2.4: Significant Figures - *CK-12 License*
    - 2.5: Rounding - *CK-12 License*
    - 2.6: Significant Figures in Addition and Subtraction - *CK-12 License*
    - 2.7: Significant Figures in Multiplication and Division - *CK-12 License*
    - 2.8: Scientific Notation - *CK-12 License*
  - 3: Unit Systems and Dimensional Analysis - *Undeclared*
    - 3.1: SI Base Units - *CK-12 License*
    - 3.2: Metric Prefixes - *CK-12 License*
    - 3.3: Length and Volume - *CK-12 License*
    - 3.4: Mass and Weight - *CK-12 License*
    - 3.5: Metric Unit Conversions - *CK-12 License*
    - 3.6: Derived Units - *CK-12 License*
    - 3.7: Dimensional Analysis - *CK-12 License*
    - 3.8: Density - *CK-12 License*
    - 3.9: Kinetic Energy - *CK-12 License*
    - 3.10: Temperature and Temperature Scales - *CK-12 License*
  - 4: Matter and Change - *Undeclared*
    - 4.1: Matter, Mass, and Volume - *CK-12 License*
    - 4.2: States of Matter - *CK-12 License*
    - 4.3: Physical Properties - *CK-12 License*
    - 4.4: Chemical Properties and Chemical Reactions - *CK-12 License*
    - 4.5: Extensive and Intensive Properties - *CK-12 License*
    - 4.6: Physical Change - *CK-12 License*
    - 4.7: Chemical Change - *CK-12 License*
    - 4.8: Recognizing Chemical Reactions - *CK-12 License*
    - 4.9: Pure Substances - *CK-12 License*
    - 4.10: Mixture - *CK-12 License*
    - 4.11: Heterogeneous Mixtures - *CK-12 License*
    - 4.12: Homogeneous Mixture - *CK-12 License*
    - 4.13: Separating Mixtures - *CK-12 License*
    - 4.14: Elements - *CK-12 License*
    - 4.15: Compounds - *CK-12 License*
    - 4.16: Chemical Symbols and Formulas - *CK-12 License*
    - 4.17: A Molecular View of Elements and Compounds - *CK-12 License*
  - 5: Atoms, Molecules, and Subatomic Particles - *Undeclared*
    - 5.1: Atom - *CK-12 License*
    - 5.2: Protons - *CK-12 License*
    - 5.3: Neutrons - *CK-12 License*
    - 5.4: Electrons - *CK-12 License*
    - 5.5: Atomic Nucleus - *CK-12 License*
    - 5.6: Atomic Number - *CK-12 License*
    - 5.7: Mass Number - *CK-12 License*
    - 5.8: Isotopes - *CK-12 License*
    - 5.9: Atomic Mass Unit - *CK-12 License*
    - 5.10: Calculating Average Atomic Mass - *CK-12 License*
    - 5.11: Ion - *CK-12 License*
  - 6: Electronic Structure and Chemical Periodicity - *Undeclared*
    - 6.1: Periodic Law - *CK-12 License*

- 6.2: Modern Periodic Table- Periods and Groups - *CK-12 License*
- 6.3: Blocks of the Periodic Table - *CK-12 License*
- 6.4: Energy Level - *CK-12 License*
- 6.5: Orbitals - *CK-12 License*
- 6.6: Electron Configurations - *CK-12 License*
- 6.7: Noble Gas Configuration - *CK-12 License*
- 6.8: Hund's Rule and Orbital Filling Diagrams - *CK-12 License*
- 6.9: Aufbau Principle - *CK-12 License*
- 6.10: Periodic Trends- Atomic Radius - *CK-12 License*
- 6.11: Metals - *CK-12 License*
- 6.12: Nonmetals - *CK-12 License*
- 6.13: Metalloids - *CK-12 License*

◦ 7: Chemical Bonds - *Undeclared*
- 7.1: Chemical Bond - *CK-12 License*
- 7.2: Electron Dot Diagrams - *CK-12 License*
- 7.3: Octet Rule - *CK-12 License*
- 7.4: Cation Formation - *CK-12 License*
- 7.5: Anion Formation - *CK-12 License*
- 7.6: Ionic Bonding - *CK-12 License*
- 7.7: Polyatomic Ions - *CK-12 License*
- 7.8: Covalent Bond - *CK-12 License*
- 7.9: Lewis Electron-Dot Structures - *CK-12 License*
- 7.10: Single Covalent Bonds - *CK-12 License*
- 7.11: Multiple Covalent Bonds - *CK-12 License*
- 7.12: Valence Electrons - *CK-12 License*
- 7.13: Coordinate Covalent Bond - *CK-12 License*
- 7.14: Resonance - *CK-12 License*
- 7.15: VSEPR Theory - *CK-12 License*
- 7.16: Molecular Shapes- No Lone Pairs on Central Atoms - *CK-12 License*
- 7.17: Molecular Shapes - Lone Pair(s) on Central Atom - *CK-12 License*
- 7.18: Bond Polarity - *CK-12 License*
- 7.19: Polar Molecules - *CK-12 License*

◦ 8: Chemical Nomenclature - *CK-12 License*
- 8.1: Molecular Formula - *CK-12 License*
- 8.2: Empirical Formula - *CK-12 License*
- 8.3: Cations - *CK-12 License*
- 8.4: Anions - *CK-12 License*
- 8.5: Transition Metal Ions - *CK-12 License*
- 8.6: The Stock System of Nomenclature - *CK-12 License*
- 8.7: Naming Binary Ionic Compounds - *CK-12 License*
- 8.8: Formulas for Binary Ionic Compounds - *CK-12 License*
- 8.9: Polyatomic Ions - *CK-12 License*
- 8.10: Ternary Ionic Compounds- Naming and Formulas - *CK-12 License*
- 8.11: Binary Molecular Compounds- Naming and Formulas - *CK-12 License*
- 8.12: Acids - Naming and Formulas - *CK-12 License*
- 8.13: Bases- Naming and Formulas - *CK-12 License*

◦ 9: The Mole - *CK-12 License*
- 9.1: Law of Definite Proportions - *CK-12 License*
- 9.2: Avogadro's Number - *CK-12 License*
- 9.3: Conversions Between Moles and Atoms - *CK-12 License*
- 9.4: Molar Mass - *CK-12 License*
- 9.5: Conversions Between Moles and Mass - *CK-12 License*
- 9.6: Conversions Between Mass and Number of Particles - *CK-12 License*
- 9.7: Avogadro's Hypothesis and Molar Volume - *CK-12 License*
- 9.8: Conversions Between Moles and Gas Volume - *CK-12 License*
- 9.9: Gas Density - *CK-12 License*
- 9.10: Mole Road Map - *CK-12 License*
- 9.11: Percent Composition - *CK-12 License*
- 9.12: Percent of Water in a Hydrate - *CK-12 License*
- 9.13: Determining Empirical Formulas - *CK-12 License*
- 9.14: Determining Molecular Formulas - *CK-12 License*

◦ 10: Chemical Calculations Involving Chemical Equations - *Undeclared*
- 10.1: Law of Conservation of Mass - *CK-12 License*
- 10.2: Word Equations - *CK-12 License*
- 10.3: Chemical Equations - *CK-12 License*
- 10.4: Balancing Equations - *CK-12 License*
- 10.5: Combination Reactions - *CK-12 License*
- 10.6: Decomposition Reactions - *CK-12 License*
- 10.7: Combustion Reactions - *CK-12 License*
- 10.8: Single Replacement Reactions - *CK-12 License*
- 10.9: Activity Series - *CK-12 License*
- 10.10: Double Replacement Reactions - *CK-12 License*
- 10.11: Everyday Stoichiometry - *CK-12 License*
- 10.12: Mole Ratios - *CK-12 License*
- 10.13: Mass-Mole Stoichiometry - *CK-12 License*
- 10.14: Mass-Mass Stoichiometry - *CK-12 License*
- 10.15: Volume-Volume Stoichiometry - *CK-12 License*
- 10.16: Mass-Volume Stoichiometry - *CK-12 License*
- 10.17: Limiting Reactant - *CK-12 License*
- 10.18: Determining the Limiting Reactant - *CK-12 License*
- 10.19: Theoretical Yield and Percent Yield - *CK-12 License*

◦ 11: States of Matter - *CK-12 License*
- 11.1: Physical Properties and Intermolecular Forces - *CK-12 License*
- 11.2: Hydrogen Bonding - *CK-12 License*

- 11.3: Kinetic Molecular Theory - *CK-12 License*
- 11.4: Gas Pressure - *CK-12 License*
- 11.5: Atmospheric Pressure - *CK-12 License*
- 11.6: Pressure Units and Conversions - *CK-12 License*
- 11.7: Average Kinetic Energy and Temperature - *CK-12 License*
- 11.8: Physical Properties and Intermolecular Forces - *CK-12 License*
- 11.9: Surface Tension - *CK-12 License*
- 11.10: Evaporation - *CK-12 License*
- 11.11: Vapor Pressure - *CK-12 License*
- 11.12: Boiling - *CK-12 License*
- 11.13: Vapor Pressure Curves - *CK-12 License*
- 11.14: Melting - *CK-12 License*
- 11.15: Sublimation - *CK-12 License*
- 11.16: Heating and Cooling Curves - *CK-12 License*
- 11.17: General Phase Diagram - *CK-12 License*
- 11.18: Phase Diagram for Water - *CK-12 License*

◦ 12: Gas Laws - *Undeclared*

◦ 13: Solutions - *CK-12 License*
- 13.1: Solute-Solvent Combinations - *CK-12 License*
- 13.2: Rate of Dissolution - *CK-12 License*
- 13.3: Saturated and Unsaturated Solutions - *CK-12 License*
- 13.4: How Temperature Influences Solubility - *CK-12 License*
- 13.5: Supersaturated Solutions - *CK-12 License*
- 13.6: Predicting Precipitates Using Solubility Rules - *CK-12 License*
- 13.7: Percent Solutions - *CK-12 License*
- 13.8: Molarity - *CK-12 License*
- 13.9: Preparing Solutions - *CK-12 License*
- 13.10: Dilution - *CK-12 License*
- 13.11: Molality - *CK-12 License*
- 13.12: Calculating Molar Mass - *CK-12 License*
- 13.13: Electrolytes and Nonelectrolytes - *CK-12 License*
- 13.14: Strong and Weak Electrolytes - *CK-12 License*
- 13.15: Molecular and Ionic Equations - *CK-12 License*
- 13.16: Net Ionic Equations - *CK-12 License*

◦ 14: Acids and Bases - *CK-12 License*
- 14.1: Properties of Acids - *CK-12 License*
- 14.2: Properties of Bases - *CK-12 License*
- 14.3: Arrhenius Acids - *CK-12 License*
- 14.4: Arrhenius Bases - *CK-12 License*
- 14.5: Brønsted-Lowry Acids and Bases - *CK-12 License*
- 14.6: Brønsted-Lowry Acid-Base Reactions - *CK-12 License*
- 14.7: Ion-Product of Water - *CK-12 License*
- 14.8: The pH Scale - *CK-12 License*
- 14.9: Calculating pH of Acids and Bases - *CK-12 License*
- 14.10: The pOH Concept - *CK-12 License*
- 14.11: Strong and Weak Acids and Acid Ionization Constant (K_a) - *CK-12 License*
- 14.12: Strong and Weak Bases and Base Ionization Constant - *CK-12 License*
- 14.13: Calculating Acid and Base Dissociation Constants - *CK-12 License*
- 14.14: Calculating pH of Weak Acid and Base Solutions - *CK-12 License*
- 14.15: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions - *CK-12 License*
- 14.16: Titration Experiment - *CK-12 License*
- 14.17: Titration Calculations - *CK-12 License*
- 14.18: Titration Curves - *CK-12 License*
- 14.19: Indicators - *CK-12 License*
- 14.20: Hydrolysis of Salts - Equations - *CK-12 License*
- 14.21: Calculating pH of Salt Solutions - *CK-12 License*

◦ Back Matter - *Undeclared*
- Index - *Undeclared*
- Glossary - *Undeclared*
- Detailed Licensing - *Undeclared*

This page titled Introductory Chemistry at Solano College - Updated 2023/04/03 is shared under a license and was authored, remixed, and/or curated by LibreTexts.