CHEM 150: EXPLORING OUR CHEMICAL ENVIRONMENT



Tara Bunag
Modesto Junior College

LibreTexts™

Textbook for Chemistry 150

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

1: Introduction to Chemistry

1.1: What is Chemistry?

1.2: Using the Scientific Method

1.3: Pure and Applied Research

1.4: Classification and Properties of Matter

1.5: Applying Properties - MSDS

1.5.01: Applying Properties - MSDS

1.5.1: Applying Properties - MSDS/SDS

1.6: All About the Elements

2: The Mathematics of Chemistry

2.1: Measuring Matter

2.2: Dimensional Analysis

2.3: Density

2.4: Temperature

2.E: The Mathematics of Chemistry (Exercises)

3: Atoms and the Periodic Table

3.1: Basic Atomic Theory

3.2: Development of the Modern Periodic Table

3.3: Subatomic Particles - Electrons, Protons, and Neutrons

3.4: Atomic Mass and Atomic Number

3.5: Isotopes

3.6: The Importance of Ions to a Chemist

3.E: Atoms and the Periodic Table (Exercises)

4: Valence Electrons and Bonding

4.1: Atomic Models of the Twentieth Century

4.2: Understanding Atomic Spectra

4.3: Atomic Spectroscopy Applications

4.4: Predicting Ion Charges

4.5: Lewis Dot and Bonding

4.6: Ionic Formula Writing

4.7: Applications and Solubility of Ionic Compounds

4.8: Covalent Bonding and Formula Writing

4.9: Free Radicals and the Environment

4.10: Free Radicals and Health

4.11: Applications and Solubility of Covalent Compounds

4.12: Drug Study

4.13: Arrhenius Acids and Bases
4.E: Valence Electrons and Bonding (Exercises)

5: Basics of Nuclear Science

5.1: The Discovery of Radiation
5.2: The Electromagnetic Spectrum
5.3: Types of Radiation
5.4: Ionizing Radiation and Non-ionizing Radiation
5.5: The Basics of Skin Cancer
5.6: Measuring Radiation
5.7: Calculating Half-Life
5.8: Using Nuclear Science to Diagnose Disease
5.9: Therapeutic Radiation
5.10: Non-medical Applications of Nuclear Science
5.E: Basics of Nuclear Science (Exercises)

6: Nuclear Weapons- Fission and Fusion

6.1: The History and Basics of Fission
6.2: The Manhattan Project - Prewar
6.3: The Manhattan Project - Labs and Fuel
6.4: The Manhattan Project - Critical Mass and Bomb Construction
6.5: Hiroshima and Nagasaki
6.6: Aftermath of Bombings
6.7: Fusion
6.8: Fallout and Weapons Testing
6.9: Additional Readings and Media Fun
6.E: Nuclear Weapons- Fission and Fusion (Exercises)

7: Nuclear Energy

7.1: Sharing and Monitoring Nuclear Technology
7.2: Generating Electricity
7.3: Powering the United States
7.4: Generating Electricity by Using Fission
7.5: United States Commercial Nuclear Power Reactors
7.6: South Carolina Energy
7.7: Three Mile Island Nuclear Accident
7.8: Chernobyl Nuclear Disaster
7.9: Fukushima Nuclear Disaster
7.10: Storing and Transporting Nuclear Waste Materials
7.11: New Generation Nuclear Reactors
7.E: Nuclear Energy (Exercises)

8: Water chemistry

8.1: Water Cycle

8.2: EPA's Primary and Secondary Drinking Water Standards

8.3: Biological Contamination of Water

8.4: Chemical Contamination of Water

8.5: Water Concentration Units

8.6: Sanitation of Drinking Water

8.7: Fluoridation of Water

8.8: Other Water treatment processes

8.9: Sanitizing water while traveling

8.10: Bottled Water Basics

8.11: Love Canal Case Study

8.12: Flint River Case Study

9: Organic Chemistry

9.1: Introduction to Organic Chemistry

9.2: Alkanes

9.3: Alkenes

9.4: Alkynes

9.5: Cyclics

9.6: Isomers

9.7: Hydrogenation

9.8: Functional Group introduction

9.9: Aromatics

9.10: Alcohols

9.11: Ethers

9.12: Carbonyl Friends

9.13: Amines and Thiols

9.E: Organic Chemistry (Exercises)

10: Pharmaceuticals

10.1: Chemotherapy

10.2: Side Effects and Warnings

10.3: Allergy and Cold Meds

10.4: Combination Medications

10.5: Stomach Medicine

10.6: DEA and Controlled Substances

10.7: Narcotics

10.8: Basics of Neurotransmitters

10.9: Antianxiety Medications

10.10: Amphetamines

10.11: Antipsychotic medication

10.12: Analgesics
10.13: Narcotic Analgesics
10.14: Antidepressant medications

11: Introduction to Chemical Reactions

11.1: Prelude to Introduction to Chemical Reactions
11.2: The Law of Conservation of Matter
11.3: Chemical Equations
11.4: Quantitative Relationships Based on Chemical Equations
11.5: Some Types of Chemical Reactions
11.E: Introduction to Chemical Reactions (Exercises)
11.S: Introduction to Chemical Reactions (Summary)

12: Quantities in Chemical Reactions

12.1: Prelude to Quantities in Chemical Reactions
12.2: The Mole
12.3: Atomic and Molar Masses
12.4: Mole-Mass Conversions
12.5: Mole-Mole Relationships in Chemical Reactions
12.6: Mole-Mass and Mass-Mass Problems
12.E: Quantities in Chemical Reactions (Exercise)
12.S: Quantities in Chemical Reactions (Summary)

13: Solids, Liquids, and Gases

13.1: Prelude to Solids, Liquids, and Gases
13.2: Intermolecular Interactions
13.3: Solids and Liquids
13.4: Gases and Pressure
13.5: Gas Laws
13.E: Solids, Liquids, and Gases (Exercises)
13.S: Solids, Liquids, and Gases (Summary)

14: Solutions

14.1: Prelude to Solutions
14.2: Solutions
14.3: Concentration
14.4: The Dissolution Process
14.5: Properties of Solutions
14.E: Solutions (Exercises)
14.S: Solutions (Summary)

15: Acids and Bases

15.1: Prelude to Acids and Bases
15.2: Arrhenius Definition of Acids and Bases

15.3: Brønsted-Lowry Definition of Acids and Bases
15.4: Water - Both an Acid and a Base
15.5: The Strengths of Acids and Bases
15.6: Buffers
15.E: Acids and Bases (Exercises)
15.S: Acids and Bases (Summary)

16: Examples of Final Project

16.1: Amphetamines and the Brain
16.2: Great Pacific Garbage Patch
16.3: Sorting Recyclable Plastics by Density
16.4: Solubility and Pollution
16.5: Radiation Contamination from Fukushima
16.6: pH of Chesapeake Bay
16.7: 18-F fluorodeoxyglucose and Atherosclerotic Plaque
Student Project

1: Unit 1

Index

Index

Glossary

This page titled CHEM 150: Exploring Our Chemical Environment is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

1: INTRODUCTION TO CHEMISTRY

What is chemistry? Simply put, chemistry is the study of the interactions of matter with other matter and with energy. This seems straightforward enough. However, the definition of chemistry includes a wide range of topics that must be understood to gain a mastery of the topic or even take additional courses in chemistry. Get ready for a fantastic journey through a world of wonder, delight, and knowledge. One of the themes of this book is “chemistry is everywhere,” and indeed it is; you would not be alive if it weren’t for chemistry because your body is a big chemical machine. If you don’t believe it, don’t worry. Every chapter in this book contains examples that will show you how chemistry is, in fact, everywhere. So enjoy the ride—and enjoy chemistry.

1.1: What is Chemistry?

1.2: Using the Scientific Method

1.3: Pure and Applied Research

1.4: Classification and Properties of Matter

1.5: Applying Properties - MSDS

1.5.01: Applying Properties - MSDS

1.5.1: Applying Properties - MSDS/SDS

1.6: All About the Elements

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

1.1: What is Chemistry?

Learning Objectives

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

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

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 1.1.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

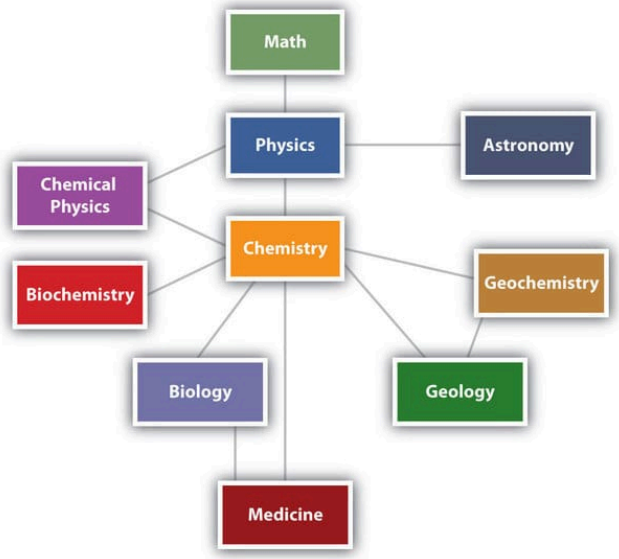


Figure 1.1.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 1.1.1 shows how many of the individual fields of science are related.

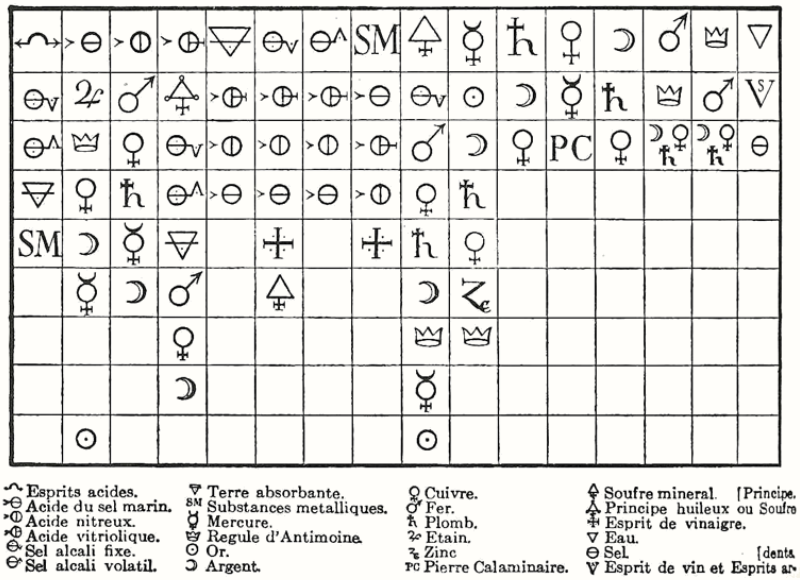
There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed

Alchemy Is in No way Chemistry!

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

Alchemy was a somewhat mystical and secretive approach to learning how to manipulate matter. Practitioners, called alchemists, thought that all matter was composed of different proportions of the four basic elements—fire, water, earth,

and air—and believed that if you changed the relative proportions of these elements in a substance, you could change the substance. The long-standing attempts to “transmute” common metals into gold represented one goal of alchemy. Alchemy’s other major goal was to synthesize the philosopher’s stone, a material that could impart long life—even immortality. Alchemists used symbols to represent substances, some of which are shown in the accompanying figure. This was not done to better communicate ideas, as chemists do today, but to maintain the secrecy of alchemical knowledge, keeping others from sharing in it.



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

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

Alchemy and the ACS (American Chemical Society)

While watching the video below and answer the following questions. This section is optional. If skipping, continue at Areas of Chemistry, 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

Questions

1. What was the chief goal of an alchemist according to the video?
2. What could the philosopher’s stone do to urine?
3. Is Alchemy a true science?
4. When urine is boiled down to a white paste, what is the name and symbol for the element that was obtained?
5. List some properties of this element that were discussed in the video.
6. Did wealthy people produce more of this element than poorer people?
7. What types of applications (applied science) did this element lead us to?

8. Instead of collecting urine, where can be obtain higher concentrations of this element?
9. The video discussed phosphoric acid (formula: H_3PO_4). Name all the elements in this compound.
10. What were the mentioned applications of phosphoric acid?
11. What are some of the organic and biochemical applications of element 13?

Areas of Chemistry

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

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon with hydrogen. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that 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.
- **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 better treatments can be developed.



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

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

Chemists at work

The American Chemical Society (ACS) has designed a series of videos illustrating the different fields that a chemist could pursue. Please watch this video and note what the chemist does to improve the world and the lives of others.

Summary

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

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

Contributors

-
- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 1.1: What is Chemistry? is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

1.2: Using the Scientific Method

Learning Objectives

- To identify the components of the scientific method

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments, which in turn lead to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.2.1).

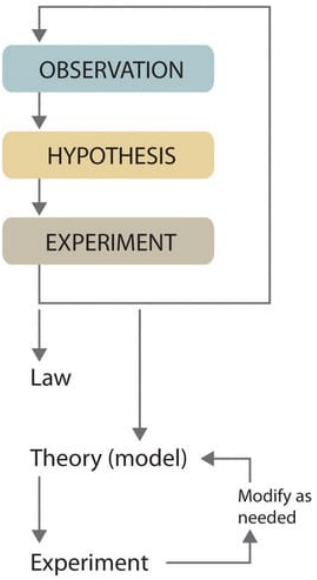


Figure 1.2.1: The Steps in the Scientific Method. Image used with permission (CC BY-SA-NC; anonymous).

Step 1: Make observations

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

Step 2: Formulate a hypothesis

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

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

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

Step 3: Design and perform experiments

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

Step 4: Accept or modify the hypothesis

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

Step 5: Development into a law and/or theory

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

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

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

Because scientists can enter the cycle shown in Figure 1.2.1 at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

A Real World Application of the Scientific Method

Observation: You enter a room and flip the light switch, but the light does not go on.

Hypothesis: Determine more than one possible hypothesis to explain your observation.

Experiments: What experiments could you design and perform to test your hypothesis?

Exercise 1.2.1

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

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. According to Albert Einstein, mass X speed of light = energy
d. When 10 g of ice were added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
e. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution:

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is an educated guess regarding the origin of birds, so it is a hypothesis.
c. This is a theory which explains an explanation of events and can be disproven at any time.
d. The temperature is measured before and after a change is made in a system, so these are quantitative observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

Exercise 1.2.2

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

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter ever to play professional basketball.
e. Limestone is relatively insoluble in water but dissolves readily in dilute acid with the evolution of a gas.
f. Gas mixtures that contain more than 4% hydrogen in air are potentially explosive.

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

qualitative observation

Answer f

quantitative observation

Summary

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

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

Contributors

-

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 1.2: Using the Scientific Method is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

1.3: Pure and Applied Research

In science, we usually talk about two types of research: pure and applied. **Pure research** focuses on answering basic questions such as, "how do gases behave?" **Applied research** would be involved in the process of developing specific preparation for a gas in order for it to be produced and delivered efficiently and economically. This division sounds like it would be easy to make, but sometimes we cannot draw a clear line between what is "pure" and what is "applied".

Examples of "Pure" Research

A lot of "pure" research is of the "what is this?" or "how does it work?" variety. The early history of chemistry contains many examples. The ancient Greek philosophers debated the composition of matter (earth? air? fire? water? all of the above?). They weren't going to do anything with their knowledge - they just wanted to know.

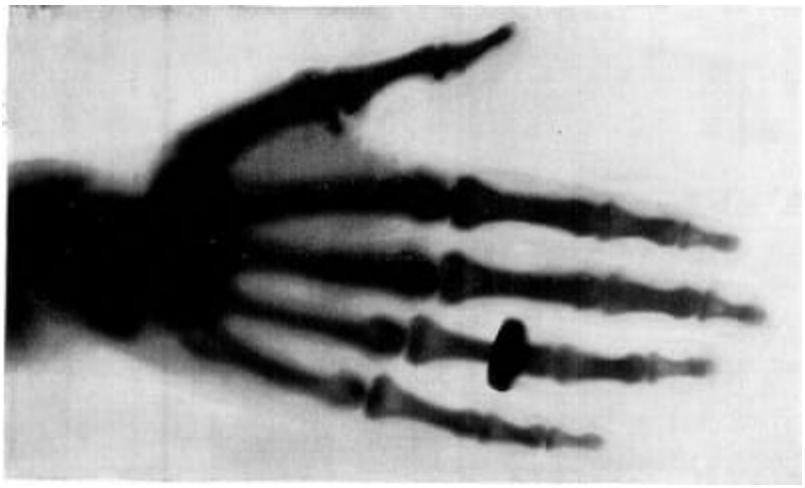


Ancient Greek philosophers."A Greek Philosopher and His Disciples", Antonio Zucchi

Studies on the elements (especially after Mendeleev's periodic table was published) were primarily "pure" research types of experiments. Does this element exist? What are its properties? The scientists did not have any practical application in mind, but were curious about the world around them.

Examples of "Applied" Research (Technology)

There is a great deal of "applied" research taking place today. In general, no new science principles are discovered, but existing knowledge is used to develop a new product. A good example of this type of research is the application of x-rays in medicine. In the last 1800's, Wilhelm Röntgen discovered how to produce x-rays by using a cathode ray type. He noted that this new ray could go through the body. Realizing this, Röntgen believed that x-rays would be very helpful in diagnosing and determining disease and injury in the body.

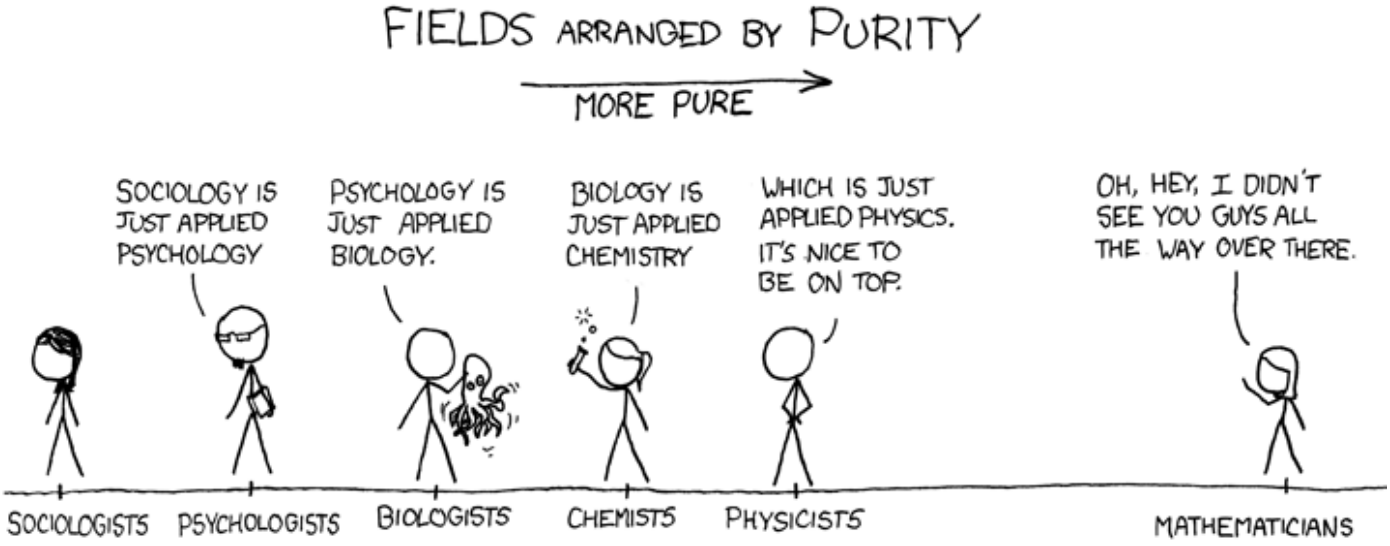


Röntgen is the father of the X-ray. while working with experiments and procedures of previous scientists, Röntgen noticed that even through secured champers for light, and unknown energy was emitted that reacted with phosphorescent materials he was using in his experiments. his discovery and subsequent papers earned him the Nobel Prize in Physics in 1901. He used his wife's hand to determine if the new rays could pass through the body.

Some "In-Between" Examples

Sometimes it is hard to differentiate between pure and applied research. What may start out as simply asking a question may result in some very useful information. If scientists are studying the biochemistry of a microorganism that causes a disease, they may soon find information that would suggest a way to make a chemical that would inactivate the microorganism. The compound could be used to learn more about the biochemistry, but could also be used to cure the disease.

Hemoglobin is a protein in red blood cells that transports oxygen in the bloodstream. Scientists studied hemoglobin simply to learn how it worked. Out of this research came an understanding of how the protein changes shape when oxygen attaches to it. This information was then applied to help patients with sickle cell anemia, a disorder caused by an abnormal hemoglobin structure that makes hemoglobin molecules clump up when oxygen leaves the protein. Basic knowledge of protein structure led to an improved understanding of a wide-spread disease and opened the door for development of treatments.



What is an applied science or research is relative perspective depending on the field. Image used with permission (CC BY-NC 2.5; xkcd.com).

Who Can Do Research?

Anyone can perform research, not just chemists, physicists, and other scientists. There are many opportunities to be a "citizen scientist." You can be a participant in research studies, work with others to design research, or even perform research as part of your work or personal life. How will you incorporate research in your life?

Summary

- Pure research focuses on understanding basic properties and processes.
- Applied research focuses on the use of information to create useful materials.
- Sometimes there is no clear line between pure and applied research.

Contributors

-

- Elizabeth R. Gordon (Furman University)

This page titled 1.3: Pure and Applied Research is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

1.4: Classification and Properties of Matter

Learning Objectives

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

Part of understanding matter is being able to describe it. One way chemists describe matter is to assign different kinds of properties to different categories.

Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. Physical properties are characteristics that **describe** matter. They include characteristics such as size, shape, color, and mass. Many of these properties can be quantitative in nature. For example, quantitative physical properties of water would be boiling point (100 °C / 212 °F) and melting point (0°C / 32 °F). Chemical properties are characteristics that describe how matter **changes** its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. Oxidation, rusting, decomposition, and inertness are chemical properties as well. Click on this video and record the physical and chemical properties of the element sodium.

Elements and Compounds

Any sample of matter that has the same physical and chemical properties throughout the sample is called a substance. There are two types of substances. A substance that cannot be broken down into chemically simpler components is an element. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a compound (Figure 1.4.1). Water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

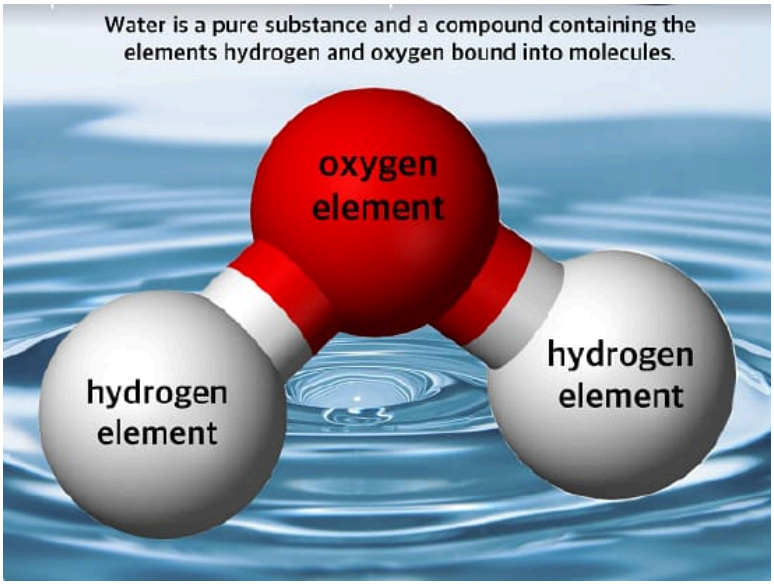


Figure 1.4.1: To a chemist, the pure substance water is a compound (it contains two kinds of atoms bound to one another in just one kind of molecule). An environmentalist might consider water "pure" even if it contains the normal amount of dissolved oxygen and carbon dioxide, but no other "pollutants". To a chemist, water containing oxygen is no longer a pure substance, but a mixture. In pure water, the ratio of hydrogen to oxygen atoms is always 2:1.

Sometimes the word *pure* is added to substance, but this is not absolutely necessary. By definition, any single substance is pure.

The smallest part of an element that maintains the identity of that element is called an atom. Atoms are extremely tiny; to make a line 1 inch long, you would need 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a molecule. Molecules are composed of atoms that are attached together and behave as a unit. Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the macroscopic view of the universe. However, scientists can also describe chemical events on the level of individual atoms or molecules, which is referred to as the microscopic viewpoint. We will see examples of both macroscopic and microscopic viewpoints throughout this book (Figure 1.4.2).

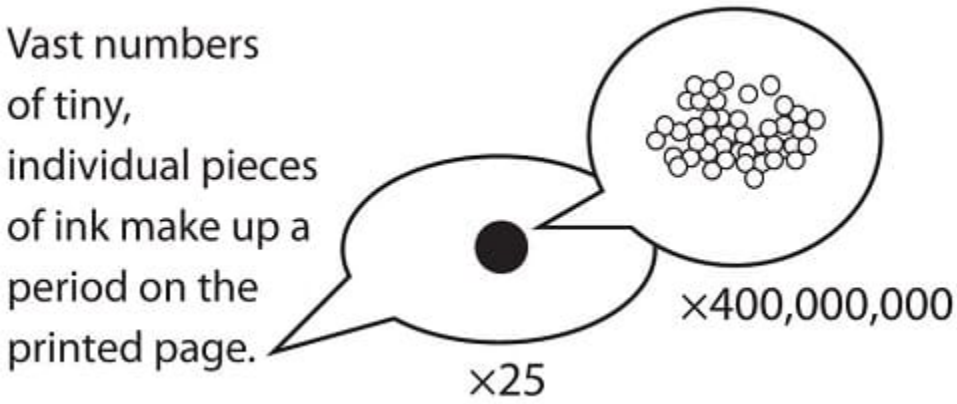


Figure 1.4.2: How Many Particles Are Needed for a Period in a Sentence?

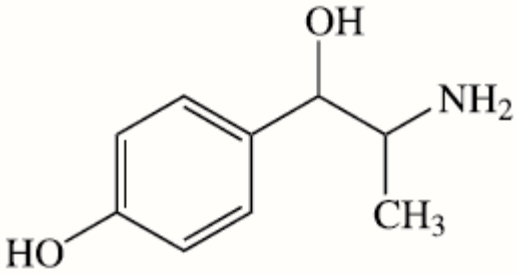
Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The ×25 and ×400,000,000 indicate the number of times the image is magnified.

Inorganic and Organic Compounds

Compounds can be further classified as being inorganic or organic. Organic compounds contain carbon, hydrogen, and usually another element from the right side of the periodic table. Be careful when thinking about organic chemicals. In the real-world, the term "organic" is typically used to describe products that have natural ingredients. In the world of science, the word "organic" is used to denote compounds that have carbon and hydrogen connections. Look at some of the structures below to get an appreciate of organic chemical structures.

Organic structures

The two molecules shown below are heavily used in the pharmaceutical industry. The compound on the left is known as hydrocodone. This is an opioid painkiller that is extremely addictive. It is regularly dispensed for surgical and chronic pain. The compound on the right is adderall. This medication is used to treat ADHD and narcolepsy. Both of these medications are controlled by the DEA (Drug Enforcement Agency) and are illegal to possess unless you have a valid prescription.



Hydrocodone skeletal.svg

Mixtures

A material composed of two or more substances is a mixture. In a mixture, the individual substances maintain their chemical identities. Many mixtures are obvious combinations of two or more substances, such as a mixture of sand and water. Such mixtures are called **heterogeneous mixtures**. In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Mixtures with a consistent composition throughout are called **homogeneous mixtures** (or **solutions**). Sugar dissolved in water is an example of a solution. A metal alloy, such as steel, is an example of a solid solution. Air, a mixture of mainly nitrogen and oxygen, is a gaseous solution.

Exercise 1.4.1

How would a chemist categorize each example of matter?

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

SOLUTION

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

Exercise 1.4.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a

Coffee could be a heterogeneous or homogeneous mixture. If insoluble parts are able to be identified, then it would be classified as being heterogeneous. If not, then coffee would be a homogeneous mixture or a solution.

Answer b

Hydrogen is an element. It can be located on the periodic table.

Answer c

An egg (assume raw) would be a heterogeneous mixture. The yolk and egg white are two different substances that can be easily seen.

Phases

Another way to classify matter is to describe it as a solid, a liquid, or a gas, which was done in the examples of solutions. These three descriptions, each implying that the matter has certain physical properties, represent the three phases of matter. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers. We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas).



Figure 1.4.3: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. Figure used with permission from Wikipedia

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a phase change, a physical process in which a substance goes from one phase to another (Figure 1.4.3). Phase changes have particular names depending on what phases are involved, as summarized in Table 1.4.1.

***Table** 1.4.1**:** Phase Changes*

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

Figure 1.4.4 illustrates the relationships between the different ways matter can be classified.

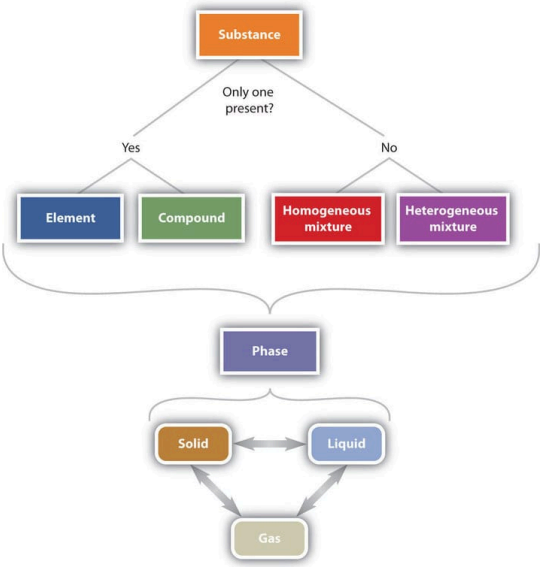


Figure 1.4.4: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties.

Exercise 1.4.3

a. Classify each as being element, compound, heterogeneous mixture, or homogeneous mixture (solution): air at Furman University, Furman Lake water, copper wire, and Furman tap water.
b. What is the difference between a heterogeneous mixture and a homogeneous mixture? Give an example of each.
c. Which is/are physical change(s): burning of fuel, digestion of food, salt dissolving in water, copper conducting electricity, and water boiling.

Answer a

Hopefully, Furman air is a homogenous mixture (solution). If there is no burning or dust build-up in the area, this is safe to assume. Furman Lake is definitely a heterogeneous mixture (in some areas it can get a bit swampy). Copper wire is an element. There are very few elements in nature. Lastly, Furman tap water is usually a homogeneous mixture. Natural and man-made chemicals are in our tap water (trust me, you want the chlorine). Our toilets do use grey water which can look heterogeneous (ok for sewer but not to drink).

Answer b

A heterogeneous mixture is obviously a mixture, such as dirt; a homogeneous mixture behaves like a single substance, such as saltwater.

Answer c

The physical changes are: salt dissolving water, copper conducting and water boiling. If composition is not changed, then this is a physical change. If there is a method to physically separate the components (like boiling, evaporation, distillation, magnetism, or filtration), then a physical change has occurred.

Key Takeaways

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

Contributors

- Anonymous
- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 1.4: Classification and Properties of Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

1.5: APPLYING PROPERTIES - MSDS

This page was auto-generated because a user created a sub-page to this page.

1.5.01: Applying Properties - MSDS

1.5.1: Applying Properties - MSDS/SDS

This page titled 1.5: Applying Properties - MSDS is shared under a license and was authored, remixed, and/or curated by LibreTexts.

1.5.01: APPLYING PROPERTIES - MSDS

This page was auto-generated because a user created a sub-page to this page.

1.5.1: Applying Properties - MSDS/SDS

This page titled 1.5.01: Applying Properties - MSDS is shared under a license and was authored, remixed, and/or curated by LibreTexts.

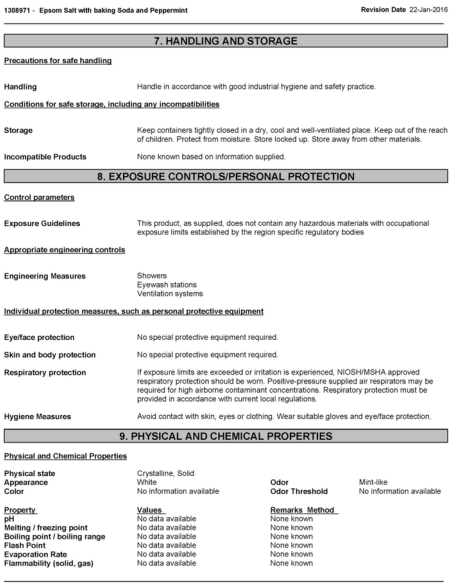
1.5.1: Applying Properties - MSDS/SDS

What is a MSDS/SDS?

A Material Safety Data Sheet (MSDS) or Safety Data Sheet (SDS) is designed to provide both workers and emergency personnel with the proper procedures for handling or working with a particular substance. MSDS's include information such as physical data (melting point, boiling point, flash point etc.), toxicity, health effects, first aid, reactivity, storage, disposal, protective equipment, and spill/leak procedures. These are of particular use if a spill or other accident occurs.

The MSDS's are available from many sources such as the chemistry store, health and safety offices, material producing companies etc. Any business or industry that has chemicals on site is required to have access to MSDS files for employers to be able to access at any time. If an employee is concerned about their exposure to chemicals, they can persuse a MSDS to obtain helpful information. Likewise, one could research household chemicals that could be dangerous to humans and animals. Many substances found around the home (pesticides, fuels, and cleaners) can be hazardous or harmless.

A partial MSDS of magnesium sulfate (bath salts used for soaking that are not illegal substances is illustrated below. This particular page only shows chemical and physical properties of this substance. This would give one an idea of what magnesium sulfate looks like and how it should react to other substances.



1308971 - **Epsom Salt with baking Soda and Peppermint** **Revision Date** 22-Jan-2016

7. HANDLING AND STORAGE

Precautions for safe handling

Handling Handle in accordance with good industrial hygiene and safety practice.

Conditions for safe storage, including any incompatibilities

Storage Keep containers tightly closed in a dry, cool and well-ventilated place. Keep out of the reach of children. Protect from moisture. Store locked up. Store away from other materials.

Incompatible Products None known based on information supplied.

8. EXPOSURE CONTROLS/PERSONAL PROTECTION

Control parameters

Exposure Guidelines This product, as supplied, does not contain any hazardous materials with occupational exposure limits established by the region specific regulatory bodies

Appropriate engineering controls

Engineering Measures Showers
Eyewash stations
Ventilation systems

Individual protection measures, such as personal protective equipment

Eye/face protection No special protective equipment required.

Skin and body protection No special protective equipment required.

Respiratory protection If exposure limits are exceeded or irritation is experienced, NIOSH/MSHA approved respiratory protection should be worn. Positive-pressure supplied air respirators may be required for high airborne contaminant concentrations. Respiratory protection must be provided in accordance with current local regulations.

Hygiene Measures Avoid contact with skin, eyes or clothing. Wear suitable gloves and eye/face protection.

9. PHYSICAL AND CHEMICAL PROPERTIES

Physical and Chemical Properties

Physical state	Crystalline, Solid		
Appearance	White	**Odor**	Mint-like
Color	No information available	**Odor Threshold**	No information available

Property	**Values**	**Remarks Method**
pH	No data available	None known
Melting / freezing point	No data available	None known
Boiling point / boiling range	No data available	None known
Flash Point	No data available	None known
Evaporation Rate	No data available	None known
Flammability (solid, gas)	No data available	None known

Other Terms Used in MSDS/SDS

Aspects of toxicity and how chemicals will affect a body can also be found in this resource. **Along with physical and chemical properties, you should be familiar with the terms below that could also be found in a MSDS/SDS document.**

- **Mutagen:** able to alter DNA and could result in carcinogenesis (chemical example would be nitrosamine)
- **Carcinogen:** able to cause cancer (this can get complicated, animal and human carcinogens both exist). There are very few chemicals classified as being human carcinogens. (benzene, radon, asbestos)
- **Neurotoxin:** affects the nervous system (chemical examples include lead and mercury)

Figure 1.5.1.1: Symptoms of lead poisoning. Image used with permission (Public Domain; Mikael Häggström).

- **Teratogen**: affects developing fetus or embryo (pharmaceutical example would be thalidomide)
- **Corrosive**: can chemically destroy things easily (acids/bases are examples)
- **Flammable/Combustible**: able to catch on fire (examples would include gasoline and acetone)

- **Volatile**: a liquid that can transform to a gas easily (examples would include gasoline and acetone)

MSDS/SDS's can be several pages long. A complete MSDS/SDS guide for acetic acid (also known as vinegar) can be viewed by clicking on this here. Although this chemical might appear dangerous, it used in the production of vinegar, ketchup, and barbeque sauce. *After looking over vinegar's entire MSDS/SDS sheet, write down any questions for your instructor regarding unfamiliar terminology.*

Exercise 1.5.1.1

After looking at the MSDS for acetic acid, answer the questions below:

a. Is this chemical a mutagen?
b. Does acetic acid react with metals?
c. What odor does acetic acid produce?
d. If this chemical is ingestion, should you induce vomiting to remove it from the body?
e. Does this chemical have a color?

Answer a

Acetic acid is not a mutagen. It will not alter DNA.

Answer b

Acetic acid reacts with metals. This is a chemical property of this substance.

Answer c

Acetic acid smells vinegar.

Answer d

Acetic acid is very corrosive and will eat biological tissues. If swallowed, vomiting should not be induced.

Answer e

This chemical is colorless

Contributors

- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)
- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 1.5.1: Applying Properties - MSDS/SDS is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

1.6: All About the Elements

Basic Arrangement of Periodic Table

In Chapter 3, we will delve more into the initial construction and modifications of today's periodic table. For now, you need to be aware that the periodic table has columns (known as **families/groups**). Horizontal rows are called **periods**. On the table below, the element symbols are either one of two letters. Previously, some of the newly discovered elements (see period seven on other tables) used three letter symbols. These three letter symbols corresponded to a Latin numbering system until the elements were given official names.

Languages of Element Names

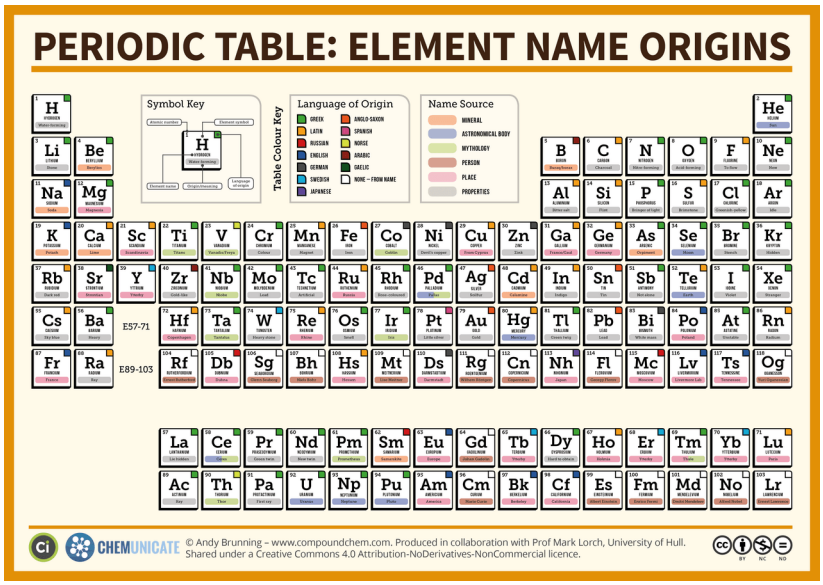
Interestingly, different languages are used on the periodic table. Depending on your background, you might be more familiar with some elemental symbols than others. *For* example, Hydrogen is represented by the symbol H. This element's name is derived from the Greek hydro, and genes which means to generate water. Other elemental symbols like copper (which has a symbol of Cu) are not so predictable. This particular element's name originated from the Latin *cuprum* (from the island of Cyprus). A variety of languages are displayed on the periodic table. For example, German roots are present in the element Tungsten. The symbol for this element is W which corresponds to the German word, wofram. . **For this particular class, you will find a list of required elements/symbols that need to be memorized. Please locate this document on the moodle page for this course.**

Meanings of Element Names

Many elements are named in honor of people. Some familiar names you might see are curium (Marie and Pierre Curie), einstenium (Albert Einstein), and medelelvium (Dimitri Mendeleev). IUPAC (International Union of Pure and Applied Chemistry) requires these types of names to be done posthumously. In other words, the scientist who discovered the element could not name it after her/himself or anyone who is still currently living. Over the years, there have two exceptions. The elements of seaborgium (Sg) and oganesson (Og) were named while Glen Seaborg and Yuri Oganessian were still living. Seaborgium was named in honor of an American nuclear chemist who synthesized the element plutonium. He is also credited with assisting the discovery of other elements like c, americium (element 95), curium (element 96), and fermium (element 100), . His accomplishments encouraged his fellow researcher, Albert Ghiorso, to name Seaborgium in honor of him. As for element 118 (oganessian), a joint research commission of Russian and American scientists named this element for Russian physcist, Yuri Oganessian. As of 2017, Oganessian is still living and is responsible for the discovery of super heavy elements. **For test 1, please research an element that has not been mentioned on this page or on your list to be memorized. By accessing this link, information about a selected an element can be obtained. You will need provide the symbol and interesting factoid for five point question.**

It is acceptable to name an element in honor of a geographical location. Some notable Earth locations would include: Am (americium, American continent), Ge (germanium, Germany), Fr (francium, France), and Eu (europium, European continent). Some scientists have ventured past Earth to name their elements in honor of other places. Examples of these would include Pu (plutonium, Pluto), U (uranium, Uranus), and Np (neptunium, Neptune).

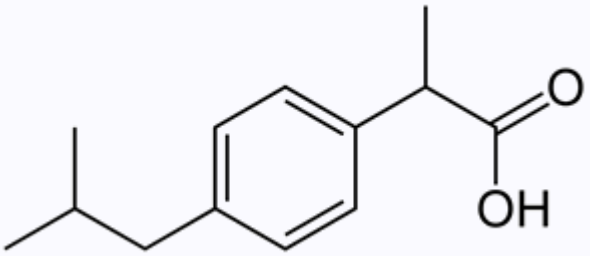
Other options for naming elements are shown in the table below. Recently, elements 113-118 received their official names. **Watch this ACS video to get some background regarding these new members of the table.**



Example 1.6.1

List the element names contained in the compounds below. Classify each compound as being organic or inorganic.

a. Na_2SO_4
b. HCl
c. $Cu(NO_3)_2 \cdot H_2O$



d.
e. CO_2

SOLUTION

a. sodium, sulfur, and oxygen. This compound is inorganic.
b. hydrogen and chlorine. This compound is inorganic
c. copper, nitrogen, oxygen, and hydrogen. This compound is inorganic.
d. The corners of each line drawing represent carbons. You can assume for the single lines, the carbons have hydrogens attached (organic chemists tend not to draw all atoms). This compound is organic and contains carbon, hydrogen and oxygen.
e. carbon and oxygen. This compound is inorganic.

Example 1.6.2

Locate and name all the elements you see below in the compounds that are given below. Circle the organic compounds.

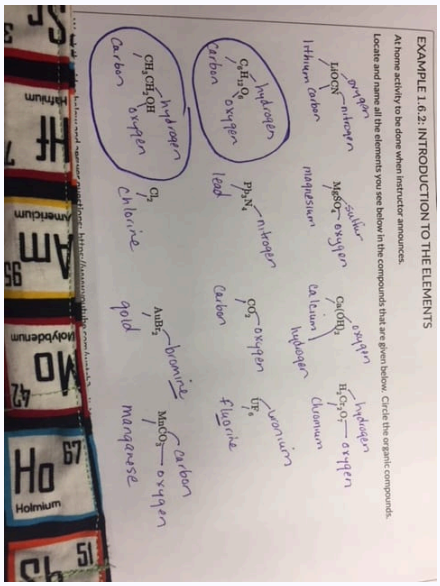
$LiOCN$	$MgSO_4$	$Ca(OH)_2$	$H_2Cr_2O_7$
$C_6H_{12}O_6$	Pb_3N_4	CO_2	UF_6
Ch_2CH_2OH	Cl_2	$AuBr_2$	$MnCO_3$

Watch the video below and answer questions: https://www.youtube.com/watch?v=ItrYpHsymfI

a. How were the last four elements made?
b. How many total elements do we have now?
c. How many elements are named in honor of female scientists?
d. What would you name your new element and what would its symbol might be?

Answer

(elements only)



Contributors

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 1.6: All About the Elements is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2: THE MATHEMATICS OF CHEMISTRY

Mathematics is used widely in chemistry and are absolutely necessary to explore important concepts in chemistry. Without some basic mathematics skills, these calculations, and therefore chemistry itself, will be extremely difficult. However, with a basic knowledge of some of the mathematics that will be used in your chemistry course, you will be well prepared to deal with the concepts and theories of chemistry.

2.1: Measuring Matter

2.2: Dimensional Analysis

2.3: Density

2.4: Temperature

2.E: The Mathematics of Chemistry (Exercises)

This page titled 2: The Mathematics of Chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2.1: Measuring Matter

Template:HideTOC

A variety of instruments are available for making direct measurements of the macroscopic properties of a chemical substance. For example, we usually measure the **volume** of a liquid sample with pipettes, burets, graduated cylinders, and volumetric flasks (Figure 2.1.1), whereas we usually measure the mass of a solid or liquid substance with a balance (Figure 2.1.2).

Instruments that Measure Volume

Many academic scientists report their volume measurements with the milliliter unit. On the other hand, analytical chemists tend to work with much smaller volumes, and they would use nanoliter and microliter samples in their laboratory. When considering the industrial realm, scientists work with larger volumes and would thus employ the liter unit of measurement.

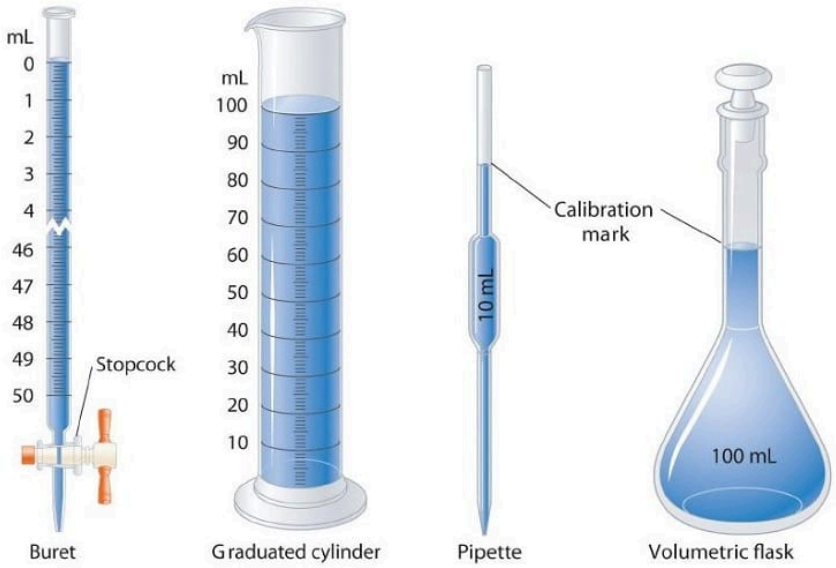


Figure 2.1.1: (left two) Graduated glassware is used to deliver variable volumes of liquid. (right two) Volumetric glassware is used to deliver (pipette) or contain (volumetric flask) a single volume accurately when filled to the mark. (CC BY-SA-NC; anonymous).

Instruments that Measure Mass

Balances are utilized to measure the **mass** of matter. Typically, high school and middle schools will use triple beam balances for laboratory activities. Colleges will rely on digital balances. For larger quantities of materials, top-loading digital balances for employed. For higher sensitivity and accuracy, some chemists use analytical balances.



Figure 2.1.2: Triple beam balance (image from: https://upload.wikimedia.org/wikiped...m_balance2.JPG. Digital top-loading balance (https://upload.wikimedia.org/wikiped...00-2M_-_01.jpg) and an analytical balance (https://upload.wikimedia.org/wikiped...-NaCl-1mol.jpg)

In the academic laboratory (like Furman University's), a chemist will commonly use the gram, or the even smaller unit, of the milligram. Industrial chemists making larger quantities of material would work with kilograms of materials. Analytical chemists (industrial or academic), environmental scientists, and toxicologists are typically more concerned with the smaller units of the metric system. For example, many chemicals found in drinking water and air are regulated in concentrations like nanograms and or micrograms.

Tools that Measure Length

Length is measured using a metric stick or a metric ruler marked in millimeters and centimeters. Most devices used to measure length (in the United States) contain both English and metric markings.

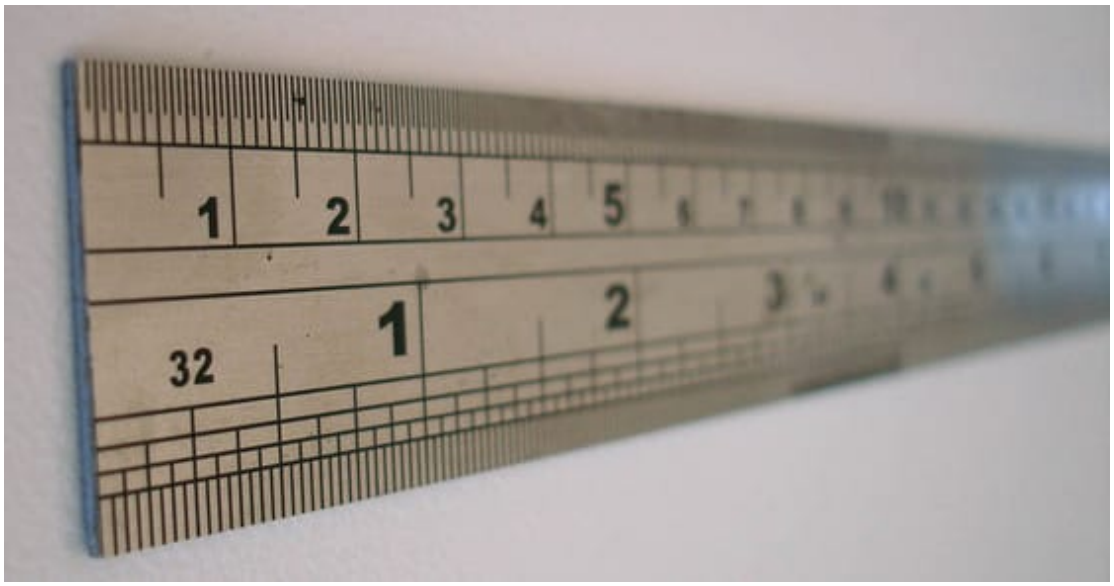


Figure 2.1.3: Metric and English measuring devices. Image used with permission (CC BY-SA 4.0; Ejay).

SI Units

All reported measurements must include an appropriate unit of measurement. To say that a substance has "a mass of 10" does not tell whether the mass was measured in grams, pounds, tons, or some other unit. For example, a nurse giving morphine to a patient needs to be specific in the unit dosages. A volume of liters would kill a patient quickly while a milliliter amount

would provide sufficient pain relief. To establish worldwide standards for the consistent measurement of significant physical and chemical properties, an international body called the General Conference on Weights and Measures devised the Système Internationale d'unités (or SI). The International System of Units is based on metric units and requires that measurements be expressed in decimal form. Table 2.1.1 lists the seven base units of the SI system; all other SI units of measurement are derived from them.

***Table** 2.1.1**:** SI Base Units*

Base Quantity	Unit Name	Abbreviation
mass	kilogram	kg
length	meter	m
time	second	s
temperature	Kelvin	K

The Metric System

***Table** 2.1.2**:** Prefixes Used with metric bases of meter, liter, and gram*

Prefix	Symbol	Value	Power of 10	Meaning
tera	T	1,000,000,000,000	10^{12}	trillion
giga	G	1,000,000,000	10^{9}	billion
mega	M	1,000,000	10^{6}	million
kilo	**k**	**1000**	**10^{3}**	**thousand**
hecto	h	100	10^{2}	hundred
deca	da	10	10^{1}	ten
—	—	1 (this is base unit which could be gram, liter, or meter)	10^{0}	one
deci	d	0.1	10^{-1}	tenth
centi	**c**	**0.01**	**10^{-2}**	**hundredth**
milli	**m**	**0.001**	**10^{-3}**	**thousandth**
micro	**μ (medical community uses mcg)**	**0.000001**	**10^{-6}**	**millionth**
nano	**n**	**0.000000001**	**10^{-9}**	**billionth**
pico	p	0.000000000001	10^{-12}	trillionth
femto	f	0.000000000000001	10^{-15}	quadrillionth



Figure 2.1.4: Pharmaceuticals use a different notation for the microgram unit than traditional chemists. Image used with permission (CC BY-SA 3.0; Crohnie).

Two methods of Memorization for the Metric System

Using negative exponents (see above) is the traditional way to memorize the metric system. Many students who have been taught outside of the United States approach the metric system this way. For their conversion factors, students who have been taught the metric system in the United States tend to use positive exponents. Either method is fine, pick one method and use it consistently.

Two different ways to memorize the metric systems (be sure to know all bolded conversion factors listed in section 2.1 of this OER)

- 1 gigabase = 1 x10^{9} base
- 1 megabase = 1x10^{6} base
- **1 kilobase = 1x10^{3}base**
- 1 decibase = 1x10^{-1}base or 1x10^{1} decibase = 1 base
- **1 centibase = 1x10^{-2}base or 1x10^{2} centibase = 1 base**
- **1 millibase = 1x10^{-3}base or 1x10^{3} millibase = 1 base**
- **1 microbase = 1x10^{-6}base or 1x10^{6} microbase = 1 base**
- **1 nanobase = 1x10^{-9}base or 1x10^{9} nanobase = 1 base**
- 1 picobase = 1x10^{-12} base or 1x10^{12} picobase = 1 base

Note: A base unit can be gram, liter, or meter

Additional Units

When traveling to different regions of the globe, it is helpful to be able to convert from English to Metric units. **The bolded conversions below may be helpful in travel and other situations. You may want to memorize them**.

Unit Conversion
2.21 pounds = 1.00 kilograms
1 mile = 1.61 kilometers

1.00 inch = 2.54 centimeters
1.00 mL = 1.00 cc (cubic centimeters) = 1.00 cm^3
1 teaspoon (tsp) = 5.00 mL
12 in = 1.00 ft.
16 oz. = 1.00 pounds

Scientific Notation

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

$$N \times 10^n \tag{2.1.1}$$

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

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

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

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases, and vice versa. The application of this rule is illustrated in Example 2.1.1.

Example 2.1.1: scientific notation

Convert each number to scientific notation.

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

SOLUTION

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

Explanation	Solution
To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 4.79 Because the decimal point was moved two places to the right, n = –2.	$0.0479 = 4.79 \times 10^{-2}$
This is usually expressed simply as 7.86. (Recall that $10^0 = 1$).	7.86×10^0
Move the decimal point four places to the left and note a positive exponent.	1.2378×10^4
Move the decimal place four places to the right a note a negative exponent.	3.2×10^{-4}
Move the decimal place one space to the left and note a positive exponent.	6.106700×10^1
Move the decimal place three places to the left and note a positive exponent.	2.002080×10^3
Move the decimal place two places to the right and note a negative exponent.	1.020×10^{-2}

Using the Scientific Calculator

Please be sure to obtain a calculator that has an exponent button on it. It will either say "exp" or have an "ee" key indicating this function. This key represents the function of 10^x. You do not need to enter a 10 for an exponent when using this particular function. Read the steps below and practice on your calculator.

Use of Calculator

To input powers of ten:

1. Locate the ee or exp button on your calculator. Do not use the carrot (looks like the ^ button).
2. Put the coefficient number in and then hit "ee" or "exp."
3. If you need to have a negative "exp," then locate the +/- button.
4. Lastly, input your power of ten. You do not need to put the actual number ten in for the exponent.

Table

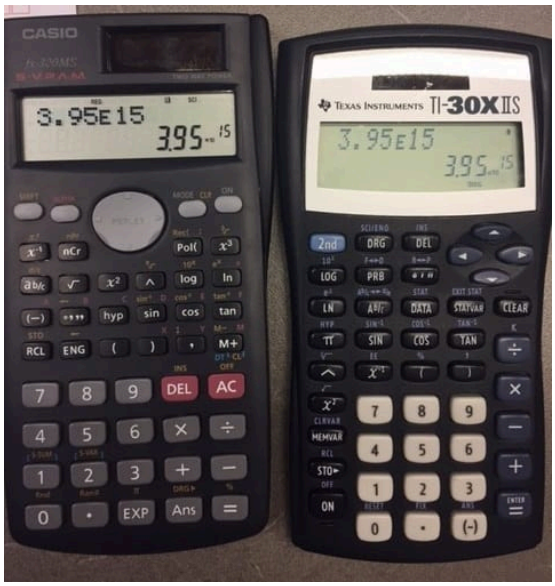
:
Examples

Number	How to Input in calculator
3.95×10^{15}	Put 3.95 in as the coefficient number. Hit the "ee" or "exp" button. Lastly, enter the 15 for the exponent. Your calculator should say $3.95e^{15}$.

Number	How to Input in calculator
9.26 x 10^{-4}	Put 9.26 in as the coefficient number. Hit the "ee" or "exp" button. Locate the small +/- button for negative (not minus function). Next, enter the number four. Your calculator should say $9.26e^{-4}$.

Suggested Calculators

This is an image of calculators you could use in this class. You will need a non-programmable, scientific calculator that does not have metric conversion functions. In this picture, you will see the correct way to input one of the numbers shown in the example above. Please take time to practice with your calculator.



Contributors

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 2.1: Measuring Matter is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2.2: Dimensional Analysis

Learning Objective

- Convert from one unit to another unit of the same type.
- Understand how to use Conversion Factors

Conversion Factors

Recall your metric table while working these types of problems.

Two methods of Memorization of Metric

Using negative exponents (see above) is the traditional way to memorize the metric system. Many students who have been taught outside of the United States approach the metric system this way. For their conversion factors, students who have been taught the metric system in the United States tend to use positive exponents. Either method is fine, just pick one way and use it consistently.

Two different ways to memorize the metric systems (be sure to know all bolded conversion factors listed in Section 2.1)

- 1 gigabase = 1 x10^{9} base
- 1 megabase = 1x10^{6} base
- **1 kilobase = 1x10^{3} base**
- 1 decibase = 1x10^{-1} base or 1x10^{1} decibase = 1 base
- **1 centibase = 1x10^{-2} base or 1x10^{2} centibase = 1 base**
- **1 millibase = 1x10^{-3} base or 1x10^{3} millibase = 1 base**
- **1 microbase = 1x10^{-6} base or 1x10^{6} microbase = 1 base**
- **1 nanobase = 1x10^{-9} base or 1x10^{9} nanobase = 1 base**
- 1 picobase = 1x10^{-12} base or 1x10^{12} picobase = 1 base

Note: A base unit can be gram, liter, or meter

We have an expression,

$$\frac{3\ ft}{1\ yd} = 1 \tag{2.2.1}$$

Equation 2.2.1 is a strange way to write 1, but it makes sense: 3 ft equal 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units. The expression in Equation 2.2.1 can be is called a conversion factor and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

To see how this happens, let us start with the original quantity: 4 yd

Now let us multiply this quantity by 1. When you multiply anything by 1, you don't change the value of the quantity. Rather than multiplying by just 1, let us write 1 as

$$\frac{3\ ft}{1\ yd} \tag{2.2.2}$$

$$4\,yd \times \frac{3\,ft}{1\,yd} \tag{2.2.3}$$

The 4 yd term can be thought of as 4 yd/1; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. In this case, what cancels is the unit *yard*:

$$4\,\cancel{yd} \times \frac{3\,ft}{1\,\cancel{yd}} \tag{2.2.4}$$

That is all that we can cancel. Now, multiply and divide all the numbers to get the final answer:

$$\frac{4 \times 3\,ft}{1} = \frac{12\,ft}{1} = 12\,ft \tag{2.2.5}$$

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is applicable to a variety of problems.

How many millimeters are in 14.66 m? To answer this, we need to construct a conversion factor between millimeters and meters and apply it correctly to the original quantity. One method would entail using the conversion factor of $1\ mm = 1 \times 10^{-3}\ m$

$$14.66\,\cancel{m} \times \frac{1\,mm}{10^{-3}\,\cancel{m}} = 14660\,mm \tag{2.2.6}$$

$$= 1.46 \times 10^{4}\,mm \tag{2.2.7}$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is meters, which we want to convert to millimeters. Because the original unit is assumed to be in the numerator, to get rid of it, we want the meter unit in the *denominator*; then they will cancel. Using the positive metric exponents, we have an alternative method to calculate the same answer as shown previously. Canceling units and performing the mathematics, we get:

$$14.66\,\cancel{m} \times \frac{1000\,mm}{1\,\cancel{m}} = 14660\,mm \tag{2.2.8}$$

$$= 1.46 \times 10^{4}\,mm \tag{2.2.9}$$

Note how m cancels, leaving mm, which is the unit of interest.

The ability to construct and apply proper conversion factors is a very powerful mathematical technique in chemistry. You need to master this technique if you are going to be successful in this and future science courses.

Example 2.2.1

a. Convert 35.9 kL to liters.
b. Convert 555 nm to meters.

Solution

a. We will use the fact that 1 kL = 1,000 L. Of the two conversion factors that can be defined, the one that will work is 1000L/1kL. Applying this conversion factor, we get:

$$35.9\,\cancel{kL} \times \frac{1000\,L}{1\,\cancel{kL}} = 35,900\,L$$

b. We will use the fact that 1 nm = 1/1,000,000,000 m, which we will rewrite as 1,000,000,000 nm = 1 m, or 10^9 nm = 1 m. Of the two possible conversion factors, the appropriate one has the nm unit in the denominator:

$$\frac{1\,m}{10^{9}\,m}$$

Applying this conversion factor, we get

$$555 \cancel{nm} \times \frac{1\,m}{10^9 \cancel{nm}} = 0.000000555\,m$$

$$= 5.55 \times 10^{-7}\,m$$

In the final step, we expressed the answer in scientific notation.

Exercise 2.2.1

a. Convert 67.08 µL to liters.
b. Convert 56.8 m to kilometers.

Answer a

6.708×10^{-5} L

Answer b

5.68×10^{-2} km

Exercise 2.2.2

How many milliliters are in 607.8 kL?

Answer

This problem involves multiple steps that require going through the metric base of liters. Think of a pathway of a conversion before setting up these types of problems. The suggested pathway would be mL ⇒ L ⇒ kL

$$607.8\,kL \times \frac{1000\,L}{1\,kL} = 6.08 \times 10^5\,L$$

$$= 6.08 \times 10^5\,L \frac{1000\,mL}{1\,L}$$

$$= 608000000\,mL$$

$$= 6.08 \times 10^8\,mL$$

Exercise 2.2.3

Kathleen Hsu Barron (Furman University, Class of 2011) is exploring Switzerland with her husband Patrick Barron (Furman University, Class of 2011). The sign she is pointing to is a typical speed in km/hr. Convert this distance to miles/hr and feet/hr.



Answer

Again, this problem is a multi-step problem. There a mutiple ways that one could proceed to solve this problem. With the coversion factors in this course, the suggested round would be miles ⇒ kilometers ⇒ meters ⇒ centimeters ⇒ inches ⇒ feet. Another method could utilize miles ⇒ feet.

18.6 miles/hr, 9.8 x 10^4 ft/hr

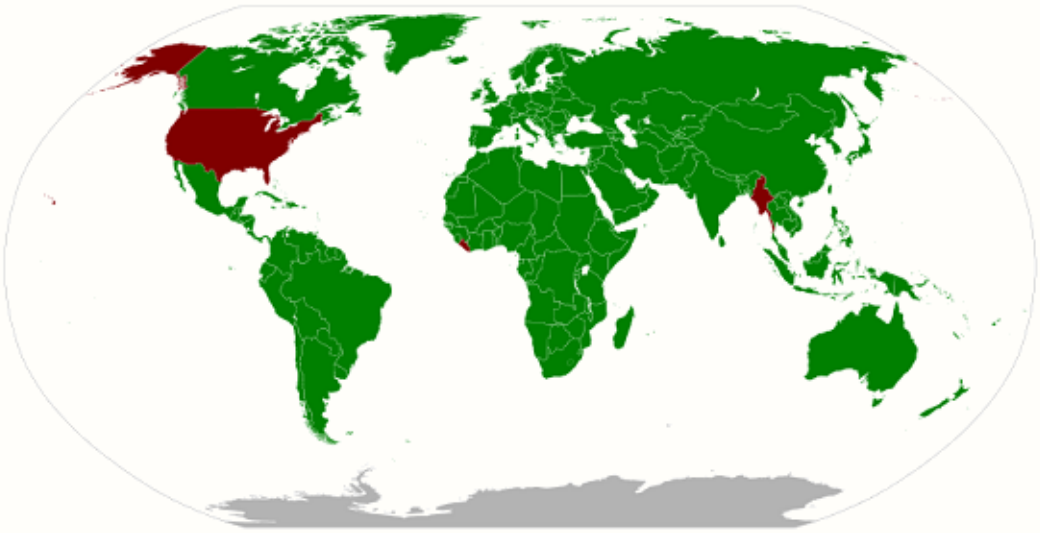
In this course, we will not worry about significant figures. You will need to read the directions on your assignment, quiz, or test and round accordingly. If you take other physical sciences, be aware of how you need to record and report the data.

Applications of conversions and the Metric System

The skill of converting units can be helpful in many aspects of your life. Measuring medications can be done precisely by using milliliters instead of teaspoons (recall 1 tsp = 5.0 mL). Of course, a pharmacist would provide you with the appropriate syringe if this was needed. Sizes of syringes can vary from 1.0 to 10.0 mL for most at home applications.

Marathons are measured using the kilometer unit. If you were raised outside the United States, you would be quite familiar with how long a kilometer truly is. To appreciate this distance, most Americans need to convert this to a mileage by dividing the distance by 1.61 km.

When traveling abroad, keep in mind the majority of countries (except Myanmar and Liberia) commonly use the metric system in day to day living. Understanding how to convert Metric to English units will help you to appreciate the measurements used by a different society.



Adoption of the metric system of measurement. Countries which have officially adopted the metric system (green) and countries which have not officially adopted the metric system (US, Myanmar, Liberia)

Why has the United States not adopted the metric system as a country? Watch this video and see if you can derive some appropriate hypotheses.

Need More Practice?

- Turn to Section 2.E of this OER and work problems #1, #3, #5, #6, and #8.

Contributors

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)

This page titled 2.2: Dimensional Analysis is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2.3: Density

Template:HideTOC

Density is a physical property that is defined as a substance's mass divided by its volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{2.3.1}$$

$$d = \frac{m}{V} \tag{2.3.2}$$

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has *derived units*, just like velocity. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances are listed in Table 2.3.1. **Memorize the density of water with its appropriate units**.

Table 2.3.1: Densities of Some Common Substances

Substance	Density (g/mL or g/cm^3)
water	**1.0**
gold	19.3
mercury	13.6
air	0.0012
cork	0.22–0.26
aluminum	2.7
iron	7.87



Figure 2.3.1: Separatory Funnel containing oil and colored water to display density differences. (CC BY-SA 3.0; PRHaney vai wikipedia).

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table 2.3.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

Start with Equation 2.3.1

$$\text{density} = \frac{m}{V}$$

and insert the relavant numbers

$$\frac{2.7g}{cm^3} = \frac{m}{7.88\,cm^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$7.88\,\cancel{cm^3} \times \frac{2.7\,g}{\cancel{cm^3}} = 21g \text{ of aluminum}$$

Example 2.3.1: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density value for mercury from Table 2.3.1 and the definition of density (Equation 2.3.1)

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

$$\frac{13.6g}{mL} = \frac{m}{44.6\,mL}$$

Remember to cross multiply here in order to isolate variable. Then, report answer with correct units.

$$44.6\,\cancel{mL} \times \frac{13.6\,g}{\cancel{mL}} = 607\,g$$

The mass of the mercury is 607 g.

Exercise 2.3.1

What is the mass of 25.0 cm^3 of iron?

Answer

Use the density value for iron from Table 2.3.1

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

$$\frac{7.87g}{cm^3} = \frac{m}{25.0\,cm^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$25.0\,\cancel{cm^3} \times \frac{7.87\,g}{\cancel{cm^3}} = 197g \text{ of iron}$$

Another way of looking at density (some students choose to perform calculations using this method)

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already

demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3\,g/mL = \frac{19.3\,g}{mL}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator:

$$d = 19.3\,g/mL = \frac{19.3\,g}{mL}$$

That is, the density value tells us that we have 19.3 grams for every 1 milliliter of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched—that is, we must take the *reciprocal* of the density. In so doing, we move not only the units but also the numbers:

$$\frac{19.3\,\cancel{g}}{mL} = \frac{45\,\cancel{g}}{V}$$

Cross multiplying denominators with numerators, we obtain the following algebraic equation.

$$19.3V = 45.9\,mL$$

then you will need to isolate the variable (volume)

$$V = \frac{45.9\,mL}{19.3}$$

After multiplication, the answer would be

$$V = 2.38\,mL$$

Example 2.3.2: Wine Cork

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/mL, what is the volume of the cork? *Regardless of the method that is used, you should still be able to obtain the same (and correct) answer.*

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find

$$\frac{0.22\,\cancel{g}}{mL} = \frac{3.78\,\cancel{g}}{V}$$

Cross multiplying denominators with numerators, we obtain the following algebraic equation.

$$0.22V = 3.78\,mL$$

then you will need to isolate the variable (volume)

$$V = \frac{3.78\,mL}{0.22}$$

so, the volume of the cork is 17.2 mL.

Exercise 2.3.2

What is the volume of 3.78 g of gold?

Answer

Before attempting this question, be sure to obtain the density of gold in the table above. If you were to need this value on a quiz or a test, then it would be provided. Once you have this value, plug it into the density equation. Next, you will need to isolate the volume variable (basic algebra). The final answer should be 0.196 cm^3.

Care must be used with density as a conversion factor. Make sure the mass units are the same or the volume units are the same, before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Using Density in Environmental Applications

Along with solubility, density can help determine how a contaminant could affect an aquatic system. For example, imagine mercury has been spilled in Furman Lake. Looking at this element's density value and comparing it to liquid water, one could determine the location of the insoluble (you would be given solubility information) mercury layer. The more dense mercury layer would reside on the bottom of Furman Lake. If one were to take a cross-section of the lake, you could see that a heterogeneous mixture would result.

In contrast, spilling ethanol (density = 0.789g/mL) would result in the formation of a homogeneous mixture. Ethanol (grain alcohol) is soluble in water. This would make it miscible (mixable to form a solution) in water and one would not be able to denote separate layers. According to the density, an alcohol layer would remain on top, but would ultimately dissolve.

Applications

What difficulties would arise from the separation and removal of contaminants?

1. Hg in Furman Lake
2. Ethanol in Furman Lake
3. oil (less dense, insoluble) in Furman Lake

Watch this video and record your observations.

1. What component was different in the two types of beverages (mass or volume)?
2. How does the above-mentioned difference affect the density equation?
3. Which beverage is denser than water?

Need More Practice?

- Turn to Section 2.E of this OER and work problems #2 and #9.

Contributors

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University, Class of 2018)

This page titled 2.3: Density is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2.4: Temperature

Learning Objectives

- Learn about the various temperature scales that are commonly used in chemistry and other cultures.
- Memorize temperature equations for **Celsius and Fahrenheit conversions**.
- Differentiate between various thermometers (alcohol or mercury-based).
- Understand how body temperature can vary.

Temperature Scales

One of the fundamental quantities in science is temperature, which is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees. There are three scales used for reporting temperatures. Figure 2.4.1 compares the three temperature scales.

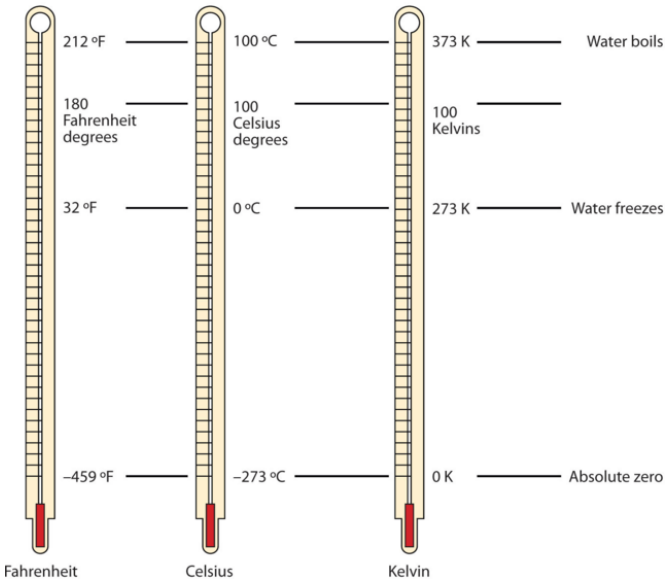


Figure 2.4.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales. These thermometers have a red-colored liquid which contains alcohol and food coloring. Silver looking thermometers contain mercury, which is a neurotoxin.

In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as "degrees Fahrenheit"). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32 °F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212 °F.

Science also uses other scales to express temperature. For example, the Celsius scale (symbolized by °C and spoken as "degrees Celsius") defines 0°C as the freezing point of water and 100 °C as the boiling point of water. This scale is divided into 100 divisions between these two landmarks and extended higher and lower as well. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}C = (^{\circ}F - 32) \times \frac{5}{9} \tag{2.4.1}$$

$$^{\circ}F = \left(^{\circ}C \times \frac{9}{5}\right) + 32 \tag{2.4.2}$$

Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-

measured temperatures with the Fahrenheit scale. (In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.)

Example 2.4.1: Conversions

a. What is 98.6°F in degrees Celsius?
b. What is 25.0°C in degrees Fahrenheit?

Solution

a. Using Equation 2.4.1, we have

$$\begin{aligned} {}^{\circ}C &= (98.6 - 32) \times \frac{5}{9} \\ &= 66.6 \times \frac{5}{9} \\ &= 37.0^{\circ}C \end{aligned}$$

b. Using Equation 2.4.2, we have

$$\begin{aligned} {}^{\circ}F &= (25.0 \times \frac{9}{5}) + 32 \\ &= 45.0 + 32 \\ &= 77.0^{\circ}F \end{aligned}$$

For more examples of how to perform these types of problems, click on this video to see your professor in action.

Exercise 2.4.1

a. Convert 0°F to degrees Celsius.
b. Convert 212°C to degrees Fahrenheit.

Answer a

–17.8°C

Answer b

414°F

The fundamental unit of temperature in SI is the Kelvin (K). The Kelvin temperature scale (note that the name of the scale capitalizes the word *Kelvin*, but the unit itself is lowercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$K =^{\circ} C + 273.15 \tag{2.4.3}$$

For most purposes, it is acceptable to use 273 instead of 273.15 in Equation 2.4.3.

Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as “two hundred ninety-five kelvin” and not “two hundred ninety-five degrees Kelvin.”

The reason that the Kelvin scale is defined this way is that there exists a minimum possible temperature called absolute zero

(zero Kelvin). The Kelvin temperature scale is set so that 0 K is absolute zero, and the temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

Example 2.4.2: Room Temperature

If the normal room temperature is 72.0°F, what is room temperature in degrees Celsius and kelvin?

Solution

First, we use Equation 2.4.1 to determine the temperature in degrees Celsius:

$$\begin{aligned} {}^{\circ}C &= (72.0 - 32) \times \frac{5}{9} \\ &= 40.0 \times \frac{5}{9} \\ &= 22.2^{\circ}C \end{aligned}$$

Then we use Equation 2.4.3 to determine the temperature in the Kelvin scale:

$$\begin{aligned} K &= 22.2^{\circ}C + 273.15 \\ &= 295.4K \end{aligned}$$

So, room temperature is about 295 K.

Health APPlication: Body temperature

Normal body temperature is defined as being 98.6°F (+/- 1.0°F). To determine body temperature, thermometers can be placed inside or on the surface of the body. The two best methods of obtaining body temperature are placing the thermometer either under the tongue or inside the rectum. Typically, children are capable of holding a thermometer in their mouths around the age of four (have fun before that age).

Fever is defined as body temperature being above 100°F (adults). High fever status occurs at and above 104°F. For adults, these adults should seek medical attention immediately if the fever exceeds 104°F. For children, these values are much lower.

Hyperthermia (inability of the body to regulate heat) occurs when normal body temperature is exceeded. Conditions that can cause hyperthermia are fever (infection), heat stroke, thyroid disorders, heart attack, or traumatic injury. Medications for cancer, arthritis, and thyroid patients can cause the body temperature to rise. Symptoms of hyperthermia include sweating, confusion, nausea, and dizziness.

Hypothermia (exposure to cold environments) occurs when normal body temperature dips below 95.0°F. When this occurs, the affected person(s) should seek immediate medical attention. During hypothermia, the body has problems producing heat. Medical conditions like diabetes, infection, and thyroid dysfunction can cause hypothermia. Watch this video of American Marines attempting to survive extreme environments. Symptoms of this condition involve shivering, confusion, and sluggish behavior.

Treating hyperthermia could involve hydrating a patient. If a fever is due to infection, analgesics (like Tylenol, Advil, aspirin, or Aleve) can help alleviate fever. Placing someone in a cool bath can also relieve symptoms.

To combat hypothermia, one must remove wet clothing, redress in warm materials, and participate in physical movement.

The Science of Cryogenics

Cryogenics is branch of physics that incorporates the production and effects of substances at temperatures ranging between -150°C to -273°C. In the late 1870's, cryogenic science began with the cooling of gaseous oxygen. At -183°C, O_2 condensed from a clear gas into a blue liquid. Changing the state of oxygen enables it to be transported easier. Both the aircraft and medical

industries rely on the liquid form of this molecule to provide breathable oxygen for pilots, travelers, and patients.



Figure 2.4.2: Liquid oxygen (O_2) exhibits a bluish hue.

Besides oxygen, other gases have been condensed as well. These permanent gases, meaning these substances normally exist in the gaseous state, include nitrogen (N_2), hydrogen (H_2), chlorine (Cl_2), and helium (He). By manipulating pressure and/or temperature, many chemists and physicists have attempted to produce the lowest, theoretical temperature possible. At absolute zero (-273°C or 0 K), the particles of a material have a minimum kinetic energy. Click on this video to determine if scientists can ever reach absolute zero in their laboratories.

Investigating a Cryogenic material: $\mathrm{N_2\,(l)}$

With a boiling point of -196°C (-321°F), liquid nitrogen has a variety of cryogenic uses. This clear, odorless liquid is used widely in the field of dermatology. For patients that have warts, skin tags (see image below), or precancerous moles, physicians can spray a selected area of skin. Typically, the unwanted skin lesion will fall off the body after being exposed to this cryogenic material. If this does not occur, then the dermatologist will use a scalpel to remove the frozen area of skin.



Figure 2.4.1: Liquid nitrogen is stored in insulated flasks called Dewars. Image courtesy of Elizabeth R. Gordon

Liquid N_2 is also utilized to freeze biological samples. Fertility clinics freeze semen, eggs, and embryos that can be used at a later date for couples who chose to undergo InVitro Fertilization (IVF). Umbilical cord blood and stem cells can be preserved in liquid nitrogen for future applications.

Culinary Arts employ liquid nitrogen for preparing various dishes and interesting cocktails. In the following video, chefs are freezing components of salads and desserts for their customers. When using liquid nitrogen for any application, chemists, chefs, and other individuals should protect their skin and eyes from the extremely cold (-196°C) temperature. Wearing thick rubber gloves and goggles can prevent frostbite. When ingesting materials that have been frozen with liquid nitrogen, there is concern that internal burns can occur. Lastly, indoor use of liquid nitrogen could induce asphyxiation. Kitchens and labs should have adequate ventilation to ensure sufficient oxygen remains in the atmosphere.

YouTube Video

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

View this YouTube Video online

Video 2.4.2*: At the Bazaar restaurant: Cooking with liquid nitrogen, more*

References

- https://www.bbc.com/news/magazine-19870668
- https://ehs.research.uiowa.edu/liqui...rogen-handling

Need More Practice?

- Turn to Section 2.E of this OER and work problems #4, #7, and #10.

Contributors

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University, Class of 2018)

This page titled 2.4: Temperature is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

2.E: The Mathematics of Chemistry (Exercises)

End of Chapter Exercises

1) Convert 45.0 cm^3 to L.

2) Acetone (CH_3COCH_3) has a density of .791g/mL. Assuming this compound is soluble in water, how many grams are present in 3.0 L of acetone. If acetone was spilled in Furman Lake, what would you expect to see and how would this be classified in the matter scheme (element,compound, heterogeneous, or homogeneous mixture?

3) Convert the distance from Sacramento to Modesto, CA. The measured length of the trip is approximately 121 kilometers. Convert this distance to meters and miles.

4) In January, the average temperature of eastern Siberia is -50°C. Calculate this temperature in Fahrenheit and Kelvin. Of the three temperature units, which one is the SI unit for temperature.

5) The mass of an adult human eyeball is 7.5 grams. Convert this value to kilograms and ounces. What type of lab equipment is used to measure mass?

6) How many milliliters and gallons are contained in a 2.0L Coca-Cola drink? (1 gallon = 3.79 L)

7) Convert your body temperature to °F and K.

8) Your cat needs .5 tsp of Tylenol. How many milliliters is this dosage?

9) The density of lead is 11.3g/mL. If you measured out .92 kg, how many milliliters would this mass displace? Based upon density only, would you expect lead to sink or float in water?

10) If you have a body temperature of 35°C, are you hypothermic or hyperthermic?

This page titled 2.E: The Mathematics of Chemistry (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3: ATOMS AND THE PERIODIC TABLE

3.1: Basic Atomic Theory

3.2: Development of the Modern Periodic Table

3.3: Subatomic Particles - Electrons, Protons, and Neutrons

3.4: Atomic Mass and Atomic Number

3.5: Isotopes

3.6: The Importance of Ions to a Chemist

3.E: Atoms and the Periodic Table (Exercises)

This page titled 3: Atoms and the Periodic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.1: Basic Atomic Theory

Learning Objectives

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

Democritus and the Greek Philosophers

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



Figure 3.1.1: Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories.

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

YouTube Video

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

View this YouTube Video online

Video 3.1.1*: Since ancient Greek times, philosophers and scientists have tried to figure out what an atom looks like. For a couple thousand years, humans could only speculate on the structure and other properties of the smallest unit of matter. It wasn't until the 1980s that chemists could see individual atoms. Bestselling author Sam Kean takes us through the nearly 2,400-year quest to see the atom in a new episode of the "Legends of Chemistry" series.*

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

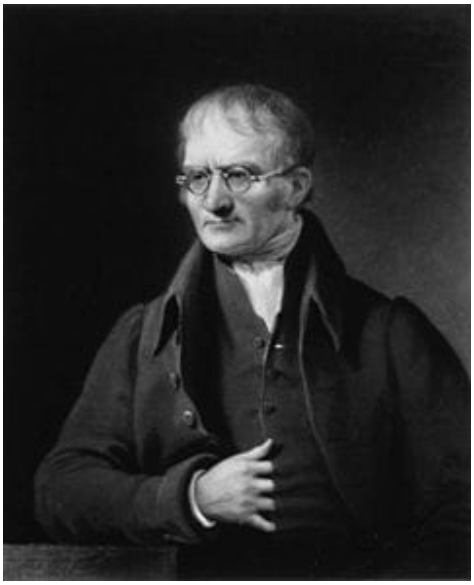


Figure 3.1.3: British physicist and chemist John Dalton (1766-1844). Unlike the greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the atom (1803). This later became known as Dalton's **atomic theory**. The published (1808) tenets of this theory were as follows:

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

Dalton's Eyes

As a scientist, John Dalton was also interested in colorblindness. Both he and his brother suffered from red/green colorblindness. At this time (1794), no other scientist had explored this medical issue. He devised theories (although not correct) as to why he could not differentiate between red and green. In addition, he named this medical condition "Daltonism." Upon his death (1844), he donated his eyes to a medical research facility in Manchester, England. Today, his eyes are still on display at the Museum of Science and Industry in Manchester, England.



Figure 3.1.3: Image taken from: https://c1.staticflickr.com/3/2823/1...37dd23af_b.jpg

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, atoms can even be used for making pretty images or as IBM research demonstrate in Video 3.1.2, control of individual atoms can be use used create animations. A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

YouTube Video

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

View this YouTube Video online

Video 3.1.2: *A Boy And His Atom: The World's Smallest Movie*

Dalton was not completely correct

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that:

1. An atom is composed of smaller particles (electrons, protons, and neutrons).
2. All atoms of an element are not identical. The existence of isotopes illustrates this phenomena.
3. Through the use of nuclear reactions, atoms of one element can be changed into atoms of another element.

The first two will be be discussed later in this chapter, while the last requires introducing "nuclear chemistry" and will be discussed in a different chapter.

References

1. https://pubs.acs.org/subscribe/archi...3chemchron.pdf

Contributors

-

This page titled 3.1: Basic Atomic Theory is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.2: Development of the Modern Periodic Table

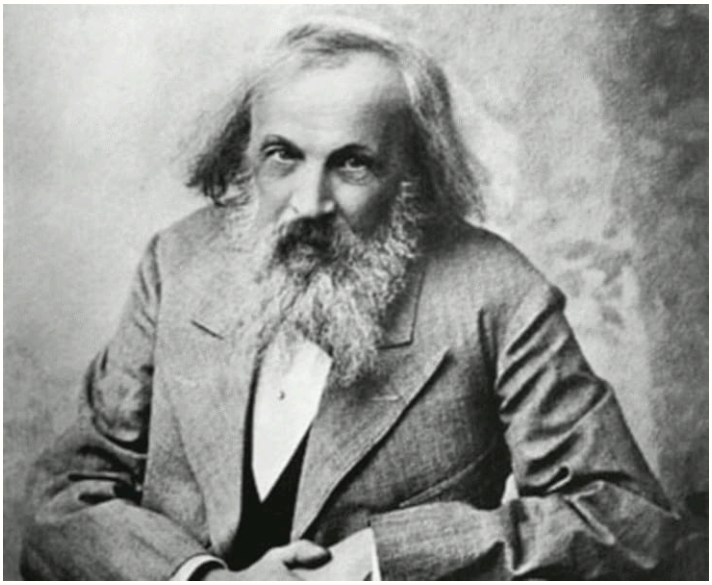
Learning Objectives

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

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

Dimitri Mendeleev (1834–1907)

The father of the modern periodic table, Dimitri Mendeleev was born in Siberia. He was the youngest of fourteen living children. Shortly after Dimtri's birth, his father lost his sight. When Dimitri was thirteen, his father died suddenly. His mother was left to manage the family's glass factory to support the family. In 1848, their glass factory burned down leaving the family with little resources. Dimitri's mother chose to take her son to St. Petersburg, Russia. Here, he would study chemistry at his father's alma mater (Main Pedagogical Institute which is not called St. Petersburg University). After several illnesses and his mother's death, he would earn his doctorate in chemistry in 1865.



From 1867 to 1890, Dimitri served as a professor of general chemistry at St. Petersburg University. In his lifetime, he published more than 400 articles and books. Two of these works were noted textbooks in general and organic chemistry.

As for his personal life, Dimitri Mendeleev married twice (his first marriage ended in a controversial divorce). Dimitri fathered a total of six children from both marriages. In 1907, he passed away from influenza complications. He was honored for his works by having element 101 (Mendelevium) named for him. Watch this video on how he constructed his periodic table.

Mendeleev's Periodic Table

Mendeleev, who first published his periodic table in 1869 (Figure 3.2.1), is usually credited with the origin of the modern periodic table. The key difference between his arrangement of the elements and that of Meyer and others is that Mendeleev did not assume that all the elements had been discovered (actually, only about two-thirds of the naturally occurring elements were known at the time). Instead, he deliberately left blanks in his table at atomic masses 44, 68, 72, and 100, in the expectation that elements with those atomic masses would be discovered. Those blanks correspond to the elements we now know as scandium, gallium, germanium, and technetium.

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	— =44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	— =68	— =72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	— =100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce= 40	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Figure 3.2.1: *Mendeleev's Periodic Table, as Published in the German Journal Annalen der Chemie und Pharmacie in 1872. The column headings "Reihen" and "Gruppe" are German for "row" and "group." Formulas indicate the type of compounds formed by each group, with "R" standing for "any element" and superscripts used where we now use subscripts. Atomic masses are shown after equal signs and increase across each row from left to right.*

The groups in Mendeleev's table are determined by how many oxygen or hydrogen atoms are needed to form compounds with each element. For example, in Group I (referred to as Gruppe 1), two atoms each element (hydrogen, lithium, potassium) bond with one atom of oxygen to form compounds. In Group VII (Gruppe VI), one atom of fluorine, F, chlorine, Cl, and bromine, Br, react with one atom of hydrogen to form respective compounds.According to Mendeleev, elements in the same column of the periodic table produced similar types of compounds. This trend could not predict the types of hydrogen/oxygen based compounds resulting from the middle section of the periodic table. Mendeleev had difficulty explaining the chemical nature of this part of the table now referred to as the transition metals.

The most convincing evidence in support of Mendeleev's arrangement of the elements was the discovery of two previously unknown elements whose properties closely corresponded with his predictions (Table 3.2.1). Two of the blanks Mendeleev had left in his original table were below aluminum and silicon, awaiting the discovery of two as-yet-unknown elements, eka-aluminum and eka-silicon (from the Sanskrit eka, meaning "one," as in "one beyond aluminum"). The observed properties of gallium and germanium matched those of eka-aluminum and eka-silicon so well that once they were discovered, Mendeleev's periodic table rapidly gained acceptance.

When the chemical properties of an element suggested that it might have been assigned the wrong place in earlier tables, Mendeleev carefully reexamined its atomic mass. He discovered, for example, that the atomic masses previously reported for beryllium, indium, and uranium were incorrect. The atomic mass of indium had originally been reported as 75.6, based on an assumed stoichiometry of InO for its oxide. If this atomic mass were correct, then indium would have to be placed in the middle of the nonmetals, between arsenic (atomic mass 75) and selenium (atomic mass 78). Because elemental indium is a silvery-white *metal*, however, Mendeleev postulated that the stoichiometry of its oxide was really In_2O_3 rather than InO. This would mean that indium's atomic mass was actually 113, placing the element between two other metals, cadmium and tin.

Table 3.2.1***:*** *Comparison of the Properties Predicted by Mendeleev in 1869 for eka-Aluminum and eka-Silicon with the Properties of Gallium (Discovered in 1875) and Germanium (Discovered in 1886)*

Property	*eka*-Aluminum (predicted)	Gallium (observed)	*eka*-Silicon (predicted)	Germanium (observed)
atomic mass	68	69.723	72	72.64
element	metal	metal	dirty-gray metal	gray-white metal
	low mp*	mp = 29.8°C	high mp	mp = 938°C
	d = 5.9 g/cm^3	d = 5.91 g/cm^3	d = 5.5 g/cm^3	d = 5.323 g/cm^3
oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
	d = 5.5 g/cm^3	d = 6.0 g/cm^3	d = 4.7 g/cm^3	d = 4.25 g/cm^3
chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$
	volatile	mp = 78°C bp* = 201°C	bp < 100°C	bp = 87°C

***mp = melting point; bp = boiling point.**

One group of elements that was absent from Mendeleev's table is the noble gases, all of which were discovered more than 20 years later, between 1894 and 1898, by Sir William Ramsay (1852–1916; Nobel Prize in Chemistry 1904). Initially, Ramsay did not know where to place these elements in the periodic table. Argon, the first to be discovered, had an atomic mass of 40. This was greater than chlorine's and comparable to that of potassium, so Ramsay, using the same kind of reasoning as Mendeleev, decided to place the noble gases between the halogens and the alkali metals.

The Role of the Atomic Number in the Periodic Table

Despite its usefulness, Mendeleev's periodic table was based entirely on empirical observation supported by very little understanding. It was not until 1913, when a young British physicist, H. G. J. Moseley (1887–1915), while analyzing the frequencies of x-rays emitted by the elements, discovered that the underlying foundation of the order of the elements was by the *atomic number*, not the atomic mass. Moseley hypothesized that the placement of each element in his series corresponded to its atomic number Z, which is the number of positive charges (protons) in its nucleus. Argon, for example, although having an atomic mass greater than that of potassium (39.9 amu versus 39.1 amu, respectively), was placed *before* potassium in the periodic table. While analyzing the frequencies of the emitted x-rays, Moseley noticed that the atomic number of argon is 18, whereas that of potassium is 19, which indicated that they were indeed placed correctly. Moseley also noticed three gaps in his table of x-ray frequencies, so he predicted the existence of three unknown elements: technetium (Z = 43), discovered in 1937; promethium (Z = 61), discovered in 1945; and rhenium (Z = 75), discovered in 1925.

H. G. J. Moseley (1887–1915)

Moseley left his research work at the University of Oxford to join the British army as a telecommunications officer during World War I. He was killed during the Battle of Gallipoli in Turkey.



Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements). The word *halogen* comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.

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

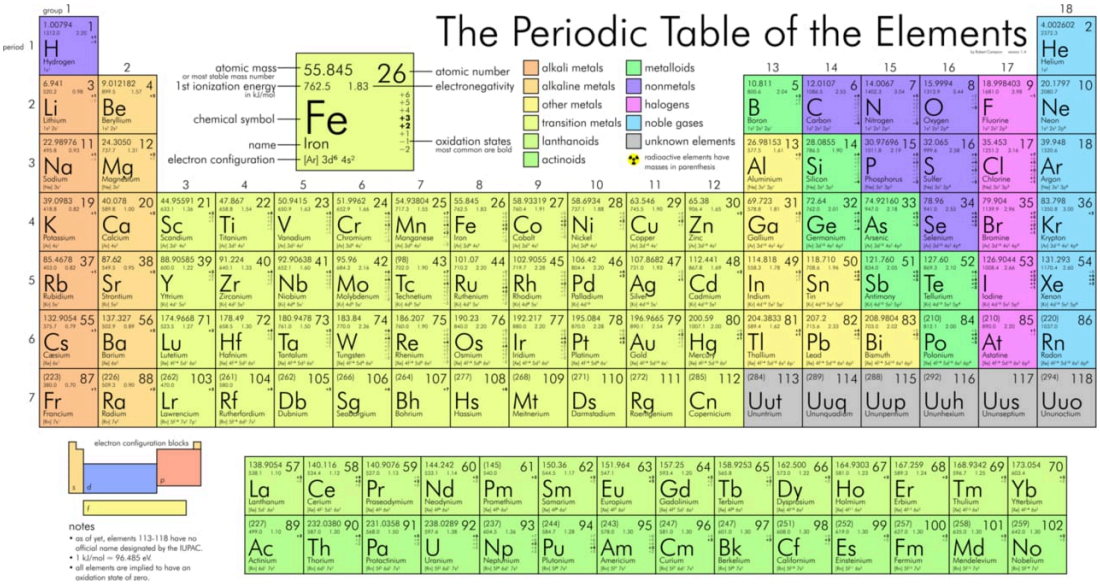


Figure 3.2.2: Types of Elements. Elements are either metals, nonmetals, or metalloid. Each group is located in a different part of the periodic table. Image used with permission (Public Domain; 2012rc).

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 3.2.3. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten

into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 3.2.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** . Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

Example 3.2.1

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

a. Se
b. Mg
c. Ge

Solution:

a. In Figure 3.2.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

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

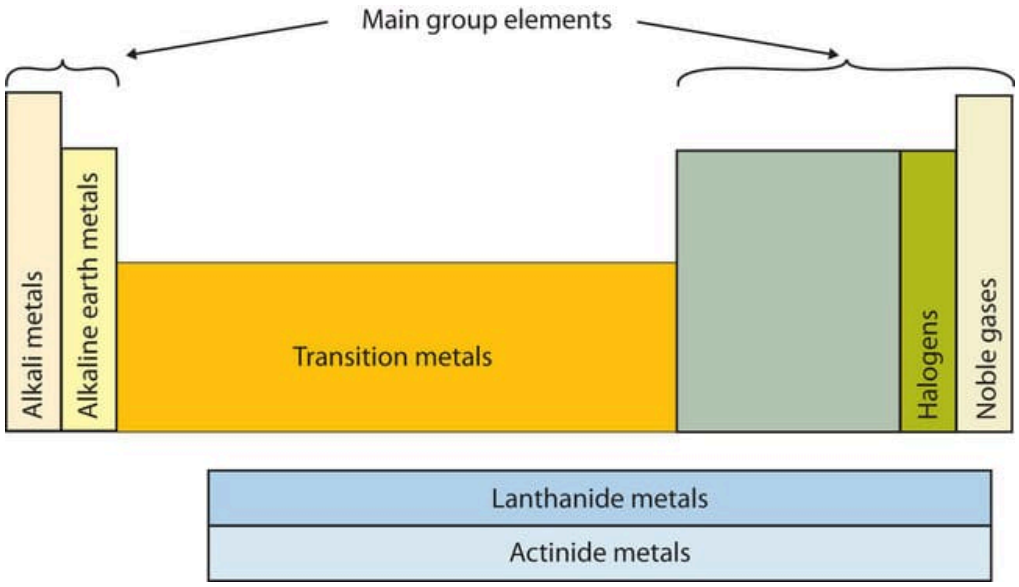


Figure 3.2.3: Some sections of the periodic table have special names. Be sure to know the section names for quizzes and tests.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins. To watch a video of your instructor presenting these groups, click on this moodle link to access.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal. The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.

Group 2: The Alkaline Earth Metals

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

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt forming," which reflects that all the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt). Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

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

Example 3.2.2: Groups

Provide the family/group names and period numbers (horizontal values)of each element.

a. Li
b. Ar
c. Ra

Solution:

a. Lithium is an alkali metal. It is located in period two.
b. Argon is a noble gas. It is located in period three.
c. Radium is an alkaline metal. It is located in period seven.

Example 3.2.3: Classification of Elements

Provide elemental names for the following combinations:

a. The alkali metal in period three.
b. The halogen in period two

c. A metalloid in period four
d. A transition metal in period three

Solution:

a. Sodium
b. Fluorine
c. Germanium or Arsenic
d. There are no transition metals in period three (gotcha!)

Summary

- The periodic table is an arrangement of the elements in order of increasing atomic number.
- Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods.
- Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.
- The elements can be broadly divided into metals, nonmetals, and metalloids.
- Semi metals exhibit properties intermediate between those of metals and nonmetals.
- Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of metalloids.
- Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable).
- Solid nonmetals are generally brittle and poor electrical conductors.
- Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

References

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

This page titled 3.2: Development of the Modern Periodic Table is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.3: Subatomic Particles - Electrons, Protons, and Neutrons

skills to Develop

- Know the basics of the experiments involving the discoveries of the three subatomic particles
- Memorize relative charge values and amu masses of the three subatomic particles.
- Realize the location of the proton,electron, and neutron inside the atomic model.

Identifying and Locating the Subatomic Particles

In 1897, the British physicist J. J. Thomson (1856–1940) proved that atoms were not the most basic form of matter. His experiments involved the use of cathode ray tubes (CRT). Under a vacuumed environment, these devices produce invisible rays that originate from a negative electrode (cathode) and continue to a positive electrode (anode). In these experiments, Thomson demonstrated that cathode rays could be deflected, or bent, by magnetic or electric fields, which indicated that cathode rays consist of charged particles . More importantly, by measuring the extent of the deflection of the cathode rays in magnetic or electric fields of various strengths, Thomson was able to calculate the mass-to-charge ratio of the particles. Because like charges repel each other and opposite charges attract, Thomson concluded that the particles had a net **negative charge**; these particles are now called **electrons**. Most relevant to the field of chemistry, Thomson found that the mass-to-charge ratio of cathode rays is independent of the nature of the metal electrodes or the gas, which suggested that electrons were fundamental components of all 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

Video 3.3.1***:*** *An instructor explains nature of a cathode ray tube and the discovery of the electron.*

During the 1910's, experiments with x- rays led to this useful conclusion: the magnitude of the **positive charge** in the nucleus of every atom of a particular element is the same. In other words, all atoms of the same element have the same number of **protons**. Furthermore, different elements have a different number of protons in their nuclei, so the number of protons in the nucleus of an atom is characteristic of a particular element. This discovery was so important to our understanding of atoms that the number of protons in the nucleus of an atom is called the atomic number. These experiments were performed by Ernest Rutherford. He is credited with defining a central location of the atom which has now been named the nucleus.

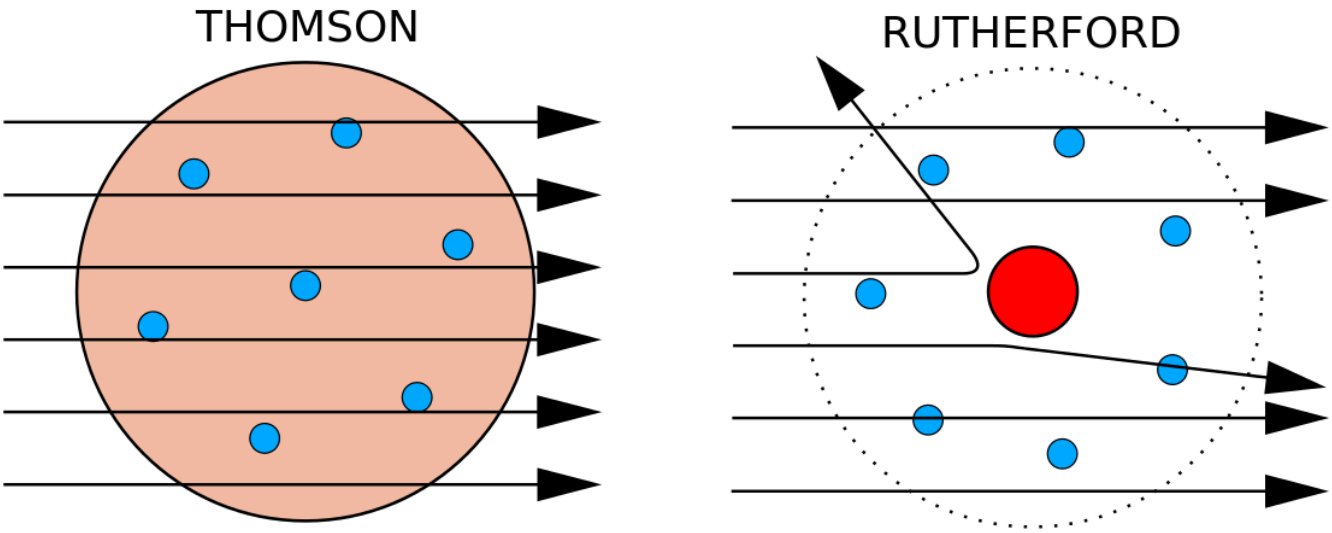


Figure 3.3.2: *Expected results: alpha particles passing through the plum pudding model of the atom with negligible deflection. Bottom: Observed results: a small portion of the particles were deflected by the concentrated positive charge of the nucleus. Image used with permissin (Public Domain).*

Rutherford was aware of the protons were inside this region. One of Ernest Rutherford's colleagues, James Chadwick, performed and studied radioactive reactions of beryllium. Upon hitting a beryllium nucleus with alpha particles, a **neutron** particle was emitted. In 1932, James Chadwick announced the existence of a third subatomic particle, the neutron. This particle has a mass of 1 atomic mass unit, but **does not have any charge**. After this discovery, nuclear science and technology started immediately.

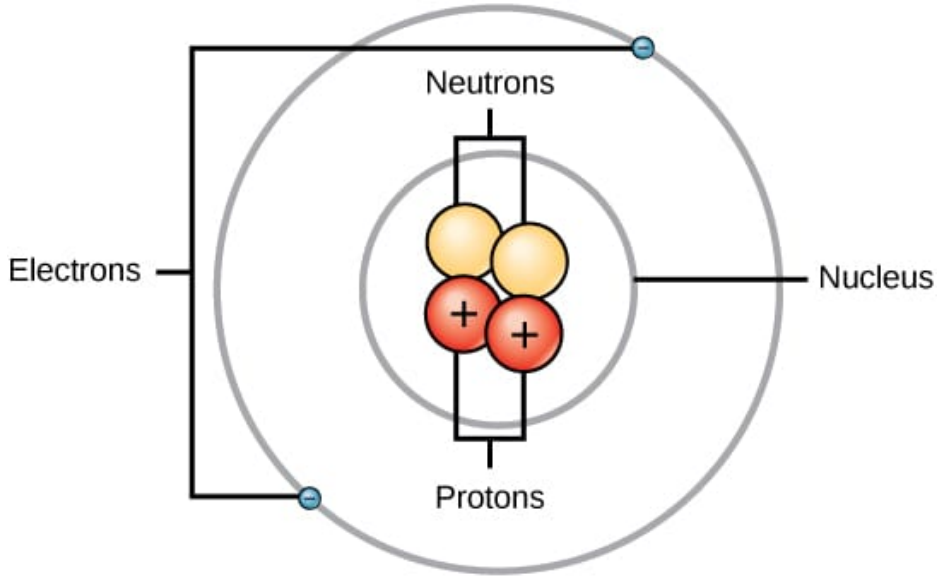


Figure 3.3.3: Elements, such as helium, depicted here, are made up of atoms. Atoms are made up of protons and neutrons located within the nucleus, with electrons in orbitals surrounding the nucleus.

Masses for the three subatomic particles can be expressed in amu (**atomic mass units**) or grams. For simplicity, we will use the amu unit for the three subatomics. Both neutrons and protons are assigned as having masses of 1 amu each. In contrast, the electron has a negligible mass of .0005 amu. This subatomic's mass is not represented in the decimal mass that is displayed on the periodic table. Only protons and neutrons contribute to an atom's mass value.

For atoms (neutral species), the number of electrons is equivalent to the number of protons. As a result, all atoms have an overall charge of zero. For discussion of ions (atoms that gain or lose electrons), look to the latter part of this chapter. Atoms that differ in neutron and mass values from the table are called isotopes. These particular species can be seen in the following section of the text.

Table 3.3.1***:*** *Properties of Subatomic Particles*

Subatomic particle	Symbol	Relative charge	Location	Amu mass
Proton	p	+1	inside the nucleus	1
Neutron	n	0	inside the nucleus	1
Electron	e^-	-1	outside the nucleus	0.0005 (~0)

Example 3.3.1

Which is true?

a. The nucleus contains protons and electrons.
b. The mass of an electron is included in the mass of the atom
c. For an atom, the number of protons = number of electrons.

SOLUTION

a. False, the nucleus contains the protons and the neutrons. Electrons are outside the core.
b. False, electrons do not contribute to the mass number of an atom.
c. True, all atoms are neutral and must have the same number of protons as electrons.

This page titled 3.3: Subatomic Particles - Electrons, Protons, and Neutrons is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.4: Atomic Mass and Atomic Number

Learning Outcomes

- Define atomic and mass numbers.
- Determine the number of protons, neutrons, and electrons in an atom.
- Identify the charge and relative mass of subatomic particles.
- Label the location of subatomic particles in the atom.
- Determine the mass of an atom based on its subatomic particles.
- Write A/Z and symbol-mass format for an atom.

Atoms are the fundamental building blocks of all matter and are composed of protons, neutrons, and electrons. Because atoms are electrically neutral, the number of positively charged protons must be equal to the number of negatively charged electrons. Since neutrons do not affect the charge, the number of neutrons is not dependent on the number of protons and will vary even among atoms of the same element.

Atomic Number

The **atomic number (represented by the letter Z)** *of an element is the number of protons in the nucleus of each atom of that element*. An atom can be classified as a particular element based solely on its atomic number. For example, any atom with an atomic number of 8 (its nucleus contains 8 protons) is an oxygen atom, and any atom with a different number of protons would be a different element. The periodic table (see figure below) 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. Next on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, beryllium atoms have four, and so on.

Since atoms are neutral, the number of electrons in an atom is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus. Helium, with two protons, will have two electrons. In the chemical classroom, the proton count will always be equivalent to an atom's atomic number. This value will not change unless the nucleus decays or is bombarded (nuclear physics).

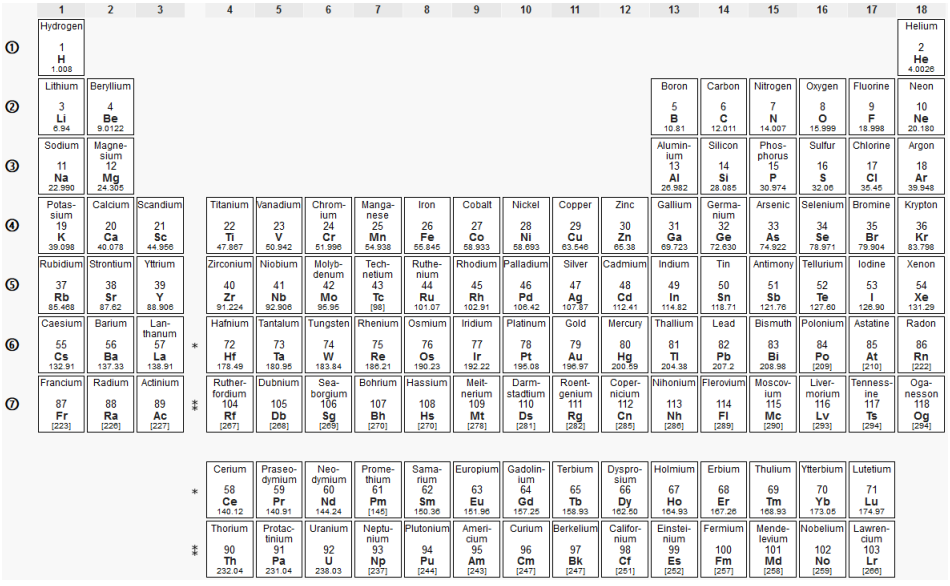


Figure 3.4.1: The periodic table of the elements. (CC BY-SA 4.0 International; DePiep via Wikipedia).

Mass Number

Experimental data 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 (represented by the letter A)** *is defined as the total number of protons and neutrons in an atom.* Consider the table below, which shows data from the first six elements of the periodic table.

Table 3.4.1: Atoms of the First Six Elements

Name	Symbol	Atomic Number (Z)	Protons	Neutrons	Electrons	Mass Number (A) (rounded to two decimals)
hydrogen	H	1	1	0	1	1.01
helium	He	2	2	2	2	4.00
lithium	Li	3	3	4	3	6.94
beryllium	Be	4	4	5	4	9.01
boron	B	5	5	6	5	10.18
carbon	C	6	6	6	6	12.01

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 a further examination of the table above will show that this is not the case. Lithium, for example, has three protons and four neutrons, giving it 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.

$$\text{Number of neutrons} = \text{ rounded mass number} - \text{atomic number} \tag{3.4.1}$$

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? To determine this, you would subtract as shown:

$$52 - 24 = 28 \text{ neutrons in a chromium atom} \tag{3.4.2}$$

The composition of any atom can be illustrated with a shorthand notation called A/Z format. Both the atomic number and mass are written to the left of the chemical symbol. The "A" value is written as a superscript while the "Z" value is written as a subscript. For an example of this notation, look to the chromium atom shown below:

$$^{52}_{24}\mathrm{Cr} \tag{3.4.3}$$

Another way to refer to a specific atom is to write the mass number of the atom after the name, separated by a hyphen. Symbol-mass format for the above atom would be written as Cr-52. In this notation, the atomic number is not included. You will need to refer to a periodic table for proton values.

Example 3.4.1

Calculate each of the three subatomic particles and give specific group or period names for each atom.

a. mercury
b. platinum
c. bromine

Solutions

a. Hg (transition metal)- has 80 electrons, 80 protons, and 121 neutrons
b. Pt (transition metal)- has 78 electrons, 78 protons, and 117 neutrons

c. Br (halogen)- has 35 electrons, 35 protons, and 45 neutrons

Example 3.4.2

Write both A/Z and symbol-mass formats for the atoms in Example 3.4.1.

Solutions

a. $^{201}_{80}\mathrm{Hg}$ and Hg-201
b. $^{195}_{78}\mathrm{Pt}$ and Pt-195
c. $^{80}_{35}\mathrm{Br}$ and Br-80

Example 3.4.3

Identify the elements based on the statements below.

a. Which element has 25 protons?
b. Which element has 0 neutrons?
c. Which element has 83 electrons?

Solutions

a. manganese

b. hydrogen

c. bismuth

Need More Practice?

- Turn to section 3.E of this OER and answer questions #1-#2, #4, and #8.

Contributors

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

This page titled 3.4: Atomic Mass and Atomic Number is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.5: Isotopes

Learning Objectives

- Understand how isotopes differ in particles and mass.
- Identify the most abundant isotope when given specific values.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Isotopes are atoms of the same element that contain different numbers of neutrons. For these species, the number of electrons and protons remain constant. This difference in neutron amount affects the atomic mass (**A**) but not the atomic number (**Z**). In a chemical laboratory, isotopes of an element appear and react the same. For this reason, it is difficult to distinguish between an atom's isotopes. In contrast, nuclear scientists can identify and separate different types of atomic nuclei (Figure 3.5.1). The technology required for this process is more sophisticated that what could be found in a typical chemical laboratory.

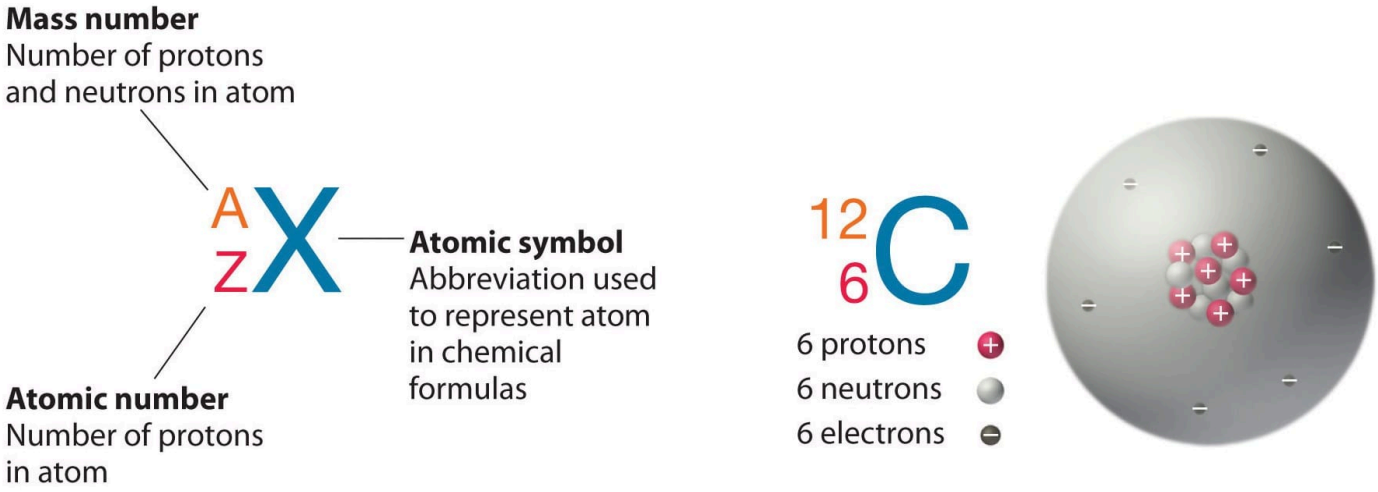


Figure 3.5.1: Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have the same number of protons, and hence the same atomic number, but different numbers of neutrons are called isotopes. All isotopes of an element have the same number of protons and electrons, which means they exhibit the same chemistry. The isotopes of an element differ only in their atomic mass, which is given by the mass number (A), the sum of the numbers of protons and neutrons. (CC BY-NS-SA; anonymous by request)

The element carbon (C) has an atomic number of 6, which means that all neutral carbon atoms contain 6 protons and 6 electrons. In a typical sample of carbon-containing material, 98.89% of the carbon atoms also contain 6 neutrons, so each has a mass number of 12. An isotope of any element can be uniquely represented as ${}^{A}_{Z}X$where X is the atomic symbol of the element. The isotope of carbon that has 6 neutrons is therefore ${}^{12}_{6}\mathrm{C}$ The subscript indicating the atomic number is actually redundant because the atomic symbol already uniquely specifies *Z*. Consequently, it is more often written as ${}^{12}\mathrm{C}$, which is read as “carbon-12.” Nevertheless, the value of Z is commonly included in the notation for *nuclear* reactions because these reactions involve changes in Z.

Most elements on the periodic table have at least two stable isotopes. For example, in addition to ${}^{12}\mathrm{C}$, a typical sample of carbon contains 1.11% ${}^{13}_{6}\mathrm{C}$, with 7 neutrons and 6 protons, and a trace of ${}^{14}_{6}\mathrm{C}$, with 8 neutrons and 6 protons. The nucleus of ${}^{14}_{6}\mathrm{C}$ is not stable, however, but undergoes a slow radioactive decay that is the basis of the carbon-14 dating technique used in archaeology. Many elements other than carbon have more than one stable isotope; tin, for example, has 10 isotopes. There are about twenty elements that exist in only one isotopic form (sodium and fluorine are examples of these). Most scientists cannot tell you how many isotopic forms exist unless they consult an isotopic table.

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure 3.5.2 compares the three isotopes of hydrogen.

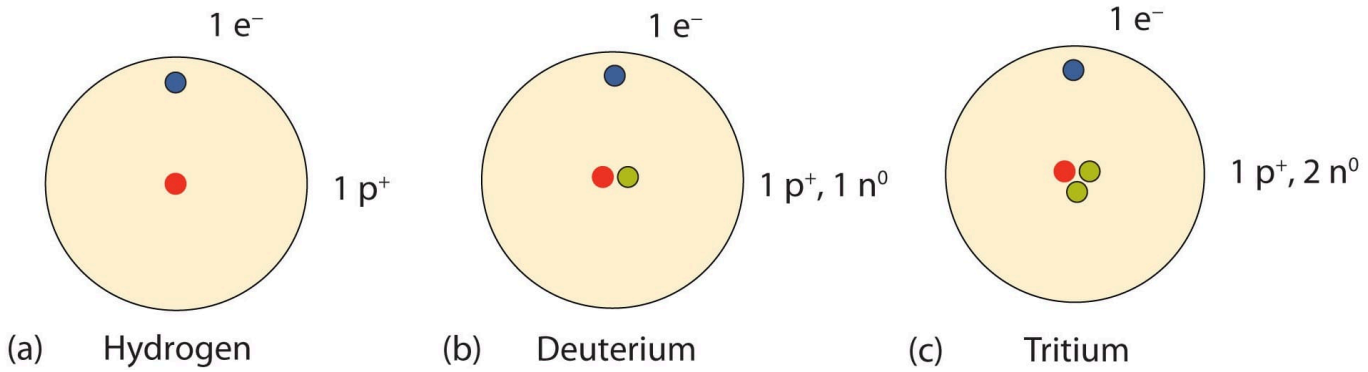


Figure 3.5.2: a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

There are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. A/Z and symbol-mass formats (refer to Section 3.4) can be used to display periodic table information. When viewing either of these two notations, isotopic differences can be obtained.

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

Look at the **A/Z formats** for the three isotopes of hydrogen in Table 3.5.1. Note how the atomic number (bottom value) remains the same while the atomic masses (top number) are varied. All isotopes of a particular element will vary in neutrons and mass. This variance in mass will be visible in the **symbol-mass format** of same isotopes as well.

Table 3.5.1

Common Name	A/Z formats	symbol-mass format	Expanded Name
Hydrogen	${}^{1}_{1}\mathrm{H}$	H-1	hydrogen-1
Deuterium	${}^{2}_{1}\mathrm{H}$	H-2	hydrogen-2
Tritium	${}^{3}_{1}\mathrm{H}$	H-3	hydrogen 3

Using a periodic table, A/Z or symbol-mass formats can be utilized to determine the amount of subatomic particles (protons, neutrons, and electrons) contained inside an isotope. When given either format, these mass values should be used to calculate the number of neutrons in the nucleus.

Exercise 3.5.1

How many neutrons are in each atom?

a. ${}^{36}_{17}\mathrm{Cl}$
b. ${}^{58}_{26}\mathrm{Fe}$
c. C-14

SOLUTION

a. In ${}^{36}_{17}\mathrm{Cl}$ there are 19 neutrons in this nucleus. To find this value, subtract A-Z or 36 – 17.
b. In ${}^{58}_{26}\mathrm{Fe}$ there are 32 neutrons i this nucleus. Again, subtract A-Z or 58-26.
c. In this example, the C-14 represents symbol-mass format. Once the atomic number is located (look at periodic table), subtract 14-6. The final answer will be 8 neutrons.

Exercise 3.5.2

For the species below, translate the A/Z format to symbol-mass. Then, calculate the number of subatomics for each atom.

a. $^{197}_{79}\mathrm{Au}$
b. $^{23}_{11}\mathrm{Na}$
c. $^{239}_{94}\mathrm{Pu}$

Answer a

Au-197 contains 79 protons, 79 electrons, and 118 neutrons.

Answer b

Na-23 contains 11 protons, 11 electrons, and 12 neutrons.

Answer c

Pu-239 contains 94 protons, 94 electrons, and 145 neutrons.

Calculating Atomic Mass

The **atomic mass** of an element is the weighted mass of all the naturally presented isotopes. On the periodic table, it is assumed that this mass has units of amu (atomic mass unit) which can be abbreviated by using the letter u. To determine the most abundant isotopic form of an element, compare given isotopes to the weighted average on the periodic table. For example, the three hydrogen isotopes (shown above) are H-1, H-2, and H-3. The atomic mass or weighted average of hydrogen is around 1.008 amu (look again to the periodic table). Of the three hydrogen isotopes, H-1 is closest in mass to the weighted average; therefore, it is the most abundant. The other two isotopes of hydrogen are quite rare, but are very exciting in the world of nuclear science.

Example 3.5.2

Identify the true statements:

a. Al-27 is more abundant than Al-25.
b. An appropriate isotope of bromine could be Br-35.
c. A chemist could easily distinguish between Cs-132 and Cs-133 by noting chemical and physical properties.
d. Most scientists know that calcium has 24 isotopes.

Answer a

This statement is true. Aluminum with a mass of 27 is closest to the mass on the periodic table. It would be more abundant than Al-25.

Answer b

This statement is false. An appropriate mass number of an isotope of bromine would be in around 80 amu (atomic mass units), not 35 (which is the atomic number).

Answer c

Chemists cannot distinguish between isotopes by looking at various properties. Most isotopes have similar solubilities,densities, and colors.

Answer d

Unless a scientist works heavily with a particular atom, he or she is not aware of the many forms that cannot exist. They would need a reference guide to know how many natural and artificial isotopes exist for a particular element.

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

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots \tag{3.5.1}$$

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

An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.

Example 3.5.3: Atomic mass of Boron

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

Solution

Boron has two isotopes. We will use the Equation 3.5.1:

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

Substitute these into the equation, and we get:

$$\begin{aligned}\text{Atomic mass} &= (0.20)\,(10) + (0.80)\,(11)\\ &= 10.8\,\text{amu}\end{aligned}$$

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

Example 3.5.4: Atomic mass of Neon

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

Solution

Neon has three isotopes. We will use the equation:

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

Substitute these into the equation, and we get:

$$\begin{aligned}\text{Atomic mass} &= (0.9092)\,(19.99) + (0.003)\,(20.99) + (0.0885)\,(21.99)\\ &= 20.17\,\text{amu}\end{aligned}$$

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

Applications of Isotopes

During the Manhattan project, the majority of federal funding was dedicated the separation of uranium isotopes. The two most common isotopes of uranium are U-238 and U-235. About 99.3% of uranium is of the U-238 variety, this form is not fissionable and will not work in a nuclear weapon or reaction. The remaining .7% is U-235 which is fissionable, but first had to be separated from U-238. This separation process is called *enrichment*. During World War II, a nuclear facility was built in Oak Ridge, Tennessee to accomplish this project. At the time, the enrichment process only produced enough U-235 for one nuclear weapon. This fuel was placed inside the smaller of the two atomic bombs (Little Boy) dropped over Japan.



Figure 3.5.3: A billet of highly enriched uranium that was recovered from scrap processed at the Y-12 National Security Complex Plant. Original and unrotated.

Uranium is a natural element that can be found in several different countries. Countries that do not have natural uranium supplies would need to obtain it from one of the countries below. Most nuclear reactors that provide energy rely on U-235 as a source of fuel. Fortunately, reactors only need 2-5% U-235 for the production of megawatts or even gigawatts of power. If the purification process exceeds this level, than it is likely a country is focusing on making nuclear weapons. For example, Manhattan Project scientists enriched U-235 up to 90% in order to produce the Little Boy weapon.

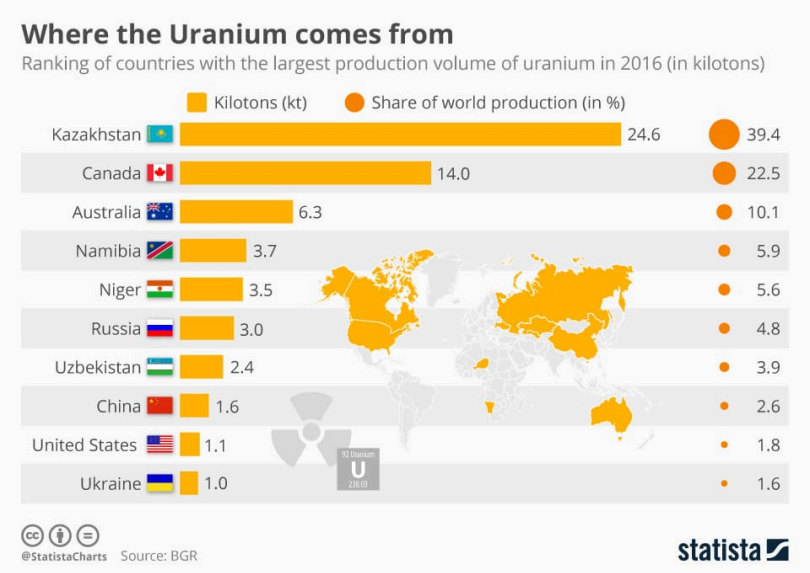


Figure 3.5.4: This infographic shows a ranking of countries with the biggest production volume of uranium in 2016 (in kilotons). (CC BY-ND; statista).

Abbreviations like HEU (highly enriched uranium) and LEU (low-enriched uranium) are used frequently by nuclear scientists and groups. HEU is defined as being over 20% pure U-235 and would not be used in most commercial nuclear reactors. This type of material is used to fuel larger submarines and aircraft carriers. If the purification of U-235 reaches 90%, then the HEU is further classified as being weapons grade material. This type of U-235 could be used to make a nuclear weapon (fission or even fusion based). As for LEU, its U-235 level would be below this 20% mark. LEU would be used for commercial nuclear reactors and smaller, nuclear powered submarines. LEU is not pure enough to be used in a conventional nuclear weapon, but could be used in a dirty bomb. This type of weapon uses conventional explosives like dynamite to spread nuclear material. Unlike a nuclear weapon, dirty bombs are not powerful enough to affect large groups of buildings or people. Unfortunately, the spread of nuclear material would cause massive chaos for a community and would result in casualties.

Summary

- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.
- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- **An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.**

Need More Practice?

- Turn to Section 3.E of this OER and work problems #3, #6, #7, and #10.

Contributors

-
- Elizabeth R. Gordon (Furman University)

This page titled 3.5: Isotopes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

3.6: The Importance of Ions to a Chemist

Template:HideTOC

Chemists appreciate isotopes and use them in basic and applied research. However, they are more concerned with the movement of electrons. When an atom gains or loses electrons, it becomes a charged species or an ion. When this occurs, the nucleus is not altered. For atoms that lose electrons, an overall positive charge will result (**#protons > #electrons**). Atoms that form these types of ions are called cations. Metal atoms (located on the left side of the periodic table) always lose electrons to become cations.

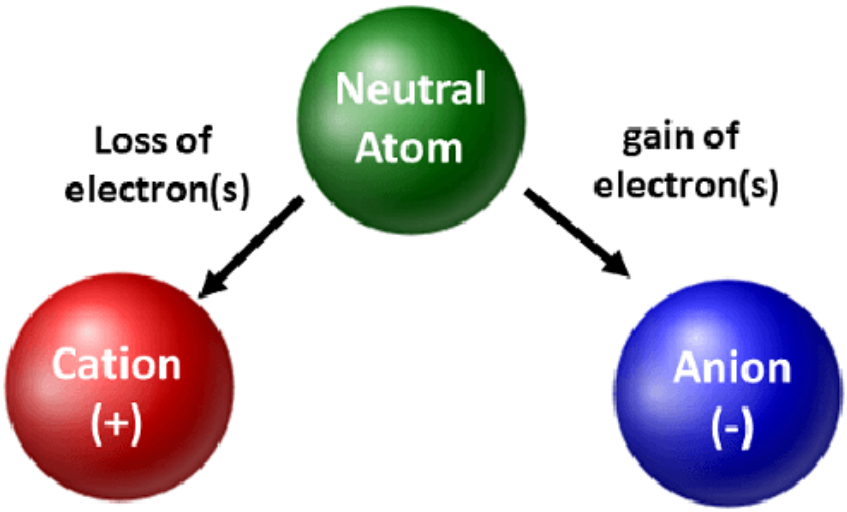


Figure 3.6.1: Neutral atoms become ions by either losing an electron (to form positively charges cations) or gaining electrons (to form negatively charged anions). Image used with permission (Public Domain; Delmar Larsen)

Unlike metal atoms, nonmetals will gain electrons to become anions. These types of ions have an overall negative charge (**#electrons > #protons**). With the exception of the noble gases, all atoms on the periodic table will lose or gain electrons to achieve electronic stability. Different types of bonding occur when atoms lose, gain, or share electrons. These types of atomic connections will be further discussed in chapter 4 of your textbook.

Interactive: Building an Atom

Build an atom out of protons, neutrons, and electrons, and see how the element, charge, and mass change. Then play a game to test your ideas!

At this point, you should be able to calculate all subatomic particles when given a specific ion charge. In addition, you should master classifications of ions (recognizing cations are positive and anions are negative). You do not need to memorize the charges, but you may learn them do to using them. Note and appreciate the ion charges for main group elements in the periodic table below. Lastly, chemists do not write a number for charges involving the number one (1). For these species, you will only see a + or - sign. Charges of higher magnitude will always have a number associated with the elemental symbol.

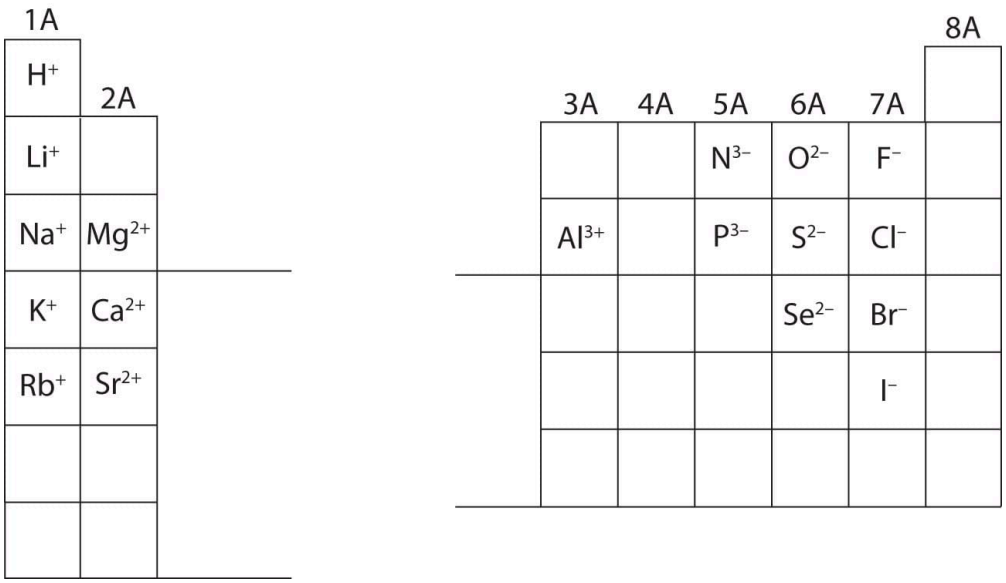


Figure 3.6.2: Periodic Table with ion charges of common main group elements. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

Example 3.6.1: IOnic Classification

Calculate the subatomic particles for the species below using the information in Figure 3.6.2. Label each as being an atom, cation, or anion. Refer to the periodic table for masses, atomic numbers, and specific ion charges.

a. Aluminum ion
b. Zirconium atom
c. Sulfur ion

Solutions

a. There are 13 protons, 14 neutrons, and 10 electrons in the aluminum ion. This ion is positively charged which means it has lost electrons and forms a cation.
b. There are 40 protons, 51 neutrons, and 40 electrons in the zirconium atom. This is an atom which has no overall charge.
c. There are 16 protons, 16 neutrons, and 18 electrons in the sulfur ion. This ion is negatively charged which means it has gained electrons and forms and anion.

Example 3.6.2: A/Z and symbol mass notations

Write A/Z and symbol mass notations for the ions/atom in the previous question.

Solutions

a. $^{27}_{13}\mathrm{Al}^{3+}$ and Al-27
b. $^{91}_{40}\mathrm{Zr}$ and Zr-91
c. $^{32}_{16}\mathrm{S}^{2-}$ and S-32

If an atom gains or loses electrons, this should be represented on the A/Z format. The symbol-mass format does not display ion charge. Remember, neutrons and protons are not affected by changes of electrons.

Need More Practice?

- Turn to Section 3.E of this OER and work problems #5 and #9.

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 3.6: The Importance of Ions to a Chemist is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

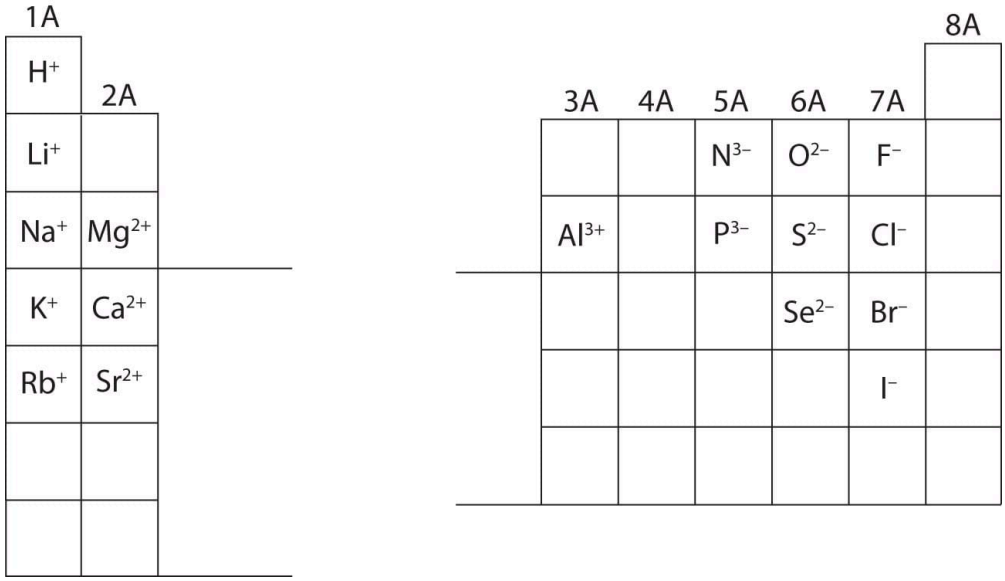
3.E: Atoms and the Periodic Table (Exercises)

Concept Review Exercises

1. Why is the atomic number so important to the identity of an atom?
2. What is the relationship between the number of protons and the number of electrons in an atom?
3. How do isotopes of an element differ from each other?
4. What is the mass number of an element?
5. Calculate the subatomic particles for
   a. Zn-67,
   b. radium,
   c. the aluminum cation,
   d. Mg-23, and
   e. the phosphorus anion.
6. Which element(s) has/have 32 neutrons: germanium, cobalt, or Mn-57?
7. Which natural isotope of platinum is most abundant?
8. Which element has 19 electrons?
9. Write A/Z format for the species in question #5.

Solutions

1. The atomic number defines the identity of an element.
2. In an electrically neutral atom, the number of protons equals the number of electrons.
3. Isotopes have different numbers of neutrons in their nuclei.
4. The mass number is the sum of the numbers of protons and neutrons in the nucleus of an atom.
5. For ion charges, refer to the table for assistance:



Species	Protons	Electrons	Neutrons
Zn-67	30	30	37 (is isotope)
Radium	88	88	138
Aluminum cation	13	10 (look up charge, 3+)	14

Species	Protons	Electrons	Neutrons
Mg-23	12	12	11 (is isotope)
Phosphorus anion	15	18 (look up charge, 3-)	16

6. Both cobalt and Mn-57 have 32 neutrons. Germanium has 41 neutrons.
7. Pt-195 is closest to the weighted average (found on periodic table) and would be the most abundant.
8. If an ion charge is not given, locate the electrons of the element by looking to the atomic number. Potassium is the only atom of these three that would have 19 electrons.
9. $^{67}_{30}\mathrm{Zn}$, $^{226}_{88}\mathrm{Ra}$, $^{27}_{13}\mathrm{Al}^{3+}$, $^{23}_{12}\mathrm{Mg}$ and $^{31}_{15}\mathrm{P}^{3-}$

This page titled 3.E: Atoms and the Periodic Table (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4: VALENCE ELECTRONS AND BONDING

Valence electrons are outer shell electrons with an atom and can participate in the formation of chemical bonds. In single covalent bonds, typically both atoms in the bond contribute one valence electron in order to form a shared pair.

4.1: Atomic Models of the Twentieth Century

4.2: Understanding Atomic Spectra

4.3: Atomic Spectroscopy Applications

4.4: Predicting Ion Charges

4.5: Lewis Dot and Bonding

4.6: Ionic Formula Writing

4.7: Applications and Solubility of Ionic Compounds

4.8: Covalent Bonding and Formula Writing

4.9: Free Radicals and the Environment

4.10: Free Radicals and Health

4.11: Applications and Solubility of Covalent Compounds

4.12: Drug Study

4.13: Arrhenius Acids and Bases

4.E: Valence Electrons and Bonding (Exercises)

This page titled 4: Valence Electrons and Bonding is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.1: Atomic Models of the Twentieth Century

Learning Objectives

- Compare different theories regarding electron arrangement within the atom.
- Understand the concept of quantized energy.
- Draw modified Bohr models for every element up through atomic number 20.
- Identify the valence electrons in a Bohr model.
- Realize the Roman numeral group numbers on the periodic table represent valence electrons.

Plum Pudding and Bohr Models

In the early twentieth century, scientists attempted to construct plausible models of the atom. One of first theories was proposed by J.J. Thomson, the physicist who discovered the electron. Thomson believed that the atom contained both positive and negative charges. He was not aware of the existence of a third neutrally charged particle (neutron). Using a plum pudding (today's comparison could be chocolate chip cookie dough) dessert for a visual representation, Thomson stated that the positive mass could be represented by the actual pudding component of the dessert. The negative species inside each atom would be scattered randomly like the plums in pudding (or chocolate chips in cookie dough).



Figure 4.1.1: Plum pudding **Figure** 4.1.2**:** Thomson's depiction of a plum pudding atom.

By 1911, Ernest Rutherford discovered a centrally located positive core of the atom. Rutherford's nucleus would house the positive protons while the negative electrons would remain outside this core. Chemists and physicists immediately abandoned the plum pudding model and looked to describe electron placement in more detail. In 1915, the Danish physicist Niels Bohr proposed a new model of the atom that involved electrons orbiting the nucleus. Stationary states or energy levels would be fixed distances from the nucleus (see below). In this model, these energy levels or shells would be represented by the letter "n."

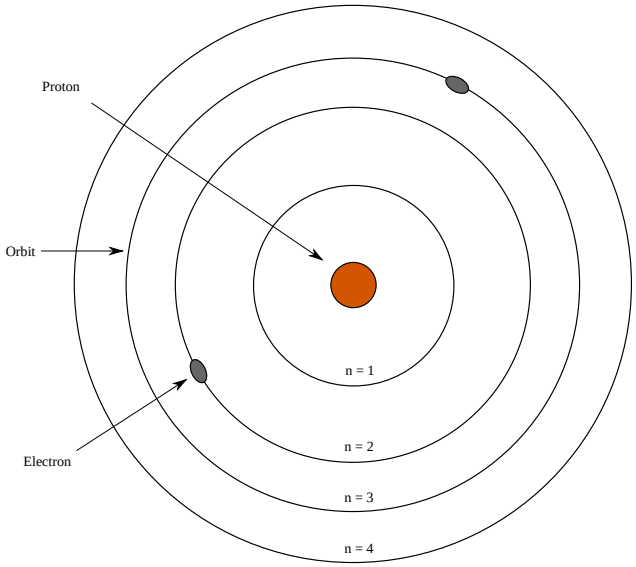


Figure 4.1.3: The Bohr model of the atom illustrating levels of electrons.

This planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system. Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy and would eventually crash into the nucleus. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to be replaced.



Figure 4.1.4: Scientists who attended the 1927 Fifth Solvay International Conference on Electrons and Photons. Niels Bohr is seated in the second row on the far right. Image taken from: https://www.flickr.com/photos/iharsten/8710094055

Energy Levels

Bohr's electrons reside in energy levels or shells. These levels are denoted with the letter "n" and are associated with a precise energy value (units would be in Joules). Shells (or levels) closer to the nucleus would require smaller amounts of energy for an electron while orbits farther from the nucleus would require the electrons to have a greater amount of energy.

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. To imagine this phenomenon, note the ladder in the image below. Ladders contain rungs that are spaced at specified distances

from the ground. On the ladder, an individual must remain on one of the rungs. He or she cannot step between the rungs and remain on the ladder successfully. The lowest rung required the least amount of energy for a person to exert. Moving to higher levels on the ladder would involve inputting more energy or effort.

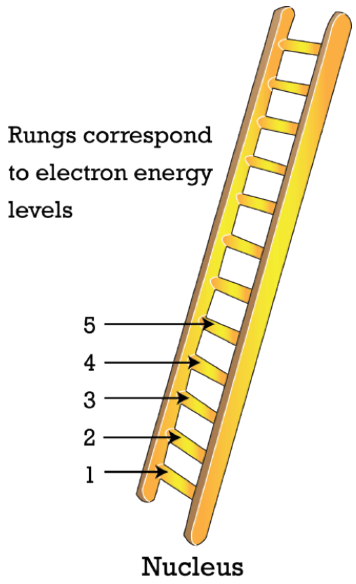


Figure 4.1.5: The concept of energy levels can be view within the context of a ladder. Each allowed electronic state has an energy. For an atom, the rungs are not equally spaced.

In the Bohr atom, the rungs represent energy levels. Electrons remain in these levels until they absorb or emit a quantum of energy. When an electron gains energy, then it will move further away from the nucleus to a higher energy level. Upon emission of energy (see the red arrow in the image below), the electron will fall back to a lower energy level. The energy absorbed or emitted during these transitions is quantized.

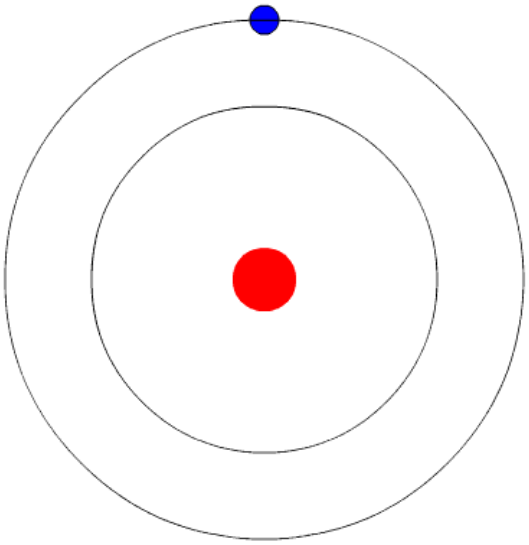


Figure 4.1.6: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. Image used with permission (CC BY-SA 3.0; Kurzon).

Drawing Bohr Models

Although not completely correct, the Bohr model can give us a visual representation of electronic levels in atoms. Bohr used the formula $2n^2$ to calculate maximum number of electrons that could be housed in an energy level or shell. In this formula, "n" represents the particular energy level. When substituting an integer above the number zero, one can obtain maximum electrons

for each shell. For example, the first energy level would be n =1 and this level could hold $2(1)^2 = 2$ maximum electrons. The second energy level could accommodate $2(2)^2 = 8$ electrons. In drawing Bohr diagrams, different amounts of shells are required to store a specific atom's electrons. Looking at a Bohr diagram for the carbon atom, one can see that two energy levels/shells are required for its six electrons.

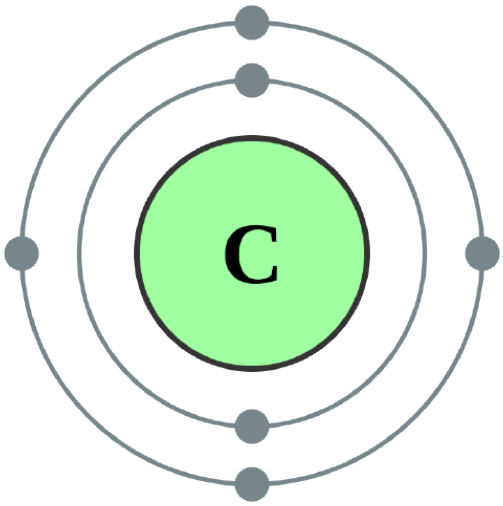


Figure 4.1.7: Electron shell diagram for carbon, the 6th element in the periodic table of elements. Image used with permission (CC BY-SA 2.0(UK); Pumbaa)

Bohr diagrams for a variety of elements are illustrated in chart below . When determining the number of electrons of an atom, remember to refer to the atomic number (the whole number) on the periodic table.

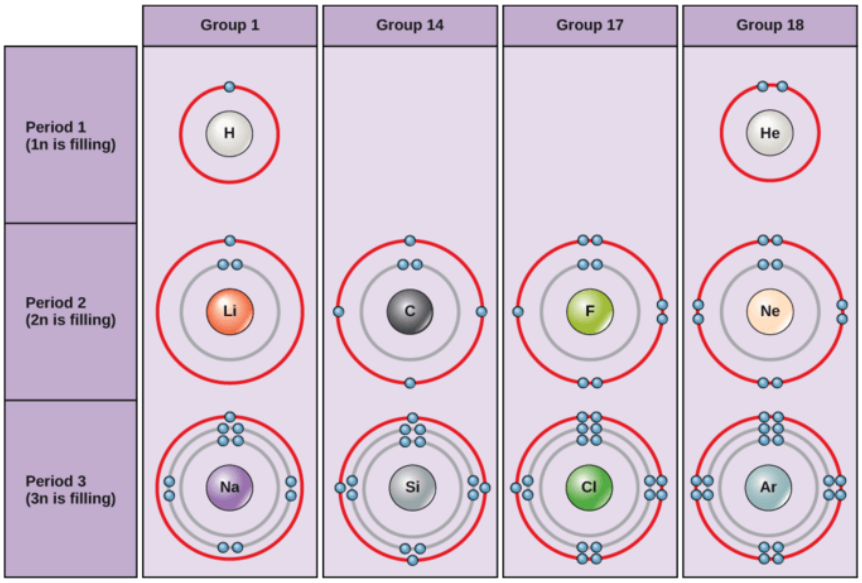


Figure 4.1.8: Bohr models of different groups/families of elements from the periodic table.

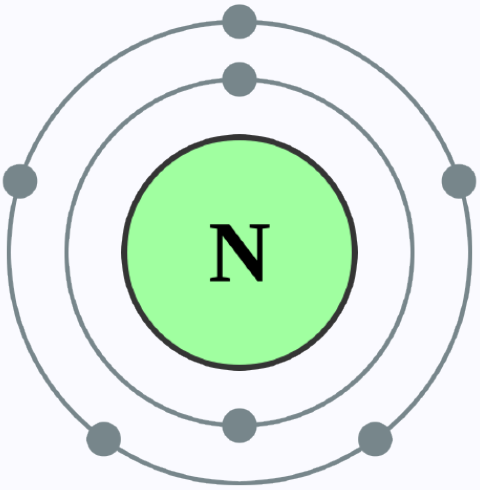
In this course, we will be using a **modified Bohr model**. In other words, we will limit each shell after n=1 to a maximum of eight total electrons. We will also be limited in drawing Bohr diagrams of only the first twenty elements on the periodic table. This process will allow us to identify the **valence** or active electrons of these particular elements. Watch this video of your professor drawing Bohr models and identifying valence electrons.

Example 4.1.1: Drawing BOhr Diagrams

Draw modified Bohr diagrams for

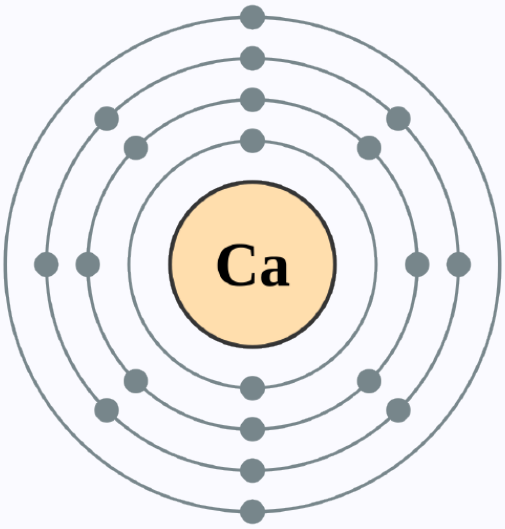
a. nitrogen
b. calcium
c. sulfur

Answer a



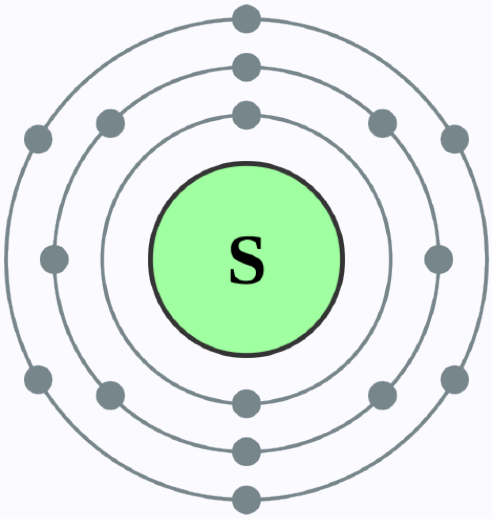
Bohr model of nitrogen: Image used with permission (CC BY-SA 2.0 uk; *Greg Robson*);

Answer b



Bohr model of calcium: Image used with permission (CC BY-SA 2.0 uk;*Greg Robson*);

Answer b



Bohr model of sulfur: Image used with permission (CC BY-SA 2.0 uk; Greg Robson).

Valence electrons are located in the highest energy level of an atom. When drawing a Bohr diagram, the valence electrons would be present in the outermost electronic level/shell (furthest away from the nucleus). An atom can have from one to eight valence electrons. In the bonding process, atoms will either gain, lose, or share valence electrons to become more stable. A quick method of checking valence electrons is to refer to an element's group number (in Roman numerals) on the periodic table. Even if you cannot draw a Bohr diagram (Bohr's are limited to elements with atomic numbers (Z)< 21), you can still determine valence electrons by looking to the group number.

Example 4.1.2: Identifying Valence Electrons

From the Bohr Diagrams constructed in Example 4.1.1, how many valence electrons are on each atom.

Answer a

Five.

Answer b

Two

Answer c

Six

Quantum Theory

The Bohr Theory failed to explain why the electron did not spiral into the nucleus after a certain amount of revolutions around the nucleus. In addition, this theory had difficulty diagramming multi-electron systems (atoms or ions with more than one electron). For these reasons, scientists turned to the quantum mechanical theory. Since 1927, this has been the working theory that most chemistry courses use (we will not use this theory in our course). For a more detailed explanation of quantum mechanics, click here.

Summary

- The Bohr model postulates that electrons orbit the nucleus at fixed energy levels.
- Orbits further from the nucleus exist at higher energy levels.
- When electrons return to a lower energy level, they emit energy in the form of light.
- Bohr's model suggests each atom has a set of unchangeable energy levels and electrons in the electron cloud of that atom must be in one of those energy levels..

Vocabulary

- Energy levels: Possible orbits that an electron can have in the electron cloud of an atom.
- Quantized: requiring or possessing specific amounts of energy.

Contributors

- Elizabeth R. Gordon (Furman University)
-
- Isabella Quiros (Furman University)

This page titled 4.1: Atomic Models of the Twentieth Century is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.2: Understanding Atomic Spectra

Learning Objectives

- Differentiate between a continuous and discontinuous spectrum
- Classify all types of radiation as being ionizing or nonionizing.
- Understand that ionizing radiation can knock electrons off DNA.
- Know that radiation of short wavelength is toxic.
- Understand the processes of absorption and emission.
- Explain the science behind fireworks.
- Realize that chemists and other scientists use AA/AE in quantitative and qualitative analyses

Continuous Spectrum

A rainbow is an example of a continuous spectrum. Here, the colors displayed are within the visible spectrum (between 380-760 nm). Light in this wavelength range is visible to the naked eye. Unlike the visible spectrum, light that is of different wavelengths (see the electromagnetic spectrum below) is not visible. Looking at Figure 4.2.1, note the different areas of the light (or electromagnetic spectrum). Areas of light that possess short wavelengths are located on the left of the spectrum. All wavelengths from UV (ultraviolet) to γ (gamma) range have the potential to ionize tissues and/or DNA. As a result, individuals who have been exposed to large amounts of these types of radiation in acute time periods could develop cancer. In contrast, visible light (see the rainbow area in Figure 4.2.1) and radiation on the right side of the spectrum has longer wavelengths and does not have the potential to ionize tissues and/or DNA. Visible, infrared (labeled IR), microwave, and radio waves are classified as being nonionizing radiation and have not been linked to cancer.

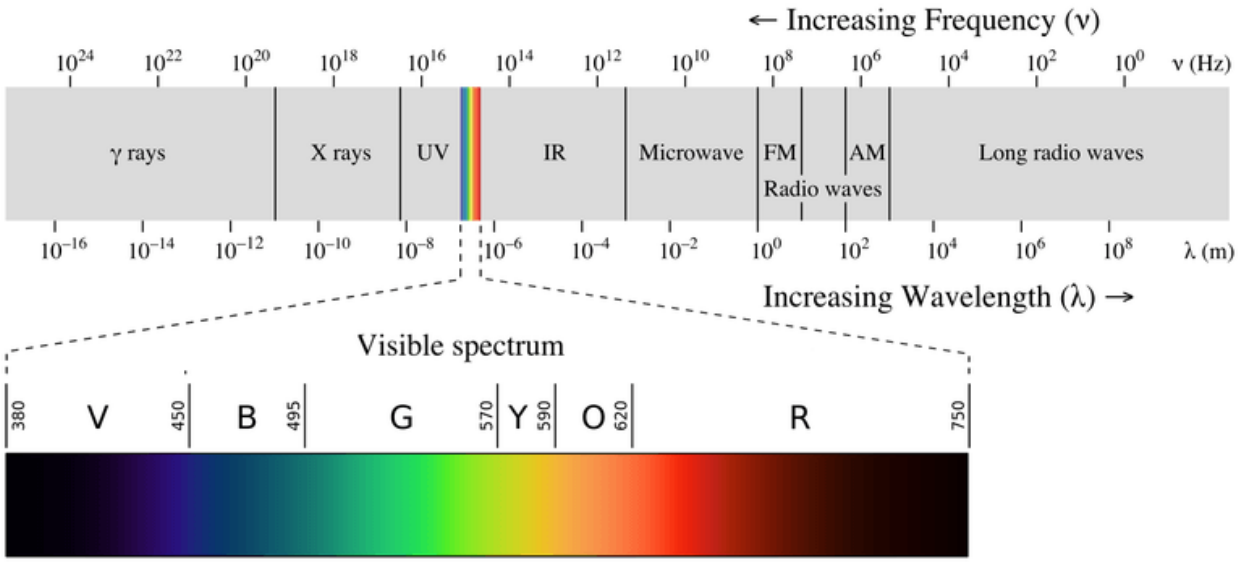


Figure 4.2.1: The electromagnetic spectrum is an example of a continous spectrum. It contains various frequencies (cycle/time) and wavelengths. Image used with permission (CC BY-SA 3.0; Philip Ronan, Gringer).

Please be aware of what types of light are ionizing and nonionizing. Do not memorize the wavelengths or frequencies of the electromagnetic spectrum. It is important that you know that the electromagnetic spectrum is an example of a continuous spectrum.

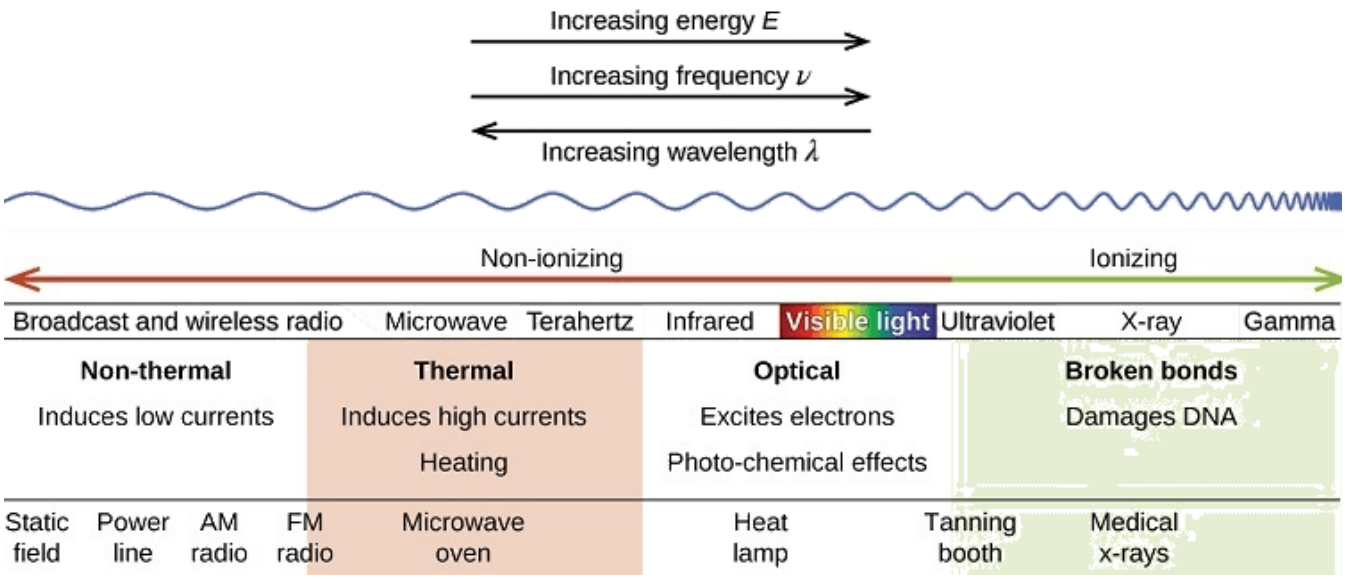


Figure 4.2.2: *Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing. Image used with permission (CC BY-SA, OpenStax).*

Examples of Radiation

For each section of the electromagnetic spectrum, be sure to know a source or application of each type (example for radio waves would be cell phones). Also, also have a good understanding of how the UV portion can be expanded. Review UV information, by clicking on this link. Please do not memorize the ABCDE's of cancer nor the types of skin cancer that occur.

Atomic Emission Spectra

The electrons in an atom tend to be arranged in such a way that the energy of the atom is as low as possible. The **ground state** of an atom is the lowest energy state of the atom. When those atoms are given energy, the electrons absorb the energy and move to a higher energy level. These energy levels of the electrons in atoms are quantized, meaning again that the electron must move from one energy level to another in discrete steps rather than continuously. An **excited state** of an atom is a state where its potential energy is higher than the ground state. An atom in the excited state is not stable. When it returns back to the ground state (Figure 4.2.3), it releases the energy that it had previously gained often in the form of electromagnetic radiation (although it can be released via heat).

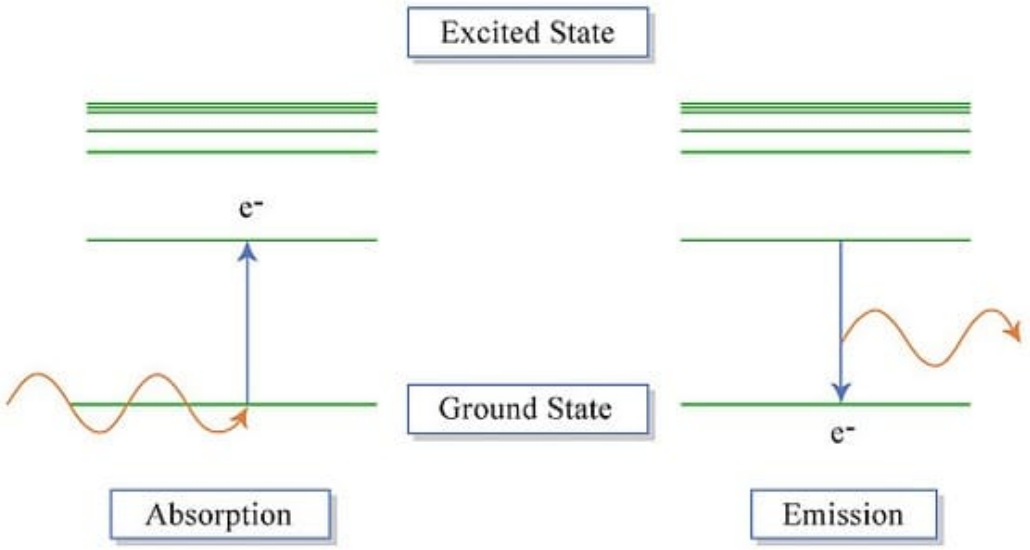


Figure 4.2.3: Atomic absorption and emission in excited and ground states. Image used with permission (CC BY-SA-NC 4.0; MIT OpenCourseWare).

Atoms can gain energy to induce these transitions from various sources. The gases in the image below have been excited with the use of electrical current. The atoms in each of these noble gases produce distinctive colors that can be used to identify the elements (Figure 4.2.4). Each of these species contains a different number of electrons that can undergo different types of excitations. In turn, each gas produces a signature color.

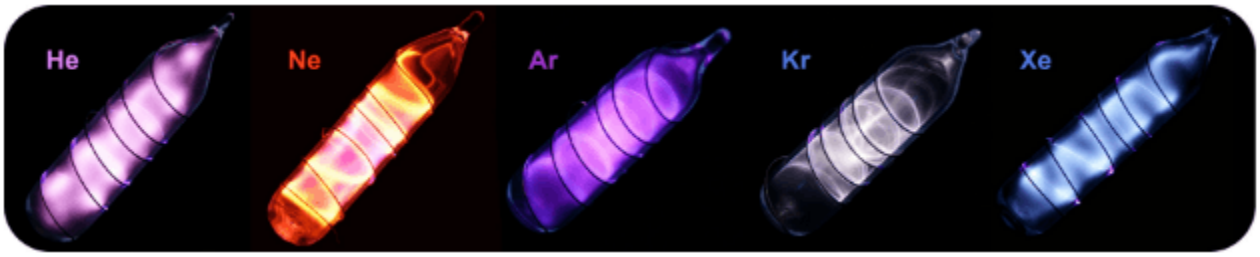


Figure 4.2.4: *Gas discharge tubes are enclosed glass tubes filled with a gas at low pressure through which an electric current is passed. Electrons in the gaseous atoms first become excited, and then fall back to lower energy levels, emitting light of a distinctive color in the process. Shown ar gas discharge tubes of helium, neon, argon, krypton, and xenon.*

Flames can be utilized to excite atoms as well. During a flame test experiment, metal chlorides are directly placed into a flame. The intense heat will promote the metal's electrons to an excited state. Upon emission, this extra energy is released in the form of visible light. If a reference panel is provided, the flame color can be used to identify a metal atom (Table 4.2.1).

YouTube Video

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

View this YouTube Video online

***Video** 4.2.1**:** Burning Salts/Flame Test demonstration: Sully Science*

Table 4.2.1***:*** *Flame colors corresponding to elements*

Metal Ion	Flame Color		Metal Ion	Flame Color
Li	Red		K	Light Purple
Na	Orange		Ba	Light Green
Ca	Orange-Red		Cu	Blue
Sr	Red			

Fireworks are similar to flame test experiments. Firework manufacturers select certain metal atoms to produce desired colors for these devices. Individual detonators will explode the metal compounds to emit radiant colors of light. Watch the video below to see the largest aerial firework shell explode over the United Arab Emirates on December 31, 2017. While watching the video, compare the colors you see to the chart above to identify the atoms that were used in the display.

YouTube Video

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

View this YouTube Video online

Video 4.2.2***:*** *Guinness World Record For World's Largest Aerial Firework Shell*

Discontinous Spectra

If the light emitted from the excited atoms is viewed through a prism, then individual patterns of lines will be produced. These lines are called spectra and correspond to fingerprint wavelengths (symbol for wavelength is λ) for a specific element. Figure 4.2.5 illustrates how the light from excited electrons can be diffracted to produce line spectra for the elements of hydrogen, helium, and iron. The specific elements produce wavelengths within the visible spectrum (between 400-700 nm) and can be seen by the naked eye. In order to obtain the numerical wavelengths (in nanometers), one would need to employ some type of detector.

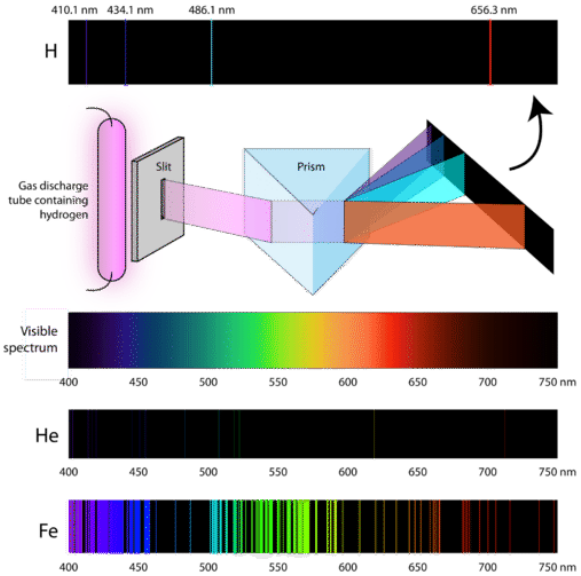


Figure 4.2.5: When light from a hydrogen gas discharge tube is passed through a prism, the light is split into four visible lines. Each of these spectral lines corresponds to a different electron transition from a higher energy state to a lower energy state. Every element has a unique atomic emission spectrum, as shown by the examples of helium (He) and iron (Fe).

Atomic Absorption Spectra

In addition to emission studies, chemists will also use atomic absorption to identify and quantify. Noting the energy changes from ground to excited states, chemists can obtain another type of discontinuous spectrum (see image below). Once again, a fingerprint wavelength pattern is produced that can be used to identify an atom. Academia and Industry could employ either an AA (atomic absorption) or AE (atomic emission) spectrometer to analyze the atoms within a sample.

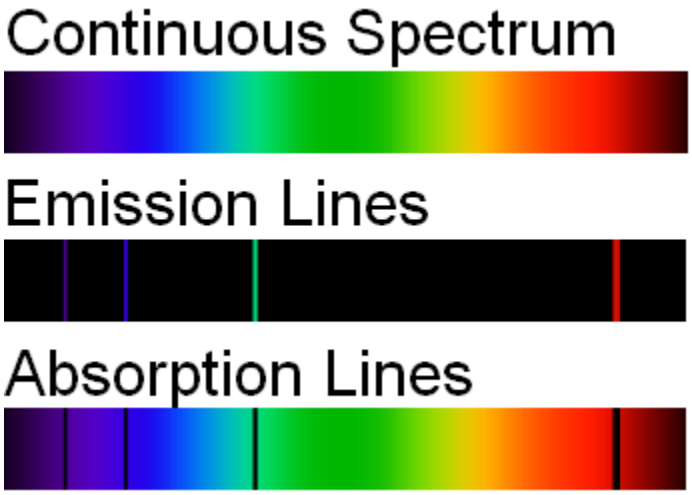


Figure 4.2.6: Spectral lines. Image used with permission (Public Domain; Jhausauer).

Example 4.2.1:

Which statements is/are true?

a. Microwaves can give you cancer.
b. When exploding fireworks, you will see the color when the metal atoms absorb energy from the detonater.
c. When using an AA or an AE spectrometer, a scientist will obtain a line spectra, not continuous spectra.
d. The flame test gives specific colors for metal atoms.
e. Gamma rays are ionizing and can cause cancer.
f. In the light spectrum, you can see IR and visible with only your eyes.
g. Continuous spectra can be used to identify an atom.

Answer a

False, they do not possess enough energy to knock electrons off tissues or DNA

Answer b

False, the color is given off during the emission or falling back process.

Answer c

True, scientists need those fingerprint wavelengths to identify atoms.

Answer d

True, the metals are the species that become excited and eventually emit the visible light.

Answer e

True, ultraviolet, gamma, and x-ray are powerful forms of radiation that can cause cancer.

Answer f

False, you can only see images in the visible spectrum (people who are colorblind are even more limited).

Answer g

False, chemists need those specific wavelengths to identify atoms, not a spectrum that shows a continuous flow of wavelengths.

Summary

Atomic emission spectra are produced when excited electrons return to the ground state.

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

Vocabulary

- Emission spectrum (or atomic spectrum): The unique pattern of light given off by an element when it is given energy
- Ground state: to be in the lowest energy level possible.
- Excited state: to be in a higher energy level.

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)
-

This page titled 4.2: Understanding Atomic Spectra is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.3: Atomic Spectroscopy Applications

Learning Objectives

- Understand why samples need to be digested for atomic absorption.
- Know what type of elements are detected by using AA or AE
- Recall practical applications of AA or AE (what businesses use these methods)

Sample Preparation for Atomic Absorption or Atomic Emission

Chemists employ atomic spectrophotometers to determine the identity and/or concentration of a metallic species. Samples must be in liquid form before being placed in the analyzer. In order to do this, a chemist will use certain acids (nitric or hydrochloric) to extract the metallic component of a sample. This process is called digestion and will not affect the analysis of the sample.



Figure 4.3.1: *A chemist adding acid to soil samples. After the digestion process, samples will be analyzed for phosphorus and potassium. Image taken from: https://upload.wikimedia.org/wikiped...le_acidity.jpg*

Once the sample has transformed into liquid form, it enters the spectrophotometer by a capillary tube. Then, it is atomized, or transformed into fine mist, by a flame. A light source (cathode lamp) bombards the misted sample. The metallic atoms inside the sample gain energy from the light source and jump to excited states. At different intervals of time, the excited atoms release energy and fall back to their ground states. This quantized energy that is involved in the absorption (AA) or emission (AE) process corresponds to a fingerprint wavelength pattern that can be used to identify a specific metal atom.

YouTube Video

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

View this YouTube Video online

Video 4.3.1**:** A combination of animation, commentary, and application to supplement the lab experience with FAA

spectroscopy.

Applications

AA or AE spectrometers have been used in many different industrial and academic settings. For example, a medical laboratory could detect the type and amount of toxic metals that could be present in patient's urine or blood. Environmental scientists could monitor metal pollutants in soil and water. The pharmaceutical industry uses these machines to determine if a metal catalyst after a drug has been purified. Lastly, the mining industry would utilize these devices to detect the quantity and presence of precious metals like gold and silver.



Figure 4.3.2: An atomic absorption spectrometer. Image taken from: https://upload.wikimedia.org/wikiped...ridesystem.jpg

Academic and industrial laboratories consider several factors before selecting spectrometer. For example, some labs require a large number of samples to be analyzed in one day while other labs might use the instrument occasionally. Specific analyzers might require a smaller amount of sample. In addition, a toxicology lab might need an instrument to report values in micro or nanograms. Lastly, academic (more often than industrial) facilities must consider the cost of the analyzer before purchasing.

Other options besides flame based AA/AE are graphite furnace (GFAA) and inductively coupled plasma (ICP-AE or ICP-OES). With the GFAA method, a furnace is used to excite the metallic sample. This technique can detect lower concentrations of metal than a flame AA/AE. Also, smaller volumes of sample are needed for a GFAA. ICP machines utilize a plasma (gas with charged ions) for the excitation process. Some of these instruments can directly analyze solid samples. Unfortunately, ICP-AE instruments are slower in producing data. Peruse Table 4.3.1 and note the general differences between the instruments.

***Table** 4.3.1**:** Common instrument attributes for spectrometers.*

Instrument	Levels of detection	Volume of sample required	Elements analyzed per unit of time	Range of cost
Flame AA/AE	High ppb	> 5.0 mL	1 element per 3-10 seconds	$15,000-$25,000
Graphite Furnace AA	Mid ppb	< 5.0 mL	1 element per 2-3 minutes	$40,000-$70,000
ICP-AE	Mid ppb	> 5.0 mL	73 different elements per minute	$70,000-$100,000

Units of Low Concentration: Parts per billion (ppb)

The letters ppb represent a concentration parts per billion. This translates to the metal atom being one part while the sample contains a billion parts. Imagine measuring one inch in 16,000 miles or counting one cent in $10 million dollars. These two

values are comparable to one ppb. In water chemistry, ppb is often referred to as µg of contaminant per liter of a sample. Table 4.3.1 above shows that a flame AA/AE detects only large amounts of ppb. This would be the least sensitive technique of the three listed.

Exercise 4.3.1: Analysis of Materials Using AA and AE

Click on this link about Forensic science and bushfires in Australia and answer the questions below.

a. In this article, what material are scientists analyzing? Why are they doing this?
b. What method-AA or AE are they using the determine the metals of the interest?
c. What metals does this article mention?
d. Why are they using XRD in this article?

Answer a

They are comparing different types of matches (unburned and burned) for arson investigations (wildfires)

Answer b

They used ICP-AE (you just need to be concerned with the AE part).

Answer c

Mg, Al, Ca, Fe, Zn, and Ba

Answer d

They used XRD (x-ray diffraction) to determine which burned matches corresponded to the unburned matches (comparing different brands).

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.3: Atomic Spectroscopy Applications is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.4: Predicting Ion Charges

Learning Objectives

- Recall how to locate valence electrons for all atoms (except transition and inner transition metals).
- Realize that metals will lose electrons to form cations and nonmetals gain electrons to form anions.
- Predict charges (using the modified Bohr model, Z < 20) for the following groups of elements: alkali, alkaline and halogens.
- Predict charges (using the modified Bohr model) for the following elements: phosphorus, nitrogen, sulfur, oxygen, and aluminum.
- Memorize charges for silver and zinc
- Label your periodic table with predicted charges from Bohr model.

Earlier, we located the valence electrons for elements Z < 20 by drawing modified Bohr structures. We can obtain these values quicker by referring to the roman numeral numbers above each family on the periodic table. The total number of valence electrons for an atom can vary between one and eight. If an element is located on the left side of the table (metal) and has less than three valence electrons, it will lose its valence in order to become stable and achieve an octet. In contrast, elements on the right side of the table (nonmetals) will gain up to eight electrons to achieve octet status.

Cations

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

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-} \tag{4.4.1}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure 4.4.1 is a graphical depiction of this process.

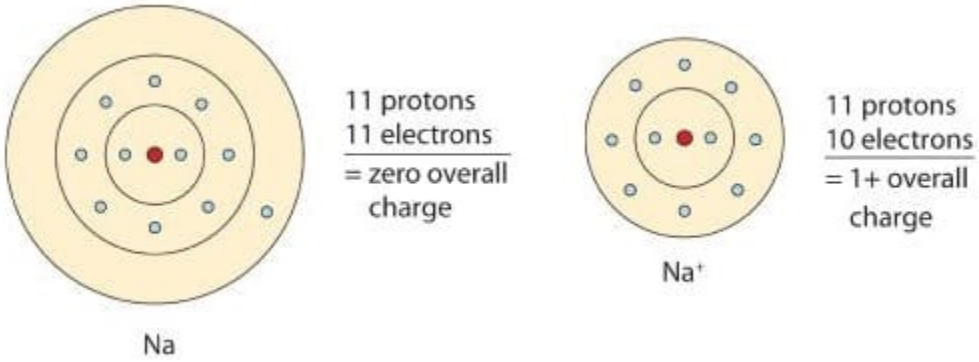


Figure 4.4.2: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Atoms vs. Cations

While they are atoms, most metals are visible. You are aware that lead is a gray, solid metal.



A piece of lead, cut through, is silvery for a short time, before the surface oxidizes. This was the left piece of the previous image, it was cut with a sharp knife, using a hammer. Image used with permission (CC By 3/0; http://images-of-elements.com/lead.php)

When lead becomes a cation, it will not be visible and can become water soluble. If lead ions enter a drinking water system, one would not be able to detect them visually. For years, lead pipes were installed to transport water to consumers in many countries including the United States. If corrosive water were to enter a lead-based pipe, it would produce soluble lead ions. These charged species could be ingested unknowingly by residents of the community. Watch the video below and note how lead has affected the city of Flint, Michigan.

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

Follow-up questions

1. Why was the drinking water source changed from Lake Huron to Flint River?
2. Did the Flint River contain lead? If not, how did the lead enter the drinking water?
3. How did the residents describe their drinking water? Please realize the discoloration of the water is due to excess iron and not lead.
4. How can lead damage the body? Who does it affect the most?
5. What could have been added to the water before releasing it to the pipes? Why did you think officials didn't add this chemical?
6. What were all the contaminants this video mentioned?
7. How did boiling the water help and hurt the residents of Flint, Michigan?

Anions

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

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

$$e^- + Cl \longrightarrow Cl^- \tag{4.4.2}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure 4.4.2 is a graphical depiction of this process.

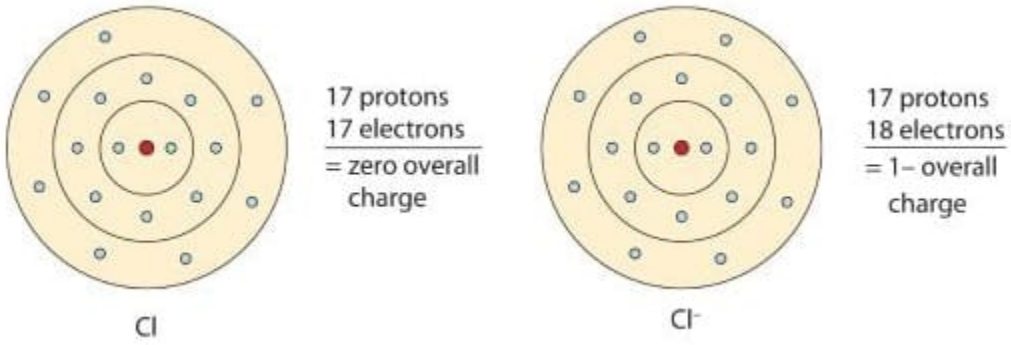
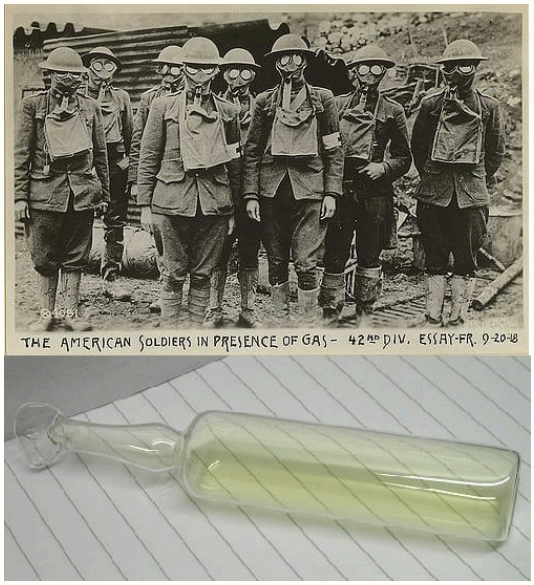


Figure 4.4.3: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1– charge.

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

Anions Vs. Atoms

Like cations, anions could be come soluble in a liquid. This would make the negatively charged species invisible to the naked eye. In this example, we will compare the element chlorine (which exists as Cl_2) to the chlorine ion (Cl^-). Chlorine is a toxic,yellow, and corrosive gas. This element was one of the first chemical weapons and was used during WWI. As an ion, chloride is used to disinfect water supplies (pools and drinking water). This colorless ion can still be corrosive, but not to the degree that chlorine gas would be.



(top) American soldiers in presence of gas (Reeve 037283), National Museum of Health and Medicine. Image used with permission (CC BY 2.0; National Museum of Health and Medicine. (bottom) Chlorine gas in an ampoule. Image used with permission (CC BY-SA 3.0; W. Oelen).

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1– charge. Figure 4.4.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table.

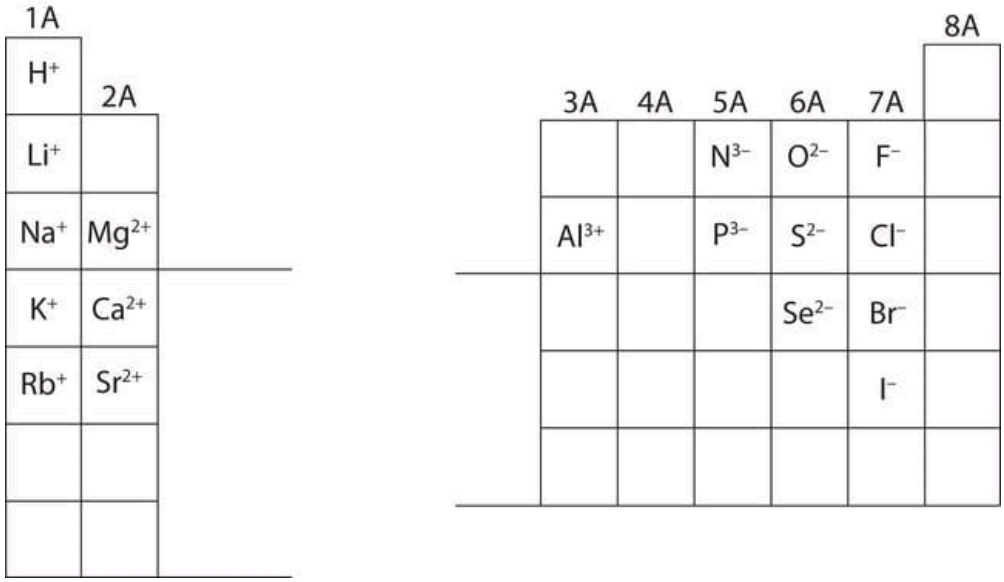


Figure 4.4.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge. Be sure to memorize silver (Ag^+) and zinc (Zn^{2+}) charges in addition to this chart.

Example 4.4.1:

Will the species below lose or gain electrons?

a. aluminium
b. halogens
c. metals in general
d. silver
e. oxygen

Answer a

Aluminum is a metal that will always lose three electrons.

Answer b

The halogens all have seven valence electrons. Each one of these elements wants to gain one electron to achieve an octet.

Answer c

Metals will always lose electrons to form cations.

Answer d

Silver is a metal that will always lose one electron. Commit this to memory!

Answer e

Oxygen is a nonmetal that has six valence electrons. It will gain two electrons to achieve an octet.

Contributors

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.4: Predicting Ion Charges is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.5: Lewis Dot and Bonding

Learning Objectives

- Realize that valence electrons can be represented by dots.
- Know that there are two types of bonding: ionic and covalent.
- Know that ionic bonding involves a metal and a nonmetal with a transfer of electrons.
- Know that covalent bonding involves two nonmetals or a metalloid with a nonmetal sharing electrons.

Bonding (Ionic and Covalent Basics)

When two atoms approach each other, they have the potential to bond (or connect). If a **metal and a nonmetal interact,** then an **ionic bond** will result. These types of bonds involve the metal donating it(s) valence electron(s) to a nonmetal. As the electronic transfer occurs, both atoms will achieve more stabile confirmations. The end result will be a less reactive compound. These type of species are composed of both cations and anions. In addition, they are crystalline and solid in nature. A few examples of real-world ionic compounds would be NaCl (table salt) and NaF (active ingredient in toothpaste).



Figure 4.5.1: NaCl crystals. Image used with permission (Public Domain; NASA).

If two nonmetals interact, then a covalent bond will result. The connection that forms is due to each atom sharing it(s) valence electron(s). In this type of bonding, each atom (except for H, B, and Be) obtains an octet and becomes stable. Metalloids and nonmetals can bond to form covalent compounds as well. Solid covalent compound are not crystalline in texture. Unlike ionic compounds. covalent compounds can be liquids or even gases.



Figure 4.5.2: Covalent compounds NO_2 and N_2O_4 contained in glass bottles. Image taken from: https://commons.wikimedia.org/wiki/File:NO2-N2O4.jpg

Table 4.5.1***:*** *Comparing ionic and covalent compounds*

Ionic Bonding	Covalent Bonding
Metal/nonmetal combination	nonmetal/nonmetal or metalloid/nonmetal combination
Metal loses valence electrons and nonmetal gains up to eight.	Both elements share valence electrons to achieve stability
Individual species have cation/anion charges	No charges are present in these compounds
All ionic compounds are solids	Covalent compounds can be solid, liquid, or gas.
Examples could include: Li_2CO_3, Fe_2O_3, and $MgSO_4$	Examples include: CO_2, H_2O, and CH_3CH_2OH

Lewis Symbols

At the beginning of the 20th century, an American physical chemist G. N. Lewis (1875–1946) devised a system of symbols—now called *Lewis electron dot symbols* (often shortened to *Lewis dot symbols*) that can be used for predicting the number of bonds formed by most elements in their compounds. Each Lewis dot symbol consists of the chemical symbol for an element surrounded by dots that represent its valence electrons. Lewis knew that incorporating models into his teaching would enable students to visualize chemical bonding easier. Gilbert Lewis was known for his interactive teaching methods and his interest in student success. Unfortunately, his nervous personality limited his ability to lecture in front of large groups of people.

Regarding his research, Lewis constructed his own theory to explain the nature of acids and bases. He was the first to synthesize D_2O (heavy water) and wrote numerous textbooks for his courses. In World War I, he served as a Major in the Gas Service unit of the American Army. He trained over 200 soldiers a week on how to adequately protect themselves from gas warfare. From these efforts, he helped reduce fatalities due to chemical gas exposures and was awarded the Distinguish Service Medal in 1922.

Lewis was nominated over thirty times for the Nobel Prize. He never received this award and does not have an element named in honor of him. For more on Gilbert Lewis, click on this link.

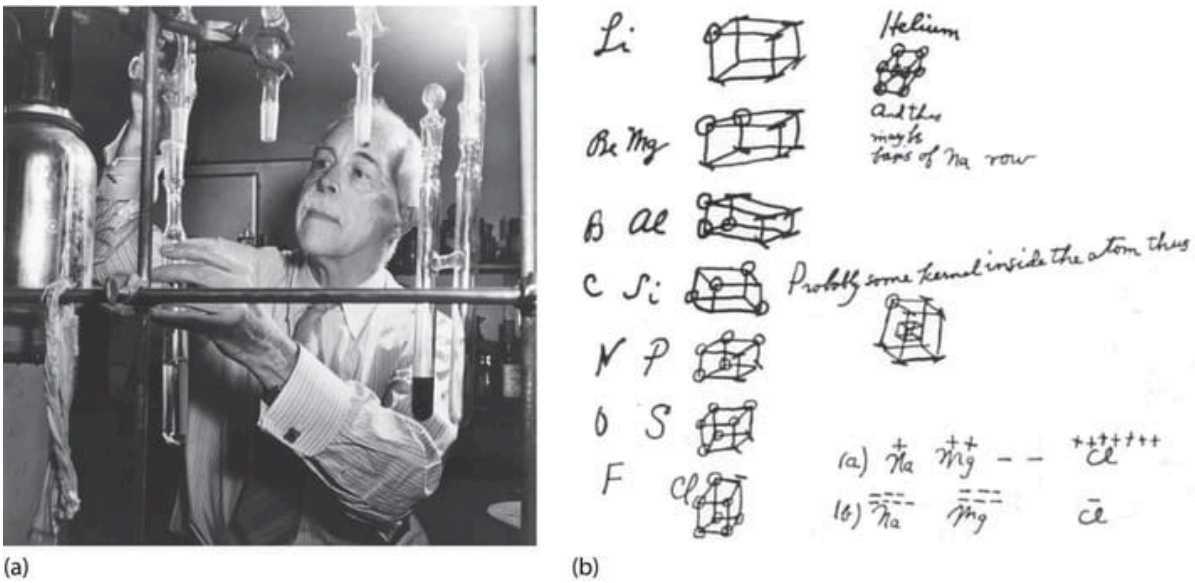
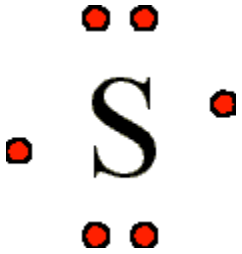


Figure 4.5.3: G. N. Lewis and the Octet Rule. (a) Lewis is working in the laboratory. (b) In Lewis's original sketch for the octet rule, he initially placed the electrons at the corners of a cube rather than placing them as we do now.

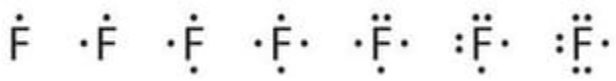
Lewis Dot symbols

- convenient representation of valence electrons
- allows scientists to keep track of valence electrons during bond formation
- consists of the chemical symbol for the element plus a dot for each **valence** electron

To write an element's Lewis dot symbol, we place dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. For example, the element sulfur has six valence electrons (note roman numeral above group on the periodic table) and its Lewis symbol would be:



Fluorine, for example, has seven valence electrons, so its Lewis dot symbol is constructed as follows:



The number of dots in the Lewis dot symbol is the same as the number of valence electrons, which is the same as the last digit of the element's group number in the periodic table. Lewis dot symbols and electron configurations (not a topic for this course) for the elements in period 2 are given in Table 4.5.2.

Table 4.5.2***:*** *Lewis Dot Symbols for the Elements in Period 2*

Group	1	2	13	14	15	16	17	18
Electron Configuration	$[He]2s^1$	$[He]2s^2$	$[He]2s^22p^1$	$[He]2s^22p^2$	$[He]2s^22p^3$	$[He]2s^22p^4$	$[He]2s^22p^5$	$[He]2s^22p^6$
Lewis Dot Symbol	Li ·	·Be·	· B ·	· C ·	· N ·	: O ·	: F ·	:Ne:

Ionic compounds are produced when a metal bonds with a nonmetal. Stability is achieved for both atoms once the transfer of electrons has occurred. The image below shows how sodium and chlorine bond to form the compound sodium chloride. Unlike a sodium atom, the resulting compound is not explosive and less corrosive than chlorine. Ionic bonding can be viewed by noting the donation of valence electrons from a metal atom to a nonmetal atom by using the Bohr model. From this theory, we will move on the Lewis structure with an understanding that metals will always lose valence electrons and nonmetals will gain up to eight in order to form stable compounds.

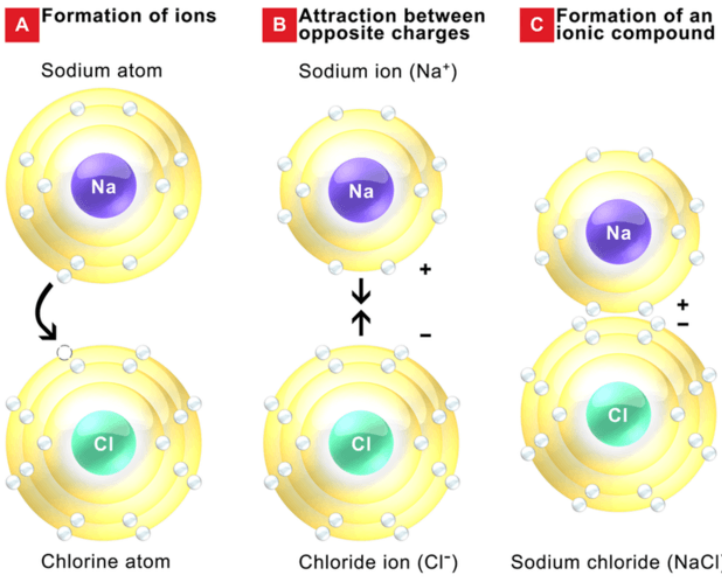


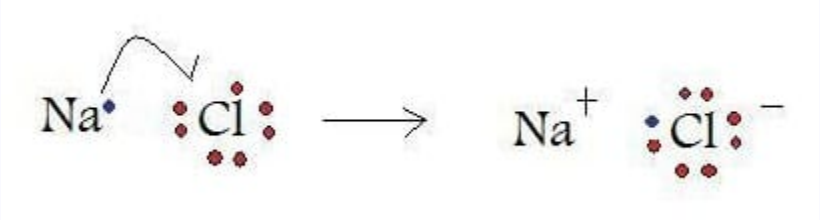
Figure 4.5.4: Formation of an ionic bond between sodium and chlorine. Image courtesy of By BruceBlaus - Own work, CC BY-SA 4.0, https://commons.wikimedia.org/w/inde...curid=44968540

Rules for Drawing IOnic Lewis structures

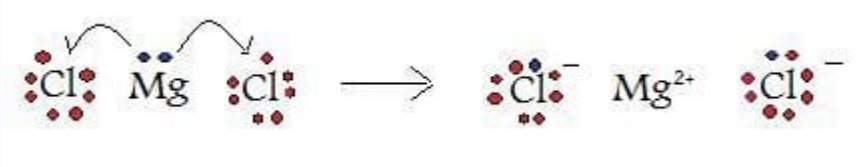
When drawing Lewis structures for compounds formed from the combination of a metal/nonmetal, use the list of rules shown below:

a. Locate the elements of interest and recall how many valence electrons each species has.
b. Draw dots around each element individually.
c. To obtain successful octets, the metal must transfer all of its valence electrons to the nonmetal.
d. At this point, the metal should have no electrons around it. In addition, the nonmetal will have an octet. If this does not occur, then refer to step e.
e. Additional atoms of the same two elements may be brought in to accomplish rules a-c.
f. **Cation and anion charges must be on Lewis structure to receive full credit (this is an ionic compound).**
g. To check structure, quickly write a formula for the ionic compound. This involves crossing down the charges diagonally and losing signs. You must reduce subscripts if they are divisible by a factor.

Example 4.5.1: Chloride Salts



In this example, the sodium atom is donating its 1 valence electron to the chlorine atom. This creates a sodium cation and a chlorine anion. Notice that the net charge of the resulting compound is 0. Only one atom of each element is needed to obtain a stable compound.



In this example, the magnesium atom is donating both of its valence electrons to chlorine atoms. Each chlorine atom can only accept 1 electron before it can achieve an octet; therefore, 2 atoms of chlorine are required to accept the 2 electrons donated by the magnesium. Notice that the net charge of the compound is 0. This Lewis structure requires two anions per one cation to produce a stable compound.

For more rigorous examples, access the CHM101 Furman moodle page (chapter 4 documents), to view a lightboard video of your instructor.

Contributors

- Mike Blaber (Florida State University)
- Wikipedia
- National Programme on Technology Enhanced Learning (India)
- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.5: Lewis Dot and Bonding is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.6: Ionic Formula Writing

Learning Objectives

- Write formulas for binary ionic compounds
- Relate the Lewis structure to the chemical formula for binary ionic compounds.
- Commit bolded polyatomic formulas to memory.
- Write formulas for ternary (polyatomic) ionic compounds.

Ionic Formulas

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a cation and an anion in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is simple: the total positive charge must balance the total negative charge. Because the charges on the ions are characteristic, sometimes we have to have more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

Figure $4.6.1$: Image taken from: https://upload.wikimedia.org/wikiped...my_Bonding.JPG

For example, consider the ionic compound between Na^{+} and Cl^{-}. Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions:

1. write the formula for the cation first and the formula for the anion next, but
2. cross the charges down diagonally.

Thus, for the compound between Na^{+} and Cl^{-}, we have the ionic formula NaCl (Figure $4.6.1$). The formula $\mathrm{Na_2Cl_2}$ also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we do not write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}, again we need only one of each ion to balance the charges. By convention, the formula is MgO.



Figure $4.6.2$: NaCl is the formular for Table Salt. The ionic compound NaCl is very common. Red rock salt from the Khewra Salt Mine in Pakistan. Image used with permission (CC BY-SA 4.0; Hubertl).

Rules for Writing Ionic Formulas

Recall ionic compound can be of the following combination: metal/nonmetal or ammonium/nonmetal

1. Recall the cation charge. Superscript this charge to the right of the cation (Ca^{2+}). If metal is multicharged metal (like most transitions and metals below metalloid staircase), then charge is indicated by roman numeral (Lead IV would translate to Pb^{4+}).
2. Recall the anion charge. Superscript this charge to the right of the anion (S^{2-}). If polyatomic ion is present, then put poly in parentheses and superscript charge to the right of this species.
3. Cross charges down diagonally and lose signs. Subscripts that are divisible by the same number must be reduced (ionics are simple compounds). Do not change what is inside a poly. You may keep your parentheses around polyatomic ions.

* Binary compound names will end with ide (except cyanide and hydroxide). This means the anion will come from the periodic table.

** Compounds that contain polyatomic ions will end in ate/ite or the two weird ides (cyanide/hydroxide). You will need to memorize the polyatomic ion table below.**

Example 4.6.1:

Write the formulas for the combinations of ions below:

a. Ca^{2+} and Cl^{-}
b. Al^{3+} and F^{-}
c. Li^{+} and O^{2-}

Solution

a. The proper ionic formula is $CaCl_2$.
b. The formula for this compound is AlF_3.
c. After crossing charges diagonally down, the formula will be Li_2O.

Exercise 4.6.2

Write the proper ionic formulas for each of the two given ions.

a. Fe^{2+} and S^{2-}
b. Fe^{3+} and S^{2-}

Answer a

FeS, Subscripts must be reduced for ionic compounds. If this is not done, then the answer will be incorrect.

Answer b

Fe_2S_3, These subscripts cannot be reduced because they are not divisible by the same number.

There also exists a group of ions that contain more than one atom. These are called **polyatomic ions**. Table 4.6.1 lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as ***oxyanions***. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 4.6.1***:*** *Common Polyatomic Ions*

Name	Formula and Charge		Name	Formula and Charge
ammonium	$(NH_4)^+$		nitrate	$(NO_3)^-$
acetate	$(C_2H_3O_2)^-$		nitrite	$(NO_2)^-$
bicarbonate (hydrogen carbonate)	$(HCO_3)^-$		carbonate	$(CO_3)^{2-}$
hydroxide	$(OH)^-$		sulfate	$(SO_4)^{2-}$
chlorate	$(ClO_3)^-$		sulfite	$(SO_3)^{2-}$
hypochlorite	$(ClO)^-$		phosphate	$(PO_4)^{3-}$
cyanide	$(CN)^-$		phosphite	$(PO_3)^{3-}$

You will need to commit Table 4.6.1 to memory for future tests and quizzes. There are more polyatomic ions than the ones listed in the table above. You will be provided with a list of these additional polyatomic ions for use on tests/quizzes. In our class, this list (List O' Polyatomics) will have the ones you are required to memorize in bold print.

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* the parentheses. Thus, the formula between calcium ions, Ca^{2+}, and nitrate ions, $NO3^-$, is properly written Ca(NO3)2, not CaNO32 or CaN2O6. In our course, I encourage you to keep your polyatomics insides parentheses at all times. Refer back to your rules for writing ionic formulas.

Example 4.6.4:

After writing these formulas, watch your instructor teach more examples of this concept by clicking on this link.

a. ammonium sulfide
b. aluminum phosphate
c. iron II phosphite

Solution

a. Ammonium is a polyatomic ion (cation). These species will go first in the formula. Take ammonium $(NH_4)^+$ and combine it with sulfide S^{2-}(ide = periodic table, except for hydroxide/cyanide). While keeping the polyatomic formula within parentheses, cross the ion charges down and lose signs. The resulting formula will be **$(NH_4)_2S$**.
b. Aluminum is Al^{3+} and phosphate is $(PO_4)^{3-}$. While keeping the polyatomic ion inside parentheses, cross charges down and reduce subscripts outside. The final formula for the compound will be **$Al(PO_4)$**. For more advanced chemistry courses, students are encouraged to drop the parentheses if a subscript is not present once charges have been crossed down. I am fine with you always keeping parentheses.
c. Transition metal charges (except for silver and zinc) will always be noted by a Roman numeral. This charge is always positive for the metal ion. Place this charge to the top right corner of the metal atom, Fe^{2+}. Next, recall that phosphite is the polyatomic, $(PO_3)^{3-}$. Cross the charges down diagonally and keep the polyatomic inside the parentheses. The final result will be **$Fe_3(PO_3)_2$**.

Exercise 4.6.4

Write the formulas for the following ionic compounds:

a. ammonium bicarbonate
b. cobalt III nitrite
c. lead IV oxide
d. potassium hydroxide
e. silver phosphide
f. mercury I sulfite

Answer a

$(NH_4)(HCO_3)$, there are no numbers outside the parentheses so you could write NH_4HCO_3 for the answer. Do not combine the hydrogen atoms in the formula.

Answer b

$Co(NO_2)_3$, the parentheses must be kept for this answer.

Answer c

PbO_2, after the charges have been crossed, subscripts must be reduced by the division of common factor

Answer d

$K(OH)$, you could drop ()

Answer e

Ag_3P, phosphide is from the periodic table, not the polyatomic table. Know your endings!!

Answer f

$Hg_2(SO_3)$, you could drop ()

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.6: Ionic Formula Writing is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.7: Applications and Solubility of Ionic Compounds

Learning Objectives

- Read the solubility of an ionic compound by using a provided table.
- Use the abbreviations (aq) and (s) appropriately.
- Realize soluble compounds will form homogeneous mixtures in water.
- Realize insoluble compounds will form heterogeneous mixtures in water.
- Know applications of the compounds contained on this page.
- Be able to match polyatomic names with their uses/properties shown in the chart of this section.

Determining Ionic Solubility

Inorganic compounds can be dissolved to sanitize drinking water. Water distributors may use a variety of methods to disinfect water. Chemical disinfection of water typically involves adding a form of chlorine. This can be done by either bubbling chlorine gas into a system or adding soluble inorganic chlorinated compounds. Using chlorine gas involves a great bit of skill due to its toxic nature. This yellow-green gas is denser than air and has previously been used as a chemical weapon. On the other hand, soluble chlorinated compounds are more stable and less toxic to handle. Regardless of its form, chlorine can be quite corrosive.



Figure 4.7.1***:*** *Chlorine gas in a Florence flask and concentrated liquid solutions for use in swimming pool. Image used with permission (CC BY-SA 4.0; gas flask - Larenmclane and pool chlorine - Maksym Kozlenko)*

In order to kill bacteria and viruses, the chlorinated compound must be soluble in water. Chlorine-based ions that become soluble can deactivate most all microorganisms except protozoans (examples would be giardia and cryptosporidium). In determining whether an inorganic chlorine compound will be soluble, chemists refer to a solubility table. The table below shows the solubilities of various compounds, in water, at a pressure of 1 atm and at room temperature (~293 K). Any box that reads "aq" results in an aqueous product in which no solid remains, while "s" indicated that the compound will not dissolve in water. Boxes marked "d" mean that the compound decomposes on reacts with water.

Solubility can be altered by manipulating temperature. Typically, many insoluble compounds can be dissolved at higher temperatures; however, this trend does not apply to every compound. For example, for Flint River residents, boiling their drinking water increased the solubility of the lead compounds that were already present. This action made the lead more concentrated and produced higher ppb levels. Please do not memorize the solubility table for this class. A chart will always be accessible for tests and quizzes.

Table 4.7.1***:*** *Solubility chart of ionic compounds in water (s = solid, aq = aqueous/soluble, and d= decomposes)*

Ion	acetate	carbonate	chlorate	chloride	fluoride	hydroxide	nitrate	nitrite	phosphate	sulfate	sulfide
aluminum	s	s	aq	aq	s	s	aq		s	aq	d
ammonium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
barium	aq	s	aq	aq	s	aq	aq	aq	s	s	s
calcium	aq	s	aq	aq	s	s	aq	aq	s	s	d
cobalt II	aq	s	aq	aq	s	s	aq		s	aq	s
iron III	s	s	s	aq	aq	s	aq		s	aq	s
copper II	aq	s	aq	aq	s	s	aq		s	aq	s
lead II	aq	s	aq	s	s	s	aq	aq	s	s	s
lithium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
magnesium	aq	s	aq	aq	s	s	aq	aq	s	aq	-------
nickel II	aq	s	aq	aq	aq	s	aq		s	aq	s
potassium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
silver	s	s	aq	s	aq	s	aq	s	s	s	s
sodium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
zinc	aq	s	aq	aq	aq	s	aq		s	aq	s

Encountering Ionics in the Home

There are many real-world applications of ionic compounds. For example, sodium chloride (table salt) is used to season food. Consuming large amounts of sodium ion (regardless if it is from sodium chloride, sodium glutamate, or sodium benzoate) can lead to cardiovascular and kidney problems. The American Heart Association recommends that individuals limit their sodium intake to less than 2300 mg per day. Watch the video below to obtain more information about sodium:

YouTube Video

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

View this YouTube Video online

Video 4.7.1***:*** *Where do Americans get most of their sodium? According to a study published in the journal Circulation, restaurant and processed foods accounted for nearly 71 percent of consumer's dietary sodium in a study of three U.S. regions.*

People who still want to season their food without sodium can opt to use potassium chloride, a table salt substitute. Unfortunately, this compound can taste extremely bitter. For this reason, some table salt substitutes will contain a mixture of both of these compounds.

Another ionic compound that can be commonly found is calcium carbonate. This compound is also known as chalk and is contained in antacids like Tums and Rolaids. The carbonate portion of this formula helps to neutralize excess stomach acid to help soothe the stomach. When purchasing these over the counter medications, remember to buy generic forms of these products.



Figure 4.7.2: Bottle of Antacid tablets. Image used with permission (CC BY-SA 3.0; Midnightcomm).

These type of antacids are quick acting and should be used for short-term stomach problems. If you experience chronic reflux or upset stomach, a health care professional would suggest a different type of medication (like Nexium, Zantac, or Prilosec) that does not neutralize excess stomach acid. Instead, these three medications deactiviate enzymes that produce the excess stomach acid.

Antacids that contain carbonates and bicarbonates will produce carbon dioxide gas inside the body. Unless an antigas ingredient is added in these products, then the patient will feel quite gaseous. Other side effects that can occur from taking these medications would be diarrhea (**if magnesium is present**) or constipation (**if aluminum is present**).

Two common compounds used to sanitize are ammonium hydroxide and sodium hypochlorite. Ammonia has a pungent odor and a basic pH (value greater than 7.0). Derivatives of ammonia are used as fertilizers and smelling salts (stimulant). When working with ammonia-based compounds, individuals need to be aware of its corrosive nature. In addition, this compound can be extremely deadly to people who have lung impairments like asthma or COPD.



Figure 4.7.3: Generic smelling salts can be used to revive a person who has fainted. Image taken from: https://www.flickr.com/photos/28581290@N08/6333856944

The common name for sodium hypochlorite is bleach. It is less volatile than ammonia but still has a distinct smell. This chemical is used to remove stains from clothing and can be used to disinfect water. Both ammonia and bleach are effective in killing most

germs. However, these two chemicals should never be combined to make a new household cleaner. The resulting product of this reaction would be chlorine gas.

Real-world applications of specific polyatomic ions

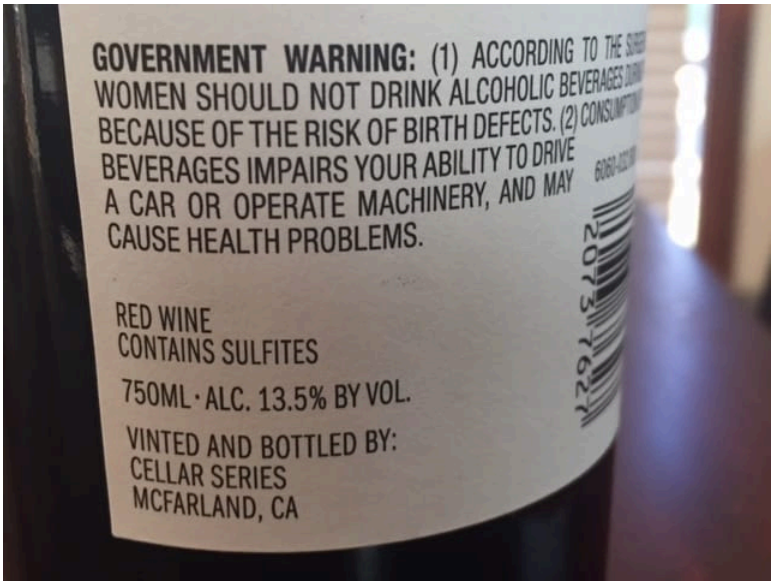


Figure 4.7.4: This red wine contains sulfite preservatives. Image made by Elizabeth R. Gordon

Ionic compounds that contain certain polyatomics can have a variety of applications. The table below gives a sample of where you might encounter polyatomic ions. For this class, it is important to match a polyatomic with its application or stated property.

Polyatomic Ion	Uses/Facts
Carbonates and Bicarbonates CO_3^{2-}/HCO_3^-	Short-term antacids, act as bases (high pH)
Nitrates NO_3^-	Fertilizers, explosives, and preserved meats (mutagen and probable carcinogen)(a suspected cause of migraines)
Nitrites NO_2^-	Preservative for meats (mutagen and probable carcinogen)(a suspected cause of migraines)
Phosphates PO_4^{3-}	Fertilizers (gives the green color to plants, if not present in sufficient amounts plants will die)
Alkali Hydroxides OH^-	Very corrosive bases, found in dry batteries, soaps, Drano, and oven cleaner
Sulfites SO_3^{2-}	Preservative found in dried fruit and red wine (possible allergen)
Cyanides CN^-	Poison used in Tylenol Murders and Jim Jones cases. Smells like almonds

Click on the video below to watch a short production of how cyanide affected the United States in 1982. After viewing, note how drug packaging changed following this event.

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

Figure 4.7.5***:*** *Tylenol 1982 Murders - PR Strategy*

Exercise 4.7.1

a. When a compound dissolves in water, what abbreviation do you place beside the chemical formula?
b. Which compounds (when dissolved in individual containers) would produce heterogeneous mixtures in water: silver sulfate, iron III chlorate, lithium hydroxide, magnesium phosphate, or lead II nitrate.
c. Which compounds could be used to preserve meats: KNO_3, KOH, or K_2CO_3?
d. Write all the chemical formulas for the aqueous compounds in #2.
e. Which compound in question c could be used as an antacid?

Answer a

Place the abbreviation (aq) to the right of the chemical formula to indicate compound is water soluble.

Answer b

Iron III chlorate and magnesium phosphate are both insoluble compounds. These ionic compounds would form heterogeneous mixtures with water.

Answer c

KNO_3 contains a nitrate which could be used as a chemical preservative.

Answer d

LiOH(aq), $Al_2(SO_4)_3$ (aq), and $Pb(NO_3)_2$(aq) are all water soluble.

Answer e

K_2CO_3 contains a carbonate which can safely neutralize excess stomach acid.

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.7: Applications and Solubility of Ionic Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.8: Covalent Bonding and Formula Writing

Learning Objectives

- To know what types of elements bond to form covalent compounds.
- To appreciate how atoms share their valence electrons in covalent bonding.
- Memorize numerical prefixes used in covalent nomenclature
- Understand that covalent compound subscripts are never reduced.

Electron Sharing

Previously, we discussed ionic bonding where electrons can be transferred from one atom to another so that both atoms have an energy-stable outer electron shell. Because most filled electron shells have eight electrons in them, chemists called this tendency the octet rule. However, there is another way an atom can achieve a full valence shell: atoms can *share* electrons. This type of bonding would be a covalent bond. Two combinations of atoms can produce this type of bonding: nonmetal/nonmetal or metalloid/nonmetal. **In this class, we will not discuss the option of metallic bonding which is a form of covalent bonding.**

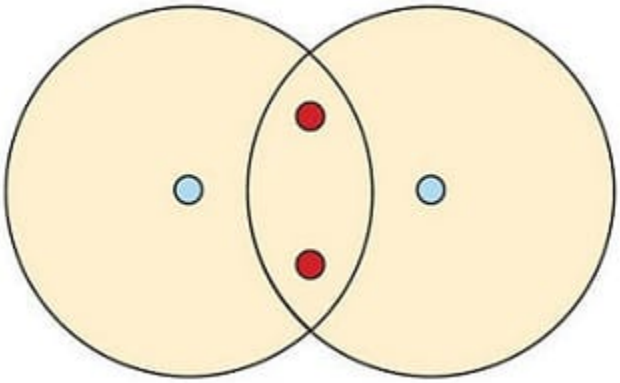


Figure 4.8.1: Sharing is caring, especially for atoms that participate in covalent bonding. Image used with permission (CC BY-2.0; Carlos Maya).

This concept can be illustrated by using two hydrogen atoms, each of which has a single electron in its valence shell. (For small atoms such as hydrogen atoms, the valence shell will be the first shell, which holds only two electrons.) We can represent the two individual hydrogen atoms as follows:



In contrast, when two hydrogen atoms get close enough together to share their electrons, they can be represented as follows:



By sharing their valence electrons, both hydrogen atoms now have two electrons in their respective valence shells. Because each valence shell is now filled, this arrangement is more stable than when the two atoms are separate. The sharing of electrons between atoms is called a covalent bond, and the two electrons that join atoms in a covalent bond are called a bonding pair of electrons. A discrete group of atoms connected by covalent bonds is called a molecule—the smallest part of a compound that retains the chemical identity of that compound.

Chemists frequently use Lewis diagrams to represent covalent bonding in molecular substances. For example, the Lewis diagrams of two separate hydrogen atoms are as follows:



The Lewis diagram of two hydrogen atoms sharing electrons looks like this:



This depiction of molecules is simplified further by using a dash to represent a covalent bond. The hydrogen molecule is then represented as follows:



Remember that the dash, also referred to as a single bond, represents a *pair* of electrons.

The bond in a hydrogen molecule, measured as the distance between the two nuclei, is about 7.4×10^{-11} m, or 74 picometers (pm; 1 pm = 1×10^{-12} m). This particular bond length represents a balance between several forces: the attractions between oppositely charged electrons and nuclei, the repulsion between two negatively charged electrons, and the repulsion between two positively charged nuclei. If the nuclei were closer together, they would repel each other more strongly; if the nuclei were farther apart, there would be less attraction between the positive and negative particles.

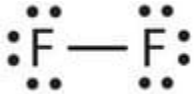
Fluorine is another element whose atoms bond together in pairs to form *diatomic* (two-atom) molecules. Two separate fluorine atoms have the following electron dot diagrams:



Each fluorine atom contributes one valence electron, making a single bond and giving each atom a complete valence shell, which fulfills the octet rule:



The circles show that each fluorine atom has eight electrons around it. As with hydrogen, we can represent the fluorine molecule with a dash in place of the bonding electrons:



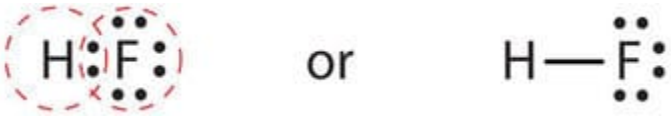
Each fluorine atom has six electrons, or three pairs of electrons, that are not participating in the covalent bond. Rather than being shared, they are considered to belong to a single atom. These are called nonbonding pairs (or **lone pairs**) of electrons.

Covalent Bonds between Different Atoms

Now that we have looked at electron sharing between atoms of the same element, let us look at covalent bond formation between atoms of different elements. Consider a molecule composed of one hydrogen atom and one fluorine atom:



Each atom needs one additional electron to complete its valence shell. By each contributing one electron, they make the following molecule:



In this molecule, the hydrogen atom does not have nonbonding electrons, while the fluorine atom has six nonbonding electrons (three lone electron pairs). The circles show how the valence electron shells are filled for both atoms.

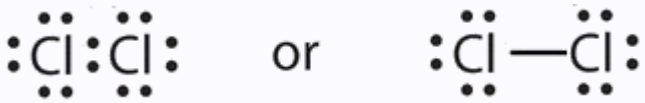
Example 4.8.1

Draw the Lewis diagram for each compound.

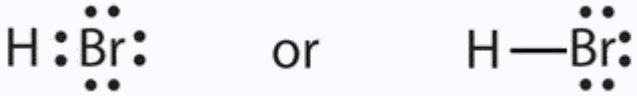
a. a molecule composed of two chlorine atoms
b. a molecule composed of a hydrogen atom and a bromine atom

SOLUTION

a. Chlorine has the same valence shell electron configuration as fluorine, so the Lewis diagram for a molecule composed of two chlorine atoms is similar to the one for fluorine:



b. Bromine has the same valence shell electron configuration as fluorine, so the Lewis diagram for a molecule composed of a hydrogen atom and a bromine atom is similar to that for hydrogen and fluorine:

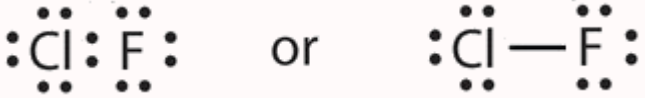


Exercise 4.8.1

Draw the Lewis diagram for each compound (show both formats).

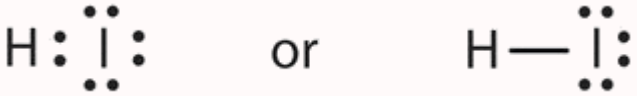
a. a molecule composed of one chlorine atom and one fluorine atom
b. a molecule composed of one hydrogen atom and one iodine atom

Answer a



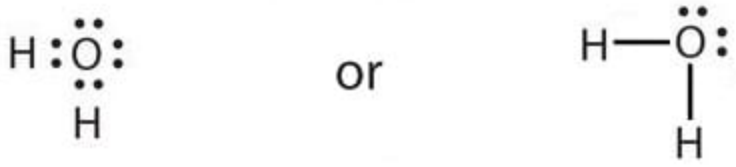
Chlorine monofluoride is a volatile interhalogen compound.

Answer b



HI is used in organic and inorganic synthesis as one of the primary sources of iodine.

Larger molecules are constructed in a similar fashion, with some atoms participating in more than one covalent bond. For example, water, with two hydrogen atoms and one oxygen atom can be represented as follows:



and methane (CH_4), with one carbon atom and four hydrogen atoms, can also be represented as follows:

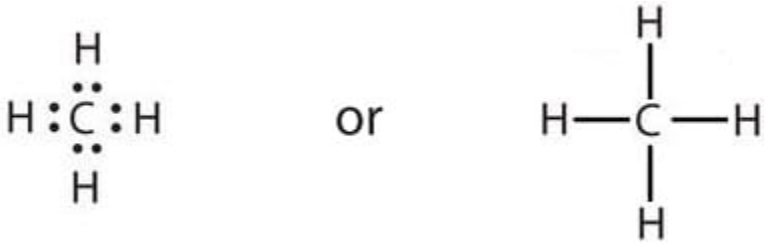


Figure 4.8.1 shows the number of covalent bonds various atoms typically form.

Rules for drawing Covalent Lewis Structures

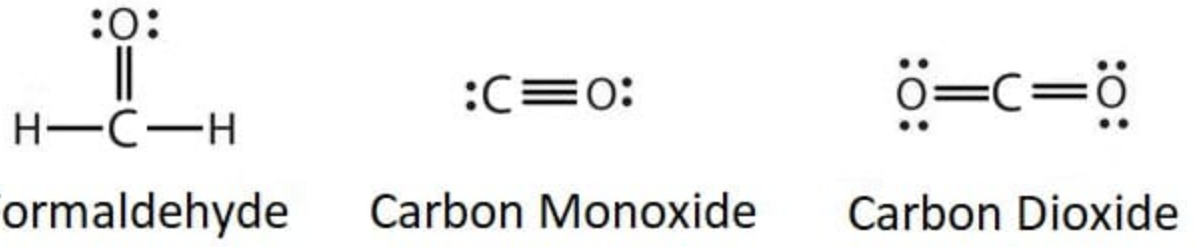
These apply to either nonmetal/nonmetal or metalloid/nonmetal combinations.

1. Tally up all valence electrons that are to be used. Remember valence electrons are the roman numerals above the element.
2. Determine which atom is to be central. Carbon is always central and Hydrogen is a terminal atom. If these atoms are not present, then go with the atom you have least of in the formula.
3. Connect each atom to the central atom using single bonds first.
4. For simple molecules or ions (only have two atoms), try single bonds, then double, and finally triple to achieve octets around each atom (except for B, Be, and H)
5. For more complex molecules or ions (3 or more atoms), try one multiple bond at a time if single bonds achieve octets.
6. My series for molecules/ions that need more than single bonds can be seen below (keeping in mind H, B, and Be only have single bonds)
    - Single/single ($-$, $-$) ,
    - then single/double ($-$, $=$),
    - then double/double ($=$, $=$),
    - and if that does not work try triple/single ($\equiv$, $-$)
7. In our class, we do not use multiple bonds for halogens.

Double and Triple Bonds

In covalent bonding, using double (four total electrons) or triple (six total electrons) is an option in order to achieve a stable octet. If a single bond does not work, then you may incorporate a multiple bond. Please note the rules above on which atoms will never achieve octets (hydrogen and boron). Although it is mathematically feasible, halogens will not use multiple bonds to become stable either.

Let's look at a few molecules that utilize multiple bonds:



Atoms that could employ multiple bonds would be oxygen, nitrogen, sulfur, phosphorus, and carbon. If an octet can be achieved with a single bond, do not incorporate multiple bonds into the structure.



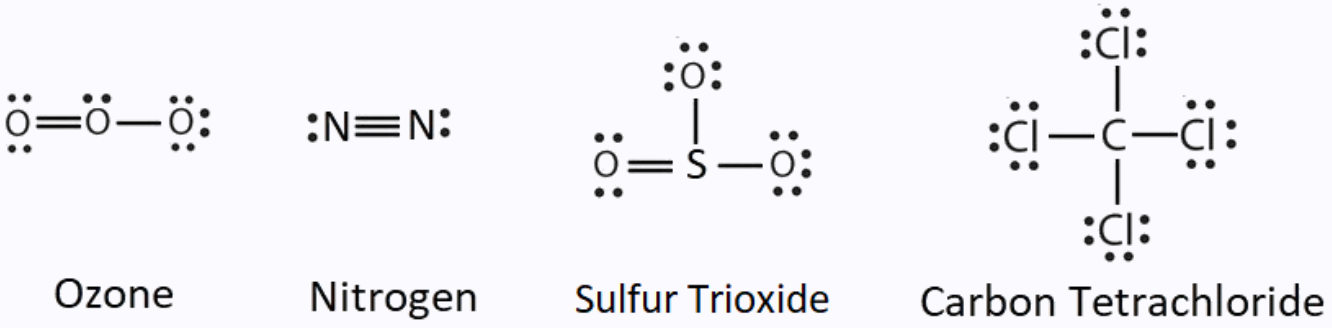
Multiple bonds involve more that one type of connection. For this reason, it is more difficult to break or cleave a multiple bond than it would be a single bond or a bond of a lower order (when comparing the same connected atoms). For example, the triple bond found in carbon monoxide is harder to break than one of the double bonds found in carbon dioxide. When reading about smog production or ozone depletion, consider the bond strengths of the chemicals involved in this environmental problems.

Example 4.8.1

Before drawing the structures below, access the Furman University CHM101 moodle documents, Here, you will see a lightboard video of your instructor drawing Lewis structures for a variety of covalent compounds. Then, try to construct structures of the other molecules in this example.

a. O_3 (ozone)
b. N_2 (nitrogen)
c. SO_3 (sulfur trioxide)
d. CCl_4 (carbon tetrachloride)

Solutions



Formula Writing for Covalent Compounds

Covalent compounds use prefixes to indicate the numbers of atoms in their formulas. In order to write correct formulas, you will

need to **memorize** Table 4.8.1 below.

Table 4.8.1***:*** *Prefix to occurrence correlations.*

Prefix name	Number of Atoms
Mono	1
Di (not bi)	2
Tri	3
Tetra (not quad)	4
Penta	5
Hexa	6
Hepta (not septa)	7
Octa	8
Nona	9
Deca	10

Rules for Writing Covalent formulas

1. There are no charges involved with covalent compounds.
2. Interpret the prefixes to represent the number of atoms or the subscript.
3. Do not cross or reduce subscripts (unlike ionic compounds).
4. The formula should reflect the name of the compound.

Example 4.8.1

Write formulas for the following covalent compounds:

a. dinitrogen pentoxide
b. boron trihydride
c. sulfur hexafluoride
d. carbon tetrachloride
e. silicon dioxide
f. dihydrogen dioxide

Answer a

N_2O_5

Answer b

BH_3

Answer c

SF_6

Answer d

CCl_4

Answer e

SiO_2

Answer f

H_2O_2, don't reduce subscripts in covalent compounds

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

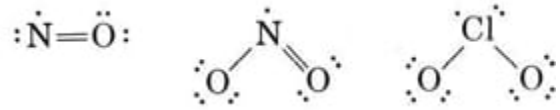
This page titled 4.8: Covalent Bonding and Formula Writing is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.9: Free Radicals and the Environment

Learning Objectives

- Recognize the structures of free radicals.
- Understand how free radicals affect the environment and biological systems.
- Know specific compounds that form photochemical smog.
- Understand how photochemical smog affects the environment and people.
- Realize that ozone in the stratosphere protects the earth from UV rays.
- Explain the basics of the Chapman Cycle.
- Provide sources of chemicals that deplete the ozone layer at an accelerated rate.
- Know where areas of depleted ozone are on the planet.
- Differentiate between CFCs, HCFC, HFC, and halons.
- Know the meaning and environmental impact of ODS and GWP.
- Appreciate the Montreal Protocol.

The majority of molecules or complex ions discussed in general chemistry courses are demonstrated to have *pairs* of electrons. However, there are a few stable molecules which contain an odd number of electrons. These molecules, called ***"free radicals"***, contain at least one unpaired electron, a clear violation of the octet rule. Free radicals play many important roles a wide range of applied chemistry fields, including biology, medicine, and environmental chemistry. Three well-known examples of such molecules are nitrogen monoxide, nitrogen dioxide, and chlorine dioxide. The most plausible Lewis structures for these molecules are



In this course, you need to be concerned with identifying free radical species. It is not necessary for you to know where to place the spare electron. In addition, please focus on the problems with free radicals that are discussed in the sections below.

Photochemical Smog

The two nitrogen-containing compounds above are classified as being $\mathrm{NO_{X}s}$. These free radicals are produced from high-temperature combustion of fuel by vehicles and industry. $\mathrm{NO_{X}s}$ can also be produced by natural processes. For example, lightning and volcanic eruptions produce these species. These gases do differ from one another. Nitrogen monoxide (also known as nitric oxide) is a colorless and non-flammable gas. In contrast, nitrogen dioxide is a reddish-orange gas that is extremely flammable. Nitrogen monoxide has a slight odor, while nitrogen dioxide has a very sharp, acrid smell.

Once $\mathrm{NO_{X}s}$ are produced, they can react with ultraviolet light to produce more free radicals. After a series of reactions, these species produce ground-level ozone ($\mathrm{O_3}$) which is referred to as smog. In regards to the environment, smog reduces visibility and smog precursors $\mathrm{NO_{X}s}$ and $\mathrm{SO_{X}s}$ react with water to produce acid rain. As for humans, $\mathrm{NO_{X}s}$ can affect lung, heart, and immunity functioning. Photochemical smog is more of an issue in highly populated cities that rely heavily on transportation sources and electrical power plants. Fortunately, the United States Environmental Protection Agency (EPA) regulates $\mathrm{NO_{X}s}$ amounts for vehicles and power plants.



Figure 4.9.2***:*** *Arial photograph of photochemical smog in Manhattan, New York. Documerica was a program sponsored by the Environmental Protection Agency to photographically document subjects of environmental concern in America during the 1970s.*

Consumers of American cars can review smog ratings before purchasing a vehicle. Cars and trucks produce other chemicals besides NOxs that contribute to smog formation. For a basic rating on smog-producing ability, click here and input your vehicle's information. On this website, smog ratings are assigned to mobile sources (cars, trucks, sports utility vehicles) from one to ten. A low score indicates that the vehicle produces a large number of smog precursors.

Protective Ozone (O^3)

Ozone or O_3 can be harmful or helpful depending on where it is located. In the troposphere (see image below), ozone is referred to as smog. This particular species can be formed from free radicals. Previously, you have learned the harmful effects of being exposed to ground level ozone. However, ozone acts as a protective shield in the stratosphere. Of the three forms of ultraviolet radiation, ozone completely screens the Earth's surface from UV-C radiation. Portions of UV-B (causes more burning of skin) and UV-A (causes more tanning of skin) still reaches the terrestrial environment.

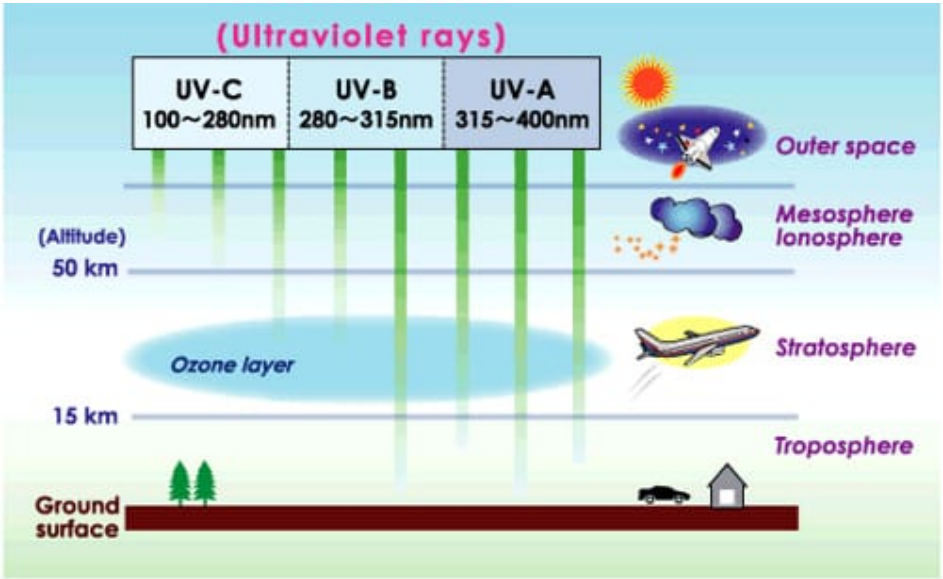


Figure 4.9.3: Levels of the atmosphere and location of the O_3 layer. Image is taken from https://commons.wikimedia.org/wiki/F...5%E0%B8%81.jpg

If all the ozone in the atmosphere were combined, it would have a thickness of less than 0.5 cm (.20 inches). Ozone is most concentrated in the lowest 15 km portion of the stratosphere. NASA and other atmospheric scientists utilize the Dobson scale to represent the concentration of ozone molecules. One Dobson Unit (DU) is the number of ozone molecules that could create a 0.01 mm layer of thickness (conditions of 0°C and 1.0 atm). NASA reports the average thickness of the ozone layer to be **300 DU** which translates to **3 mm**. A depleted ozone layer would register as being equal to or less than **100 DU** or a thickness of **1 mm.** The NASA image below depicts a purplish-blue region that corresponds to the largest area of ozone depletion. First detected in 1985, the ozone hole has grown to a maximum of 7.6 million square miles (or 2.5 times the size of the United States. NASA reported this peak value on September 11, 2017.

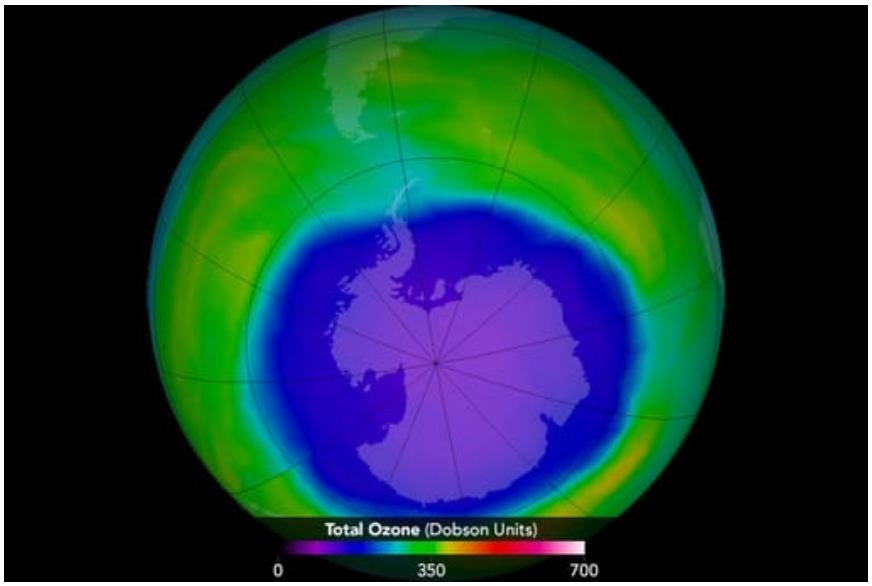


Figure 4.9.4: NASA image of Dobson units that correlates to O_3 concentration. Image used with permission (Public Domain; NASA)

Ozone is made and decomposed naturally via the Chapman cycle. In the stratosphere, ultraviolet light breaks the double bond of the diatomic oxygen molecule. From this reaction, spare oxygen (O) atoms encounter more diatomic oxygen (O_2) molecules to produce ozone. Ultraviolet light can attack ozone molecules BY CLEAVING ONE oxygen bond in the structure. This natural cycle of ozone production and decomposition repeats in the Chapman Cycle.

Figure 4.9.5: Production and decomposition of ozone via the Chapman Cycle. The image is taken from https://upload.wikimedia.org/wikiped..._cycle.svg.png

Understanding Stratospheric Ozone

Watch The following video 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

a. What is the formula for ozone? what types of bonds does this molecule have (you will need to draw)?
b. How many ozone molecules are present per 10 million molecules of air?
c. What percentage of UV rays does ozone layer absorb?
d. What chemicals are responsible for depleting ozone?
e. What are the uses of the chemicals in question #4?

f. When UV light hits ozone-depleting chemicals, what atom is liberated? How many valence electrons does this atom have?
g. Where are the areas on the earth that have ozone holes or patches?
h. Why is the largest area of depletion over Antartica?
i. At this time, is the ozone hole growing or shrinking?
j. What protocol banned the production of ozone-depleting chemicals?

Answer a

Ozone (O_3), which this molecule has a single and a double bond. The single bond is much weaker and longer. It would be easier to break this bond.

Answer b

3

Answer c

98%

Answer d

CFCs are organic compounds that contain different combinations of carbon, chlorine, fluorine, and sometimes hydrogen. The chlorine portion is particularly toxic to the ozone layer.

Answer e

Chlorine gas or chlorine that is used to disinfect drinking water is not a CFC.

Answer f

When ultraviolet light hits a CFC, chlorine is freed and becomes a free radical. This reactive species attacks the ozone molecule at the single bond location. Chlorine atom has seven valence electrons.

Answer g

Both poles contain areas that are lacking in O_3.

Answer h

The CFCS used in northern hemisphere eventually migrated to the cold areas of Antarctica. The temperatures in this area catalyze (speed up) the free radical reactions. This produces more chlorine atoms that will attack ozone molecules.

Answer i

The hole is shrinking!

Answer j

Montreal Protocol

Causes of Depletion

Before the 1920's, the mechanism of chemical refrigeration required such toxic chemicals as ammonia or sulfur dioxide. Frigidaire scientists focused on synthesizing safer compounds to be used in this industry. This research lead to the production of CFCs or compounds that contained a combination of chlorine, carbon, and fluorine. Chemists found these colorless liquids to be very volatile and inert. By the 1970's, the chemical industry was producing one million tons a year of these chemicals to be used as refrigerants, propellants, and foaming agents.



Figure 4.9.6: Asthma inhaler with CFC propellant. Image used with permission (Public Domain; US Govnerment).

Perhaps the most common CFC made was dichlorodifluoromethane. This CFC is also known as Freon-12 and was used as a refrigerant for cars, commercial, and industrial air conditioning systems. Today, the United States is no longer able to make this compound and distributors must have a Freon license to distribute existing supplies. Even though you might use the term "freon" to describe air conditioning coolant, you are not using this particular chemical due to environmental restrictions.

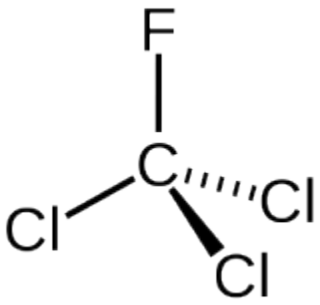


Figure 4.9.7: Chemical structure of Freon-12.

In the early to late 1970's, many international studies determined CFCs and other halons (once used as fire retardants), particularly those containing bromine, contributed to the destruction of the ozone layer. It was not until the mid-1980's that an international treaty was formulated to reduce the production of certain ODS (ozone depleting substance). In 1987, over 190 countries signed the Montreal Protocol. Upon signature, countries would agree to reduce and then eliminate ODS chemicals that contained chlorine or bromine. Countries that could not afford to find suitable replacements for ODS would be given time extensions and funding.



Figure 4.9.8: Label indicating ozone friendly packaging. Picture courtesy of Elizabeth R. Gordon

Operation Cool breeze

Even though freon production in the United States has halted, this chemical is still used in refrigeration systems that cannot switch to ozone-friendly replacements. This has opened an illegal smuggling of freon into countries that have signed the Montreal Protocol. In the United States, freon smuggling is an environmental crime. Violators that have been prosecuted have paid fines and/or served time in prison. Read the article in this link and record main ideas regarding EPA/s Operation Cool Breeze.

The United States and several other countries replaced many CFCs with HCFC compounds. By adding hydrogen, this allowed a CFC compound to retain its chlorine atoms. In other words, ultraviolet light would not be able to break the C-Cl bond releasing a chlorine free radical. Unfortunately, HCFCs were found to have higher global warming potentials (GWP) and would need to be phased out as well. For most countries, this HCFC reduction would occur in 2010. During this time, countries substituted HFCs for both HCFC and CFC compounds. HFCs have the lowest OD potential but still, have high GWP. Today, countries are focusing on finding more environmentally friendly replacements for these chemicals as well. Keep in mind, each chemical substitution that has been mentioned is costly and still affects the environment.

Table 4.9.1: Comparing CFCs, HCFCs, HFCs, and halons.

Chemicals Formula	Classification	Atmospheric lifetime (years)	ODP	GWP
CF_2Cl_3	CFC	69	1.0	4660
CF_2Cl_2	CFC	195	1.0	10,200
CF_3Br	Halon	65	13	6300
CF_2ClBr	Halon	20	5.0	1750
$CHClF_2$	HCFC	13	.055	1760
$CHCl_2CF_3$	HCFC	1.4	.02	80
CH_2F_2	HFC	5	0	677
CHF_2CF_3	HFC	29	0	3170

Free radical formation from the cleavage (breaking) of the C-Cl bond by ultraviolet light can be seen in the reactions below. It is **not** important for you to memorize these series of reactions. However, it should be noted that chlorine and bromine cause ozone depletion when they are freed from ozone-depleting substances.

Ultraviolet radiation, denoted by $h\nu$, breaks the carbon chlorine bond to form two free radicals

$$\mathrm{CF_2Cl_2 + h\ \nu \rightarrow CF_2Cl^{\cdot} + Cl^{\cdot}} \tag{4.9.1}$$

Chlorine atom being highly reactive reacts with ozone ($\mathrm{O_3}$).

$$\mathrm{Cl^{\cdot} + O_3 \ \rightarrow ClO^{\cdot} + O_2} \tag{4.9.2}$$

The monoxide of chlorine further reacts with another molecule of $\mathrm{O_3}$

$$\mathrm{ClO^{\cdot} + O_3 \ \rightarrow Cl^{\cdot} + 2\,O_2} \tag{4.9.3}$$

The chlorine atom so obtained reacts with another ozone molecule. Hence, steps 4.9.2 and 4.9.3 are repeated again and again and, leads to the depletion of concentration of ozone.

The Montreal Protocol

International policy efforts to restrict production of ozone depleting CFCs culminated in the 1987 treaty known as the **Montreal Protocol** in which signing nations agreed to cut CFC production in half by 1998. At least five follow-up agreements since then helped to deepen the cuts, advanced timetables for compliance, and addressed additional ozone-depleting substances such as halons, methyl chloroform, carbon tetrachloride, and hydrochlorofluorocarbons (HCFCs). Most countries around the world have phased out production of the substances covered by the agreements and industry has been able to shift to safer alternative chemicals. As a result, there's evidence that the Antarctic ozone hole has stopped growing worse, although recovery is not expected anytime soon. Phasing out CFCs and HCFCs is also beneficial in protecting the earth's climate, as these substances are also very damaging greenhouse gases.

As part of the United States' commitment to implementing the Montreal Protocol, the U.S. Congress amended the Clean Air Act (section 6.7), adding provisions for protection of the ozone layer. Most importantly, the amended Act required the gradual end to the production of chemicals that deplete the ozone layer. The Clean Air Act amendments passed by Congress requires the Environmental Protection Agency (EPA) to develop and implement regulations for the responsible management of ozone-depleting substances in the United States. Under the Clean Air Act, EPA has created several regulatory programs to address numerous issues, including:

- ending the production of ozone-depleting substances,
- ensuring that refrigerants and halon fire extinguishing agents are recycled properly,
- identifying safe and effective alternatives to ozone-depleting substances,
- banning the release of ozone-depleting refrigerants during the service, maintenance, and disposal of air conditioners and other refrigeration equipment,
- requiring that manufacturers label products either containing or made with the most harmful ozone depleting substances.

References

1. https://www.nap.edu/read/5744/chapter/5#23
2. https://www.epa.gov/ozone-layer-prot...ing-substances

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)
- The Montreal Protocol contributed by Caralyn Zehnder et al.

This page titled 4.9: Free Radicals and the Environment is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.10: Free Radicals and Health

Learning Objectives

- Realize how free radicals can affect your health.
- Name some health conditions and diseases that can result from excessive free radical exposure.
- Provide examples of foods and beverages that contain antioxidants.
- Recognize chemicals in skin care products that can slow the effects of aging

Free radicals can be generated from a variety of processes. For example, combustion of any material will produce $OH^{\cdot}$ (not hydroxide), $NO^{\cdot}$ free radicals. Once inhaled, OH can enter the body and affect cells. Watch the video below to see how free radicals affect biological systems.

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

How do antioxidants keep you healthy? Our body is under constant attack from oxidative stress. By wrong lifestyle choices, we make billions of free radicals in our body leading to oxidative stress. This can also increase when we are physically and/or emotionally stressed. Oxidative stress occurs when there is an imbalance of free radicals and antioxidants (too many free radicals and too few antioxidants). It is a condition that results when our body does not have enough antioxidants to counter free radicals in our body.

Free Radicals and Disease

Sources of free radicals: need this pic http://www.thepathmag.com/what-are-free-radicals/

Free radical exposure has been linked to cancer and a number of autoimmune diseases. One way of fighting free radicals is to consume or apply antioxidants. By providing sources of electrons, antioxidants can detour free radicals from attacking healthy cells in the body. The word wall below displays a variety of foods that are rich in antioxidants. Foods and beverages that are rich in vitamins A, C, and E provide sources of electrons for free radicals. Minerals like selenium are also antioxidants. For a more comprehensive list of foods are labeled as being antioxidants, click here.

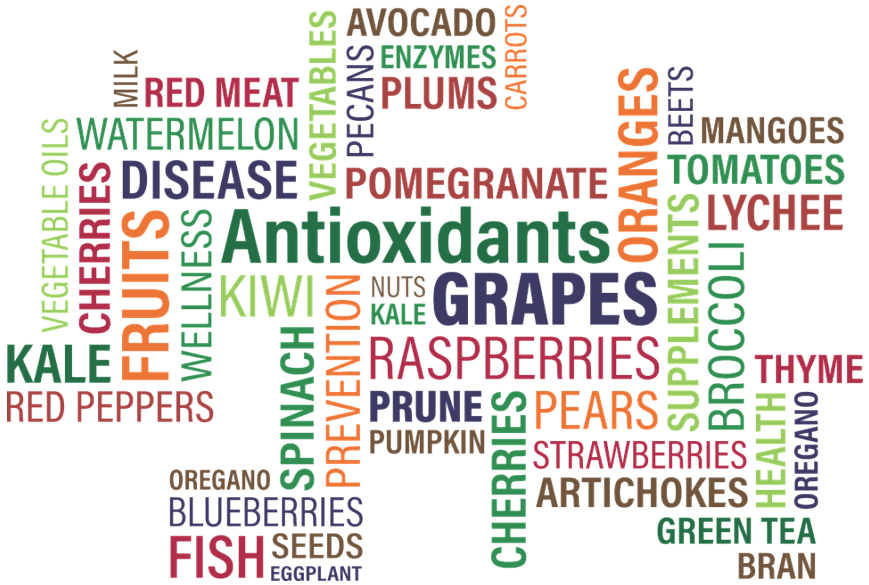


Figure 4.10.9: Word wall art displaying antioxidants. Image is taken from https://pixabay.com/en/cloud-concept...dants-18527

Free Radicals and Aging

Some studies have linked biological free radical production to aging. In a cell's mitochondria, reactive oxygen species are naturally made. These substances damage fats, proteins, and mitochondria DNA. During this process, more free radicals are produced in the cell. Two of these destructive free radicals are superoxide (O_2-) and nitrogen monoxide (NO). These radicals have been linked to cancer, skin aging, heart disease, stroke, and other chronic diseases.

By eating more fruits and vegetables, many nutritionists believe the free radical effect can be reduced in regards to chronic disease. As for aging, many scientists believe that not just one theory can explain why aging occurs. For care of skin, dermatologists recommend that you reduce UV exposure. Ionizing radiation is a source of free radicals that can be minimized. Once the skin has been exposed to UV radiation, there is very little that can be done to counteract its effects.



Figure 4.10.10: Skin comparison of a younger to an older person. Images are taken from: https://commons.wikimedia.org/wiki/F...Hand_aging.jpg and Elizabeth R. Gordon

Dermatologists and nutritionists recommend a diet rich in antioxidants (vitamins C, E, selenium, and coenzyme Q) to protect skin from damaging free radicals. Lotions, moisturizers, and sunscreens can contain these ingredients as well.

Skin Care products

Watch the following video to understand how to care for your skin

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 are the purposes of the Epidermis and dermis in your skin?
2. What do GAGS do:
3. How can UVB help the body?
4. What do the letters ROS represent?
5. How do free radicals affect DNA, proteins, Fats, and the Dermis?
6. What types of chemicals help repair damaged skin?
7. When should the chemicals in question #6 not be used?
8. How does Salicylic acid make skin appear more youthful?
9. What is the purpose of applying hyaluronic (it is a humectant) acid to skin?
10. Why are peptide products useless?
11. According to this video, what is the best anti-aging product you can purchase?

Figure 4.10.11 *Retinol*

Manufacturers of anti-aging lotions or creams might use the scientific chemical names for the antioxidant ingredients. When purchasing these products, remember to translate them accordingly. Consumers often spend excessive amounts of money on anti-aging creams and lotions. The best way to reduce advanced aging is to protect your skin from ultraviolet exposure. In addition, selecting a skin-care product that contains a combination of some of the antioxidants below is helpful. Lastly, there is no need to pay a large amount of money for an anti-aging product.

Table 4.10.2***: Antioxidants and their effect on skin***

Antioxidant	Effect on skin
Vitamin E (also known as Alpha-lipoic acid, 3-5%)	protects skin from sun damage
Vitamin A (also known as Retinoic acid, 0.01-0.1%)	treats fine wrinkles, age spots, and rough skin
Flavonoids (extracts from green tea or chocolate)	reverses sun damage
Vitamin B_6	treats fine wrinkles and protects skin from sun damage
Lycopene (a carotene that is present in tomatoes, watermelon, asparagus, and grapefruit)	protects skin from sun damage

References

1. https://www.nap.edu/read/5744/chapter/5#23
2. https://www.epa.gov/ozone-layer-prot...ing-substances

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.10: Free Radicals and Health is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.11: Applications and Solubility of Covalent Compounds

Learning Objectives

- Know the basic difference between the terms polar and nonpolar.
- Apply the like dissolves like solubility rule for covalent compounds
- Understand why molecular shapes are important to pharmaceutical chemists and biologists.
- Appreciate how medications/vitamins need to be soluble in the body.
- Pick a vitamin and memorize highlighted information
- Realize that molecules are three dimensional in nature.
- Know the applications of covalent compounds in this section

The Solubility of Covalent Compounds

Unlike ionic solubility, covalent compound solubility cannot be determined by a table. Instead, structures and three-dimensional shapes must be drawn. Once a correct geometry has been determined, the compound would be classified as being polar on nonpolar. Polar species are soluble in water, while nonpolar species are soluble in oils and fats. Covalent solubility uses the like dissolves like rule. This means that substances with the same type of polarity will be soluble in one another. Moreover, compounds with differing polarities will be insoluble in one another.

Oil and water form a heterogeneous mixture due to their differing polarities; these substances are immiscible (not mixable) (Figure $4.11.1a$). In contrast, alcohol dissolves in water to form a homogeneous mixture (Figure $4.11.1b$).



Figure $4.11.1$: Two phases (water and oil) in the same state of aggregation (liquid). Images used with permission (CC BY-SA 4.0; Victor Blacus and Elizabeth R. Gordon, respectively L to R)

In this class, we will not explore molecular geometries that are used to determine polarity. Instead, the polarity of a substance will be provided. It is important to remember that water is polar and oil/fat is nonpolar. If the polarity of a substance is given, you should be able to classify it as being water or oil/fat soluble.

Phet Simulation: Molecular Shapes

PhET Simulation

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

View this PhET Simulation online

Template:HideTOC

In determining polarity, chemists look to the power of atom's' nucleus. The protons from an atom's nucleus are capable of attracting another atom's electrons. Within a covalent bond, valence electrons are pulled toward's an atom that has a more powerful nucleus. This pull is called electronegativity. If different atoms are connected in a bond, then one tends to be more electronegative than the other. Molecules that have an overall pull in one direction are labeled as being polar species. Look at the structure of water that is shown below. This structure of this molecule shows the bonding electrons being pulled towards oxygen. Therefore, oxygen has a more powerful nucleus than hydrogen. Water's bent molecular shape does not cancel out the individual dipole pulls. As a result, water has an overall pull and is classified as being polar.

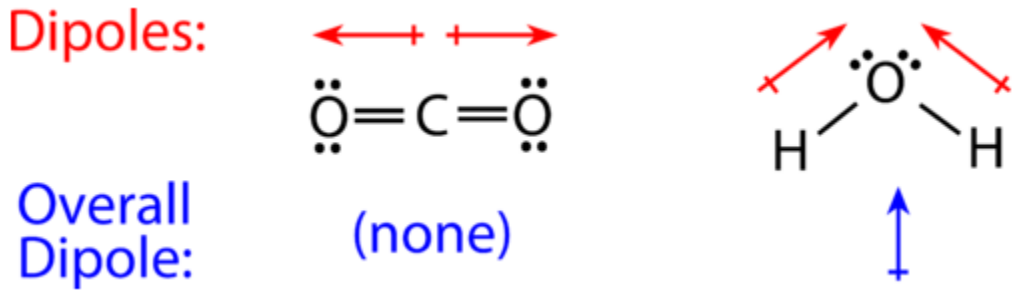


Figure 4.11.2: *The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.*

Carbon dioxide also has local dipoles (arrows) that pull in opposing directions. This molecule does not have an overall pull in one direction and is classified as being nonpolar. Some other molecules are shown in the figure below. The top three assymetrical molecules are all polar. They all have dipoles (pulls) that do not cancel. All of these molecules would be soluble in water. The bottom three molecules are nonpolar. These are symmetrical molecules that have dipoles that cancel. Both of these molecules would be oil or fat soluble.

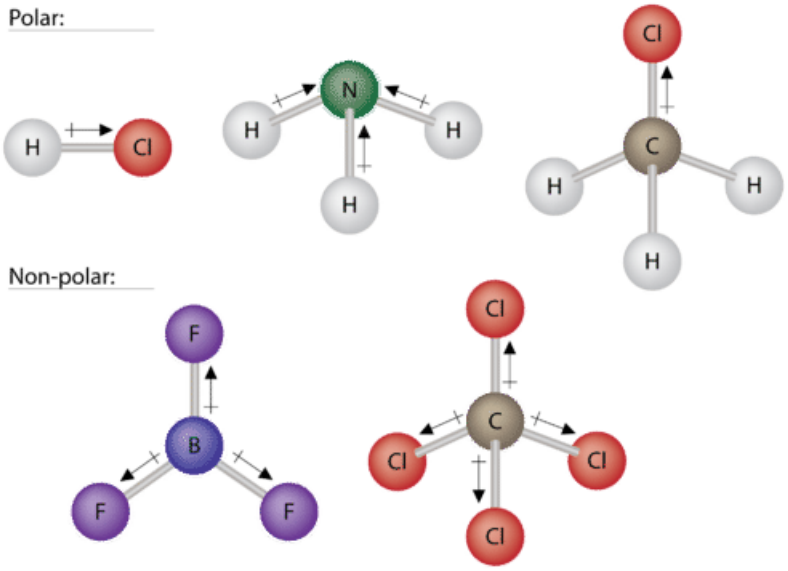


Figure 4.11.3: Some examples of polar and nonpolar molecules based on molecular geometry.

Covalent solubility is important in the pharmaceutical industry. If a medication is not water soluble, then it will not dissolve in the bloodstream and react in the active site of the body in a timely and potent fashion. Watch the video below to obtain a basic understanding of how ibuprofen travels through the body to reduce pain and/or inflammation.

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 polarity of vitamins can affect how long they remain in the body. Vitamins B and C are both water soluble and remain in the body for a short period of time. Our bodies must intake these vitamins more frequently. On the other hand, vitamins A, D, E, and K are all fat or oil soluble. Nutrients that are fat soluble remain in the body longer. These vitamins accumulate easily and can be toxic if recommended doses are surpassed.

Exercise

Pick one vitamin from the table below and memorize the common/chemical names, one source, and one function. In addition, note the solubility (water or fat) of your selected vitamin.

Vitamins with their sources and functions. Image taken from: https://upload.wikimedia.org/wikiped...mins_Table.png

Vitamin	Origination	Functions
A (retinol)	Ingested in a precursor form. Found in animal sources such as milk and eggs. Also found carrots and spinach (contain pro vitamin A carotenoids).	Vitamin A is a fat-soluble vitamin. It helps cells differentiate, also lowering your risk of getting cancer. Vitamin A helps to keep vision healthy. It is required during pregnancy. Vitamin A also influences the function and development of sperm, ovaries and placenta and is a vital component of the reproductive process.
B1 (thiamine)	Found in wheat germ, whole wheat, peas, beans, enriched flour, fish, peanuts and meats.	Vitamin B1 is a water-soluble vitamin that the body requires to break down carbohydrates, fat and protein. The body needs vitamin b in order to make adenine triphosphate (ATP). Vitamin B1 is also essential for the proper functioning of nerve cells.
B2 (riboflavin)	Found in milk cheese, leafy green vegetables, liver, soybeans yeast and almonds. Exposure to light destroys riboflavin.	Vitamin B2 is a water-soluble vitamin that helps the body process amino acids and fats. Activated vitamin B6 and folic acid helps convert carbohydrates to adenosine triphosphate (ATP). Sometimes vitamin B2 can act as an antioxidant.
B3 (niacin)	Found in beets, brewer's yeast, beef liver, beef kidney, pork, turkey, chicken, veal, fish, salmon, swordfish, tuna, sunflower seeds, and peanuts.	Vitamin B3 is required for cell respiration and helps release the energy in carbohydrates, fats, and proteins. It helps with proper circulation and healthy skin, functioning of the nervous system, and normal secretion of bile and stomach fluids. It is used in the synthesis of sex hormones, treating schizophrenia and other mental illnesses, and as a memory-enhancer.
C	Found citrus fruits such as oranges, grapefruit and lemon.	Vitamin C is an essential water-soluble vitamin. It is needed to make collagen. Vitamin C also aids in the formation of liver bile which helps to detoxify alcohol and other substances. Evidence indicates that vitamin C levels in the eye decrease with age and this may be a cause of cataracts. Vitamin C has been reported to reduce activity of the enzyme, aldose reductase, which helps protect people with diabetes. It may also protect the body against accumulation or retention of the toxic mineral, lead.
D	Produced by the human body during exposure to the ultraviolet rays of the sun.	Vitamin D is a fat-soluble vitamin that helps maintain blood levels of calcium. Vitamin D is necessary for healthy bones and teeth. Vitamin D plays a role in immunity and blood cell formation and also helps cells differentiate this lowers your chance of getting cancer.
E	Found in vegetable oils, nuts, and green leafy vegetables. Fortified cereals are also an important source of vitamin E in the United States.	Vitamin E is an antioxidant that protects cell membranes and other fat-soluble parts of the body, such as LDL cholesterol (the "bad" cholesterol), from damage.
K	Found in kale, collard greens, spinach, mustard greens, turnip greens and Brussels sprouts. Also found vegetable oils such as soybean, canola, cottonseed, and olive.	Vitamin K by helping transport Ca, vitamin K is necessary for proper bone growth and blood coagulation.
Folic Acid	Found in many vegetables including, broccoli, peas, asparagus, spinach, green leafy types. Also found in fresh fruit, liver and yeast.	Folic acid reduces the risk of a child having spine and brain defects called neural tube defects (NTDsing)
B12	Found in meat, fish, eggs and milk but not in vegetables.	Vitamin B12 is used in helping anemia.
B6 (pyridoxine)	Found in cereals, yeast, liver, and fish.	Vitamin B6 is a coenzyme in amino acid synthesis.

Applications of Covalent Compounds

Carbon dioxide Sulfur dioxide/Sulfur trioxide Dihydrogen Dioxide

Dihydrogen monosulfide Carbon monoxide Carbon tetrahydride (methane)

Contributors

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)

This page titled 4.11: Applications and Solubility of Covalent Compounds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

4.12: Drug Study

Drug Design

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

4.13: Arrhenius Acids and Bases

Learning Outcomes

- Understand that all Arrhenius Acids produce H^+ or protons in water.
- Understand that all Arrhenius Bases produce OH^- or hydroxide in water.
- Write formulas for binary and ternary Arrhenius Acids.
- Write formulas for Arrhenius Bases.
- Appreciate the pH scale.
- Know applications of selected Arrhenius Acids and Bases.
- Describe the properties of acids and bases.

pH Scale

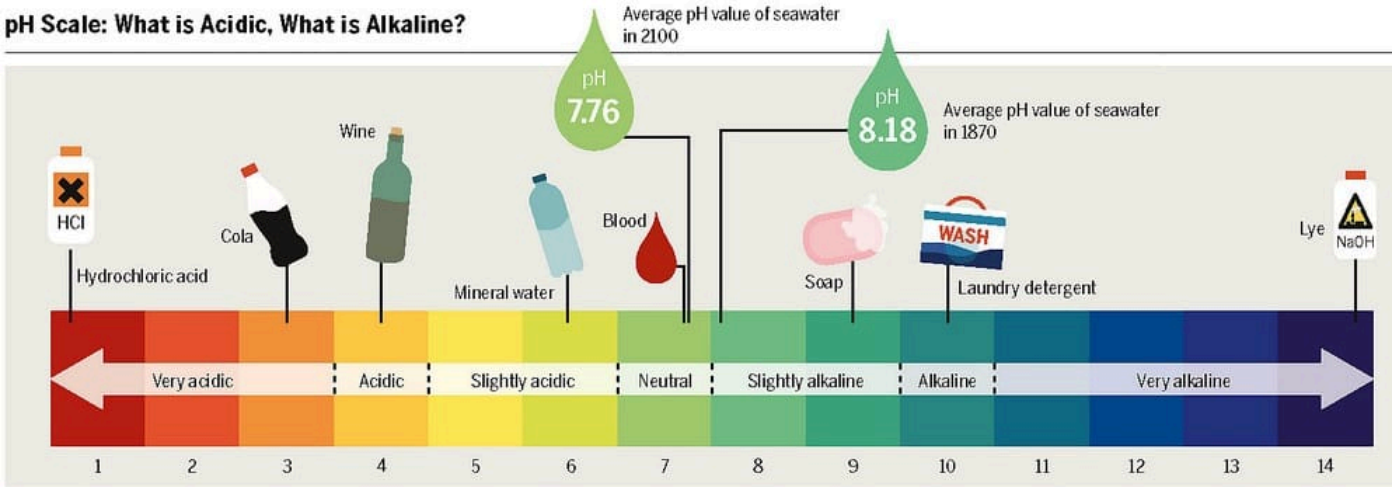


Figure 4.13.0: Average pH value of seawater in 1870 and 2100 (estimated). The difference may seem small, but the decline in the pH value from 1970 to 2100 would mean a 170 % increase in the acidity. Much small changes already post problems fro many sea creatures. pH value of various substances (Hydrochloric acid, Cola, Sopa, Laundry detergent, and more). Image used with permission (CC BY-SA 2.0; Generic; Heinrich-Böll-Stiftung).

Arrhenius Definitions

The 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 which ionizes to yield hydrogen ions* (H^+) *in aqueous solution*. An **Arrhenius base** *is a compound which ionizes to yield hydroxide ions* (OH^-) *in aqueous solution*.

Acid Properties

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 (see figure below). All carbonated beverages contain carbonic acid. This chemical easily decomposes to produce bubbly carbon dioxide and water. Acids can be commonly found in many places. For example, your stomach uses gastric acid (hydrochloric) to dissolve food. Vinegar is a mixture of water and acetic acid. This solution preserves condiments like ketchup and barbeque sauce.

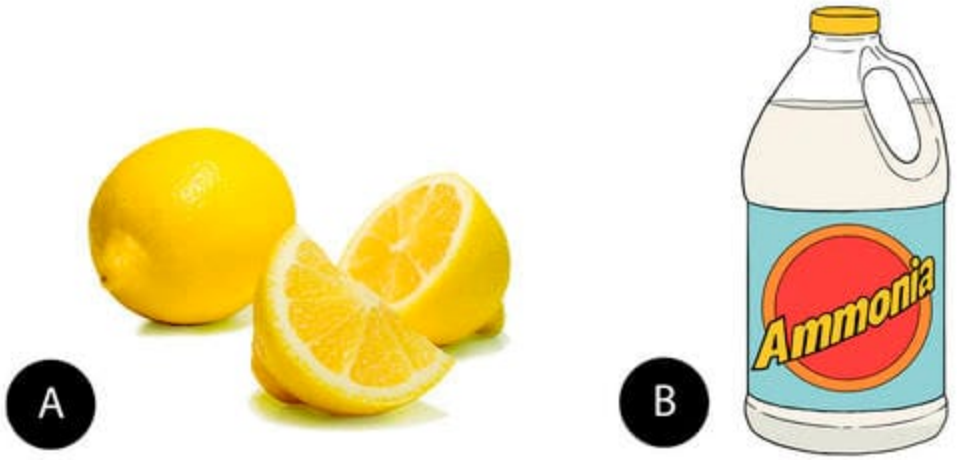


Figure 4.13.1: (A) Lemons and other citrus fruits contain citric and ascorbic acids. (B) Ammonia is a base that is present in many household cleaners.

Acids are a distinct class of compounds because of the properties of their aqueous solutions. Those properties are outlined below.

1. Aqueous solutions of acids are electrolytes, meaning that they conduct an electrical current. Some acids are strong electrolytes because they ionize completely in water. Other acids are weak electrolytes which partially ionize 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 indicators .Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein remains colorless.
4. Acids react with some metals to yield hydrogen gas.
5. Acids (pH < 7.0) react with bases (pH > 7.0) to produce a salt and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The resulting mixture will have a more neutral pH.

Acids in your world

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

Formula writing and Applications of Arrhenius Acids

If the acid name begins with the prefix " hydro" then it is a binary acid. These types of acids involve anions from the periodic table. For " hydro" acids, perform the steps below:

1. Start the formula with a proton (H^+) and refer to the periodic table for the anion species.
2. Once charges from step #1 have been obtained, cross the charges down diagonally and remove signs (+/-)
3. Acid formulas will need to display the state of matter (aq) beside the chemical formula.

If the acid name does not begin with the prefix "hydro," then it is a polyatomic or ternary acid.

1. Start the fomula with a proton (H^+) and note the ending of the acid name. Acids that have "ic" endings will use the "ate" polyatomic ion. In contrast, acids that have "ous" endings will use the "ite" polyatomic ion.
2. Cross the proton and polyatomic charges down and remove signs (+/-).
3. Acid formulas will need to display the state of matter (aq) beside the chemical formula.

Table 4.13.1: Common Acids

Acid Name	Formula	Real-world Applications
hydrochloric acid	HCl	Contained in stomach acid and liquid toliet bowl cleaner
nitric acid	HNO_3	Contained in chemical bombs and fertilizers.
sulfuric acid	H_2SO_4	Wet batter acid, chemical bomb, and fertilizers.
phosphoric acid	H_3PO_4	Preservative in dark colored colas.
acetic acid	CH_3COOH	Active ingredient of vinegar solution
hypochlorous acid	$HClO$	Acid used to sanitize water.
carbonic acid	H_2CO_3	Acid contained in all carbonated beverages

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 H_2SO_4), or triprotic (such as H_3PO_4).

Bases Properties

Bases have properties that mostly contrast with those of acids.

1. Arrhenius bases that are soluble in water can conduct electricity.
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. Red litmus turns blue in the presence of a base (see figure below), while phenolphthalein turns pink.
4. Some bases react with metals to produce hydrogen gas.
5. Acids (pH < 7.0) react with bases (pH > 7.0) to produce a salt and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The resulting mixture will have a more neutral pH.

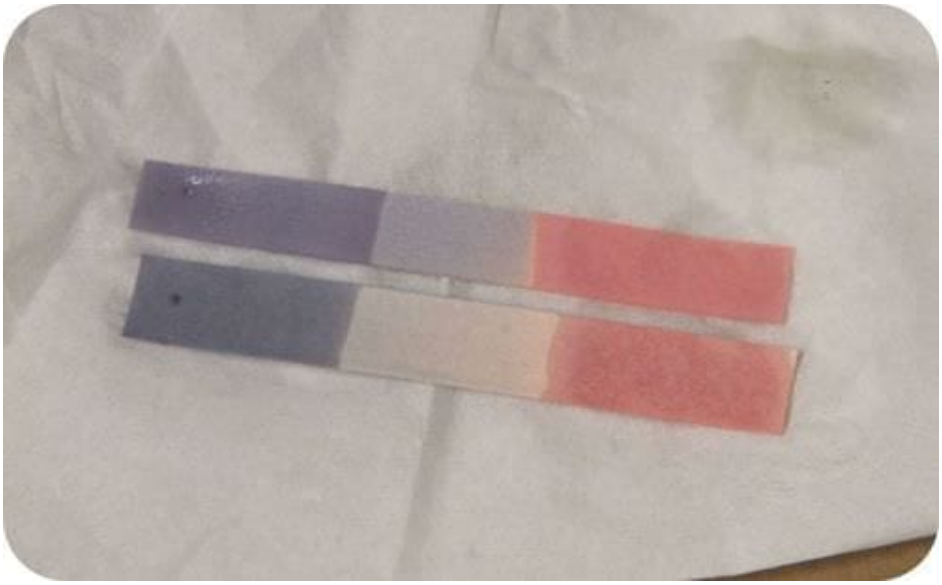


Figure 4.13.2: *Litmus paper has be treated with the plant dye called litmus. It turns red in the presence of an acid and blue in the presence of a base.*

Applications of Arrhenius Bases

Bases are ionic compounds which yield the hydroxide ion (OH^-) upon dissociating in water. The table below lists several of the more common bases.

Table $4.13.2$*: Common Bases*

Base Name	Formula	Real-world Applications
Sodium hydroxide	\(\ce{NaOHDra	
Potassium hydroxide	KOH	
Magnesium hydroxide	$Mg(OH)_2$	
Calcium hydroxide	$Ca(OH)_2$	
Ammonium hydroxide	NH_4OH	

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.

$$NaOH\,(s) \overset{H_2O}{\rightarrow} Na^+\,(aq) + OH^-\,(aq) \tag{4.13.1}$$

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 $4.13.4$: This foot has severe burns due to prolonged contact with a solution of sodium hydroxide, also known as lye.

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 (NH_3). An aqueous solution of ammonia turns litmus blue, reacts with acids, and displays various other properties that are common for bases. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by the Danish chemist Johannes Brønsted (1879 - 1947) and the 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*. Because the most common isotope of hydrogen consists of a single proton and a single electron, a hydrogen ion (in which the single electron has been removed) is commonly referred to as a proton. As a result, acids and bases are often called 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 define as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

Contributors

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

This page titled 4.13: Arrhenius Acids and Bases is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

4.E: Valence Electrons and Bonding (Exercises)

Exercises

1: Identify whether each compound(s) is/are ionic.

1. NaI
2. Au_2CO_3
3. N_2O
4. SiO_2

Answer

Only answers (a) and (b) are ionic compounds. Both formulas begin with a metal (atom from the left side of the periodic table) ion.

2: Identify whether each compound(s) is/are covalent.

a. SiH_3
b. C_6H_5Cl
c. $KC_2H_3O_2$
d. $Ca(OH)_2$

Answer

Only answers (a) and (b) are covalent compounds. Both formulas begin with nonmetals and/or metalloids (atoms from the right side of the table)

3: Write the formula for each ionic compound.

a. magnesium sulfite
b. aluminum nitride
c. mercury II hydroxide
d. ammonium oxide

Answer

a. $MgSO_3$ b. AlN c) $Hg(OH)_2$ d) $(NH_4)_2O$

4: Write the formula for each covalent compound.

a. iodine trichloride
b. disulfur dibromide
c. arsenic trioxide
d. xenon hexafluoride

Answer

a. ICl_3 b. S_2Br_2 c. AsO_3 d. XeF_6

5: Which covalent species would be considered free radicals?

a. hydrogen monoxide
b. carbon dioxide
c. phosphorus disulfide

d. ammonia (NH_3)

Answer

a. silicon tetrafluoride
b. nitrogen dioxide
c. carbon disulfide
d. diphosphorus pentoxide

Answers

1.

2.

3.
4.

9. a. ICl_3
   b. S_2Br_2
   c. AsO_3
   d. XeF_6

11. H_2O and NH_3 (water and ammonia) (answers will vary)

This page titled 4.E: Valence Electrons and Bonding (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5: BASICS OF NUCLEAR SCIENCE

Nuclear physics is the field of physics that studies atomic nuclei and their constituents and interactions. Discoveries in nuclear physics have led to applications in many fields. This includes nuclear power, nuclear weapons, nuclear medicine and magnetic resonance imaging, industrial and agricultural isotopes, ion implantation in materials engineering, and radiocarbon dating in geology and archeology. Such applications are studied in the field of nuclear engineering.

5.1: The Discovery of Radiation

5.2: The Electromagnetic Spectrum

5.3: Types of Radiation

5.4: Ionizing Radiation and Non-ionizing Radiation

5.5: The Basics of Skin Cancer

5.6: Measuring Radiation

5.7: Calculating Half-Life

5.8: Using Nuclear Science to Diagnose Disease

5.9: Therapeutic Radiation

5.10: Non-medical Applications of Nuclear Science

5.E: Basics of Nuclear Science (Exercises)

This page titled 5: Basics of Nuclear Science is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.1: The Discovery of Radiation

Learning Objectives

- Understand the difference between induced and natural radioactivity
- Know which elements have unstable nuclei.
- Correlate the relationship between binding energy and stability.

From Section 3.3, we saw how J. J. Thomson discovered the electron by using cathode ray tubes (CRT). In November of 1895, a German physicist named Wilhelm Roengten used this same technology to discover the x-ray. Roengten's ray originated from the CRT and reacted with a barium platinocyanide screen to fluoresce. After seeing this, Roengten place objects between this ray and the screen. He noted that the ray would still penetrate substances and leave an image on photographic film.

When naming this type of radiation, Roentgen reflected back to his experiments. He could not confidently explain his observations. For this reason, he decided to call this new type of radiation the x-ray. Throughout his lifetime, he continue to explore this technology and even used his wife as a test subject (refer to chapter 1 for her x-ray image). By never patenting this invention, Roengten freely shared the x-ray with other researchers and medical professionals. In 1901, he was awarded the Nobel Prize in physics for his work with the x-ray.

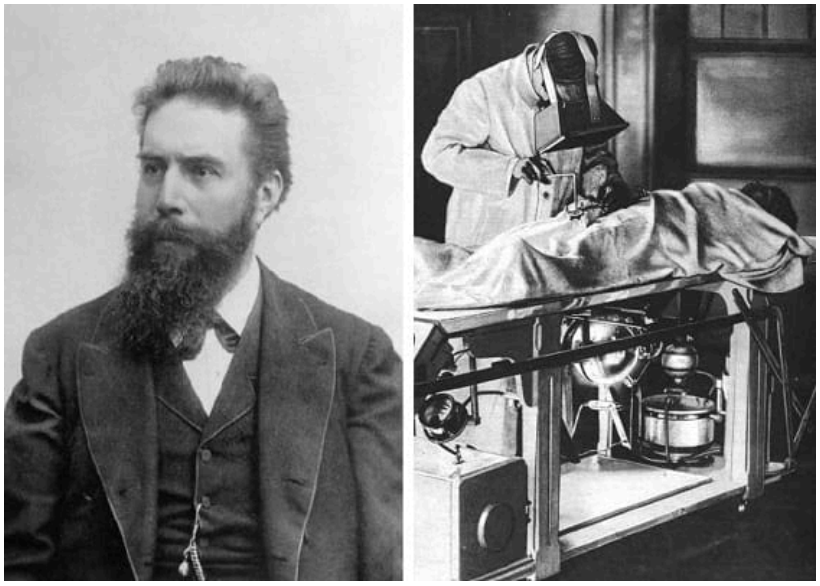


Figure 5.1.1: (left) Photo of Wilhelm Conrad Roengten, the discoverer of x-rays. (right) Surgical operation during World War I using a fluoroscope that used x-rays to find embedded bullets

As with any new technology, people looked for ways to apply the x-ray to day to day life. Thomas Edison, the inventor of the light bulb, thought the average person should have an x-ray machine in their home. He designed x-ray machines to be smaller and portable. Unfortunately, many of his technicians died from radiation poisoning.

The Harmful Effects of X-rays

In November 1896, an article entitled "the harmful effects of X-rays" was published in Nature. The witness was an X-ray demonstrator during the summer in London. He therefore paid for himself throughout the summer at the rate of several hours per day of exposure. He testified: "In the first two or three weeks I felt no inconvenience, but after a while appeared on the fingers of my right hand many dark spots which pierced under the skin, and gradually they became very painful; the rest of the skin was red and strongly inflamed My hand was so bad that I was constantly forced to bathe it in very cold water An ointment momentarily calm the pain but the epidermis had dried up had become hard and yellow like parchment and

completely insensible, so I was not surprised when my hand began to peel. "

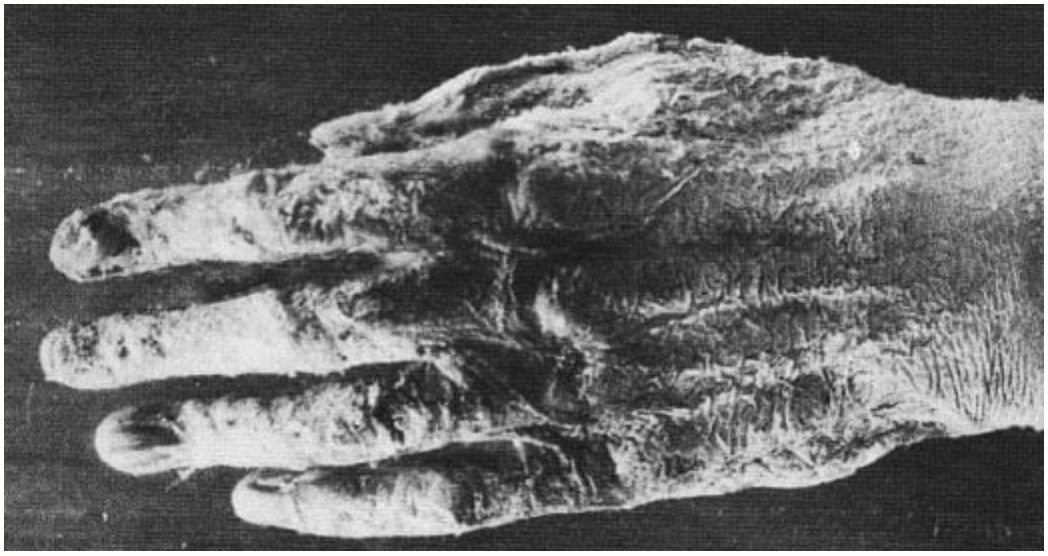


Figure 5.1.2: Radiation poisoning due to over exposure of x-rays.

"Soon the skin and nails fall off and the fingers swelled, the pain remaining constant. I lost the skin of my right and left hands, and four of my nails have disappeared from the right hand and two left and three others were ready to fall off. During more than six weeks I was unable to hold anything in my right hand and I can not hold a pen since the loss of my nails ..."

During Edison's time, people would host x-ray parties in their homes. The host of these gatherings would allow guests to x-ray different parts of their body. People would leave these events with their framed souvenirs. X-rays were even used to measure foot size. Devices like the fluoroscope were placed in shoe stores. Technicians, consumers, and observers could look in oculars while the foot was being x-rayed. Unnecessary exposure to radiation was not considered as a hazard during these times. In this era, most believed this was the most accurate way to measure the foot.

YouTube Video

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

View this YouTube Video online

Video 5.1.1***:*** *The shoe fitting fluoroscope was a common fixture in shoe stores during the 1930s, 1940s and 1950s. The first fluoroscopic device for x-raying feet may have been created during World War I to eliminating the need for patients to remove their boots, to speed up the processing of the large number of injured military personnel who were seeking help. After the war the device was modified the device for shoe-fitting and showed it for the first time at a shoe retailers convention in Boston in 1920.*

Radioactivity

When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same x-rays. Becquerel placed some of his rock crystals on top of a well-covered photographic plate and sat them in the sunlight. The sunlight made the crystals glow with a bright fluorescent light, but when Becquerel developed the film he was very disappointed. He found that only one of his minerals, a uranium salt, had fogged the photographic plate. He decided to try again, and this time, to leave them out in the sun for a longer period of time. Fortunately, the weather did not cooperate and Becquerel had to leave the crystals and film stored in a drawer for several cloudy days. Before continuing his experiments, Becquerel decided to check one of the

photographic plates to make sure the chemicals were still good. To his amazement, he found that the plate had been exposed in spots where it had been near the uranium containing rocks and some of these rocks had not been exposed to sunlight at all. In later experiments, Becquerel confirmed that the radiation from the uranium had no connection with light or fluorescence, but the amount of radiation was directly proportional to the concentration of uranium in the rock. Becquerel had discovered **radioactivity**.

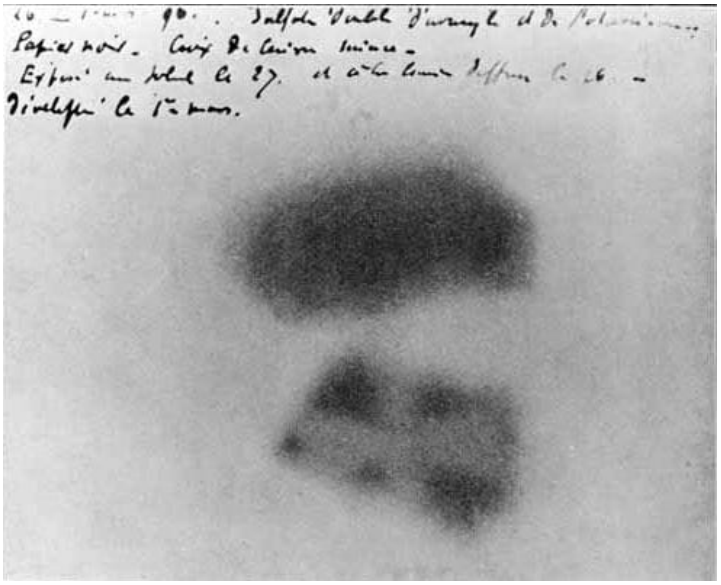


Figure 5.1.3: Image of Becquerel's photographic plate which has been fogged by exposure to radiation from a uranium salt. The shadow of a metal Maltese Cross placed between the plate and the uranium salt is clearly visible (Public Domain).

The Curies and Radium

One of Becquerel's assistants, a young Polish scientist named Maria Sklowdowska (to become Marie Curie after she married Pierre Curie), became interested in the phenomenon of radioactivity. With her husband, she decided to find out if chemicals other than uranium were radioactive. The Austrian government was happy to send the Curies a ton of pitchblende from the mining region of Joachimstahl because it was waste material that had to be disposed of anyway. The Curies wanted the pitchblende because it was the residue of uranium mining. From the ton of pitchblende, the Curies separated $0.10\ \text{g}$ of a previously unknown element, radium, in the form of the compound radium chloride. This radium was many times more radioactive than uranium.

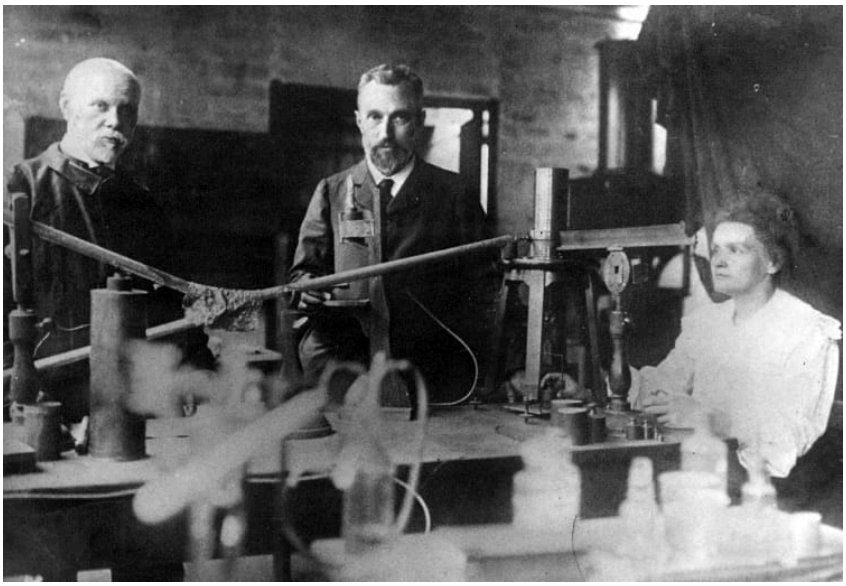


Figure 5.1.4: Marie (left) and Pierre Curie (middle) with Henri Becquerel (right) shared the 1903 Nobel Prize

By 1902, the world was aware of a new phenomenon called radioactivity and of new elements which exhibited natural radioactivity.

For this work, Becquerel and the Curies shared the 1903 Nobel Prize in Physics. These three researchers continued to work with hazardous radioactive materials. They experienced various ailments included burns and weakness by touching these substances. In 1906, Pierre Curie accidently fell under a horse drawn cart. He immediately died from him wounds and left Marie with two small children (Irene, 9 years old and Eve, 2 years old). After his death, Marie continued her research by identifying a new element, Polonium. In 1911, she was awarded the Nobel Prize in chemistry for the discoveries of radium and polonium. As of today, Marie Curie is the only person ever to have received two Nobel Prizes in the sciences.

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

Figure 5.1.2***:*** *Everyone knows that winning the Nobel Prize is a big deal, but why do we even have a Nobel Prize? And why does it matter?*

Further experiments provided information about the characteristics of the penetrating emissions from radioactive substances. It was soon discovered that there were three common types of radioactive emissions. Some of the radiation could pass easily through aluminum foil while some of the radiation was stopped by the foil. Some of the radiation could even pass through foil up to a centimeter thick. The three basic types of radiation were named alpha, beta, and gamma radiation. The actual composition of the three types of radiation was still not known.

Eventually, scientists were able to demonstrate experimentally that the alpha particle, α, was a helium nucleus (a particle containing two protons and two neutrons), a beta particle, β, was a high speed electron, and gamma rays, γ, were a very high energy form of light (even higher energy than x-rays).

Unstable Nuclei May Disintegrate

A nucleus (with one exception, hydrogen-1) consists of some number of protons and neutrons pulled together in an extremely tiny volume. Since protons are positively charged and like charges repel, it is clear that protons cannot remain together in the nucleus unless there is a powerful force holding them there. The force which holds the nucleus together is generated by **nuclear binding energy**.

A nucleus with a large amount of binding energy per nucleon (proton or neutron) will be held together tightly and is referred to as stable. These nuclei do not break apart. When there is too little binding energy per nucleon, the nucleus will be less stable and may disintegrate (come apart). Such disintegrations are referred to as **natural radioactivity**. It is also possible for scientists to smash nuclear particles together and cause nuclear reactions between normally stable nuclei. These disintegrations are referred to as **artificial radioactivity**. None of the elements above #92 on the periodic table occur on earth naturally; they are all products of artificial radioactivity (man-made).

When nuclei come apart, they come apart violently accompanied by a tremendous release of energy in the form of heat, light, and radiation. This energy comes from some of the nuclear binding energy. In nuclear changes, the energy involved comes from the nuclear binding energy. However, in chemical reactions, the energy comes from electrons moving energy levels. A typical nuclear change (such as fission) may involve millions of times more energy per atom changing compared to a chemical change (such as burning)!

Need More Practice?

- Turn to Section 5.E of this OER and answer questions #5 and #7.

Contributors

- Wikibooks
-
- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

This page titled 5.1: The Discovery of Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.2: The Electromagnetic Spectrum

Learning Objectives

- Know the difference between wave and particle radiation
- Understand the relationships of wavelength, frequency, and energy
- Calculate frequency and wavelength of radiation
- Know the order of the electromagnetic spectrum (names, not numbers)

Roentgen, Becquerel and the Curies experimented with radiation. Using machines, Roengten's focused on the production X-rays, which is a wave-type of radiation. Unlike Roengten, Henri Becquerel studied the natural radioactivity of elements. His uranium experiments yielded the production of alpha (particle-type) and gamma (ray-type) radiations. Becquerel's two graduate students, Marie and Pierre Curie, isolated two radioactive elements. Both radium and polonium emit alpha and beta particles upon decay. In this section, wave form radiation will be explored. This type of radiation has an extremely wide range of wavelengths, frequencies, and energies.

Properties of Waves

A wave is a periodic oscillation that transmits energy through space. Anyone who has visited a beach or dropped a stone into a puddle has observed waves traveling through water (Figure 5.2.1). These waves are produced when wind, a stone, or some other disturbance, such as a passing boat, transfers energy to the water, causing the surface to oscillate up and down as the energy travels outward from its point of origin. As a wave passes a particular point on the surface of the water, anything floating there moves up and down.

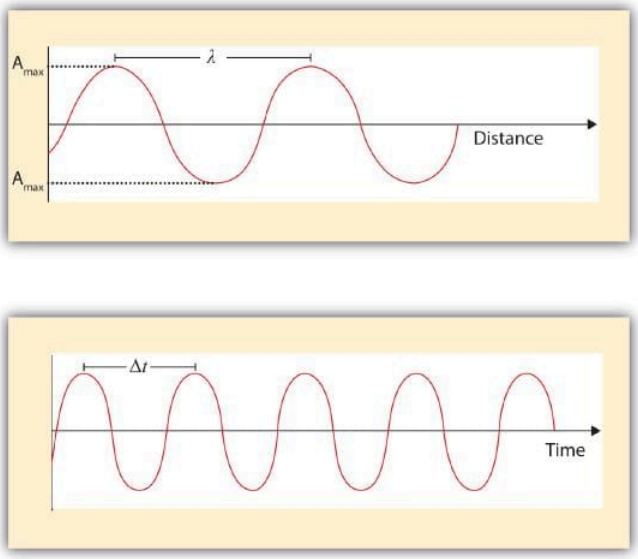


Figure 5.2.1: Characteristics of Light Waves. Light acts as a wave and can be described by a wavelength λ and a frequency ν.

Waves have characteristic properties (Figure 5.2.1). Waves are periodic, meaning, they repeat regularly in both space and time. The **distance** between two corresponding points in a wave—between the midpoints of two peaks, for example, or two troughs—is the **wavelength** (**λ**), distance between two corresponding points in a wave—between the midpoints of two peaks or two troughs. λ is the lowercase Greek lambda, and ν is the lowercase Greek nu. Wavelengths are described by a unit of distance, typically meters. The **frequency** (**ν**) **is the number of oscillations (i.e., of a wave) that pass a particular point in a given period of time**. The unit for frequency is per second (1/s = s^{-1}), which is equivalent to the SI unit of hertz (Hz). Amplitude, or vertical height of a wave, is defined as half the peak-to-trough. As the amplitude of a wave with a given frequency increases, so does its energy.

Electromagnetic Radiation

Energy that is transmitted, or radiated, through space in the form of periodic oscillations of electric and magnetic fields is known as electromagnetic radiation. (Figure 5.2.2). Some forms of electromagnetic radiation are shown in Figure 5.2.4. In a vacuum, all forms of electromagnetic radiation—whether microwaves, visible light, or gamma rays—travel at the speed of light (c), which is the speed with which all forms of electromagnetic radiation travel in a vacuum, a fundamental physical constant with a value of 2.99792458 × 10^8 m/s (which is about **3.00 ×10^8 m/s** or 1.86 × 10^5 mi/s). This is about a million times faster than the speed of sound.

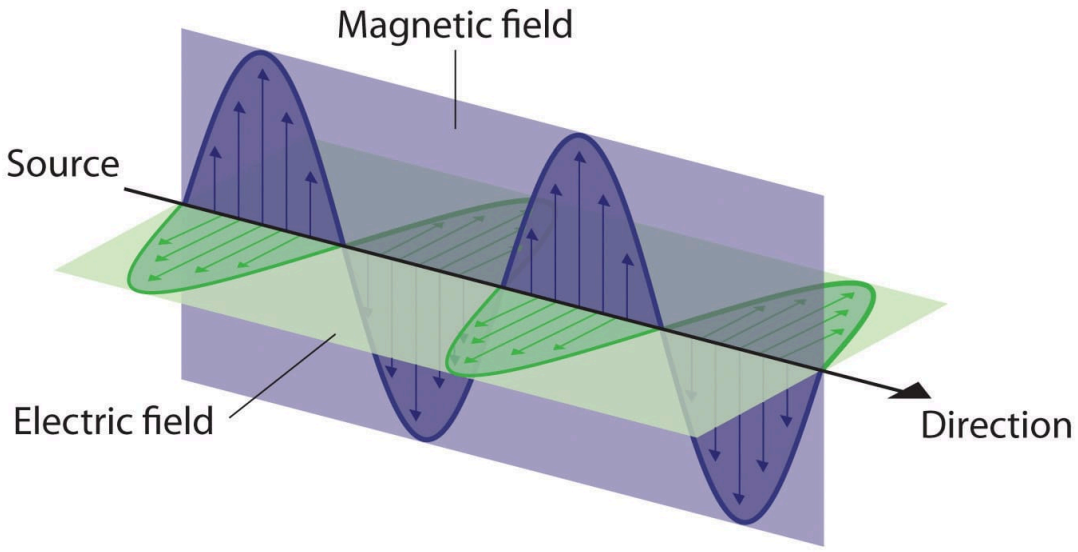


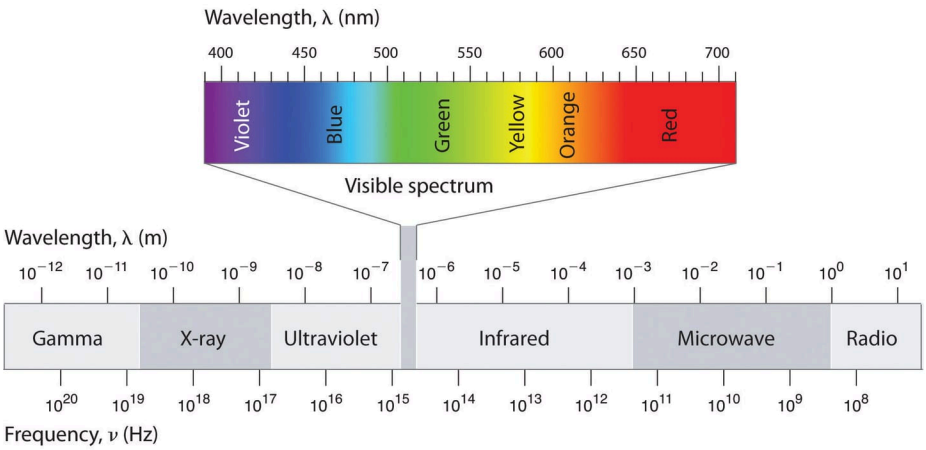
Figure 5.2.2: The Nature of Electromagnetic Radiation. All forms of electromagnetic radiation consist of perpendicular oscillating electric and magnetic fields.

Because the various kinds of electromagnetic radiation all have the same speed (*c*), they differ in only wavelength and frequency. As shown in Figure 5.2.3 and Table 5.2.1, the wavelengths of familiar electromagnetic radiation range from 10^1 m for radio waves to 10^{-12} m for gamma rays, which are emitted by nuclear reactions. Viewing the equation below, we can see how frequency of electromagnetic radiation is inversely proportional to its wavelength:

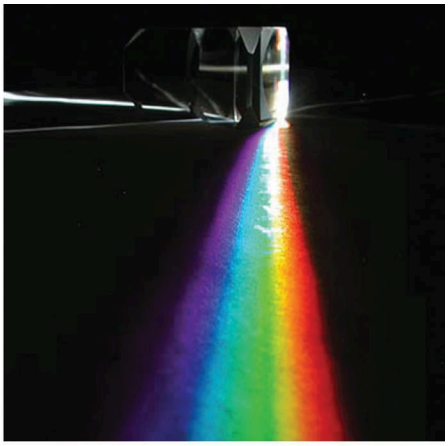
$$c = \lambda\nu$$
$$\nu = \frac{c}{\lambda} \tag{5.2.1}$$

For example, the frequency of radio waves is about 10^8 Hz, whereas the frequency of gamma rays is about 10^{20} Hz. Visible light, which is electromagnetic radiation that can be detected by the human eye, has wavelengths between about 7 × 10^{-7} m (700 nm, or 4.3 × 10^{14} Hz) and 4 × 10^{-7} m (400 nm, or 7.5 × 10^{14} Hz). Note that when frequency increases, wavelength decreases; c being a constant stays the same. Similarly when frequency decreases, the wavelength increases.

Please memorize equation 5.2.1 and the speed of light (with units). In addition, it is important to know which side of the electromagnetic spectrum is deadly.



(a)



(b)

Figure 5.2.3: The Electromagnetic Spectrum. (a) This diagram shows the wavelength and frequency ranges of electromagnetic radiation. The visible portion of the electromagnetic spectrum is the narrow region with wavelengths between about 400 and 700 nm. (b) When white light is passed through a prism, it is split into light of different wavelengths, whose colors correspond to the visible spectrum.

Table* 5.2.1*: Common Wavelength Units for Electromagnetic Radiation

Unit	Symbol	Wavelength (m)	Type of Radiation
picometer	pm	10^{-12}	gamma ray
nanometer	nm	10^{-9}	x-ray
micrometer	μm	10^{-6}	infrared
millimeter	mm	10^{-3}	infrared
centimeter	cm	10^{-2}	microwave
meter	m	10^{0}	radio

Light also behaves like a package of energy. It turns out that for light, the energy of the "package" of energy is proportional to its frequency.

$$E \propto \nu \tag{5.2.2}$$

$$E \propto \frac{1}{\lambda} \tag{5.2.3}$$

Whereas visible light is essentially harmless to our skin, ultraviolet light, with wavelengths of ≤ 400 nm, has enough energy to cause severe damage to our skin in the form of sunburn. Because the ozone layer absorbs sunlight with wavelengths less than 350 nm, it protects us from the damaging effects of highly energetic ultraviolet radiation.

In this course, we will not do energy calculations. You should know the relationship between frequency and energy. Also, you show realize that short wavelength radiation is associated with a high energy.

The energy of electromagnetic radiation increases with increasing frequency and decreasing wavelength.

Example 5.2.1

What is the frequency of light if its wavelength is 5.55×10^{-7} m?

Solution

We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^{8} m/s = \left(5.55 \times 10^{-7} m\right) \nu \tag{5.2.4}$$

We divide both sides of the equation by 5.55 × 10^{-7} m and get

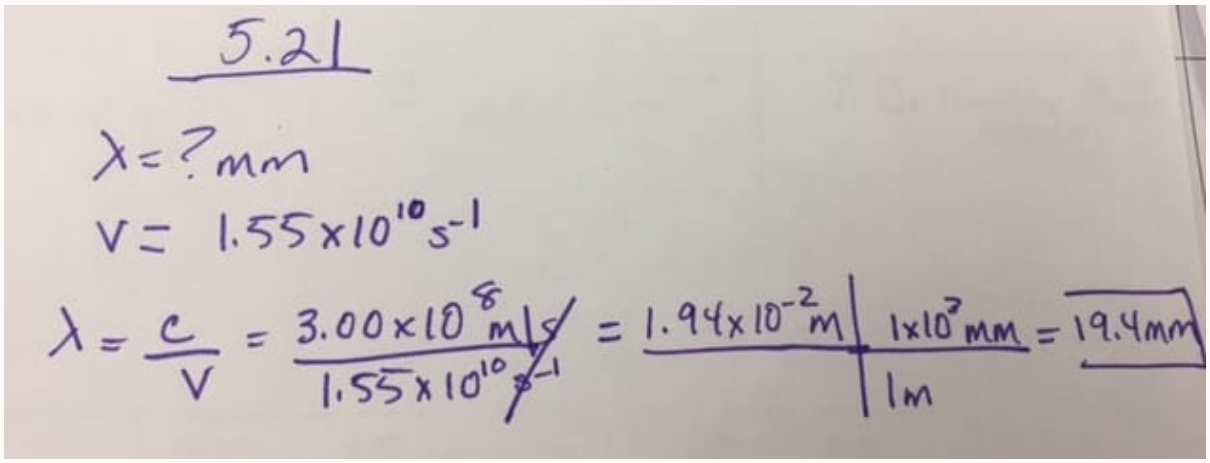
$$\nu = 5.41 \times 10^{14} s^{-1} \tag{5.2.5}$$

Note how the m units cancel, leaving s in the denominator. A unit in a denominator is indicated by a –1 power—s^{-1}—and read as “per second.”

Exercise 5.2.1

What is the wavelength (in mm) of light if its frequency is 1.55 × 10^{10} s^{-1}?

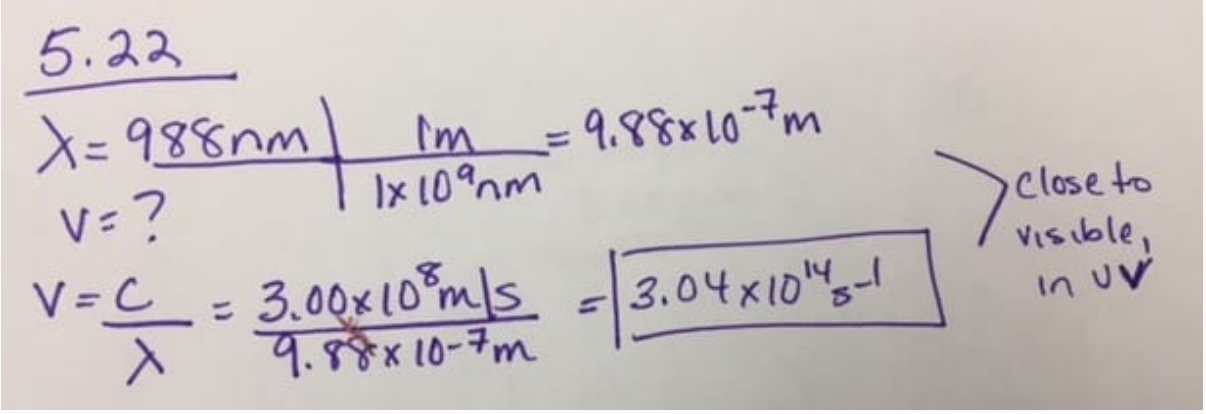
Answer



Example 5.2.2

Calculate the frequency of radiation if its wavelength is 988 nm. Where does this radiation appear in the electromagnetic spectrum?

Answer



Need More Practice?

- Refer to section 5.E of this OER and work problem #4. Refer to the electromagnetic spectrum to locate the type of radiation.

Contributors

-
- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

This page titled 5.2: The Electromagnetic Spectrum is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.3: Types of Radiation

Learning Objectives

- Express the changes in the atomic number and mass number of a radioactive nuclei when particle or ray is emitted.
- Write and balance nuclear reactions when given symbol mass format.
- Appreciate a decay series for radioactive elements.

Many nuclei are radioactive; that is, they decompose by emitting particles or rays and in doing so, become a different nucleus. In our studies up to this point, atoms of one element were unable to change into different elements. That is because in all other types of changes we have talked about only the electrons were changing. In these changes, the nucleus, which contains the protons which dictate which element an atom is, is changing. All nuclei with 84 or more protons are radioactive and elements with less than 84 protons have both stable and unstable isotopes. All of these elements can go through nuclear changes and turn into different elements.

In natural radioactive decay, three common emissions occur. When these emissions were originally observed, scientists were unable to identify them as some already known particles and so named them

- alpha particles (α),
- beta particles, (β), and
- gamma rays (γ)

using the first three letters of the Greek alphabet. Some later time, alpha particles were identified as helium-4 nuclei, beta particles were identified as electrons, and gamma rays as a form of electromagnetic radiation like x-rays except much higher in energy and even more dangerous to living systems.

This table has four columns and seven rows. The first row is a header row and it labels each column: “Name,” “Symbol(s),” “Representation,” and “Description.” Under the “Name” column are the following: “Alpha particle,” “Beta particle,” “Positron,” “Proton,” “Neutron,” and “Gamma ray.” Under the “Symbol(s)” column are the following: “ superscript 4 stacked over a subscript 2 H e or lowercase alpha,” “superscript 0 stacked over a subscript 1 e or lowercase beta,” “superscript 0 stacked over a positive subscript 1 e or lowercase beta superscript positive sign,” “superscript 1 stacked over a subscript 1 H or lowercase rho superscript 1 stacked over a subscript 1 H,” “superscript 1 stacked over a subscript 0 n or lowercase eta superscript 1 stacked over a subscript 0 n,” and a lowercase gamma. Under the “Representation column,” are the following: two white sphere attached to two blue spheres of about the same size with positive signs in them; a small red sphere with a negative sign in it; a small red sphere with a positive sign in it; a blue spheres with a positive sign in it; a white sphere; and a purple squiggle ling with an arrow pointing right to a lowercase gamma. Under the “Description” column are the following: “(High-energy) helium nuclei consisting of two protons and two neutrons,” “(High-energy) elections,” “Particles with the same mass as an electron but with 1 unit of positive charge,” “Nuclei of hydrogen atoms,” “Particles with a mass approximately equal to that of a proton but with no charge,” and “Very high-energy electromagnetic radiation.”

Figure 5.3.1: Although many species are encountered in nuclear reactions, this table summarizes the names, symbols, representations, and descriptions of the most common of these. **Please commit this chart to memory along with A/Z formats for the three Hydrogen Isotopes {H-2 = deuterium (d) and H-3 = tritium (t)}**

Alpha Decay

The nuclear disintegration process that emits alpha particles is called alpha decay. An example of a nucleus that undergoes alpha decay is uranium-238. The alpha decay of U-238 is

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th} \tag{5.3.1}$$

In this nuclear change, the uranium atom ($^{238}_{92}\mathrm{U}$) transmuted into an atom of thorium ($^{234}_{90}\mathrm{Th}$) and, in the process, gave off an alpha particle. Look at the symbol for the alpha particle: $^{4}_{2}\mathrm{He}$. Where does an alpha particle get this symbol? The bottom number in a

nuclear symbol is the number of protons. That means that the alpha particle has two protons in it which were lost by the uranium atom. The two protons also have a charge of $+2$. The top number, 4, is the mass number or the total of the protons and neutrons in the particle. Because it has 2 protons, and a total of 4 protons and neutrons, alpha particles must also have two neutrons. Alpha particles always have this same composition: two protons and two neutrons.

Another alpha particle producer is thorium-230.

$$^{230}_{90}\mathrm{Th} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{226}_{88}\mathrm{Ra} \tag{5.3.2}$$

Alpha decays occur with radioactive isotopes of radium, radon, uranium, and thorium.

Beta Decay

Another common decay process is beta particle emission, or beta decay. A beta particle is simply a high energy electron that is emitted from the nucleus. It may occur to you that we have a logically difficult situation here. Nuclei do not contain electrons and yet during beta decay, an electron is emitted from a nucleus. At the same time that the electron is being ejected from the nucleus, a neutron is becoming a proton. It is tempting to picture this as a neutron breaking into two pieces with the pieces being a proton and an electron. That would be convenient for simplicity, but unfortunately that is not what happens; more about this at the end of this section. For convenience sake, though, we will treat beta decay as a neutron splitting into a proton and an electron. The proton stays in the nucleus, increasing the atomic number of the atom by one. The electron is ejected from the nucleus and is the particle of radiation called beta.

To insert an electron into a nuclear equation and have the numbers add up properly, an atomic number and a mass number had to be assigned to an electron. The mass number assigned to an electron is zero (0) which is reasonable since the mass number is the number of protons plus neutrons and an electron contains no protons and no neutrons. The atomic number assigned to an electron is negative one (-1), because that allows a nuclear equation containing an electron to balance atomic numbers. Therefore, the nuclear symbol representing an electron (beta particle) is

$$^{0}_{-1}\mathrm{e} \text{ or } {}^{0}_{-1}\beta$$

Thorium-234 is a nucleus that undergoes beta decay. Here is the nuclear equation for this beta decay.

$$^{234}_{90}\mathrm{Th} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{234}_{91}\mathrm{Pa} \tag{5.3.3}$$

Beta decays are common with Sr-90, C-14, H-3 and S-35.

Gamma Radiation

Frequently, gamma ray production accompanies nuclear reactions of all types. In the alpha decay of U-238, two gamma rays of different energies are emitted in addition to the alpha particle.

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th} + 2{}^{0}_{0}\gamma \tag{5.3.4}$$

Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown. Nuclear reactions produce a great deal more energy than chemical reactions. Chemical reactions release the difference between the chemical bond energy of the reactants and products, and the energies released have an order of magnitude of 1×10^{3} $\mathrm{kJ/mol}$. Nuclear reactions release some of the binding energy and may convert tiny amounts of matter into energy. The energy released in a nuclear reaction has an order of magnitude of 1×10^{18} $\mathrm{kJ/mol}$. That means that nuclear changes involve almost **a million times more energy** per atom than chemical changes!

Common gamma emitters would include I-131, Cs-137, Co-60, and Tc-99.

Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown.

The essential features of each reaction are shown in Figure 5.3.2.

Decay Type	Radiation Emitted	Generic Equation	Model
Alpha decay	${}^{4}_{2}\alpha$	${}^{A}_{Z}X \longrightarrow {}^{A-4}_{Z-2}X' + {}^{4}_{2}\alpha$	Parent → Daughter + Alpha Particle
Beta decay	${}^{0}_{-1}\beta$	${}^{A}_{Z}X \longrightarrow {}^{A}_{Z+1}X' + {}^{0}_{-1}\beta$	Parent → Daughter + Beta Particle
Gamma emission	${}^{0}_{0}\gamma$	${}^{A}_{Z}X^{*} \xrightarrow{\text{Relaxation}} {}^{A}_{Z}X' + {}^{0}_{0}\gamma$	Parent (excited nuclear state) → Daughter + Gamma ray

Figure 5.3.2***:*** *Three most common modes of nuclear decay*

"Nuclear Accounting"

When writing nuclear equations, there are some general rules that will help you:

- The sum of the mass numbers (top numbers) on the reactant side equal the sum of the mass numbers on the product side.
- The atomic numbers (bottom numbers) on the two sides of the reaction will also be equal.

In the alpha decay of ${}^{238}\mathrm{U}$ (Equation 5.3.1), both atomic and mass numbers are conserved:

- mass number: $238 = 4 + 234$
- atomic number: $92 = 2 + 90$

Confirm that this equation is correctly balanced by adding up the reactants' and products' atomic and mass numbers. Also, note that because this was an alpha reaction, one of the products is the alpha particle, ${}^{4}_{2}\mathrm{He}$.

Note that both the mass numbers and the atomic numbers add up properly for the beta decay of Thorium-234 (Equation 5.3.3):

- mass number: $234 = 0 + 234$
- atomic number: $90 = -1 + 91$

The mass numbers of the original nucleus and the new nucleus are the same because a neutron has been lost, but a proton has been gained and so the sum of protons plus neutrons remains the same. The atomic number in the process has been increased by one since the new nucleus has one more proton than the original nucleus. In this beta decay, a thorium-234 nucleus has one more proton than the original nucleus. In this beta decay, a thorium-234 nucleus has become a protactinium-234 nucleus. Protactinium-234 is also a beta emitter and produces uranium-234.

$$ {}^{234}_{91}\mathrm{Pa} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{234}_{92}\mathrm{U} \tag{5.3.5}$$

Once again, the atomic number increases by one and the mass number remains the same; confirm that the equation is correctly balanced. Before starting the problems, below preview this video of your instructor balancing nuclear reactions.

Example 5.3.1

Complete the following nuclear reaction by filling in the missing particle.

$$^{210}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + ? \tag{5.3.6}$$

Solution:

This reaction is an alpha decay. We can solve this problem one of two ways:

Solution 1: When an atom gives off an alpha particle, its atomic number drops by 2 and its mass number drops by 4 leaving: $^{206}_{84}\mathrm{Po}$. We know the symbol is Po, for polonium, because this is the element with 84 protons on the periodic table.

Solution 2: Remember that the mass numbers on each side must total up to the same amount. The same is true of the atomic numbers.

- Mass numbers: $210 = 4 + ?$
- Atomic numbers: $86 = 2 + ?$

We are left with $^{206}_{84}\mathrm{Po}$

Example $5.3.2$

Write each of the following nuclear reactions.

a. Carbon-14, used in carbon dating, decays by beta emission.
b. Uranium-238 decays by alpha emission.

Solution:

a) Beta particles have the symbol $^{0}_{-1}\mathrm{e}$. Emitting a beta particle causes the atomic number to increase by 1 and the mass number to not change. We get atomic numbers and symbols for elements using our periodic table. We are left with the following reaction:

$$^{14}_{6}\mathrm{C} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{14}_{7}\mathrm{N} \tag{5.3.7}$$

b) Alpha particles have the symbol $^{4}_{2}\mathrm{He}$. Emitting an alpha particle causes the atomic number to decrease by 2 and the mass number to decrease by 4. We are left with:

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th} \tag{5.3.8}$$

Decay Series

The decay of a radioactive nucleus is a move toward becoming stable. Often, a radioactive nucleus cannot reach a stable state through a single decay. In such cases, a series of decays will occur until a stable nucleus is formed. The decay of U-238 is an example of this. The U-238 decay series starts with U-238 and goes through fourteen separate decays to finally reach a stable nucleus, Pb-206 (Figure 17.3.3). There are similar decay series for U-235 and Th-232. The U-235 series ends with Pb-207 and the Th-232 series ends with Pb-208.

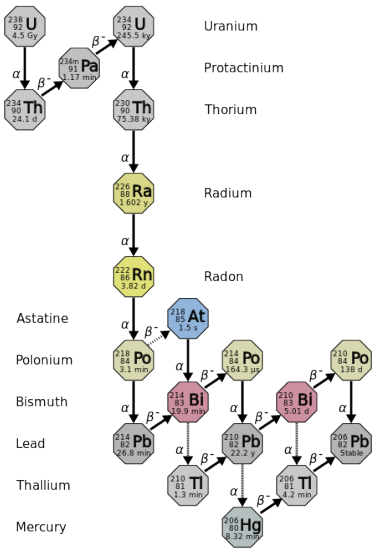


Figure 17.3.3: Uranium-238 Decay chain. Image used with permission (CC-BY-3.0 Tosaka)

Several of the radioactive nuclei that are found in nature are present there because they are produced in one of the radioactive decay series. For example, there may have been radon on the earth at the time of its formation, but that original radon would have all decayed by this time. The radon that is present now is present because it was formed in a decay series (mostly by U-238).

Need More Practice?

- Turn to Section 5.E of this OER and work problems #1 and #2.

Contributors

-
- Emma Gibney (Furman University)

This page titled 5.3: Types of Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.4: Ionizing Radiation and Non-ionizing Radiation

Skills to Develop

- Compare qualitatively the ionizing and penetration power of alpha particles (α), beta particles (β), and gamma rays (γ).
- Describe the biological impact of ionizing radiation.
- Know the most common source of background radiation and how to avoid it.

The Ionizing and Penetration Power of Radiation

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All radioactive nuclides emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause heating, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure 5.4.1).

The ability of radiation to damage molecules is analyzed in terms of what is called **ionizing power**. When a radiation particle interacts with atoms, the interaction can cause the atom to lose electrons and thus become ionized. The greater the likelihood that damage will occur by an interaction is the ionizing power of the radiation. Ionizing radiation could affect either the whole body (somatic damage) and/or eggs and sperm (genetic damage).

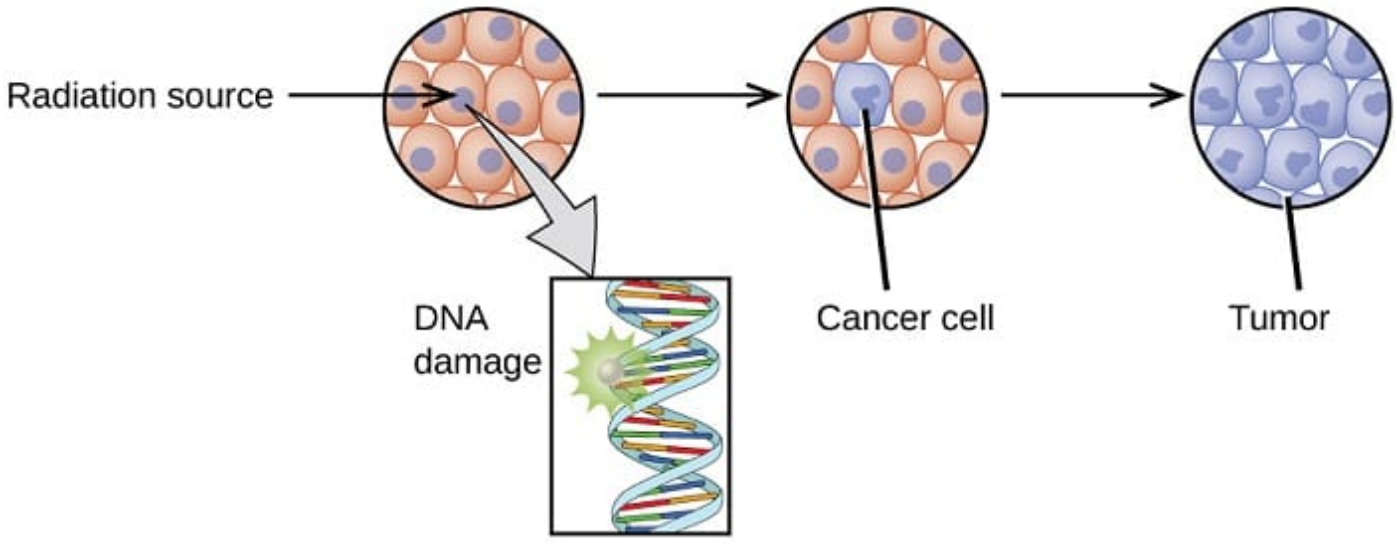


Figure 5.4.1: Radiation can harm biological systems by damaging the DNA of cells. If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer.

Much of the threat from radiation is involved with the ease or difficulty of protecting oneself from the particles or the gamma ray. How thick of a wall do you need to hide behind to be safe? The ability of each type of radiation to pass through matter is expressed in terms of **penetration power.** The more material the radiation can pass through, the greater the penetration power and the more dangerous they are. In general, the greater mass present the greater the ionizing power and the lower the penetration power.

When comparing the most common forms of ionizing radiation (alpha, beta, and gamma), alpha particles have the greatest mass. Alpha particles have approximately four times the mass of a proton or neutron and approximately ~8,000 times the mass of a beta particle (Figure 5.4.1). Because of the large mass of the alpha particle, it has the highest ionizing power and the greatest

ability to damage tissue. That same large size of alpha particles, however, makes them less able to penetrate matter. They collide with molecules very quickly when striking matter, add two electrons, and become a harmless helium atom. Alpha particles have the least penetration power and can be stopped by a thick sheet of paper or even a layer of clothes. They are also stopped by the outer layer of dead skin on people. This may seem to remove the threat from alpha particles but only from external sources. In a situation like a nuclear explosion or some sort of nuclear accident where radioactive emitters are spread around in the environment, the emitters can be inhaled or taken in with food or water and once the alpha emitter is inside you, you have no protection at all.

Figure 5.4.1***:*** *Properties of the three Primary Radiation Types*

Particle	Symbol	Mass	Penetrating Power	Ionizing Power	Shielding
alpha	α	4 amu	Very Low	Very High	Paper and Skin
beta	β	1/1837 amu	Intermediate	Intermediate	Aluminum
gamma	γ	0 (energy only)	Very High	Very Low	2 inches lead

Beta particles are much smaller than alpha particles, and therefore,they have much less ionizing power (less ability to damage tissue), but their small size gives them much greater penetration power. Most resources say that beta particles can be stopped by a one-quarter inch thick sheet of aluminum. Once again, however, the greatest danger occurs when the beta emitting source gets inside of you.

Gamma rays are not particles but a high energy form of electromagnetic radiation (like x-rays except more powerful). Gamma rays are energy that has no mass or charge. Gamma rays have tremendous penetration power and require several inches of dense material (like lead) to shield them. Gamma rays may pass all the way through a human body without striking anything. They are considered to have the least ionizing power and the greatest penetration power.

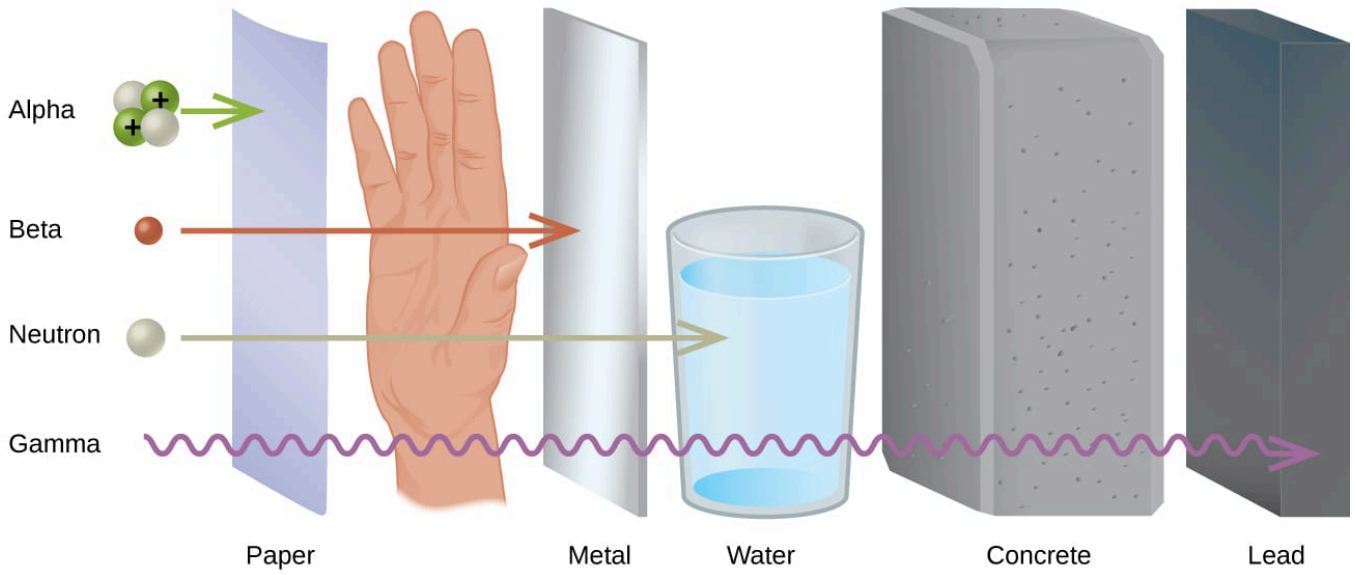


Figure 5.4.2: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma. Image used with permission (CC BY-SA, OpenStax).

The safest amount of radiation to the human body is zero. It is not possible to be not exposed to ionizing radiation so the next best goal is to be exposed to as little as possible. The two best ways to minimize exposure is to limit time of exposure and to increase distance from the source. The image below summarizes the key concepts of ionization and penetration abilities of alpha, beta, and gamma radiation.

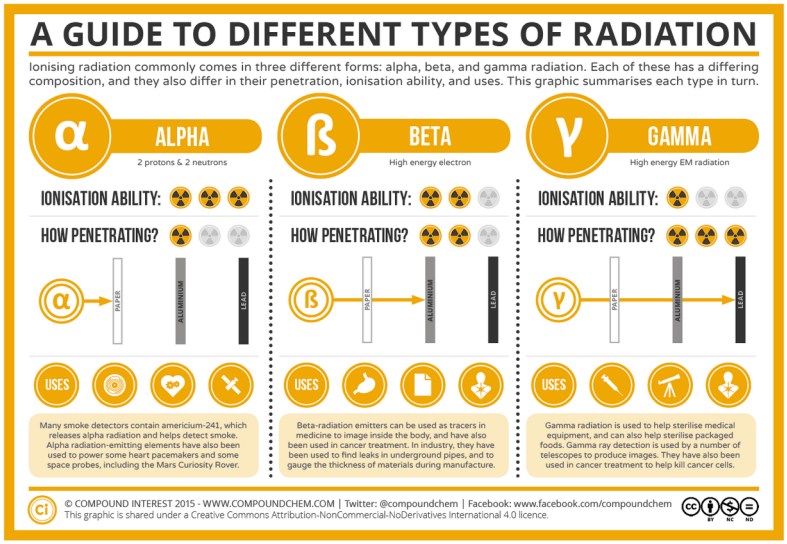


Figure 5.4.3: Ionization and penetrations abilities of alpha, beta, and gamma radiation. Image used with permission (CC BY-NC-ND, CompoundChem.com).

Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules, for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation (Figure 5.4.2).

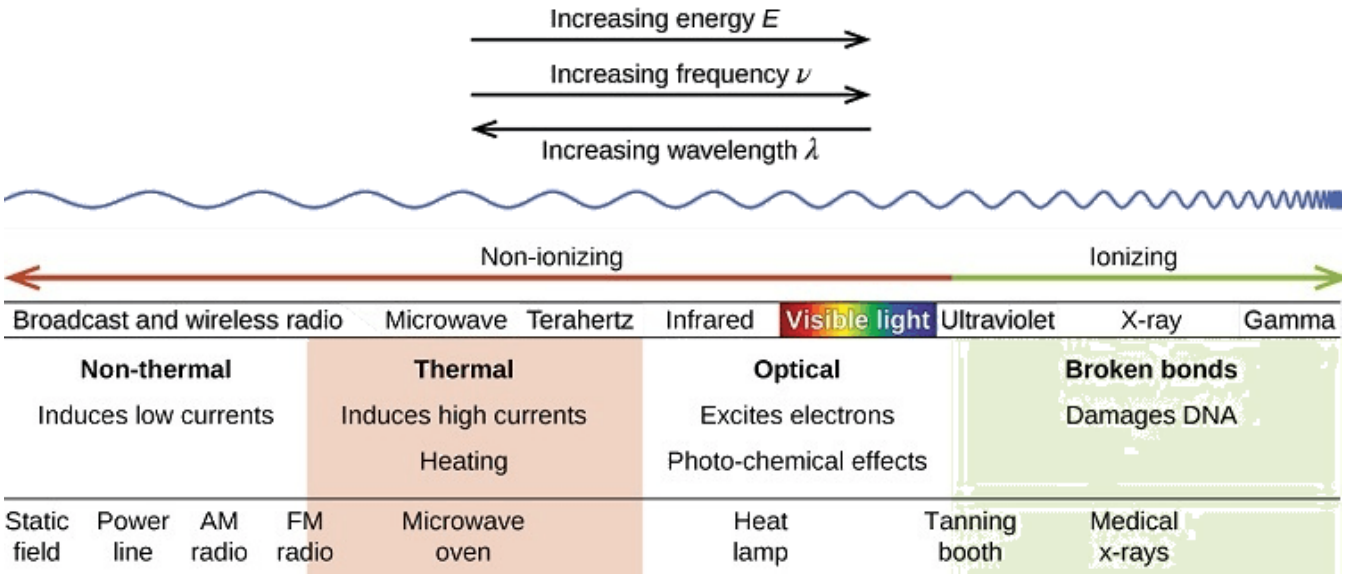


Figure 5.4.4: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing. Image used with permission (CC BY-SA, OpenStax).

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day at the beach in the sun), a large amount of nonionizing radiation is necessary before dangerous levels are reached. Forms of nonionizing radiation include wave-like radiation shown on the left side of the image. This type of radiation would include the visible spectrum through radio waves.

Everyday Exposure to Radiation

Natural radiation provides the majority of exposure to the average person. Looking at the pie chart below, the largest sources of exposure to radiation is from radon gas (Rn-222). This isotope is an α emitter with a half–life of 3.82 days. Radon is produced through the radioactive decay of U-238, which is found in trace amounts in soil and rocks. In the environment, radon concentrations can vary depending upon a geographical location. Once the soil of a particular region is disturbed, this element can escape and cause serious health issues. Please be aware that radon, and not nuclear reactors, affect more people in the United States.

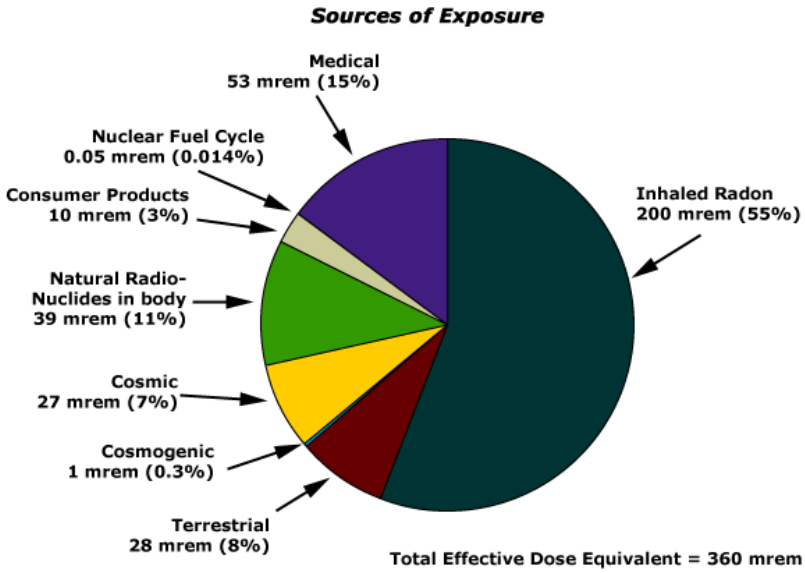


Figure 5.4.5: Sources of exposure to radiation. Image used with permission from Nuclear Science Division of Lawrence Berkeley Natioanl Lab and with CPEP (the Contemporary Physics Education Project).

Radon gas escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 5.4.5). Once airborne, the radon enters the body through inhalation or ingestion. Through alpha emission, Rn-222 decays to produce large particles. These radioactive species travel to the respiratory tract where they will ionize lung tissue. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

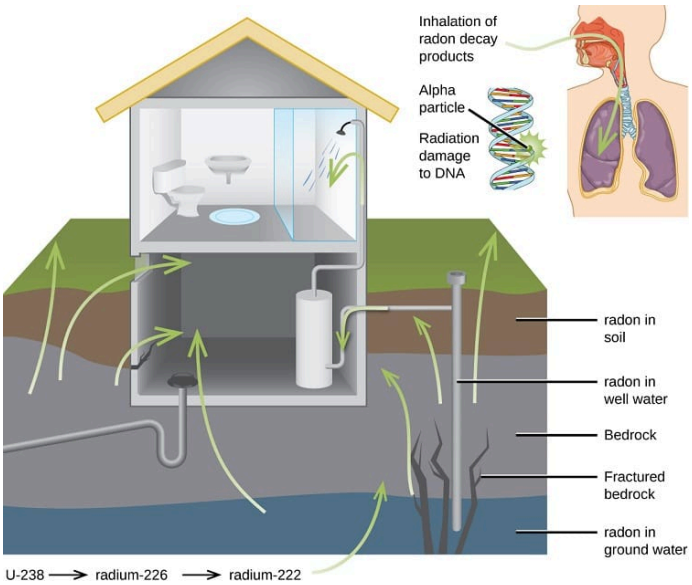


Figure 5.4.6: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinder block foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts depending on where you live. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the levels found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. The Environmental Protection Agency(EPA) tests homes throughout the United States. This agency classifies their test results into three different zones. Using color-coded keys, an individual can access their probability of being exposed to this deadly isotope. The EPA recommends an individual alter their home if the radon level exceeds 4 pCi/L. Click on this link to see your state and even county's radon level. Red-shaded areas of the EPA map indicate that tested radon levels exceed the 4 pCi/L concentration.

Radon tests can be purchased through the EPA or at most hardware stores. If you are a resident of South Carolina, you can request a free radon test kit through DHEC (Department of Health and Environmental Control). If a consumer finds excessive radon in their home, then alterations to the existing construction can be made. Some of these changes include installing barriers between the soil and the home and placing ventilation on the group floor to release radon outside. For more information regarding these types of alterations, click on this link.

Other sources of radiation in consumer products include smoke detectors, antique watches/clocks, and older ceramics/glass. Smoke detectors emit alpha particles of Am-241 (Americium). Antique watches and clocks use H-3 (tritium), Pm-147 (promethium), or Ra-226 (radium) as a fluorescent light source. Antique ceramics (dating before 1970) could contain U-238 if color is an orangey-red. If antique glassware is yellow or green in color, it could contain U-238 as well. Glassware with this isotope will glow under a black light.



Figure 5.4.7: Vaseline glass contains natural or depleted uranium. Image courtesy of Elizabeth R. Gordon

Concrete, fertilizers,kitty litter, and even food (bananas and salt substitutes) can contain trace amounts of different types of radioactivity. K-40 (potassium) is present in all of these substances.

Need More Practice?

- Turn to Section 5.E of this OER and answer questions #3, #6, #9, and #10.

References

1. https://www.epa.gov/radon

2. https://www.orau.org/ptp/collection/consumer%20products/consumer.htm

Contributors

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

This page titled 5.4: Ionizing Radiation and Non-ionizing Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.5: The Basics of Skin Cancer

Exposure to Ultraviolet Radiation

Of all the types of ionizing radiation, people come into more contact with ultraviolet radiation. Sources of this radiation would include the sun, ultraviolet lights, and tanning beds. The sun produces three different forms of ultraviolet rays (UVA, UVB and UVC). Fortunately, the deadliest of these three rays (UVC) never reaches the earth's surface and is absorbed by Earths' ozone layer. However, both UVA and UVB reach the Earth's surface to affect living systems.

Figure 5.5.1: Effect of Light on skin

UVA is produced naturally by the sun and artificially by lamps and tanning beds. For years, sunbathers believed that UVA provided only a youthful glow. Now, the American Academy of Dermatologists has condemned the use of tanning beds. Scientific research has correlated certain types of skin cancers (mainly basal and squamous cell) with the use of these devices. Besides skin cancer, tanning beds cause premature aging and damage to your eyes. If a tanning bed is not properly sanitized between uses, skin diseases can be transmitted from one consumer to another.

Recognizing Suspicious Skin Changes

According to the American Academy of Dermatologists, skin cancer is the most common form of cancer. Exposure to ultraviolet rays or arsenic can cause skin diseases. Additionally, some individuals may be genetically predisposed to developing this illness (check out your family tree). Approximately 10,000 Americans a day will be diagnosed with one of the three different forms of skin cancer. Millions of Americans have already been diagnosed with the most common form of skin cancer, basal cell carcinoma. This type of cancer grows very slowly and can spread to other parts of the body. Basal cell carcinoma frequently appears on the head and neck of a patient. Rarely, basal cell carcinoma can eventually affect other areas of the body like skin or bone.

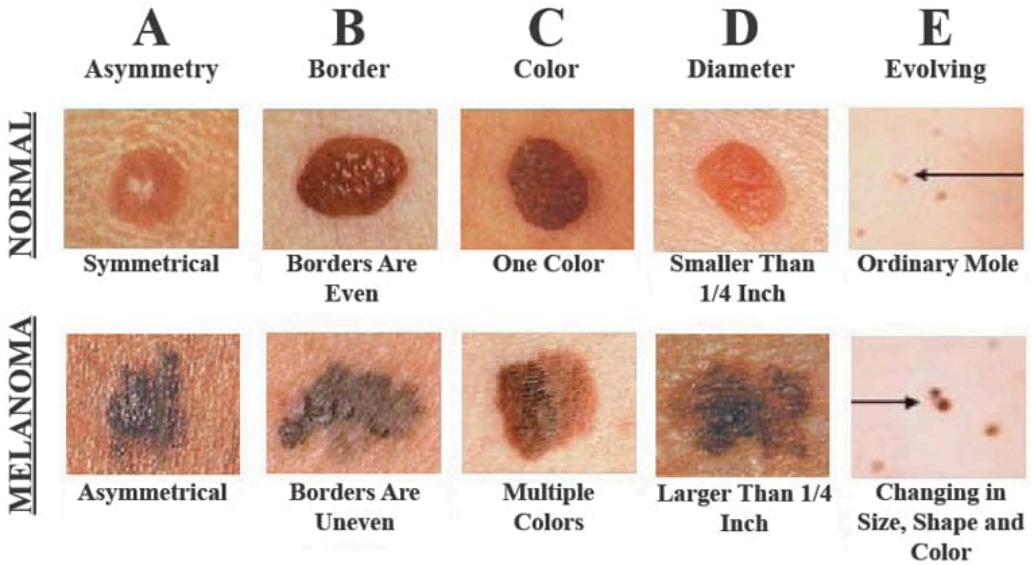


Figure 5.5.2: *ABCDEs of Skin Cancer. A brown spot on the skin is likely to be a harmless mole, but it could be a sign of skin cancer. Unlike moles, skin cancers are generally asymmetrical, have irregular borders, may be very dark in color, and may have a relatively great diameter.*

Squamous cell carcinoma appears more commonly on the face, neck, lips and hands of person. For females, squamous typically presents on the lower leg portion. This form of skin cancer is more likely to metastasize (spread) than basal cell carcinoma. If caught early and treated, patients can easily survive both basal and squamous cell carcinomas.

Melanoma, the most serious form of skin cancer, is the least common of all skin cancers. This type of cancer usually manifests on the trunks of males and the legs of females. It can spread easily to the brain, liver, bones, abdomen, or lymph notes. If detected

early, survival rates are good and life expectancy will be normal.

Figure 5.5.2 compares normal pigmentations to moles that have changed into melanoma. Monthly reviews of skin lesions and seeking medical advice is key in surviving this disease. Be sure to note these visual characteristics (the ABCDE's) of skin cancer for future evaluations (life and class).

Sunblocks and Sunscreens

UVB radiation in sunlight allows the skin to produce vitamin D. This vitamin prevents bone disorders like rickets and osteoporosis (brittle bone disease). The American Academy of Dermatology suggests vitamin D be obtained through foods or nutritional supplements. Excessive exposure to UV can be damaging and the pigment **melanin**, deposited in cells at the base of the epidermis, helps to protect the underlying layers of the skin from this damage. Melanin also colors the skin and variations in the amount of melanin produces colors from pale yellow to black. The darker the skin tone, the more melanin one has, and the less likely skin cancer will occur.



Figure 5.5.3: The author's great, great- grandmother was diagnosed with skin cancer on her nose during the 1950's.

Excess exposure to the sun can cause sunburn. This is common in humans, but light skinned animals like cats and pigs can also be sunburned, especially on the ears. Skin cancer can also result from excessive exposure to the sun. As holes in the ozone layer increase exposure to the sun's UV rays, so too does the rate of skin cancer in humans and animals.

Sunscreens and sunblocks are designed to protect skin from ultraviolet rays. *Sunblocks* contain inorganic ingredients like zinc oxide or titanium dioxide. These chemicals act as UV filters by reflecting the sun's UV rays. Sunblocks can have grainy textures due to the inorganic components. The thick nature of a sunblock can make it difficult to spread evenly on the skin. *Sunscreens* contain organic compounds like oxybenzone, avobenzone, homosalate, and octinoxate (Figure 5.5.4). By absorbing ultraviolet rays, these compounds decompose and give off heat. Sunscreens apply smoother than sunblocks. Often, manufacturers will combine sunscreen and sunblock ingredients to make their products.

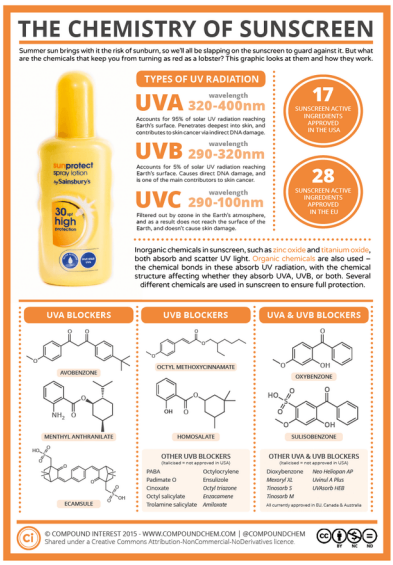


Figure 5.5.4: Effect of Light on skin

Sun protection factor (SPF) measures a product's protection from UVB rays. SPF does not quantify protection from UVA radiation. The American Academy of Dermatology recommends you select a sunscreen or sunblock with a minimum SPF factor of 30. This SPF value means a lotion can filter out 97% of UVB rays. Moving to a SPF of 50 will only filter out 1 more percentage of UVB rays. Increasing a SPF factor past 30 really does very little in shielding skin from UVB radiation.

Products that protect skin from UVA must be labeled as being broad spectrum. Sunblocks provide UVA and UVB protection, but sunscreens can vary on what they can screen. Every two hours, sunscreens/sunblocks should be reapplied to the skin. Sweating and swimming can remove sunscreen/sunblock products. No sun products are waterproof, but some are labeled as being water resistant. Consumers are encouraged to reapply these products every two hours as well.

Clothing and Other Products (need student writer?)

UPF-Ultraviolet protection factor for clothing. Blocks UVA and UVB (unlike SPF notation). Want: dark, loose, polyesters, silk. Skin Cancer Foundation Seal of Excellent Approval (UPF = 50, blocks 98% of UVA and UVB). Hats etc.

Contributors

- Elizabeth R. Gordon (Furman University)

- Template:ContribLawson
- Emma Gibney (Furman University)

This page titled 5.5: The Basics of Skin Cancer is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.6: Measuring Radiation

Skills to Develop

- Define units for measuring radiation exposure
- Explain the operation of common tools for detecting radioactivity
- List common sources of radiation exposure in the United States

Several different devices are used to detect and measure radiation, including Geiger counters, scintillation counters (scintillators), and radiation dosimeters (Figure 5.6.1). Probably the best-known radiation instrument, the Geiger counter (also called the Geiger-Müller counter) detects and measures radiation. Radiation causes the ionization of the gas in a Geiger-Müller tube. The rate of ionization is proportional to the amount of radiation. These type of devices measure alpha and beta radiation quite well. If a Geiger counter is altered, it can detect gamma and x-rays. A scintillation counter contains a scintillator (a material that emits light (luminesces) when excited by ionizing radiation) and a sensor that converts the light into an electric signal. This instrument can effectively detect gamma, x-rays and beta particles. Radiation dosimeters also measure ionizing radiation and are often used to determine personal radiation exposure. Commonly used types are electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

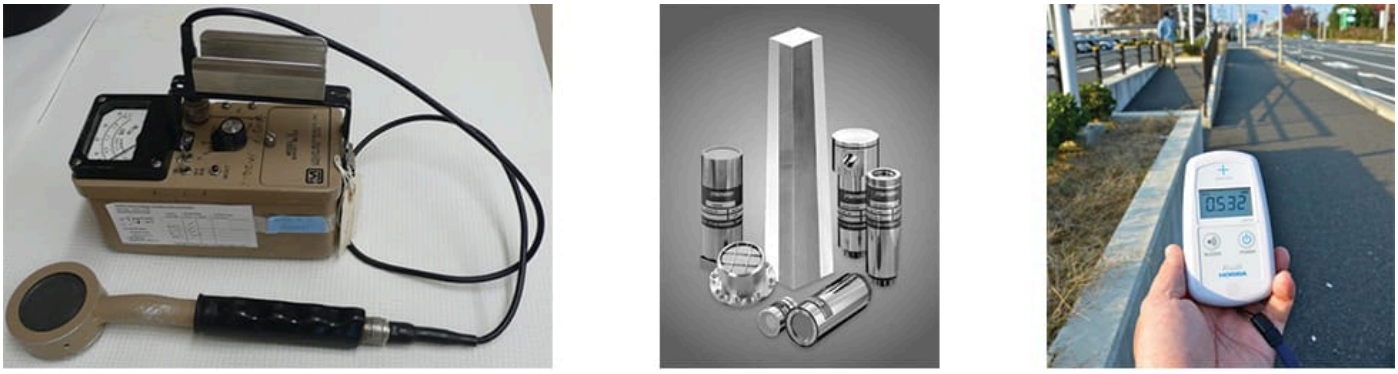


Figure 5.6.1: Devices such as (a) Geiger counters, (b) scintillators, and (c) dosimeters can be used to measure radiation. (credit c: modification of work by "osaMu"/Wikimedia commons).

A variety of units are used to measure different aspects of radiation (Table 5.6.1). The becquerel (Bq), and the curie (Ci) are used to describe rate of radioactive decay. Named in honor of Henri Becquerel and the the Curies, these units do not describe the damaging effects of radiation. Early medical applications used the units of of rad or gray (Gy) to indicate energy emitted onto an absorbing material. Today, most industries rely on the sievert (Sv) or rem. These two units measure tissue damage caused by radiation. **Looking at the chart below, please note the different units and what they describe. In CHM 101, it is not necessary for you to perform conversions (with the exception of rem/sV) with units that involve activities or absorbed doses.**

Table 5.6.1***:*** *Units Used for Measuring Radiation*

Measurement Purpose	Unit	Quantity Measured	Description
activity of source	becquerel (Bq)	radioactive decays or emissions	amount of sample that undergoes 1 decay/second

Measurement Purpose	Unit	Quantity Measured	Description
	curie (Ci)		amount of sample that undergoes 3.7×10^{10} decays/second. The body normally contains .1μCi of C-14 radiation. 1 Ci is a huge amount of radiation.
absorbed dose	gray (Gy)	energy absorbed per kg of tissue	1 Gy = 1 J/kg tissue
	radiation absorbed dose (rad)		1 rad = 0.01 J/kg tissue Acute radiation sickness occurs before 100 rad.
biologically effective dose	sievert (Sv)	tissue damage	Sv = RBE × Gy
	roentgen equivalent for man (rem)		Rem = RBE × rad An average American is exposed to around 360 mrem of radiation/yr. Acute radiation poisoning can occur at levels as low as 25 rems. Chernobyl liquidators were exposed to between 7.5×10^5 and 1.3×10^6 mrems of radiation.

The roentgen equivalent for man (rem) is the unit for radiation damage that is used most frequently in medicine (100 rem = 1 Sv). For testing or therapy, radiation values will typically be given in mrem (millirem) or mSv (millisievert). Knowing how to convert these units will enable you to compare them to given standards. In addition, understanding the exposure limits in mrem, mSv, Sv, and uSv can give you an idea of the level of toxicity involved in nuclear reactor accidents or nuclear weapons.

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

Effects of Long-term Radiation Exposure on the Human Body

It is impossible to avoid some exposure to ionizing radiation. We are constantly exposed to background radiation from a variety of natural sources, including cosmic radiation, rocks, medical procedures, consumer products, and even our own atoms. We can minimize our exposure by blocking or shielding the radiation, moving farther from the source, and limiting the time of exposure.

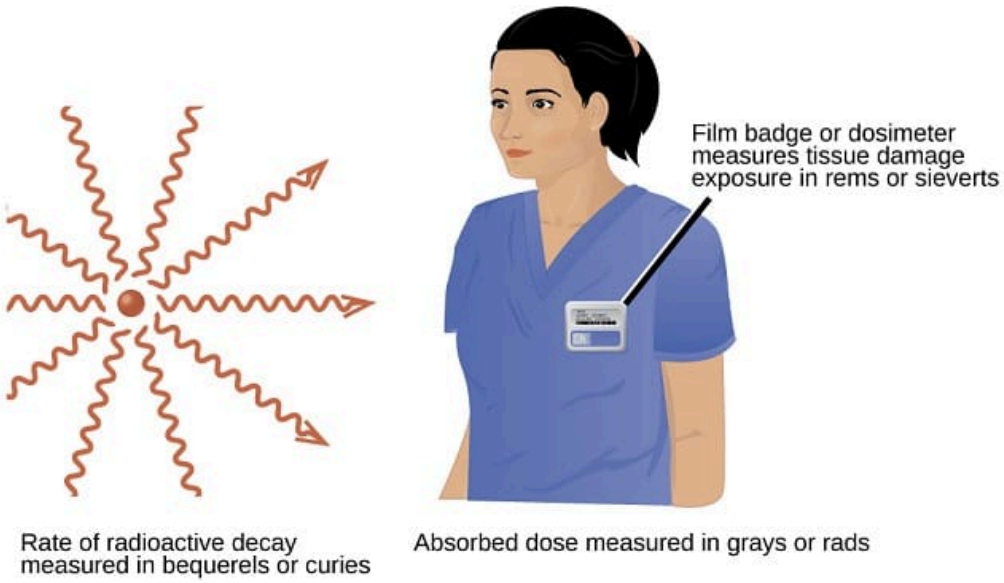


Figure 5.6.2: Different units are used to measure the rate of emission from a radioactive source, the energy that is absorbed from the source, and the amount of damage the absorbed radiation does. Image used with permission (CC BY-SA 3.0; OpenStax).

As shown in Table 5.6.2, the average person is exposed to background radiation from a variety of sources. Radon gas represents the greatest source of ionizing radiation. Indoor exposure to this alpha (α) emitter could increase the likelihood of developing lung cancer.

Table* 5.6.2*: *Average Annual Radiation Exposure (Approximate)*

Source	Amount (mrem)
radon gas	200
medical sources	53
radioactive atoms in the body naturally	39
terrestrial sources (soil)	28
cosmic sources	28
consumer products	10
nuclear energy	0.05
Total	358

Radiation and Flying

Flying from New York City to San Francisco adds 5 mrem (or 0.005 rem) to your overall radiation exposure because the plane flies above much of the atmosphere, which protects us from most cosmic radiation.

The actual effects of radioactivity and radiation exposure on a person's health depend on the type of radioactivity, the length of exposure, and the tissues exposed. Table 5.6.3 lists the potential threats to health at various amounts of exposure over short periods of time (hours or days). An average person receives around 400 mrems of radiation each year. This amount is equivalent to 0.400 rems and 4 mSV. Table 5.6.3 shows the acute toxicity level of radiation in the unit rem. Comparing yearly radiation to this table, one can get an idea of how low daily exposure levels are. Additionally, one can understand the radioactive exposure of the victims of Hiroshima and Nagasaki. Studies involving these groups of people show once exposure limits exceed 100 rem, it is most likely a person will develop some form of cancer.

Table 5.6.3*: Effects of Short-Term Exposure to Radioactivity and Radiation*

Exposure (rem)	Effect
1 (over a full year)	No detectable effect
~20	Increased risk of some cancers
~100	Damage to bone marrow and other tissues; possible internal bleeding; decrease in white blood cell count
200–300	Visible "burns" on skin, nausea, vomiting, and fatigue
>300	Loss of white blood cells; hair loss
~600	Death

Radiation Exposure

This 11 minute video shows how radiation can be measured. In addition, the scientists report their findings with various radiation units. While watching the video, answer the questions below to gain a better understanding of radiation around the world.

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 5.6.4***: Who is the most radioactive?***

Substance or Location	Amount of Radiation in μSv/hr
Banana	

Substance or Location	Amount of Radiation in μSv/hr
Range of background radiation	
Hiroshima today (event: _____)	
Uranium mine	
Trinity Site (event: _____)	
Range of radiation while flying	
Chernobyl (event: _____)	
Walking in Fukushima (event: _____)	
Pripyat, Ukraine hospital maximum reading	
CT Scan	
Fukushima resident (lifetime exposure)	
Astronaut	
Smoker	

1.Does a Geiger counter measure nonionizing radiation?

2.At an acute level, how many Sieverts (Sv) of radiation will kill a person?

3. Fill in the missing information of the chart above (events and values):

4. What two places in Marie Curie's lab are still radioactive? What type of radiation is still lingering?
5. Why did the residents of Chernobyl and Fukushima have black bags on the sides of their highways?
6. In this video, what individuals are exposed to the most radiation in μSv/hr?

References

1. 1 Source: US Environmental Protection Agency

Contributors

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

This page titled 5.6: Measuring Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.7: Calculating Half-Life

Learning Objectives

- Describe what is meant by the term half-life and what factors affect half-life.
- Calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Calculate the age of a material based upon its half-life.
- Describe how carbon-14 is used to determine the age of carbon containing objects.
- Give examples of other isotopes used in radioactive dating.
- Appreciate the half-life of isotopes involved in nuclear weapons and reactors.

Rate of Radioactive Decay

During natural radioactive decay, not all atoms of an element are instantaneously changed to atoms of another element. The decay process takes time and there is value in being able to express the rate at which a process occurs. A useful concept is **half-life (symbol is $t_{1/2}$)**, which is the time required for half of the starting material to change or decay. Half-lives can be calculated from measurements on the change in mass of a nuclide and the time it takes to occur. The only thing we know is that in the time of that substance's half-life, half of the original nuclei will disintegrate. Although chemical changes were sped up or slowed down by changing factors such as temperature, concentration, etc, these factors have no effect on half-life. Each radioactive isotope will have its own unique half-life that is independent of any of these factors.

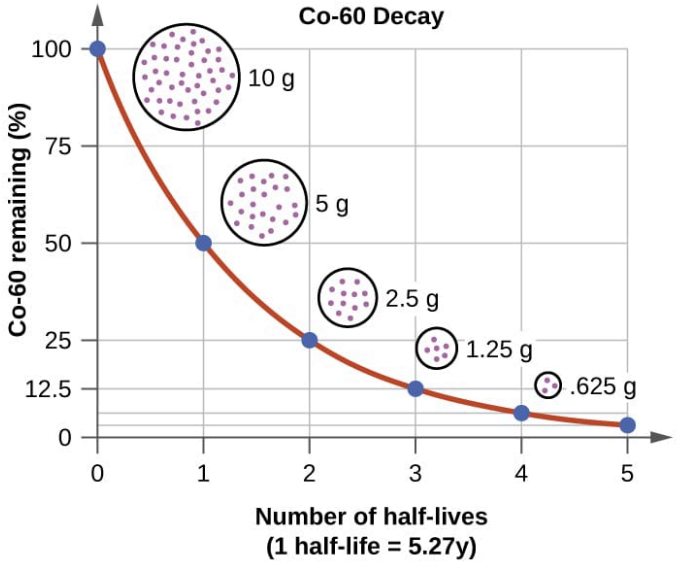


Figure 5.7.1: For cobalt-60, which has a half-life of 5.27 years, 50% remains after 5.27 years (one half-life), 25% remains after 10.54 years (two half-lives), 12.5% remains after 15.81 years (three half-lives), and so on. Image used with permission (CC-BY 4.0; OpenStax).

The half-lives of many radioactive isotopes have been determined and they have been found to range from extremely long half-lives of 10 billion years to extremely short half-lives of fractions of a second. The table below illustrates half-lives for selected elements. In addition, the final elemental product is listed after the decal process. Knowing how an element decays (alpha, beta, gamma) can allow a person to appropriately shield their body from excess radiation.

Table* 5.7.1*: Table of Selected Half-lives

Element	Mass Number (A)	Half-life	Element	Mass Number (A)	Half Life
Uranium	238	4.5 Billion years	Californium	251	800 years
Neptunium	240	1 hour	Nobelium	254	3 seconds

Element	Mass Number (A)	Half-life	Element	Mass Number (A)	Half Life
Plutonium	243	5 hours	Carbon	14	5730 years
Americium	245	25 minutes	Carbon	16	740 milliseconds

The quantity of radioactive nuclei at any given time will decrease to half as much in one half-life. For example, if there were 100 g of Cf-251 in a sample at some time, after 800 years, there would be 50 g of Cf-251 remaining. After another 800 years (1600 years total), there would only be 25 g remaining.

Remember, the half-life is the time it takes for half of your sample, no matter how much you have, to remain. Each half-life will follow the same general pattern as Cf-251. The only difference is the length of time it takes for half of a sample to decay.

Interactive Simulation: Visualizing Half-Life

Click on this interactive simulation to visualize what happens to a radioisotope when it decays. and learn about different types of radiometric dating, such as carbon dating. Understand how decay and half life work to enable radiometric dating. Play a game that tests your ability to match the percentage of the dating element that remains to the age of the object.

There are two types of half-life problems we will perform. One format involves calculating a mass amount of the original isotope. Using the equation below, we can determine how much of the original isotope remains after a certain interval of time.

$$\text{how much mass remains} = \frac{1}{2^n}(\text{original mass}) \tag{5.7.1}$$

n is the number of half-lives.

Example 5.7.1

If there are 60 grams of Np-240 present, how much Np-240 will remain after 4 hours? (Np-240 has a half-life of 1 hour)

Solution

$$\text{how much mass remains} = \frac{1}{2^4}(60\,\text{grams})$$

After 4 hours, only 3.75 g of our original 60 g sample would remain the radioactive isotope Np-240.

Example 5.7.2

A sample of Ac-225 originally contained 8.0 ug. After 720 hours, how much of the original Ac-225 remains? The half-life of this isotope is 10 days.

Solution

To determine the number of half-lives (n), both time units must be the same.

$$720\,\cancel{hours} \times \frac{1\,day}{\cancel{24\,hours}} = 30\,days$$

$$n = 3 = \frac{30\,days}{10\,days}$$

$$\text{how much mass remains} = \frac{1}{2^3}(8.0\,ug)$$

After 720 hours, 1.0 ug of the material remains as Ac-225

Radioactive Dating

Radioactive dating *is a process by which the approximate age of an object is determined through the use of certain radioactive nuclides*. For example, carbon-14 has a half-life of 5,730 years and is used to measure the age of organic material. The ratio of carbon-14 to carbon-12 in living things remains constant while the organism is alive because fresh carbon-14 is entering the organism whenever it consumes nutrients. When the organism dies, this consumption stops, and no new carbon-14 is added to the organism. As time goes by, the ratio of carbon-14 to carbon-12 in the organism gradually declines, because carbon-14 radioactively decays while carbon-12 is stable. Analysis of this ratio allows archaeologists to estimate the age of organisms that were alive many thousands of years ago.

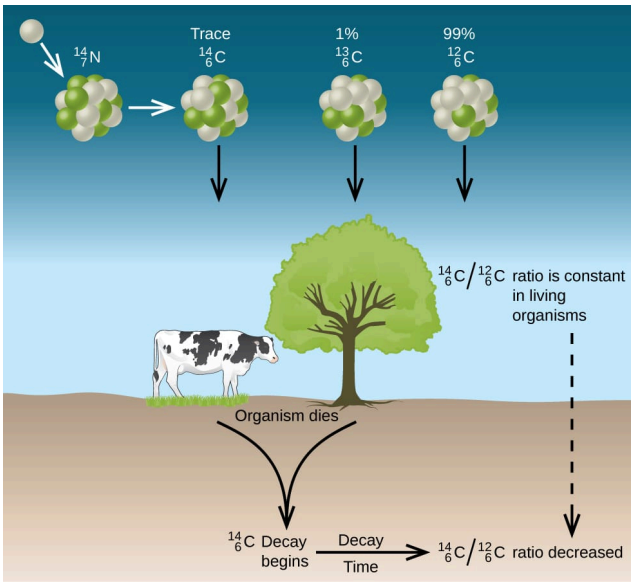


Figure 5.7.2: Along with stable carbon-12, radioactive carbon-14 is taken in by plants and animals, and remains at a constant level within them while they are alive. After death, the C-14 decays and the C-14:C-12 ratio in the remains decreases. Comparing this ratio to the C-14:C-12 ratio in living organisms allows us to determine how long ago the organism lived (and died). Image used with permission (CC-BY 4.0; OpenStax).

C-14 dating does have limitations. For example, a sample can be C-14 dating if it is approximately 100 to 50,000 years old. Before or after this range, there is too little of the isotope to be detected. Substances must have obtained C-14 from the atmosphere. For this reason, aquatic samples cannot be effectively C-14 dated. Lastly, accuracy of C-14 dating has been affected by atmosphere nuclear weapons testing. Fission bombs ignite to produce more C-14 artificially. Samples tested during and after this period must be checked against another method of dating (isotopic or tree rings).

To calculate the age of a substance using isotopic dating, use the equation below:

$$\text{how old (time)} = n \times t_{1/2} \tag{5.7.2}$$

n is the number of half-lives and $t_{1/2}$ is the half-life (in time).

Example 5.7.3

How long will it take for 18.0 grams of Ra-226 to decay to leave a total of 2.25 grams? Ra-226 has a half-life of 1600 years.

Solution

18.0g ⇒ 9.0g ⇒ 4.5g ⇒ 2.25g, this is three half-lives

$$\text{how old (time)} = 3 \times 1600\,years$$

This decay process takes 4800 years to occur.

Radioactive Dating Using Nuclides Other than Carbon-14

Radioactive dating can also use other radioactive nuclides with longer half-lives to date older events. For example, uranium-238 (which decays in a series of steps into lead-206) can be used for establishing the age of rocks (and the approximate age of the oldest rocks on earth). Since U-238 has a half-life of 4.5 billion years, it takes that amount of time for half of the original U-238 to decay into Pb-206. In a sample of rock that does not contain appreciable amounts of Pb-208, the most abundant isotope of lead, we can assume that lead was not present when the rock was formed. Therefore, by measuring and analyzing the ratio of U-238:Pb-206, we can determine the age of the rock. This assumes that all of the lead-206 present came from the decay of uranium-238. If there is additional lead-206 present, which is indicated by the presence of other lead isotopes in the sample, it is necessary to make an adjustment.

Potassium-argon dating uses a similar method. K-40 decays by positron emission and electron capture to form Ar-40 with a half-life of 1.25 billion years. If a rock sample is crushed and the amount of Ar-40 gas that escapes is measured, determination of the Ar-40:K-40 ratio yields the age of the rock. Other methods, such as rubidium-strontium dating (Rb-87 decays into Sr-87 with a half-life of 48.8 billion years), operate on the same principle. To estimate the lower limit for the earth's age, scientists determine the age of various rocks and minerals, making the assumption that the earth is older than the oldest rocks and minerals in its crust. As of 2014, the oldest known rocks on earth are the Jack Hills zircons from Australia, found by uranium-lead dating to be almost 4.4 billion years old.

An ingenious application of half-life studies established a new science of determining ages of materials by half-life calculations. For geological dating, the decay of U-238 can be used. The half-life of U-238 is 4.5×10^9 years. The end product of the decay of U-238 is Pb-206. After one half-life, a 1.00 gram sample of uranium will have decayed to 0.50 grams of U-238 and 0.43 grams of Pb-206. By comparing the amount of U-238 to the amount of Pb-206 in a sample of uranium mineral, the age of the mineral can be estimated. Present day estimates for the age of the Earth's crust from this method is at 4 billion years. This radioactivity approach can be used to detecting fake wine vintages too.

Isotopes with shorter half-lives are used to date more recent samples. For example, tritium ($t_{1/2}$ =12.3 years) can date samples within an age range of 1-100 years. Chemists and geologists use tritium dating to determine the age of water (ocean and fresh). In addition, tritium dating can be useful in determining the age of wines and brandies.

Summary and Vocabulary

The half-life of an isotope is used to describe the rate at which the isotope will decay and give off radiation. Using the half-life, it is possible to predict the amount of radioactive material that will remain after a given amount of time. C-14 dating procedures have been used to determine the age of organic artifacts. Its half-life is approximately 5700 years.

- *Background radiation*: Radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes. These natural sources of radiation account for the largest amount of radiation received by most people.
- *Half-life*: The half-life of a radioactive substance is the time interval required for a quantity of material to decay to half its original value.

Contributors

-

- Emma Gibney (Furman University)

This page titled 5.7: Calculating Half-Life is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.8: Using Nuclear Science to Diagnose Disease

Learning Objectives

- Understand the difference between diagnostic and therapeutic radiation.
- Know what testing involve nuclear chemistry.
- Calculate dosages of radiation received with a nuclear test.
- Compare nuclear testing dosages to toxic concentrations.

Diagnostic medical applications involve testing for a disease or condition. In nuclear medicine, this could involve using x-rays, CT or PET scans, or isotopic studies. The radiation involved for each of these types of tools will vary in mrem or mSv amounts.

X-Ray Imaging

X-Ray imaging is one of the most basic and routine forms of imaging within modern medicine. Using electromagnetic radiation, scientists and doctors have the ability to visualize internal situations. X-Ray imaging takes advantage of the fact that dense structures (i.e., bone) absorb more x-rays than the softer tissues that surround it. By projecting radiation onto special film paper, the denser substance casts a shadow where less x-rays pass through the body, and a contrasting image is produced. The radiation emitted by an X-Ray is of the wavelength of 10 to 0.01 nanometers (nm). Because some parts of the body may be thicker than others, X-Rays are used in "soft" (10 to 0.10 nm) and "hard" (0.10 to 0.01 nm) dosages or radiation.

X-rays are produced by passing an electron beam through a tungsten anode in a vacuum. The electromagnetic waves that are released can be focused and adjusted according to what is imaged. X-Rays can image in very high definition at a low cost, but they provide poor contrast in soft tissues. Contrast agents can also be added to the body prior to imaging, which allows different tissues to absorb more radiation, creating a better contrasting image. This allows doctors and scientists to work with much more specific images within their diagnosis or research (i.e., the ability to determine different aspects of soft-tissue).

Projectional radioagraphy (Figure 5.8.1) can be done standing or laying down. To obtain clearer images, a patient needs to be as still as possible. A technician may ask to take a big breath and hold it right before pushing the scan button. This allows the patient to keep their chest and abdomen in position and will allow a clearer image to be captured. Breathing will cause movement which will result in a blurred image.

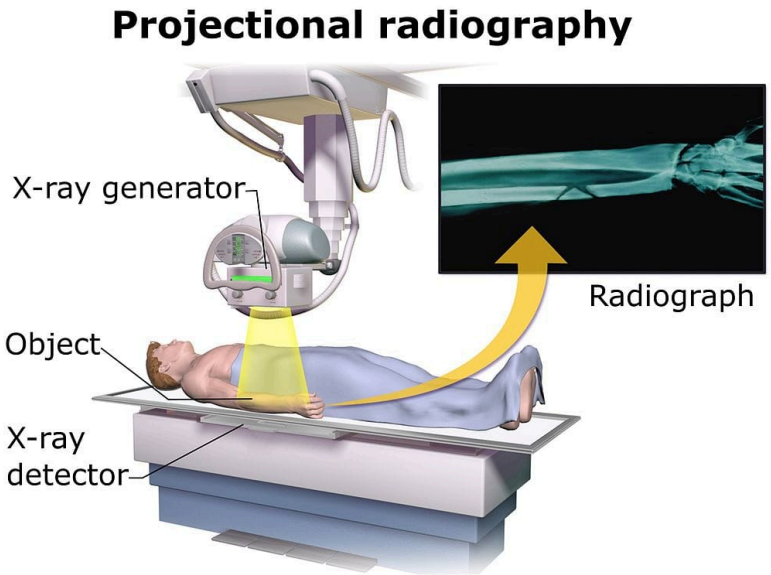


Figure 5.8.1: *Acquisition of projectional radiography, with an X-ray generator and a detector. Image used with permission (CC BY-SA 4.0; Blausen Medical).*

Before having an x-ray or CT scan, patients should remove all metallic clothing or jewelry. Metals can distort x-ray images easily.

Sometimes, lead aprons are provided to patients. The lead aprons shields a patient's reproductive organs from ionizing radiation. X-rays and CT scans can harm a developing embryo or fetus. For this reason, female patients are required to sign a liability wavier stating they are not pregnant at the time of testing. Other options like MRI or ultrasounds (nonionzing techniques) are available for these patients.

Diagnostic x-rays differ greatly in millisievert (mSV) dosages. Table 5.8.1 illustrates a few of these values. Using this table, you can calculate rem and mrem amounts and compare it to toxic values shown in Section 5.6.

***Table** 5.8.1**:** Diagnostic x-rays levels*

Type of X-ray Exam	Diagnostic Reference level (DRL) in mSv
Chest	0.96
Abdomen	7.10
C-spine	3.56
T-sping	8.00
L-Spline	12.0
Skull	6.00
Lumb	7.21
Hip	7.21
Source: Sonawane AU[1], Shirva VK, Pradhan AS., Radiat Prot Dosimetry. 2010 Feb;138(2):129-36. doi: 10.1093/rpd/ncp235.	

Example 5.8.1: Chest X-ray

A chest x-ray yields a dosage of 6 mrems of radiation to a patient. Calculate this dosage in rems and millisieverts (mSv). Compare these values to your annual background dose (around 360 mrems) and cancer dose (100 rems).

Solution

For the first calculation, recall the metric prefix milli from Chapter 2.

$$\frac{6\ \cancel{mrem}}{1}\ \frac{1\,rem}{1000\ \cancel{mrem}} = 6 \times 10^{-3}\,rem$$

For the second calculation, you will need to insert the 100 rem = 1 Sv factor. Then, you will need to use your milli metric prefix again.

$$\frac{6 \times 10^{-3}\ \cancel{rem}}{1}\ \frac{1\ \cancel{Sr}}{100\ \cancel{rem}}\ \frac{1000\,mSr}{1\ \cancel{Sr}} = 6 \times 10^{-2}\,mSr$$

The dose of 6 mrem dose is significantly lower than the level of background radiation. In addition, the calculated $6 \times 10^{-3}\ rem$ value is well below the 100 rem dosage that is associated with development of cancer.

Example 5.8.2: Mammogram

Your grandmother/mother goes to get her yearly mammogram (x-ray image that helps detect breast cancer). The radiation level for this scan is .36mSv. Convert this value to mrem and rem. Compare calculated values to background and cancer causing levels (shown in previous problem).

Figure 5.8.2: Sharon Grubbs, 2d Medical Group mammography technologist, demonstrates the mammography machine. Image used with permission (Public Domain; U.S. Air Force photo by Staff Sgt. Mike Andriacco)

$$\frac{0.36\ \cancel{mSv}}{1}\ \frac{1\ \cancel{Sv}}{1000\ \cancel{mSv}}\ \frac{100\,rem}{1\ \cancel{Sr}} = 3.6 \times 10^{-2}\,rem$$

$$\frac{3.6 \times 10^{-2}\ \cancel{rem}}{1}\ \frac{1000\,mrem}{1\ \cancel{rem}} = 36\,mrem$$

36 mrems is less than an average person's background radiation of 360 mrems. In addition, this value is 96.4 rems below the cancer level of 100 rems.

CT Scanning

Imagine a patient enters the hospital complaining of severe chest pain. After consulting the patient, the doctors decide that the source of the patient's problems is a constricted artery. However, the human body contains several long arteries. The doctors could use surgery to find the source of the clot, but this could take many hours and with no definite guarantee of success. This is exactly where radiation helps medical professionals bridge the gap that once existed. Computerized Axial Tomography (CT) scans use X-ray particles along with advanced computer technology in order to produce highly detailed cross sectional images of the body. Using these scans, medical professionals can accurately find clots and other other harmful medical conditions.

CT/CAT scans have the same fundamentals of x-rays, but they produce 3-D images using tomography. Using digital geometric processing, the CT scan creates a 3-D image compiling a large number of 2-D x-rays. These 2-D images are thin cross sections of the body taken from a single axis of rotation. A CT scanner consists of a table where the patient lies on, an x-ray device which rotates at high speed around the patient taking hundreds of cross sections of the body, and a computer with appropriate software to compile and render the 3-D image. The table in Figure 1 moves through the tube, allowing the machine to take hundreds of pictures from head to toe.



Figure 5.8.3: PET/CT-System with 16-slice CT; the ceiling mounted device is an injection pump for CT contrast agent. Image used with permission (Public Domain; Hg6996).

Similar to x-ray, contrast agents are typically added to the body before imaging occurs so that different tissues are more defined. Due to the fact that many more images are being taken in order to form a complete three dimensional image, the radiation dose is much higher. While a chest x-ray is around 0.10 mSv or radiation, a chest CAT scan subjects the patient to about 8 mSv of radiation; over twice the amount that the body would absorb from its environment in a year. For this reason, it is uncommon and unsafe to have multiple CAT scans taken unless completely necessary.

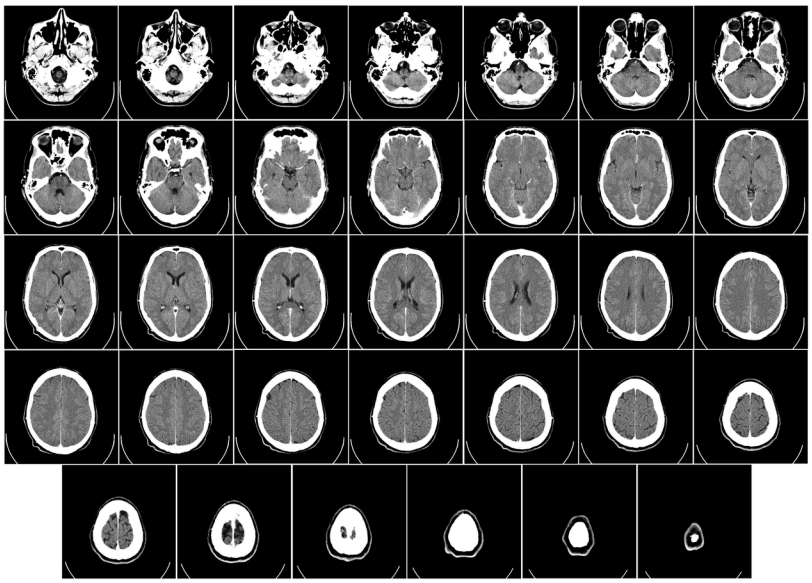


Figure 5.8.4: CT scan of the human brain. These images assist in diagnosing migraines, brain tumors, strokes, bleeding, and head injuries. http://upload.wikimedia.org/wikipedia/commons/5/50/Computed_tomography_of_human_brain_-_large.png

Example 5.8.3: Body CT Scan

Your doctor suspects you have appendicitis and orders a CT scan. This type of evaluation exposes you to 12 msV of radiation. Convert the given mSv amount to rem and mrem. Lastly, compare this value to background and cancer mrem/rem values.

$$\frac{12\ \cancel{mSr}}{1}\ \frac{1\ \cancel{Sr}}{1000\ \cancel{mSr}}\ \frac{100\,rem}{1\ \cancel{Sr}} = 1.2\,rem$$

$$\frac{1.2\ \cancel{rem}}{1}\ \frac{1000\,mrem}{1\ \cancel{rem}} = 1200\,mrem$$

PET Scanning

Positron Emission Tomography or PET scan is a type of nuclear medicine imaging. Depending on the area of the body being imaged, a radioactive isotope is either injected into a vein, swallowed by mouth, or inhaled as a gas. When the radioisotope is collected in the appropriate area of the body, the gamma ray emissions are detected by a PET scanner (often called a gamma camera) which works together with a computer to generate special pictures, providing details on both the structure and function of various organs. Watch this informational video on how this technique works. PET scans are used to:

- Detect cancer
- Determine the amount of cancer spread
- Assess the effectiveness of treatment plans
- Determine blood flow to the heart muscle
- Determine the effects of a heart attack
- Evaluate brain abnormalities such as tumors and memory disorders
- Map brain and heart function

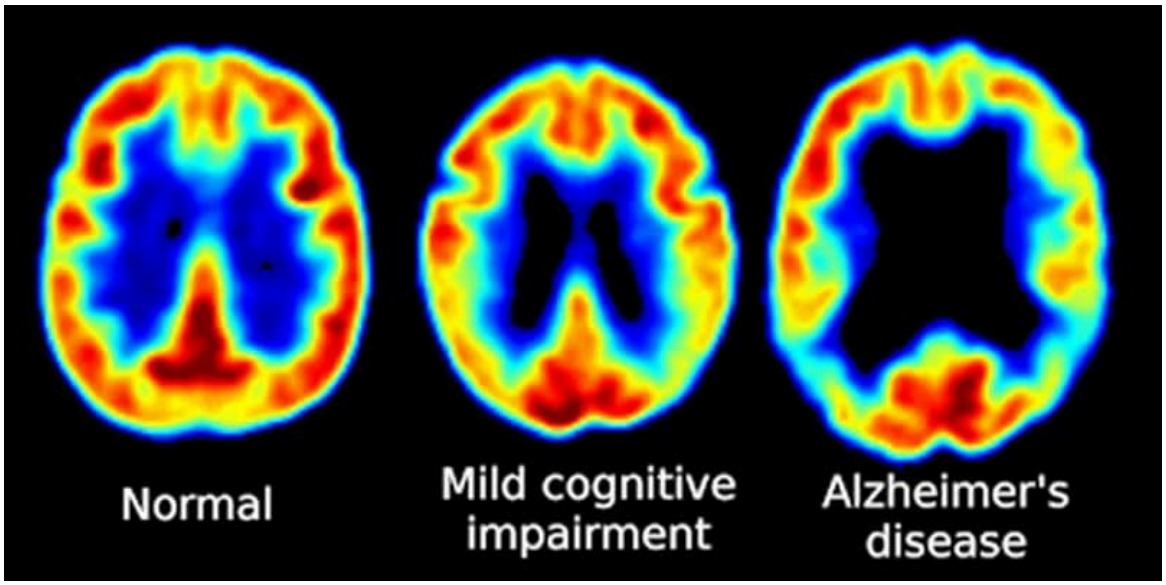


Figure 5.8.5: Mild cognitive impairment (MCI) is a state between normal ageing and dementia, where someone's mind is functioning less well than would be expected for their age. This image is for illustrative purposes only. Image used with permission (Public Domain; Center For Functional Imaging, Lawrence Berkeley National Laboratory. Alzheimer's Disease Neuroimaging Initiative (ADNI).

PET Scanning, is used to image the physiological aspects of the body rather than the anatomy. It images the function of the body rather than the form, such as where tagged molecules go and how they are used. For instance, if you were to image the brain of a deceased person, nothing would show up on a PET scan opposed to a CAT scan, as the brain is no longer functional. Pet Scanning is very useful in imaging tumors, which can be done when patients are injected with certain tracers. Often times PET scanners are used in collaboration with CAT scanners to create a composite image that shows both the function and form of the body. The animation below is a whole-body PET scan using the radioisotope of ^{18}F ($t_{1/2}$ = 110 min). Using this tracer, doctors can determine if cancer has metastasized by looking at the metabolic activity of glucose.

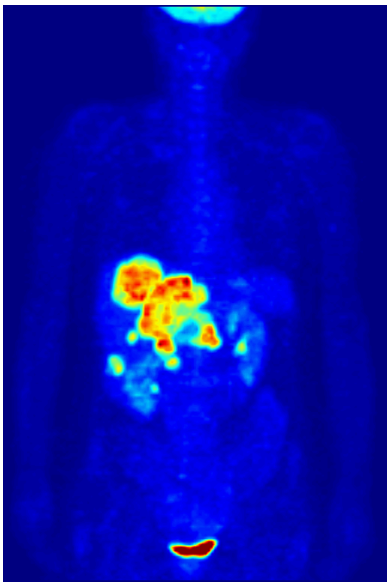


Figure 5.8.6: A whole body analysis by a PET scan. Image used with permission (Public Domain; Jens Maus, http://jens-maus.de/)

Other Isotopic Tests

Radioisotopes have revolutionized medical practice, where they are used extensively. More than 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States. Four typical examples of radioactive tracers used in medicine are technetium-99 ($^{99}_{43}\mathrm{Tc}$), thallium-201 ($^{201}_{81}\mathrm{Tl}$), iodine-131 ($^{131}_{53}\mathrm{I}$), and sodium-24 ($^{24}_{11}\mathrm{Na}$). Damaged tissues in the heart, liver, and lungs absorb certain compounds of technetium-99 preferentially. After it is injected, the

location of the technetium compound, and hence the damaged tissue, can be determined by detecting the γ rays emitted by the Tc-99 isotope. Thallium-201 (Figure 5.8.1) becomes concentrated in healthy heart tissue, so the two isotopes, Tc-99 and Tl-201, are used together to study heart tissue. Iodine-131 concentrates in the thyroid gland, the liver, and some parts of the brain. It can therefore be used to monitor goiter and treat thyroid conditions, such as Grave's disease, as well as liver and brain tumors. Salt solutions containing compounds of sodium-24 are injected into the bloodstream to help locate obstructions to the flow of blood.

Small doses of I-131 (too small to kill cells) are used for purposes of imaging the thyroid. Once the iodine is concentrated in the thyroid, the patient lays down on a sheet of film and the radiation from the I-131 makes a picture of the thyroid on the film. The half-life of iodine-131 is approximately 8 days so after a few weeks, virtually all of the radioactive iodine is out of the patient's system. During that time, they are advised that they will set off radiation detectors in airports and will need to get special permission to fly on commercial flights.

Some isotopes that are used to diagnose diseases are shown in Table 5.8.2. **Be sure to memorize one symbol-mass format and application for the test on this unit**. All of these nuclear isotopes release one form of ionizing radiation (either/add particle or ray). In addition, each isotopic application would involve a specific amount of mrem/mSv radiation.

***Table** 5.8.2**:** Selected radioisotopes used in diagnostic nuclear medicine*

Symbol-mass	Half-Life ($t_{1/2}$)	Application
Xe-133	5.27 days	Lung imaging
H-3	12.26 years	Analyzing total body water
Tl-201	73 hours	Stress tests for heart problems
Fe-59	44.5 days	Detection of anemia
Gd-153	242 days	Analyzing bone density
Cr-51	27.8 days	Determining blood volume
C-11	20.4 minutes	Brain scans
Tc-99m	6.0 hours	Heart, lung, kidney, bone marrow, brain, or bone marrow imaging
Pu-238	86 years	Powering pacemakers
I-131	8.0 days	Imaging Thyroid

Keep in mind that x-rays, CT scans, PET scans, and isotopic studies involve ionizing radiation. In contrast, MRI (magnetic resonance imaging) and ultrasounds do not utilize ionizing forms of radiation.

I like this picture-https://www.reddit.com/r/coolguides/comments/cv5mwx/brain_imaging/

Sources

1. http://cerncourier.com/cws/article/cern/66177 (radioisotope information)
2. https://en.wikipedia.org/wiki/Nuclear_medicine (radioisotope information)

Contributors

- Elizabeth R. Gordon (Furman University)
- Mike Reed (UC Davis), Eugene Kwon (UC Davis)
-
- Emma Gibney (Furman University)

This page titled 5.8: Using Nuclear Science to Diagnose Disease is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.9: Therapeutic Radiation

Learning Objectives

- Understand the difference between diagnostic and therapeutic radiation.
- Know the types of cancer treatments available.
- Compare/contrast external beam therapy and brachytherapy.
- Realize the side effects of radiation therapy and medications available.

There are many techniques used to treat cancer. Surgery can be used to remove cancerous tumors inside or on the body. With chemotherapy, ingested or injected chemicals are used to kill rapidly dividing cells (cancerous and noncancerous). For more information regarding chemotherapy, please access this link. Other cancer treatment methods include immunotherapy, stem cell replacement, hormone therapy, and targeted therapy.

Radiation therapy and Chemotherapy: Two different treatment procedures

Patients diagnosed with cancer might be required to do chemotherapy or radiation therapy. Sometimes, both of these methods are used for a patient. In this picture, a female patient is receiving chemotherapy through an IV. She is participating in cooling therapy while receiving her treatment. By placing her hands and feet in cooling devices, this will lower her chances of losing her finger and toenails. Cold cap therapy is also now available for chemotherapy patients. Wearing this type of device could enable a patient to keep his/her hair during chemotherapy.



Figure 5.9.1: *Image used with permission (CC BY 2.0; Jenny Mealing).*

Current therapeutic radiation applications involve the use of gamma, x-rays, or protons. Recently, some research facilities are investigating the use of alpha and beta tagged molecules to kill cancer cells. These radioisotopes will first locate a cancer related molecule on a tumor cell. Then, the alpha or beta tagged species will inject its radiation into the tumor. Sr-89 (beta emitter) and Ra-223 (alpha emitter) have been used in clinical research trials of certain types of bone cancers.

Radiation Therapy is used as a treatment to control malignant cells within cancer patients. Oncologists (specialists that deal with cancer) utilize radiation frequently to help slow or cure the spread of cancer within individuals. Radiation is specifically applied to malignant tumors in order to shrink them in size. Medical professionals, mainly radiation oncologists, administer a variety of dosages to patient, contingent to the patients current health, as well as other treatments such as chemotherapy, success of surgery, etc.

External Beam Therapy (Photon and Proton Therapy)

External Beam Therapy (EBT) is a method of delivering a high energy beam of radiation to the precise location of a patient's tumor. These beams can destroy cancer cells and with careful planning, NOT kill surrounding cells. The concept is to have several beams of radiation, each of which is sub-lethal, enter the body from different directions. The only place in the body where the beam would be lethal is at the point where all the beams intersect. Before the EBT process, the patient is three-dimensionally mapped using CT scans and x-rays. The patient receives small tattoos to allow the therapist to line up the beams exactly. Alignment lasers are used to precisely locate the target. The radiation beam is usually generated with a linear accelerator. The video below illustrates the basic preparation and administration of external beam therapy.

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

Targeting Cancer - Radiation Therapy Treatment Process

Photon EBT utilizes either x-ray or gamma rays. An x-ray source would require a linear accelerator to produce high energy electrons. In contrast, a gamma source incorporates a radioactive isotope (like Co-60). Keep in mind both of these technologies use ionizing radiation. As a result, cancer patients must be monitored throughout their life to ensure they do not develop other cancers, like leukemia. EBT is used to treat the following diseases as well as others:

- Breast cancer
- Colorectal cancer
- Head and neck cancer
- Lung cancer
- Prostate cancer

The majority of radiation patients receive photon EBT. The smaller size of this machine makes this therapy an option for all sizes of hospitals and cancer treatment centers. Photon EBT equipment costs approximately three million dollars. The size and the price of this technology enables smaller facilities to keep their patients closer to home during treatment

Another method of radiation treatment involving protons is not as commonly used in the United States. This type of therapy requires a cyclotron to generate proton beams (recall, a proton is an ionized H-1 isotope). Unlike x-ray or gamma rays (photon therapy), protons are extremely heavy. Today's proton facilities occupy large plots of land to house (the size of a football field that is three stories high) the medical cyclotron. With a price tag of 150-200 million, these facilities are sporadically located in the United States.



Figure 5.9.2: Current Proton Therapy Centers in the US. The world's first hospital-based proton therapy center was a low energy cyclotron centre for ocular tumours at the Clatterbridge Centre for Oncology in the UK, opened in 1989, followed in 1990 at the Loma Linda University Medical Center in Loma Linda, California. Later, The Northeast Proton Therapy Center at Massachusetts General Hospitalwas brought online, and the HCL treatment program was transferred to it during 2001 and 2002. By 2010 these facilities were joined by an additional seven regional hospital-based proton therapy centers in the United States alone, and many more worldwide. See Wall Street Journal article for information on the details of opening up a center.

At this time, research facilities are working on miniturazing proton generators. An ideal technology would reduce the cost from hundreds of millions of dollars to about twenty to thirty million per device. This would make proton therapy more available and convenient for patients.

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

Proton Therapy versus Radiation Therapy

There are many advantages to choosing proton therapy over photon therapy. Unlike photon radiation, proton beams will only penetrate to the depth of the tumor and not pass through the entire body. This reduces the overall toxicity dose. In addition, fewer treatments are required for proton therapy patients than photon therapy.Unfortunately, proton therapy is more expensive than photon therapy and less common. Once approved by a facility and medical insurance, a patient may have to move temporarily to a larger city to receive treatment. Some forms of cancer have never been clinically treated with proton therapy (namely, breast cancer). Patients desiring proton therapy might not be able to receive type of therapy due to lack of research.

Brachytherapy

The Greek term brachy means to travel a short distance. This type of radiation involves placing ionizing pellets(seeds) or rods directly at the tumor. Photons (in the form of x-ray or gamma ray) are produced inside the body and will penetrate throughout this particular area localizing the radiation. Pellets are surgically implanted while rods can be temporarily inserted to produce radiation internally. With pellets/seeds, the patient will remain radioactive as long as these devices remain inside the body. People undergoing this type of radiotherapy need to be aware of their constant emission of radiation. Radiation that is administered through rods connected to a photon device will disperse energy immediately and not leave the patient radioactive.

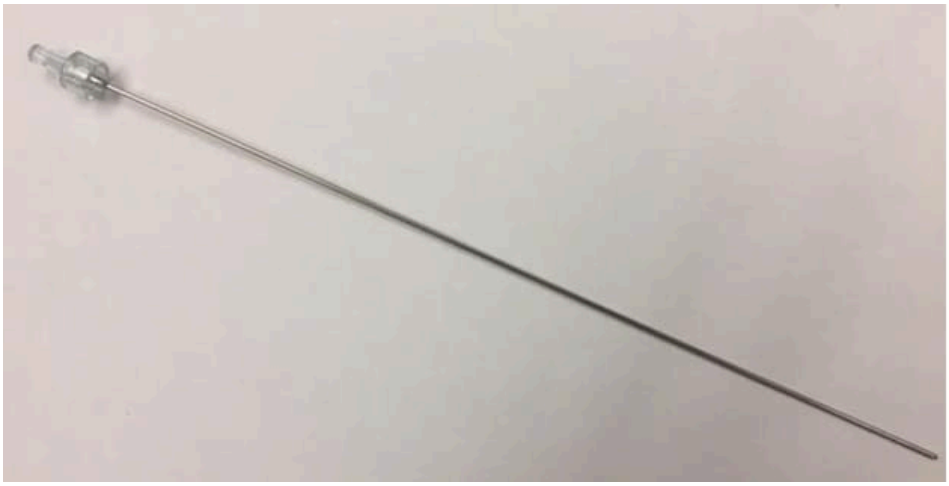




Figure 5.9.4: The picture on the left shows a brachytherapy rod. This device would be connected to a gamma source and connected inside a human body. The second images shows radioactive pellets or seeds that can be surgically implanted near a tumor. Note how small the pellets are when compared to an American penny. Image taken from: https://c1.staticflickr.com/8/7254/7...74d83b7058.jpg

Brachytherapy is widely used in the treatment of cancers involving reproductive organs. Because the radiation is isolated internally, patients are less likely to experience side effects when receiving this type of treatment. Cancers that have been treated with brachytherapy are shown below:

- Prostate
- Breast
- Esophageal
- Lung
- Uterine
- Anal/Rectal
- Sarcomas
- Head and neck

***Table** 5.9.1: Commonly used radiation sources (radionuclides) for brachytherapy.*

Radionuclide	Type	Half-life
Cesium-131 (^{131}Cs)	Electron Capture, ε	9.7 days
Cesium-137 (^{137}Cs)	β^-- particles, γ-rays	30.17 years
Cobalt-60 (^{60}Co)	β^-- particles, γ-rays	5.26 years
Iridium-192 (^{192}Ir)	γ-rays	73.8 days
Iodine-125 (^{125}I)	Electron Capture, ε	59.6 days
Palladium-103 (^{103}Pd)	Electron Capture, ε	17.0 days
Ruthenium-106 (^{106}Ru)	β^-- particles	1.02 years
Radium-226 (^{226}Ra)	β^-- particles	1599 years

Side Effects of Radiation Therapy

Patients receiving radiation therapy can experience a variety of side effects. For example, sterility could occur if reproductive organs are irradiated. Skin that has been irradiated can appear dry and feel itchy. Some patients will loose sensation in the irradiated area. Radiation can affect the production of white and red blood cells. A reduction of white blood cells results in immunity disorders. Red blood cell lose causes anemia. Gastrointestinal issues such as diarrhea and nausea are common during radiation therapy. Some patients will lose hair as well. Lastly, dry mouth and tooth decay are prevalent during radiation treatments.

Medications are available to alleviate symptoms of radiation therapy. Narcotics can be prescribed to help alleviate intense pain. Prescription medications like zofran and phenergan can help with nausea. Special mouthwashes have been formulated to reduce dry mouth and cavities.

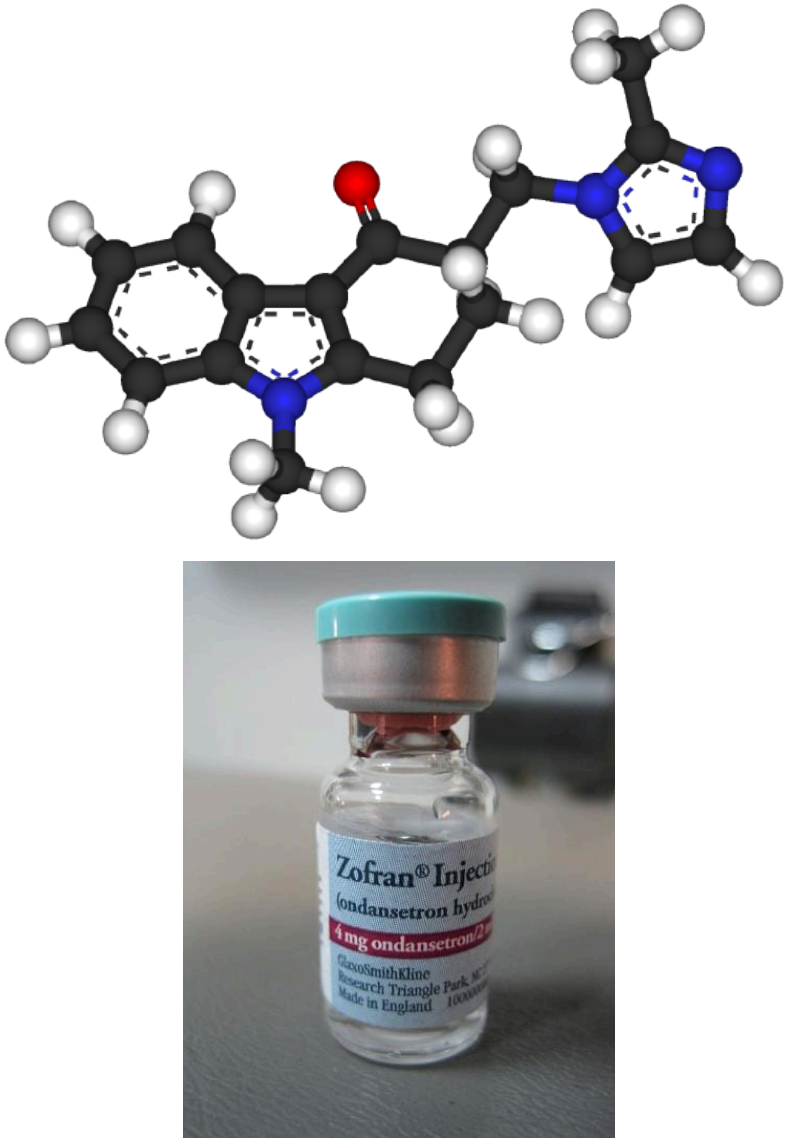


Figure 5.9.5: (left) *Zofran is a medication used to prevent nausea and vomiting caused by cancer chemotherapy, radiation therapy, or surgery. (right) A vial of Zofran 4 mg containing ondansetron for intravenous injection. This medicine is a controlled substance and is used for nausea. Images used with permision (left: Public domain; Fvasconcellos and right: CC BY 3.0; Intropin)*

Hair loss is a side effect of radiation, but only locally

Radiation therapy can cause hair loss, but hair is only lost in the area being treated. For instance, radiation to your head may cause you to lose some or all the hair on your head (even eyebrows and lashes), but if you get treatment to your hip, you won't lose the hair on your head.

References

1. https://www.cancer.gov/about-cancer/treatment/types
2. https://weillcornellgucancer.org/201...-cancer-cells/
3. http://www.proton-therapy.org/zapper.htm

4. http://www.brachytherapy.com/

Contributors

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
-
- Emma Gibney (Furman University)

This page titled 5.9: Therapeutic Radiation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.10: Non-medical Applications of Nuclear Science

Learning Objectives

- Understand how radiation is used in other types of industries.
- Realize various isotopes and wave radiation are uses as tracers and sanitizers.
- Distinguish between millimeter and back scatter technologies uses by the Transportation Security Administration (TSA).
- Know examples of foods that are irradiation in the United States.

Tracer Applications

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

Tracers are used in diverse ways to study the mechanisms of chemical reactions in plants and animals. These include labeling fertilizers in studies of nutrient uptake by plants and crop growth, investigations of digestive and milk-producing processes in cows, and studies on the growth and metabolism of animals and plants.

For example, the radioisotope C-14 was used to elucidate the details of how photosynthesis occurs. The overall reaction is:

$$6\,\mathrm{CO_2}(g) + 6\,\mathrm{H_2O}(l) \longrightarrow \mathrm{C_6H_{12}O_6}(s) + 6\,\mathrm{O_2}(g), \tag{5.10.1}$$

The process is much more complex, proceeding through a series of steps in which various organic compounds are produced. In studies of the pathway of this reaction, plants were exposed to CO_2 containing a high concentration of ${}^{14}_{6}\mathrm{C}$. At regular intervals, the plants were analyzed to determine which organic compounds contained carbon-14 and how much of each compound was present. From the time sequence in which the compounds appeared and the amount of each present at given time intervals, scientists learned more about the pathway of the reaction.

Airport Applications

Commercial airports use x-rays and CT imaging to items contained in luggage. An x-ray machine can produce images of hard and soft materials. Based upon relative densities, an x-ray image will display a different color for a different type of material. For example, foods and organic materials will appear orange on an x-ray scan. Harder materials like thicker plastics will appear green. Lastly, metals and extremely thick materials will produce a bluish color when scanned.



Figure 5.10.1: Photograph of a baggage screening x-ray display. Image used with permission (CC BY-SA; IDuke via Wikimedia).

American airports once used x-rays to scan human bodies for weapons. Up through 2013, back scatter technology was utilized to give very detailed body images. Due to the "naked body" pictures, this type of technology was banned in American airports and has now been replaced with millimeter wave generators. These machines produce cartoon like images and expose the body to high energy radio waves.

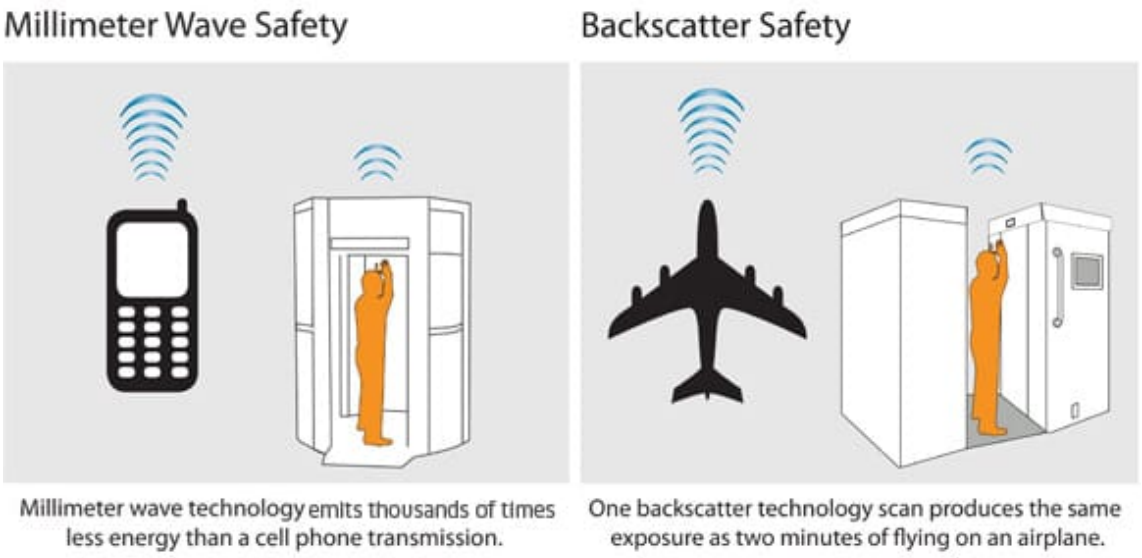


Figure 5.10.2: Different whole body scanning technologies have been used in airport security. Image taken from: https://upload.wikimedia.org/wikiped...afety_demo.jpg

Irradiation of Food and Mail

Co-60 (gamma source) and x-rays are use to irradiate many foods in the United States. Ionizing radiation can be used to kill food borne illnesses like salmonella and e coli. Irradiating food can also prolong shelf-life, delay ripening, and destroy insects. In addition, gamma/x-ray can sterilize foods making refrigeration unnecessary. Often, sterilized food is served to hospital patients who have impaired immune systems.

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

In the United States,irradiation of food is regulated by the FDA (Food and Drug Administration). Foods that have been irradiated must display the symbol called the Radura Figure 5.10.3 on their packaging. Some foods that could display this symbol are: meats (beef, pork, chicken), shrimp, lobster, fruits, vegetables, shellfish, and spices. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.

Figure 5.10.3 image taken from: https://commons.wikimedia.org/wiki/F...ura-Symbol.svg

In the fall of 2001, anthrax laced letters were sent to various new agencies and two United States senators. Anthrax is an infectious disease caused by bacteria. It exists naturally in some soils and can be isolated in a laboratory. After September 11, 2011, this substance was used as a biochemical weapon. This particular incident killed five people and injured twenty-two other individuals. Anthrax produces ulcers on the skin and causes severe breathing difficulties. Today, only members of the United States military receive vaccines for anthrax.



Figure 5.10.4: Anthrax-laced letter that was sent to Senator Tom Daschle. Image taken from: https://upload.wikimedia.org/wikiped...hle_letter.jpg

To combat this form of bioterrorism, the United States Postal Service(USPS) and the Federal Bureau of Investigation (FBI) installed x-ray generators to irradiate suspicious looking mail sent to some governmental facilities. X-rays will kill the majority of this bacteria and some viruses as well. Once irradiated, mail (packages and letters) might change in color, texture, and smell. This ionizing radiation chemically alters the composition of the mail's paper component. The x-rays do not leave the mail radioactive.

Smoke Detectors

Americium-241, an α emitter with a half-life of 458 years, is used in tiny amounts in ionization-type smoke detectors (Figure 5.10.5). The α emissions from Am-241 ionize the air between two electrode plates in the ionizing chamber. A battery supplies a potential that causes movement of the ions, thus creating a small electric current. When smoke enters the chamber, the movement of the ions is impeded, reducing the conductivity of the air. This causes a marked drop in the current, triggering an alarm.

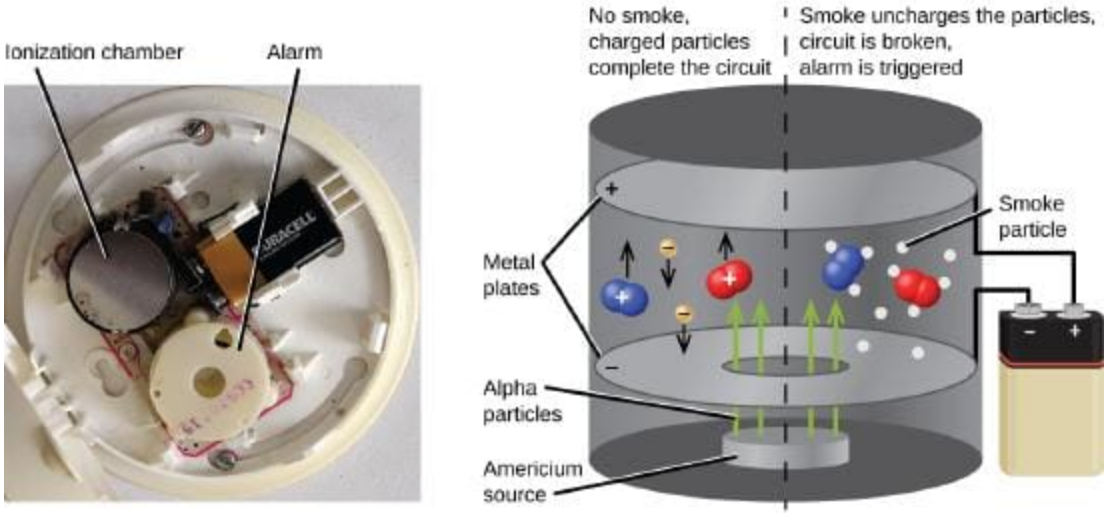


Figure 5.10.5: Inside a smoke detector, Am-241 emits α particles that ionize the air, creating a small electric current. During a fire, smoke particles impede the flow of ions, reducing the current and triggering an alarm. (credit a: modification of work by "Muffet"/Wikimedia Commons)

References

1. https://medlineplus.gov/anthrax.html
2. https://www3.epa.gov/radtown/mail-irradiation.html

References

- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

This page titled 5.10: Non-medical Applications of Nuclear Science is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

5.E: Basics of Nuclear Science (Exercises)

1) Perform a beta emission of Rb-86.

2) What is the missing piece of this nuclear reaction: n → p + ?

3) Write a reaction for the decay of Rn-222. Review your ionizing section of your text book to determine the emission path. In addition, how could you shield yourself from the radiation released from this reaction?

4) You are exposed to radiation that has a wavelength of 50 μm. What is the frequency of this radiation? Where does this radiation appear? Is this radiation ionizing or nonionizing?

5) What is the difference between the radiation from Uranium and x-rays?

6) How can you alter a house that tests over the 4pCi level for radon?

7) Give one crazy application of x-ray science.

8) Which is the heaviest form of radiation: alpha, beta, or gamma?

9) Which is the least ionizing form of radiation: alpha, beta, or gamma?

10) Is your smoke detector radioactive?

11) Fatman (Nagasaki bomb) used approximately 6.4 kg of Pu-239. This artificially produced isotope ($t_{1/2}$ = 24,100 years) undergoes alpha decay. How many kg and pounds would remain after three decay processes?

12) Can one chemically change the half-life of a material?

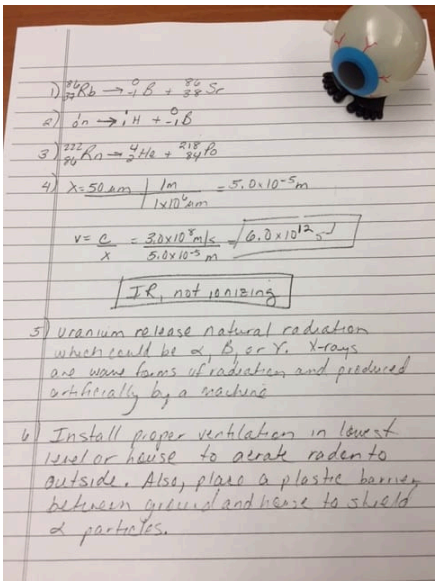
13) The $t_{1/2}$ of As-74 is 18 days, how much of a 160 mg sample remains after 2160 hours?

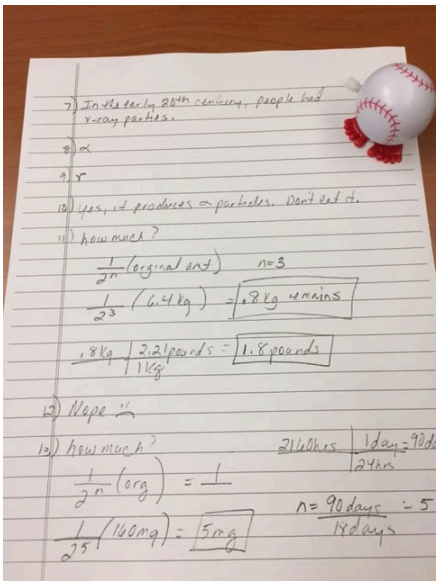
14) Write a balance reaction for the decay Ir-192. This isotope undergoes alpha and gamma emission at the same time.

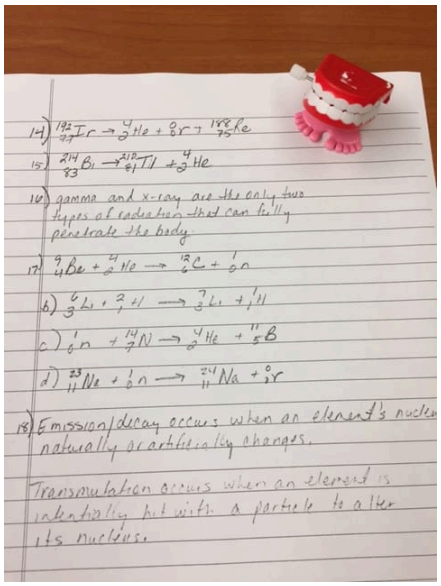
15) Bi-214 can decay to form Tl-210. Write balanced reaction for this emission process.

16) When focused from the outside of the body, which type(s) of radiation penetrate(s) the body completely? gamma, proton, x-ray, alpha, or beta

Solutions







This page titled 5.E: Basics of Nuclear Science (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6: NUCLEAR WEAPONS- FISSION AND FUSION

A nuclear weapon is an explosive device that derives its destructive force from nuclear reactions, either fission (fission bomb) or from a combination of fission and fusion reactions (thermonuclear bomb). Both bomb types release large quantities of energy from relatively small amounts of matter.

6.1: The History and Basics of Fission

6.2: The Manhattan Project - Prewar

6.3: The Manhattan Project - Labs and Fuel

6.4: The Manhattan Project - Critical Mass and Bomb Construction

6.5: Hiroshima and Nagasaki

6.6: Aftermath of Bombings

6.7: Fusion

6.8: Fallout and Weapons Testing

6.9: Additional Readings and Media Fun

6.E: Nuclear Weapons- Fission and Fusion (Exercises)

This page titled 6: Nuclear Weapons- Fission and Fusion is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.1: The History and Basics of Fission

Learning Objectives

- Understand the importance of artificial transmutation.
- Write and balance artificial transmutation reactions.
- Know the history of fission.
- Determine why only certain atoms can undergo fission.
- Know what fission is and where it is used.

Understanding the Neutron

In 1911, Ernest Rutherford applied a beam of alpha particle to a gold foil. This experiment helped him to determine the positive center of the atom, which he named the nucleus. Eight years later, this same scientist discovered the existence of proton by performing the following nuclear reaction:

$$^{14}_{7}\mathrm{N} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{17}_{8}\mathrm{O} + {}^{1}_{1}\mathrm{H} \tag{6.1.1}$$

This reaction is an example of artificial transmutation. By taking a particle and hitting any existing nucleus, a new atom will be synthesized artificially. Note how this reaction differs from a decay/emission reaction. In the latter type of reaction, only one reactant will produce at least two products. Transmutations will illustrate a minimum of two reactants and various number of products.

Example 6.1.1: artificial transmutation or Decay reactions?

Label each as being an artificial transmutation or a decay reaction.

a. $^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th}$
b. $^{198}_{78}\mathrm{Pt} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{199}_{78}\mathrm{Pt} \rightarrow {}^{199}_{79}\mathrm{Au} + {}^{0}_{-1}\beta$
c. $^{239}_{93}\mathrm{Np} \rightarrow {}^{239}_{94}\mathrm{Pu} + {}^{0}_{-1}\beta$
d. $^{11}_{6}\mathrm{C} \rightarrow {}^{11}_{5}\mathrm{B} + {}^{0}_{+1}\beta$
e. $^{27}_{13}\mathrm{Al} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{30}_{15}\mathrm{P} + {}^{1}_{0}\mathrm{n}$

Solution

a) Only reactions b and e are artificial transmutations. The other reactions are examples of decays/emissions.

Throughout the 1930's, many scientists became focused on producing new radioactive elements. Before this time period, newly discovered elements were extracted from natural ores obtained from mining. Irene (daughter of Marie and Pierre Curie) and her husband Frederic Joliot-Curie bombarded existing elements with alpha particles. These experiments yielded new radioactive isotopes by an artificial process. For their efforts, Irene and Frederic were awarded the 1935 Noble Prize in chemistry. One of their key syntheses involve the bombardment of aluminum foil with alpha particles.

$$^{27}_{13}\mathrm{Al} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{30}_{15}\mathrm{P} + {}^{1}_{0}\mathrm{n} \tag{6.1.2}$$

In their laboratory, the Joliet-Curies could not explain or identify the existence of the neutron. Their measurements showed that an additional, non-charged substance was produced in the aluminum foil reaction. They hypothesized that this unknown piece of the reaction must be a gamma ray, which has a neutral charge.



Figure 6.1.1: Irene and Frederic Joliot-Curie working in their lab to synthesize new isotopes by artificial transmutation. Image taken from: https://upload.wikimedia.org/wikiped...Curie_1935.jpg

The Significance of Artificial Transmutation

Unfortunately, the Joliot-Curies were incorrect in their analysis. Around the same time, a British scientist named James Chadwick bombarded beryllium with alpha particles to produce the species below:

$$ {}^{9}_{4}\mathrm{Be} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{12}_{6}\mathrm{C} + {}^{1}_{0}\mathrm{n} \tag{6.1.3} $$

Using this reaction, he was able to characterize and explain the existence of this new atomic subatomic particle. He noted this atomic piece was the same mass as a proton, but had no charge. In 1935, James Chadwick was awarded the Nobel Prize in physics for the discovery of the neutron.

Example 6.1.2:

Translate the following reactions to A/Z symbols. Balance these artificial transmutations by finding the missing pieces.

a. sodium-23 is hit with an alpha particle to produce a proton and another isotope
b. fluorine-19 is bombarded a neutron to produce an alpha particle and another isotope
c. uranium-238 is hit with a neutron to produce an isotope and a beta particle

Solution

a. missing isotope piece = Mg-26
b. missing isotope piece = N-16
c. missing isotope piece = Np-239

Artificial transmutation has been utilized to make new elements. All elements with atomic numbers greater than 92 are produced synthetically. These elements are called transuranium elements and have been prepared by bombarding suitable target nuclei with smaller particles. The first of the transuranium elements to be made was neptunium (Z = 93), which was synthesized in 1940 by bombarding a ^{238}U target with neutrons. As shown below, this reaction occurs in two steps. Initially, a neutron combines with a ^{238}U nucleus to form ^{239}U, which is unstable and undergoes beta decay to produce ^{239}Np:

$$ {}^{238}_{92}\mathrm{U} + {}_{0}{}^{1}\ \mathrm{n} \ \rightarrow {}^{239}_{92}\mathrm{U} \tag{6.1.4} $$

$$ {}^{239}_{92}\mathrm{U} \rightarrow {}^{239}_{93}\ \mathrm{Np} + {}^{0}_{-1}\beta \tag{6.1.5} $$

Subsequent beta decay of ^{239}Np produces the second transuranium element, plutonium (Z = 94):

$$^{239}_{93}\mathrm{Np} \rightarrow {}^{239}_{94}\mathrm{Pu} + {}^{0}_{-1}\beta \tag{6.1.6}$$

Bombarding the target with more massive nuclei creates elements that have atomic numbers significantly greater than that of the target nucleus (Table 6.1.1). Such techniques have resulted in the creation of the superheavy elements 114 and 116, both of which lie in or near the “island of stability." As of this writing, 22 transuranium elements have been produced and officially recognized by IUPAC; several other elements have formation claims that are waiting for approval.

Table 6.1.1***:*** *Preparation of Some of the Transuranium Elements*

Name	Symbol	Atomic Number	Reaction
americium	Am	95	$^{239}_{94}\mathrm{Pu} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{240}_{95}\mathrm{Am} + {}^{0}_{-1}\mathrm{e}$
curium	Cm	96	$^{239}_{94}\mathrm{Pu} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{242}_{96}\mathrm{Cm} + {}^{1}_{0}\mathrm{n}$
californium	Cf	98	$^{242}_{96}\mathrm{Cm} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{243}_{97}\mathrm{Bk} + 2{}^{1}_{0}\mathrm{n}$
einsteinium	Es	99	$^{238}_{92}\mathrm{U} + 15{}^{1}_{0}\mathrm{n} \longrightarrow {}^{253}_{99}\mathrm{Es} + 7{}^{0}_{-1}\mathrm{e}$
mendelevium	Md	101	$^{253}_{99}\mathrm{Es} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{256}_{101}\mathrm{Md} + {}^{1}_{0}\mathrm{n}$
nobelium	No	102	$^{246}_{96}\mathrm{Cm} + {}^{12}_{6}\mathrm{C} \longrightarrow {}^{254}_{102}\mathrm{No} + 4\,{}^{1}_{0}\mathrm{n}$
rutherfordium	Rf	104	$^{249}_{98}\mathrm{Cf} + {}^{12}_{6}\mathrm{C} \longrightarrow {}^{257}_{104}\mathrm{Rf} + 4\,{}^{1}_{0}\mathrm{n}$
seaborgium	Sg	106	$^{206}_{82}\mathrm{Pb} + {}^{54}_{24}\mathrm{Cr} \longrightarrow {}^{257}_{106}\mathrm{Sg} + 3\,{}^{1}_{0}\mathrm{n}$ $^{249}_{98}\mathrm{Cf} + {}^{18}_{8}\mathrm{O} \longrightarrow {}^{263}_{106}\mathrm{Sg} + 4\,{}^{1}_{0}\mathrm{n}$
meitnerium	Mt	107	$^{209}_{83}\mathrm{Bi} + {}^{58}_{26}\mathrm{Fe} \longrightarrow {}^{266}_{109}\mathrm{Mt} + {}^{1}_{0}\mathrm{n}$

Particle Accelerators: Artificial Transmutators

To perform artificial transmutation reactions, there are two large machines that can be utilized. A device called a **particle accelerator** can accelerate positively charged particles to the speeds needed to overcome the electrostatic repulsions between them and then target nuclei by using electrical and magnetic fields. Operationally, the simplest particle accelerator is the linear accelerator (Figure 6.1.2), in which a beam of particles is injected at one end of a long evacuated tube. A modern linear accelerator such as the Stanford Linear Accelerator (SLAC) at Stanford University is about 2 miles long.

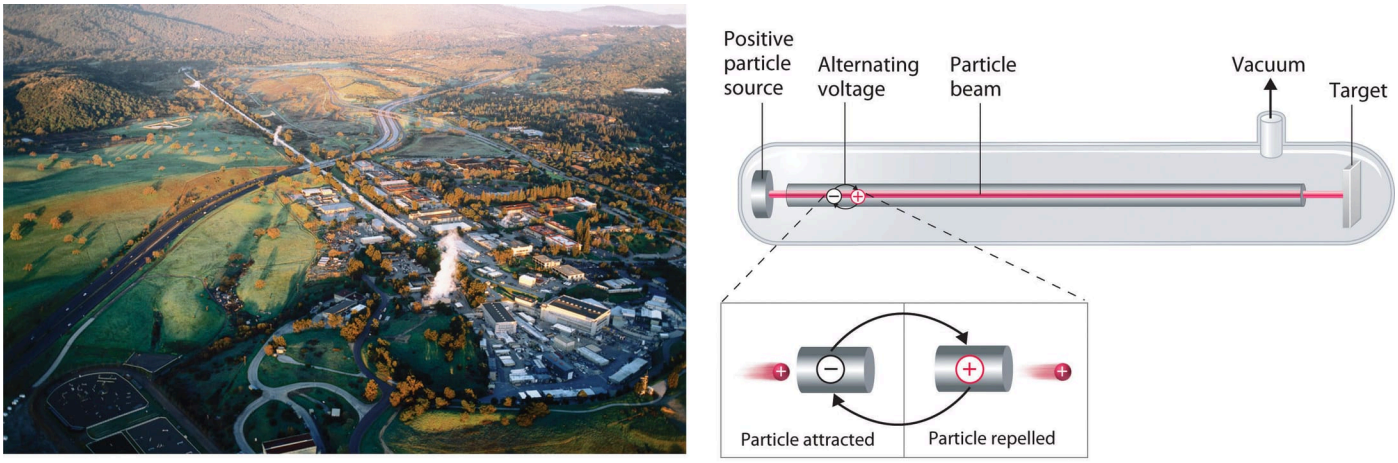


Figure 6.1.2: *A Linear Particle Accelerator. (a) An aerial view of the SLAC, the longest linear particle accelerator in the world; the overall length of the tunnel is 2 miles. (b) Rapidly reversing the polarity of the electrodes in the tube causes the charged particles to be alternately attracted as they enter one section of the tube and repelled as they leave that section. As a result, the particles are continuously accelerated along the length of the tube.*

To achieve the same outcome in less space, a particle accelerator called a **cyclotron** forces the charged particles to travel in a circular path rather than a linear one. The particles are injected into the center of a ring and accelerated by rapidly alternating the polarity of two large D-shaped electrodes above and below the ring, which accelerates the particles outward along a spiral path toward the target.

Fission and Fusion: Selecting the Isotopes

Fission is the splitting of an atomic nucleus. In nuclear weapons and reactors, neutrons hit unstable nuclei to form smaller atoms. Nuclei that are larger than Fe-56 may undergo fission. Of these elements, fission requires heavy, unstable nuclei. This means selected atoms would have low binding energies and would have large atomic masses. Figure four shows the relative binding energies for various isotopes. According to figure 4, uranium-238 and uranium-235 both have lower binding energies with heavy masses-. These two isotopes would be suitable for splitting based on these requirements. Of these two isotopes, only uranium-235 is readily fissionable. This is due to odd neutron count contributing to additional instability. Another fissionable isotope not shown in figure 4 is plutonium-239. This is a synthetic isotope produced by transmutation and decay reactions. Like uranium-235, it has low binding energy, high mass, and an odd number of neutrons. U-235 and Pu-239 are used in atomic bombs (fission based) and nuclear reactors.

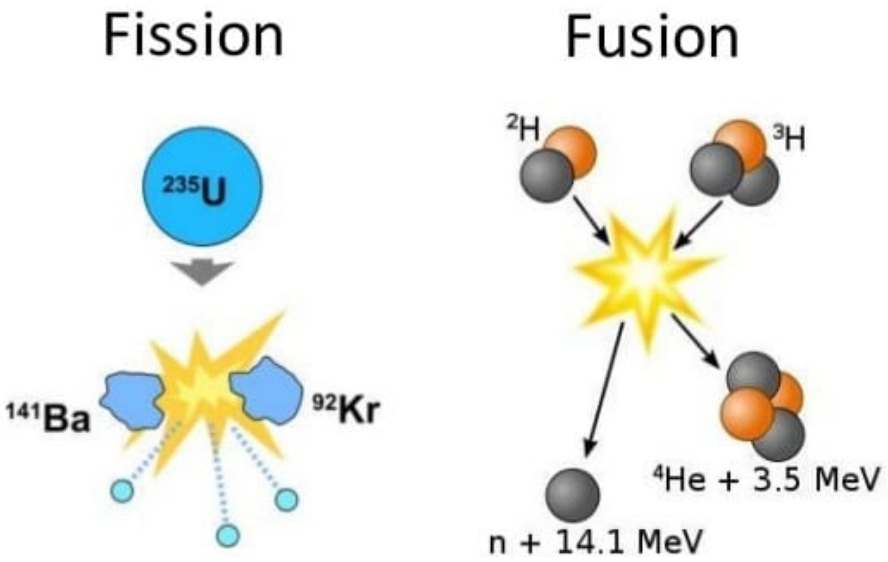


Figure 6.1.3: (left) Nuclear fission occurs with one large nuclear is split into two or more smaller nuclei. (right) Nuclear fusion happens when two small nuclei combine to make a larger nucleus.

Fusion is the bringing together of two atomic nuclei to form a larger atom. Positively charged centers of atoms make this type of reaction extremely difficult. For this reason, smaller atoms are suitable for fusion reactions. In addition, these particular isotopes need to have low binding energies in order to undergo fusion. Atoms that possess these two qualities are H-2 and H-3. For fusion to occur, extreme temperatures are required to fuse these deuterium and tritium together.

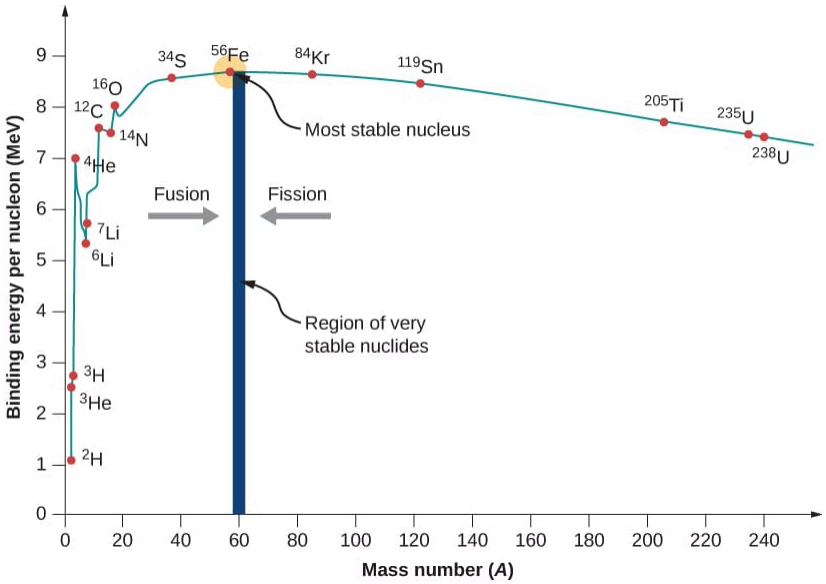


Figure 6.1.4: In this graph of binding energy per nucleon for stable nuclei, the BEN is greatest for nuclei with a mass near $^{56}\mathrm{Fe}$. Therefore, fusion of nuclei with mass numbers much less than that of Fe, and fission of nuclei with mass numbers greater than that of Fe, are exothermic processes. Image used with permission (CC BY. OpenStax).

Fission and Chain Reactions

In both fission and fusion, large amounts of energy are given off in the form of heat, light, and gamma radiation. Italian physicist, Enrico Fermi, performed the first fission reaction in 1934. He was unaware that he had split a uranium atom into two smaller nuclei. Four years later, two German scientists, Otto Hahn and Fritz Strassman, repeated Fermi's uranium reaction. While performing this experiment, two other scientists, Lise Meitner and Otto Frisch, analyzed the data. Meitner, an Austrian physicist,noted that the uranium atom had split into two smaller pieces. She used the term "fission" to describe her observations.



Austrian physicist, Lise Meitner who coined the term "fission" to explain the neutron bombardment of U-238. Image take from: https://upload.wikimedia.org/wikiped...ner12crop2.JPG

The nuclear equation illustrating this fission reaction is shown below:

$$ {}_{0}^{1}\mathrm{n} + {}_{92}^{235}\mathrm{U} \rightarrow {}_{56}{}^{141}\mathrm{Ba} + {}_{36}^{92}\mathrm{Kr} + 3\,{}_{0}^{1}\mathrm{n} \tag{6.1.7} $$

The products shown are only one of many sets of products from the disintegration of a U-235 nucleus. Over 35 different elements have been observed in the fission products of U-235. Fission has been used in nuclear weapons and powers all nuclear reactors. Approximately fifty-five countries world wide possess fission technology in the form of research or energy reactors. Less than ten of these countries have fission weapons.

On the other hand, fusion technology is not as accessible. Only five countries (five permanent members of the United Nations Security Council) have knowledge and scientific capability of constructing fusion weapons. Unlike fission reactors. there are no commercial fusion reactors. Fusing nuclei is quite difficult to perform in a safe and controlled environment. For more insight as the the differences between fission and fusion, click on the informational 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

Energy of Reactions

Although Albert Einstein was not directly involved in the construction of nuclear weapons, he did devise a theory to explain the amount of energy released during a nuclear reaction (fission or fusion). This theory is known as special relativity as is shown below:

$$ E = mc^2 \tag{5.11.7} $$

Here, the mass can be converted to significant amounts of energy in nuclear reactions. Understanding the speed of light to be 3.0×10^8 m/s can give you an appreciation for how much energy can be produced per unit of matter. Nuclear reactions require very small amounts of fissionable material to produce large amounts of energy (as in a nuclear reactor) or destruction (as in a bomb).

Summary and Vocabulary

Naturally radioactive elements exist in the earth and are either alpha or beta emitters. Artificial transmutation of elements can be accomplished by bombarding the nuclei of some elements with alpha or subatomic particles. .

- Fission: A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy.
- Fusion: A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.
- Cyclotrons and Linear Accelerators: Large pieces of equipment that are used to perform artificial transmutations of atoms.
- Research nuclear reactor: A smaller scaled reactor that is used at a academic or industrial facility.
- Commercial nuclear reactor: A larger reactor that produced energy for an area of consumers.

Sources

1. https://www.aps.org/publications/aps...icshistory.cfm

Contributors

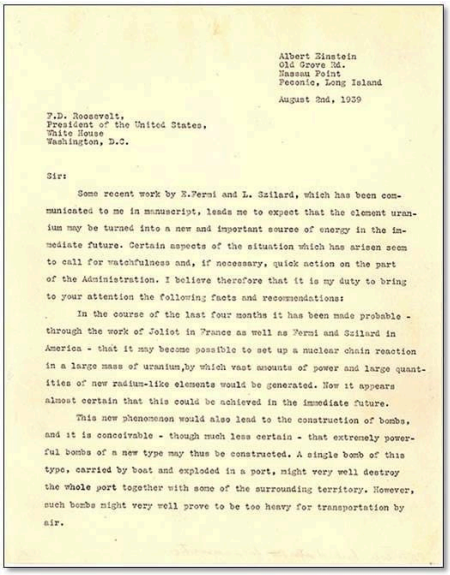
-
- Elizabeth R. Gordon (Furman University)

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 6.1: The History and Basics of Fission is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.2: The Manhattan Project - Prewar

In October of 1939, the United States became concerned about the fission research efforts occurring in Nazi Germany. To express their concerns, two recent immigrant scientists, Albert Einstein and Leo Szilard drafted a letter to President Franklin Delano Roosevelt. Within a month's time, President Roosevelt established the Advisory Committee on Uranium. This research group was provided a limited budget of approximately $150,000. Projects were to only include uranium isotopic separation. Two University of Columbia physicists were assigned accomplish this task. With a very limited budget, Enrico Fermi and Leo Szilard were expected to surpass any uranium efforts that Nazi Germany had already accomplished. Click on this link to visualize how these scientists felt about this project.



Albert Einstein
Old Grove Rd.
Nassau Point
Peconic, Long Island

August 2nd, 1939

F.D. Roosevelt,
President of the United States,
White House
Washington, D.C.

Sir:

Some recent work by E.Fermi and L. Szilard, which has been communicated to me in manuscript, leads me to expect that the element uranium may be turned into a new and important source of energy in the immediate future. Certain aspects of the situation which has arisen seem to call for watchfulness and, if necessary, quick action on the part of the Administration. I believe therefore that it is my duty to bring to your attention the following facts and recommendations:

In the course of the last four months it has been made probable - through the work of Joliot in France as well as Fermi and Szilard in America - that it may become possible to set up a nuclear chain reaction in a large mass of uranium,by which vast amounts of power and large quantities of new radium-like elements would be generated. Now it appears almost certain that this could be achieved in the immediate future.

This new phenomenon would also lead to the construction of bombs, and it is conceivable - though much less certain - that extremely powerful bombs of a new type may thus be constructed. A single bomb of this type, carried by boat and exploded in a port, might very well destroy the whole port together with some of the surrounding territory. However, such bombs might very well prove to be too heavy for transportation by air.

Figure 6.2.1: The letter written by Einstein and Szilard to President Franklin D. Roosevelt.

At this time, President Roosevelt attempted to pacify these nuclear scientists by allocating this amount of research funding. Feeling fearful, this group of researchers continually pleaded for more money and facilities where they could expand their projects. President Roosevelt denied their requests until Pearl Harbor was bombed in December of 1941. After this, Roosevelt established the Manhattan Project which was to focus on research and construction of an atomic bomb. He placed Leslie R. Groves, head of the U.S. Army Corps of Engineers, to organize and lead this undertaking.





Figure 6.2.2: General Leslie R. Groves and J. Robert Oppenheimer. Images taken from: https://upload.wikimedia.org/wikiped...lie_Groves.jpg and https://upload.wikimedia.org/wikipedia/commons/0/03/JROppenheimer-LosAlamos.jpg

General Groves assigned J. Robert Oppenheimer to lead the top secret laboratory that would construct the first atomic bomb. At the time, Oppenheimer was a theoretical physicist employed at the University of California, Berkeley. Both Groves and Oppenheimer would work together for over four years to complete this project.

Example 6.2.1:

Watch the first installment of the History Channel's Modern Marvel Series (total time: 10:59) and 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 much money was spent on the making of the atomic bomb (1940's and today's estimates)?
2. Why did Teller, Einstein, Szilard, and Bethe flee Europe?
3. What country discovered fission and which element was investigated to prove this?
4. How did Enrico Fermi escape Italty?
5. How did FDR first get introduced to the idea of atomic fission? What event encouraged him to allocate more funds to this research project?
6. Which isotope is more useful for the construction of nuclear weapons: U-235 or U-238?
7. What two processes were used for uranium enrichment? Where would these processes be done?

Contributors

- Frank A. Settle (Washington and Lee University)

This page titled 6.2: The Manhattan Project - Prewar is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.3: The Manhattan Project - Labs and Fuel

Designated Areas for Laboratories

For security purposes, the Manhattan Project selected several different geographical locations for laboratory research areas. Most employees at a research facility were not aware of the ultimate goal of this endeavor. Only selected scientists and engineers were informed of the existence of other working laboratories that would contribute to the construction of the first atomic bomb. The Manhattan project employed over 130,000 people at a cost of 2 billion dollars (today's estimate would be over 20 billion).

Different laboratory sites employed all types of workers. Engineers and scientists focused on different mechanisms of the atomic bomb. Construction workers built barracks, laboratories,schools, and community buildings. Medical personnel assisted residents with health care concerns and needs. Educators were hired for newly constructed schools. In addition, factory workers were employed to man different stations of each research facilities. ***Figure*** 6.3.1 shows the integral laboratories of the Manhattan Project. For a sample of what is was like to work at one of these settings, click on this link and watch a part or the entire interview of a female scientist who worked at Los Alamos National Laboratory.

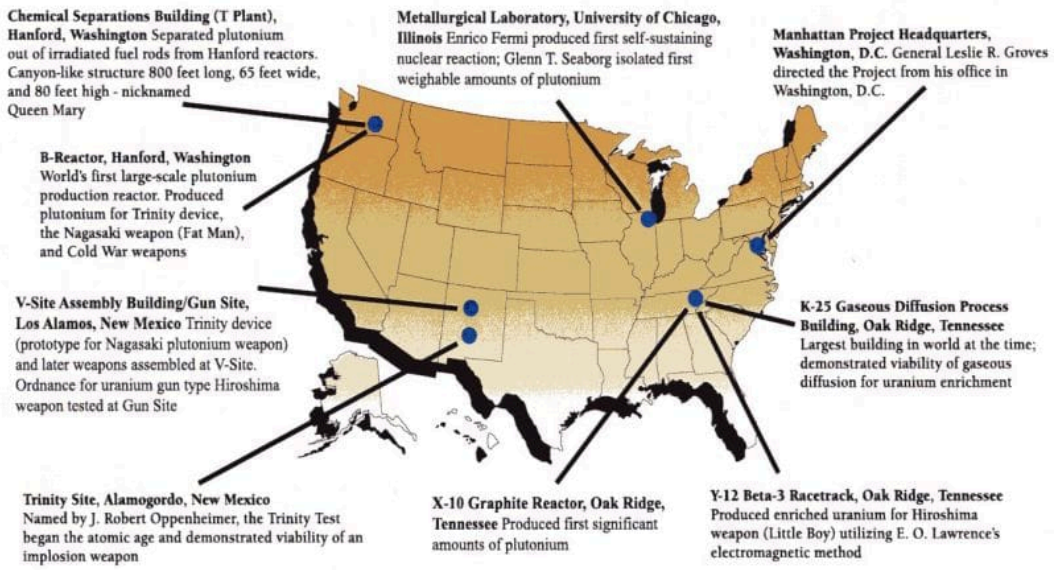


Figure 6.3.1:Various lab locations during the Manhattan Project. *Image taken from:* https://www.osti.gov/opennet/manhatt...esMapLarge.jpg

Four Major Projects

From December of 1941 to August of 1945, the labs shown in ***Figure*** 6.3.1 focused on four major problems in the construction of the first nuclear weapon. In this section of the text, we will analyze two of these issues. It is important for you to grasp the basic chemistry/physics incorporated to solve the problem. In addition, please correlate the lab responsible for solving each problem with its lead scientists. The four major obstacles that plagued Manhattan Project Scientists/ Engineers were:

1. **Obtaining and purifying U-235 to use as weapon fuel (Oak Ridge, TN by Ernest Lawrence)**
2. **Synthesizing Pu-239 to be used as an alternate weapon fuel (Hanford, WA by Glen Seaborg)**
3. **Calculating critical mass of nuclear fuel to ensure a chain reaction (University of Chicago, IL by Enrico Fermi and Leo Szilard)**
4. **Assembling and then testing nuclear device (Los Alamos, NM by J. Robert Oppenheimer)**

Project 1: Obtaining and purifying U-235

Natural uranium contains 0.7205% U-235, the fissile isotope of uranium. The remaining mass includes 99.274% U-238 and a small amount of U-234 (0.0055%). Uranium-238 does not contribute to slow neutron fission; however, it can react with neutrons

to form a fissile isotope of plutonium, Pu-239. Thus U-238 is known as a fertile material, i.e., one that can produce fissile materials. Although U-235 and U-238 are chemically identical, they differ slightly in their physical properties, most importantly mass. This small mass difference allows the isotopes to be separated and makes it possible to increase (enrich) the percentage of U-235 in uranium.

In producing U-235 for the first atomic bomb, Manhattan Project scientists considered four physical processes for uranium enrichment:

- gaseous diffusion (effusion),
- electromagnetic separation,
- liquid thermal diffusion, and
- centrifugation.

This text will focus only on the first two mentioned methods of isotopic separation. During the project, the electromagnetic separation, gaseous diffusion, and liquid thermal diffusion were employed at Oak Ridge, Tennessee. The combination of these processes formulated the fuel for the Hiroshima bomb (Little Boy). Centrifugation was abandoned because the technology and materials required to spin corrosive uranium hexafluoride with a rotator at high speeds were not practical for industrial, large-scale separations. However, advances in technology and materials make centrifugation the preferred method of enrichment today.

Gaseous Diffusion

Gaseous diffusion has been used to produce enriched uranium for use in nuclear power plants and weapons. Nuclear reactors require fuel that is 2–5% ^{235}U, and nuclear bombs need around 90% U-235 concentrations. In a gaseous diffusion enrichment plant, uranium hexafluoride (UF_6, the only uranium compound that is volatile enough to work) is slowly pumped through large cylindrical vessels called diffusers, which contain porous barriers with microscopic openings. The process is one of diffusion because the other side of the barrier is not evacuated. The $^{235}UF_6$ molecules have a higher average speed and diffuse through the barrier a little faster than the heavier $^{238}UF_6$ molecules. The gas that has passed through the barrier is slightly enriched in $^{235}UF_6$ and the residual gas is slightly depleted. The small difference in molecular weights between $^{235}UF_6$ and $^{238}UF_6$ only about 0.4% enrichment, is achieved in one diffuser (Figure 6.3.2). But by connecting many diffusers in a sequence of stages (called a cascade), the desired level of enrichment can be attained.

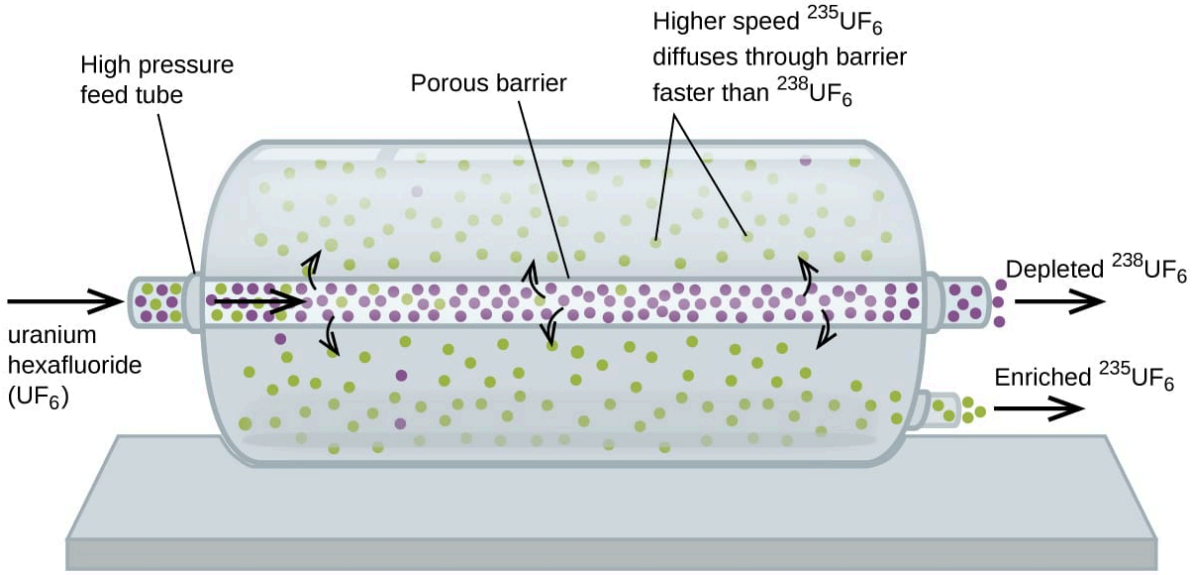


Figure 6.3.2: In a diffuser, gaseous UF_6 is pumped through a porous barrier, which partially separates $^{235}UF_6$ from $^{238}UF_6$ The UF_6 must pass through many large diffuser units to achieve sufficient enrichment in ^{235}U.

The gaseous diffusion method of isotopic separation presented many problems. First, a non-corrosive material (nickel and aluminum oxide) needed to be selected to serve as a barrier. The holes in the barrier needed to be microscopic(approximately one-millionth of an inch in diameter) and uniform in size. Gaseous diffusion plants require very large amounts of energy to alter

pressurized areas in the diffuser cascades. In addition, the entire system must be leak free; no air can be allowed in and no uranium hexafluoride (extremely corrosive gas) can escape.



Figure 6.3.3: The K25 Gaseous Diffusion Plant at Oak Ridge, Tennessee. Courtesy of Atomic Archive

The laboratory facility that would handle the large scale separation of gaseous $^{235}UF_6$ from $^{238}UF_6$ was constructed at Oak Ridge, Tennessee. This plant, built during World War II, was one-half mile long, six stories high, and covered 43 acres.

Electromagnetic Separation

The electromagnetic isotope separation (EMIS) process is based on the principle of a simple mass spectrometer, which states that a charged particle will follow a circular path when passing through a uniform magnetic field. Thus, uranium tetrachloride chloride ions containing U-235 and those containing U-238 with the same charge and kinetic energy (energy related to movement) will have slightly different paths when moving through a magnetic field. The radius of the path traversed by the heavier U-238 ions will be larger than the radius of the path of the lighter U-235 ions. This allows for the separation and collection of the isotopes in receivers (Figure 6.3.4).



Figure 6.3.4: J. Robert Oppenheimer (father of atomic bomb), Enrico Fermi (scientist who invented fission) and Ernest Lawrence (engineer who designed calutrons) discussing U-235 enrichment. Photograph of the calutron in the Y-12 facility. *Images taken from https://upload.wikimedia.org/wikiped..._Lawrence.jpeg and https://upload.wikimedia.org/wikiped...lpha_Track.jpg*

EMIS units, known as "calutrons", were operated in the Y-12 area of Oak Ridge. Because EMIS is a batch process, with each unit requiring a long time to produce small amounts of U-235, many units were used to produce the first fissionable uranium. Uranium ions for the EMIS are generated by heating solid uranium tetrachloride, UCl_4, to produce a vapor that is then bombarded with electrons to produce positively charged ions. There was a major problem with this process: less than half of the original U-235 was collected in the receivers. The rest was scattered throughout the calutron and was difficult to recover. At Oak Ridge, two groups of calutrons were used to produce weapons-grade uranium. The first group, or alpha calutrons, enriched uranium (7.0% U-235) from the gaseous diffusion plant to between 12 and 20% U-235. The second group, or beta calutrons, took the alpha product and enriched it to approximately 90% U-235 (weapons-grade uranium). As difficulties with the gaseous diffusion process were overcome, the EMIS process was phased out.

Use of enriched U-235

By the spring of 1945, Oak Ridge had shipped approximately 132 lbs. of enriched uranium (90% U-235) to Los Alamos, New Mexico; it had been produced using thermal diffusion, gaseous diffusion, and electromagnetic separation (Figure 6.3.5). This enriched uranium was used in "Little Boy", the bomb dropped on Hiroshima on August 6, 1945. With the discontinuation of thermal diffusion and electromagnetic separation after the war, gaseous diffusion became the primary method for enriching uranium in the United States. Today, most nuclear weapons use plutonium as the fissionable component.

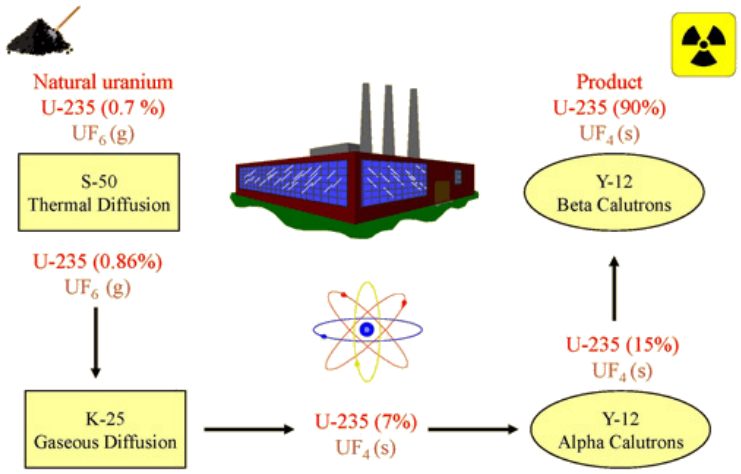


Figure 6.3.5: Methods Used to Produce Enriched Uranium for the First Atomic Bomb(Frank Settle)

Project 2: Synthesizing Pu-239

In the winter of 1940, a group of chemists (Glenn Seaborg, Edwin McMillan, Joseph Kennedy, and Arthur Wahl) bombarded U-238 with deuterons (H-2) in the University of California (Berkeley) cyclotron. They identified a new isotope of Np, which decayed by beta emission and produced a new element, plutonium.

$$^{238}_{92}\mathrm{U}_{92} + ^{2}_{1}\mathrm{H} \rightarrow ^{238}_{93}\mathrm{Np} + 2\,^{1}\mathrm{n}_{0} \tag{6.3.1}$$

$$^{238}_{93}\mathrm{Np} \xrightarrow{\tau_{½}=2.12 days} ^{238}_{94}\mathrm{Pu} + \beta^{-} \tag{6.3.2}$$

$$^{238}_{94}\mathrm{Pu} \xrightarrow{\tau_{½}=90 year} \alpha^{2+} + ^{236}_{92}\mathrm{U} \tag{6.3.3}$$

Figure 6.3.6: Glen Seaborg, the discoverer of Pu-239, standing in his laboratory. Image used with permission (Public Domain; United States Department of Energy).

The new plutonium atom was chemically different from U, Np, and other reaction products. At this time, Seaborg and Wahl could not identify which Pu isotope produced the alpha activity. It was later identified as Pu-238. In the spring of 1941, another more important isotope of plutonium, Pu-239, was produced using neutrons from the Berkeley cyclotron to target a uranium compound

surrounded by paraffin. This new Pu isotope, an alpha emitter with a half life of about 24,000 years, was separated from other reaction products using the same chemistry as that used to isolate Pu-238. However, the longer half-life of Pu-239 reduced its activity, making it more difficult to detect than Pu-238.

$$^{238}_{92}\mathrm{U} + {}^{1}\mathrm{n}_{0} \rightarrow {}^{239}_{92}\mathrm{U} + \gamma \tag{6.3.4}$$

$$^{239}_{92}\mathrm{U} \xrightarrow{\tau_{½}=23.5\ \mathrm{min}} {}^{239}_{93}\mathrm{Np} + \beta^{-} \tag{6.3.5}$$

$$^{239}_{93}\mathrm{Np} \xrightarrow{\tau_{½}=2.35\ \mathrm{days}} {}^{239}_{94}\mathrm{Pu} + \beta^{-} \tag{6.3.6}$$

$$^{239}_{93}\mathrm{Pu} \xrightarrow{\tau_{½}=24{,}110\ \mathrm{yrs}} {}^{235}_{92}\mathrm{U} + \alpha \tag{6.3.7}$$

In March 1941, Seaborg's group irradiated a sample estimated to contain 0.25 mg of Pu-239 surrounded by paraffin with neutrons produced in the cyclotron. Under these conditions, this isotope appeared to undergo fission. When the Pu was replaced with a sample containing approximately 0.5 mg U-235, the other known fissionable material, neutron-induced fission was also observed, but at a rate approximately half that of Pu-239. This discovery raised the possibility of using a controlled chain reaction to produce quantities of Pu-239 sufficient for nuclear weapons. The product Pu-239 would have to be separated from the unreacted uranium and fission products by chemical means. It now became important to investigate the chemistry of plutonium to develop large-scale separation procedures.

However, in the summer of 1942, the cyclotron was the only means of producing plutonium, and the amounts produced were so small that they could not be seen or weighed with existing balances. Calculations showed that long periods of neutron bombardment of uranium in the cyclotron would produce only a few micrograms of Pu-239, much less than that normally required to determine the physical and chemical properties of a new element. The preparation and measurement of such small quantities of plutonium required the development of "ultramicrochemical" techniques and equipment.

At the University of Chicago's Metallurgical Lab (referred to as the Met Lab), the first weighing of a plutonium compound occurred in fall of 1942. Only 2.77 micrograms of PuO_2 was isolated and measured with a balance especially designed for small masses. In November 1943, the first pure plutonium metal was chemically prepared at a temperature of 1,400° C. The plutonium metal appeared as silvery globules weighing about 3 ug each.



Figure 6.3.7: Twenty micrograms of plutonium hydroxide in a capillary tube, September 1942 (Courtesy of the University of California, Berkeley)

Although the Pu-239 isotope had the potential to be fissionable material for bombs or power generation, realization of this potential required larger amounts of this isotope. Large-scale production of Pu-239 required a controlled nuclear chain reaction of uranium, a feat that would soon be achieved by Enrico Fermi and Leo Szilard in Chicago. A facility to mass produce Pu-239 was built in Hanford, Washington. Glenn Seaborg was assigned to direct this particular laboratory. The resulting fuel would be used in the Trinity test bomb,tthe Nagasaki bomb, and two reserve weapons.

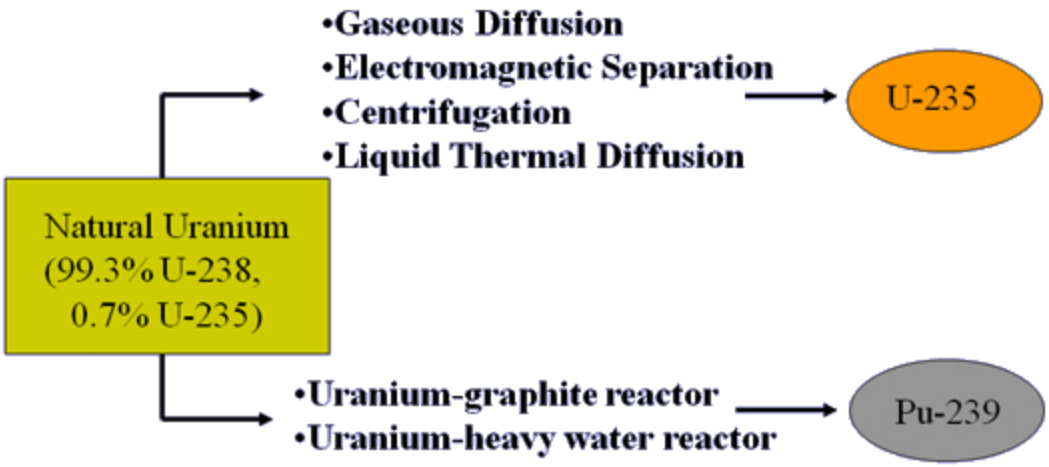


Figure 6.3.8***:*** *Routes to Fissionable Materials for Atomic Bombs*

Example 6.3.1:

Watch the second installment of the History Channel's Manhattan Project. Answer the questions below while viewing this portion of the video.

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 region of the country was selected to be the bomb design site?
2. Who did General Groves select to be the lead scientist? Briefly describe his personality.
3. How much U-235 was enriched during the Manhattan Project? What amount of U-235 was used in the construction of Little Boy?
4. What was special about the construction of Y-12? Where was this facility and what did it provide for the Manhattan Project?
5. How did Groves convince construction workers to come to Y-12?
6. Did the calutrons work perfectly when the alpha one racetrack was started?
7. What was the Gadget? What type of fuel would it use and how would it explode?
8. What element was made in 1941? What area of the country would produce this isotope in large amounts?
9. Is Pu-239 more or less likely to undergo fission than U-235?
10. What is K-25 and where is it located? What would K-25 produce?

Sources

1. http://www.history.com/topics/the-manhattan-project
2. https://www.osti.gov/opennet/manhatt...m_research.htm

Contributors

- Frank A. Settle (Washington and Lee University)

This page titled 6.3: The Manhattan Project - Labs and Fuel is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.4: The Manhattan Project - Critical Mass and Bomb Construction

Learning Objectives

- Explain the concept of critical mass for a fission reaction.
- Know the bomb mechanisms of Little Boy and Fat Man.

Determining Critical Mass

Material that can sustain a nuclear fission chain reaction is said to be fissile or fissionable. This reaction becomes self-sustaining when the number of neutrons produced by fission equals or exceeds the number of neutrons absorbed by splitting nuclei plus the number that escape into the surroundings. The amount of a fissionable material that will support a self-sustaining chain reaction is a **critical mass**.

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

In order to determine critical mass required for a nuclear chain reaction, Enrico Fermi and Leo Szilard constructed the world's first controlled fission experiment. At the University of Chicago, the physicists piled 40,000 graphite (carbon based pencil material) blocks in a 24 square foot formation (see below). Fermi and Szilard embedded the graphite with 19,000 pieces of uranium. In a squash court (similar to a racketball court) on December 2, 1942, slow neutrons bombarded the pile to produce a sustained, nuclear chain reaction. The fission reaction was detectable by using a Geiger Counter. This mini reactor did not explode or injure any scientists due to its design. From these experiments, Fermi, Szilard and their colleagues determined the critical mass of U-235 to be approximately 50 kg. In addition, they concluded that critical mass was dependent upon the type of fuel (U-235 or Pu-239), shape of fuel (round reduces amount), temperature of the reaction (higher increased neutron collisions), and density of material. Other experiments and calculations lead to the critical mass determination of Pu-239. Unlike U-235, only 10 kg of Pu-239 was required to sustain a nuclear chain reaction.

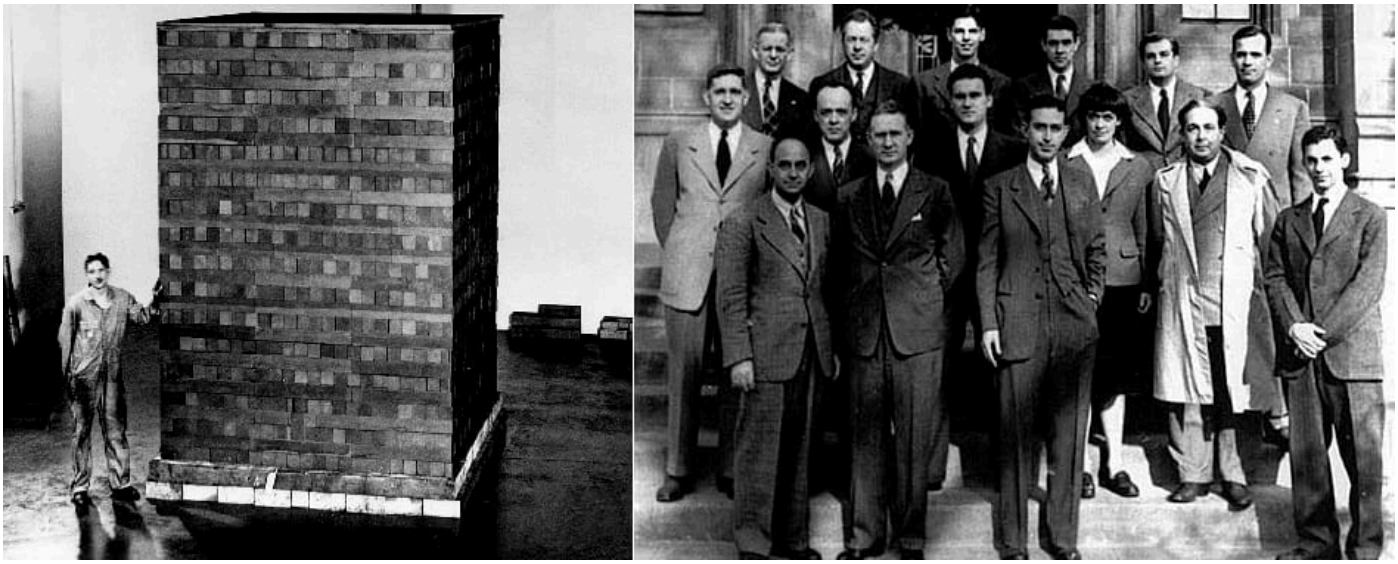


Figure 6.4.1: Chicago Pile-1 reactor built to produced sustainable nuclear fission. Chicago Pile-1 scientists standing together. Enrico Fermi is standing on bottom row, far-left. Leo Szilard is positioned to the right of the only female scientist (Leona Woods) in this particular research group. Images taken from: https://upload.wikimedia.org/wikiped...2803096%29.jpg and https://upload.wikimedia.org/wikiped...goPileTeam.png

The fission of uranium produces more neutrons than it consumes. As can be seen from Eqs. (1), for every neutron captured by a U-235 nucleus, between two and four neutrons are produced. As soon as a stray neutron bombards a U-235 nucleus, fission will take place and three neutrons will be produced. These in turn will fission three more U-235 nuclei, producing a total of nine neutrons. A third repetition will produce 27 neutrons. a fourth 81. and so on. This process (which is called a *chain reaction*) escalates very rapidly. Within a few microseconds a very large number of nuclei fission, with the release of a tremendous amount of energy, and an atomic explosion results.

$$
{}^{235}_{92}\mathrm{U} + {}^{1}_{0}n \longrightarrow \left[{}^{236}_{92}\mathrm{U}\right] \begin{cases} \longrightarrow {}^{140}_{58}\mathrm{Ba} + {}^{93}_{36}\mathrm{Kr} + 3{}^{1}_{0}n \\ \longrightarrow {}^{144}_{55}\mathrm{Cs} + {}^{90}_{37}\mathrm{Rb} + 2{}^{1}_{0}n \\ \longrightarrow {}^{144}_{54}\mathrm{Xe} + {}^{90}_{38}\mathrm{Sr} + 2{}^{1}_{0}n \\ \longrightarrow {}^{146}_{57}\mathrm{La} + {}^{87}_{35}\mathrm{Br} + 3{}^{1}_{0}n \\ \longrightarrow {}^{160}_{62}\mathrm{Sm} + {}^{72}_{30}\mathrm{Zn} + 4{}^{1}_{0}n \end{cases}
$$

There are two reasons why a normal sample of uranium metal does not spontaneously explode in this way. Natural uranium consists mainly of the isotope U-238 while the fissionable isotope U-235 comprises only 0.7 percent of the total. Most of the neutrons produced in a given fission process are captured by U-238 nuclei without any further production of neutrons. The escalation of the fission process thus becomes impossible. However, even a sample of pure U-235 will not always explode spontaneously. If it is sufficiently small, many of the neutrons will escape into the surroundings without causing further fission. The sample must exceed a **critical mass** before an explosion results. In an atomic bomb several pieces of fissionable material, all of which are below the critical mass, are held sufficiently far apart for no chain reaction to occur. When these fragments are suddenly brought together, an atomic explosion results immediately.

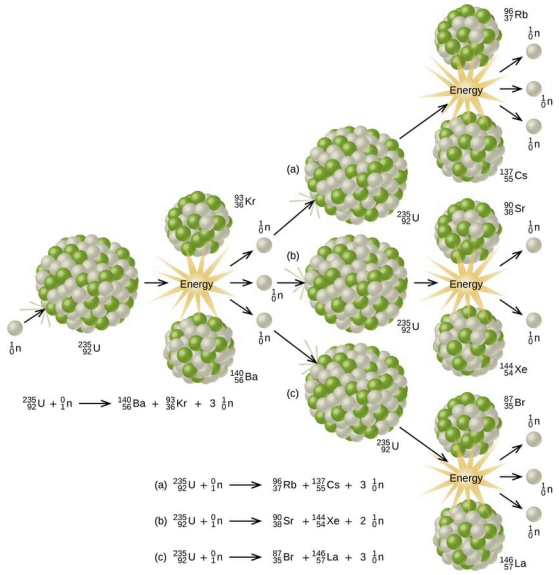


Figure 6.4.2: The fission of a large nucleus, such as U-235, produces two or three neutrons, each of which is capable of causing fission of another nucleus by the reactions shown. If this process continues, a nuclear chain reaction occurs.

Bomb Construction

At Los Alamos, New Mexico, J. Robert Oppenheimer and his research group were given the task of assembling and testing the fission bomb. Both types of fuel that were manufactured (U-235 and Pu-239) would be placed separately into two different bomb designs. The smaller of the two bombs would contain U-235 and was code named "Little Boy." The larger bomb, which housed Pu-239, was named "Fat Man." Of the two, only the Pu-239 bomb would be tested before it was dropped in wartime.

The basic design of the U-235 bomb is shown in figure. To prevent spontaneous detonation of an atomic bomb, the fissile U-235 is kept in a subcritical configuration. It is then rapidly assembled into a supercritical mass using conventional explosives. Once the bomb has achieved this mass, any neutron introduced into it will be likely to initiate a chain reaction. The mechanism for "Little Boy" was a gun that fired one subcritical piece of U-235 into another to form a supercritical mass (Figure 6.4.3). The pieces had to be assembled within a time less than the average time between appearances of spontaneous neutrons from either U-235 or cosmic radiation. A conventional explosive in an artillery barrel could fire the U-235 mass at speeds of a few millimeters per second, fast enough to prevent a fizzle caused by a spontaneous neutron setting off a premature chain reaction.

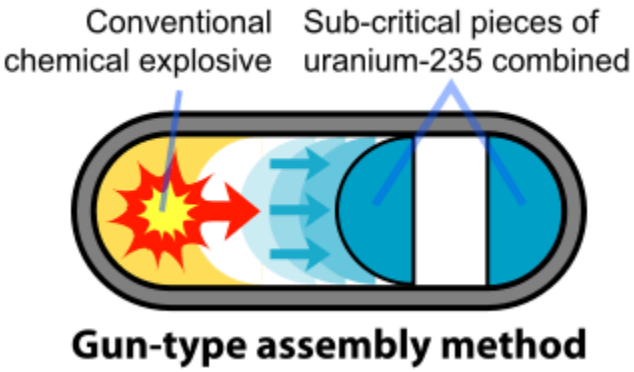


Figure 6.4.3: Little Boy: Gun-Type Device. Image used with permission (Public Domain).

Originally, the gun-type mechanism was planned for both the U-235 and Pu-239 weapons. However, a problem arose with the Pu-239 bomb that required a different assembly mechanism because of the small amount of Pu-240 that is produced with the Pu-239 in the reactor. Pu-240 emits large numbers of neutrons spontaneously: 1,030 neutrons per gram per second compared with 0.0004 neutrons per gram per second for U-235. Even at a concentration of 1% Pu-240 in the fissile Pu-239, the required mass of Pu emits 52,000 neutrons per second or one neutron every 20 microseconds. Thus, it is very probable that a neutron from Pu-240 will initiate a premature chain reaction during the critical last 100 microseconds of the critical mass in a gun-type assembly. This problem was discovered in mid 1944, well after the start of construction of the massive Hanford plutonium

production facilities.

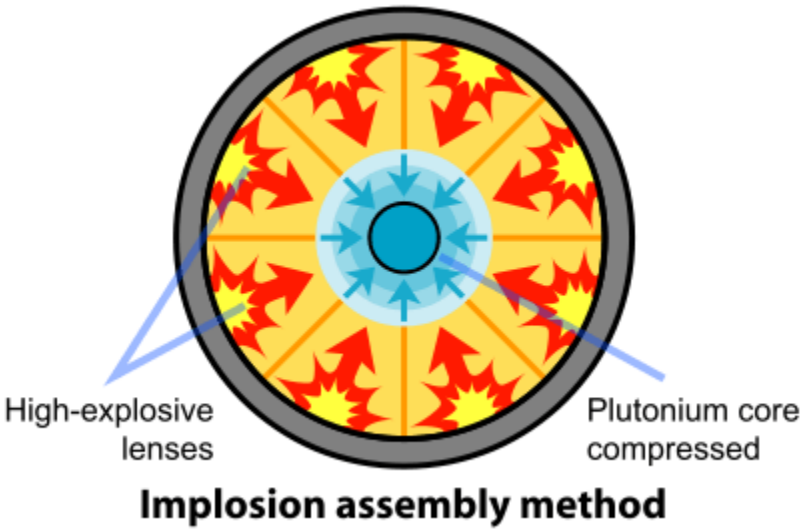


Figure 6.4.4: Fat Man: Implosion-type bomb. Image used with permission (Public Domain).

Removal of the Pu-240 was impractical. So the scientists and engineers looked for a faster method of assembling the plutonium. A mechanism based on implosion provided the solution to this problem. In this design, the fissile material is shaped into a single sphere with a mass slightly less than critical (Figure 6.4.4). Layers of carefully shaped high explosives surround the sphere. When the explosives are detonated, the force of the shock wave compresses the fissile material into a smaller volume, forming a supercritical mass. This method of assembly is much faster than the gun-type mechanism and thus eliminates the problems resulting from spontaneous neutron emission of Pu-240. The spherical mass resulted in a pumpkin-shaped weapon called "Fat Man".

Example 6.4.1

Watch the first 5:30 minutes of the third installment of the History Channel's Manhattan Project and 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. What were the problems with the two different types of fuel that were to be used in the fission weapons?
2. What was the configuration or set-up of the Pu-239 bomb?
3. The combination of electromagnetic separation and gaseous diffusion enriched enough U-235 to produce how many bombs?
4. Why was the U-235 bomb never tested?
5. When was the construction of the U-235 completed?
6. How much did Truman know about the nuclear bombs when FDR was alive?
7. Who was selected to pilot the plane for the nuclear weapons?
8. What island would serve as a base to house the nuclear bombs before delivering them to their targets?

In July 1945, the United States had enough fissile material for one uranium and two plutonium weapons. The scientists and engineers felt confident that the gun-type assembly mechanism for Little Boy would function properly. Besides, they did not have the material for a test device. They were less confident about the implosion mechanism on the plutonium weapon and felt that a test was necessary. On July 16, 1945, the first nuclear device, known as "The Gadget", was placed on a 100-foot tower and successfully detonated in the Alamogordo Desert, 200 miles south of Los Alamos (Figure 6.4.5).

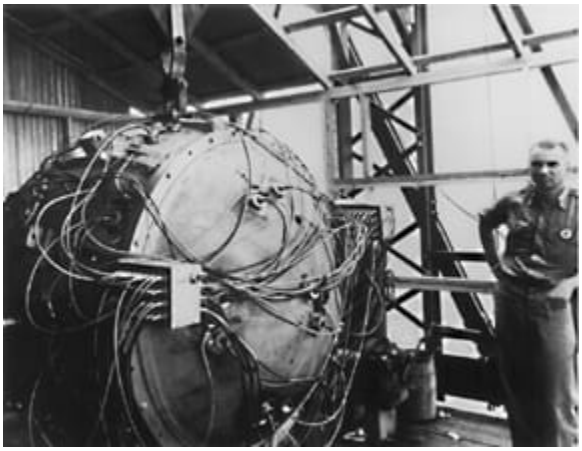


Figure 6.4.5: The Gadget at the Trinity Test Site, 1945. Courtesy of the U.S. Department of Defense

As the war with Japan continued and a costly allied invasion loomed as a real possibility, President Truman approved the use of nuclear weapons against selected Japanese targets. The U.S. Army Air Force received orders to use these weapons anytime after August 3, 1945. On August 6, "Little Boy" was dropped on Hiroshima (Figure 6.4.6).



Figure 6.4.6: Little Boy - The only U-235 weapons constructed during WWII. Courtesy of the U.S. Department of Defense

Little Boy was a uranium weapon containing 141.4 pounds of fissionable material containing 82.7% U-235. Only about two pounds fissioned, releasing an energy equivalent to 15-16,000 tons of TNT. The immediate effects of the blast killed an estimated 70,000 people, and by the end of 1945 an additional 20,000 to 70,000 deaths occurred, many due to lack of adequate medical resources.

Three days later, "Fat Man" destroyed a large part of Nagasaki (Figure 6.4.7). Fat Man contained 13.6 pounds of Pu-239, of which only 2 pounds underwent fission. The explosive yield was equivalent to about 22,000 tons of TNT. Fat man resulted in 35,000 immediate deaths. By the end of 1945, at least 70,000 people perished in this event.



Figure 6.4.7***:*** *Fat Man – One of many Pu-239 nuclear weapons constructed during WWII. Courtesy of the U.S. Department of Defense*

Example 6.4.2

Continue watching the third installment of the video and answer the questions below:

9. When and where was the first nuclear bomb tested?
10. Was the bomb placed on the ground?
11. What were scientists afraid the first nuclear bomb would do?
12. What did Edward Teller (father of fusion bomb) apply to his skin before the bomb was ignited? Why did he do this?
13. What did the bomb do to the tower and the surrounding sand?
14. In the photograph, where did General Grove focus his attention?

Sources

1. https://www.uchicago.edu/features/ho...anged_science/

Contributors

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 6.4: The Manhattan Project - Critical Mass and Bomb Construction is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.5: Hiroshima and Nagasaki

Target Cities

On May 10 and 11th, a Target Committee meet at Los Alamos National Laboratory to determine possible targets for atomic bomb. Military personnel and selected Manhattan Project scientists discussed the preparations required for the dropping of Little Boy. They mentioned height of detonation, weather and radar data, and rehearsals for the mission. Cities selected needed to be urban areas of more than three miles in diameter. In addition, bombing targets would need to be effectively damaged by a blast and they would not be damaged before being subject to nuclear warfare. The short list of targets can be seen below in order of selection:

1. Kyoto (former capital and industrial center with a population of over 1 million)
2. Hiroshima (army depot and port)
3. Yokohama (untouched industrial center)
4. Kokura (one of the largest arsenals in Japan)
5. Niigata (untouched port area, industrial center, housed oil refineries)

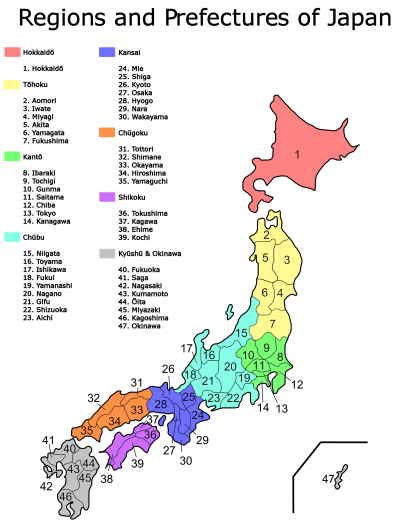


Figure 6.5.1: Map of Japan. Image taken from: https://upload.wikimedia.org/wikiped...s_of_Japan.png

Immediately, Kyoto was removed from this list. The United States Secretary of War, Henry Stimson, had honeymooned there previously and recognized the cultural significance of this area. Kyoto contained beautiful gardens, 17 World Heritage Sites, and over 2,000 Buddhist Temples and Shinto Shrines. Many committee members argued that dropping the bomb on Kyoto would have more emotional impact on the Japanese than dropping it on another city. Regardless, Kyoto was removed from the list of target sites and the bomb would be diverted to Hiroshima.

Hiroshima (August 6, 1945)

Figure 6.5.2: Crew members of the Enola Gay. Paul Tibbets (pilot) is in the center of the crew. **Figure** 6.5.3: Plane routes for the atomic bomb missions. Images taken from: https://media.defense.gov/2004/Mar/1...-9999G-005.JPG and https://upload.wikimedia.org/wikiped...on_map.svg.png

On August 6, 1945, Colonel Paul Tibbets and his 11 membered crew flew the Enola Gay from Tinian Island to Hiroshima, Japan. Previously that morning, other American weather planes had surveyed the visibility for this target. On August 6, 1945, Little Boy was dropped around 8:15 a.m. It fell for over 44 seconds and then exploded at approximately 2000 feet (epicenter) above the city of Hiroshima. During this time, two other American planes recorded scientific data (The Great Artiste) and visual images (The Necessary Evil). The bomb packed a power equivalent to 12-15 kilotons of TNT.

The Atomic Bombing of Hiroshima



"All the News That's Fit to Print"

The New York Times.

LATE CITY EDITION
Partly cloudy, less humid today. Cloudy and warm tomorrow.
Temperatures Yesterday—Max., 72; Min., 66
Sunrise today, 5:37 A. M.; Sunset, 8:06 P. M.

Copyright, 1945, by The New York Times Company.

VOL. XCIV .. No. 31,972. Entered as Second-Class Matter, Postoffice, New York, N. Y. NEW YORK, TUESDAY, AUGUST 7, 1945. THREE CENTS IN NEW YORK CITY

FIRST ATOMIC BOMB DROPPED ON JAPAN; MISSILE IS EQUAL TO 20,000 TONS OF TNT; TRUMAN WARNS FOE OF A 'RAIN OF RUIN'

HIRAM W. JOHNSON, REPUBLICAN DEAN IN THE SENATE, DIES

Isolationist Helped Prevent U. S. Entry Into League—Opposed World Charter

CALIFORNIA EX-GOVERNOR

Ran for Vice President With Theodore Roosevelt in '12—In Washington Since '17

Special to The New York Times.

WASHINGTON, Aug. 6—Senator Hiram Warren Johnson of

Jet Plane Explosion Kills Major Bong, Top U. S. Ace

Flier Who Downed 40 Japanese Craft, Sent Home to Be 'Safe,' Was Flying New 'Shooting Star' as a Test Pilot

By The United Press.

BURBANK, Calif., Aug. 6—Maj. Richard Bong, America's greatest air ace, died today in the flaming wreckage of a jet propelled fighter plane which crashed while he was testing it.

Only 24 years old, he wore twenty-six oak decorations including the nation's highest award, the Congressional Medal of Honor. He had survived countless air battles and shot down forty Japanese planes without a scratch.

The knowledge he gained in those battles was too valuable to

"The plane started to wobble up and down, then went into a left bank and hit the ground," he stated. "It exploded and burned and scattered wreckage over about a block square."

Major Bong was trying to get out of the ship when it crashed. He had released the escape hatch and was partly clear. He had pulled the ripcord to his parachute, and the silken folds lay about the body as the flames swept over it.

With a roaring sigh, the plane, like a giant blowtorch, shot over the airport just before 3 P. M. and

KYUSHU CITY RAZED

Kenney's Planes Blast Tarumizu in Record Blow From Okinawa

ROCKET SITE IS SEEN

125 B-29's Hit Japan's Toyokawa Naval Arsenal in Demolition Strike

By FRANK L. KLUCKHOHN

By Wireless to The New York Times.

MANILA, Tuesday, Aug. 7—More than 400 fighters and bomb-

REPORT BY BRITAIN

'By God's Mercy' We Beat Nazis to Bomb, Churchill Says

ROOSEVELT AID CITED

Raiders Wrecked Norse Laboratory in Race for Key to Victory

The text of Mr. Churchill's statement is on Page 8.

Steel Tower 'Vaporized' In Trial of Mighty Bomb

Scientists Awe-Struck as Blinding Flash Lighted New Mexico Desert and Great Cloud Bore 40,000 Feet Into Sky

By LEWIS WOOD

Special to The New York Times.

WASHINGTON, Aug. 6—A blinding flash many times as brilliant as the midday sun and a massive, multi-colored cloud boiling up 40,000 feet into the air accompanied the first test firing of an atomic bomb on July 16, three weeks ago today. Set in the remote desert-lands of New Mexico, the experiment was seen against a wild background where rain poured in torrents, and lightning pierced the sky up to the zero hour of the explosion at 5:30 A. M.

A steel tower from which the

Before the detonation scientists waited in tense expectancy. Minutes lengthened seemingly to hours. Lying face downward, with their feet toward the steel tower, the watchers waited, nearly breathless. They were "reaching into the unknown" and did not know what would happen.

On the instant that all was over these men leaped to their feet. The terrible tension ended, they shook hands, embraced each other and shouted in glee. Behind their tri-umph was sober consciousness of

NEW AGE USHERED

Day of Atomic Energy Hailed by President, Revealing Weapon

HIROSHIMA IS TARGET

'Impenetrable' Cloud of Dust Hides City After Single Bomb Strikes

Truman, Stimson statements on atomic bomb, Page 4.

See interactive NYTimes Story.

Before the bombing of Hiroshima, the population of the city was estimated to be around 255,000. This city was relatively flat and was densely populated. Many buildings were constructed with reinforced concrete in order to withstand earthquake damage. Hiroshima housed a military base that served as a communications center, storage point, and assembly area for the Japanese army.

At around 8:00 a.m. that morning, a Japanese radar operator detected three planes flying near Hiroshima. These planes were American weather planes noting visibility for the bombing that would occur. Due to the small number of planes, Hiroshima citizens were encouraged but not required to seek shelter. At 8:15 a.m., Little Boy detonated producing a blinding flash and a booming sound.

Due to crosswinds, the Enola Gay missed the original target of the Aioi Bridge and directly hit the Shima Surgical Clinic. Total destruction occurred within one mile of the epicenter (ground beneath = hypocenter). All buildings, with the exception of reinforce concrete, were completely leveled. Extending to 5700 feet from the epicenter, most steel framed buildings showed extensive damage. At approximately 6600 feet from the epicenter, multi-storied brick buildings had some degree of wreckage. In the city, over 69 % of the buildings were destroyed from the blast. Overhead electrical installations (power lines) were destroyed at a distance of over 5500 feet from the epicenter.

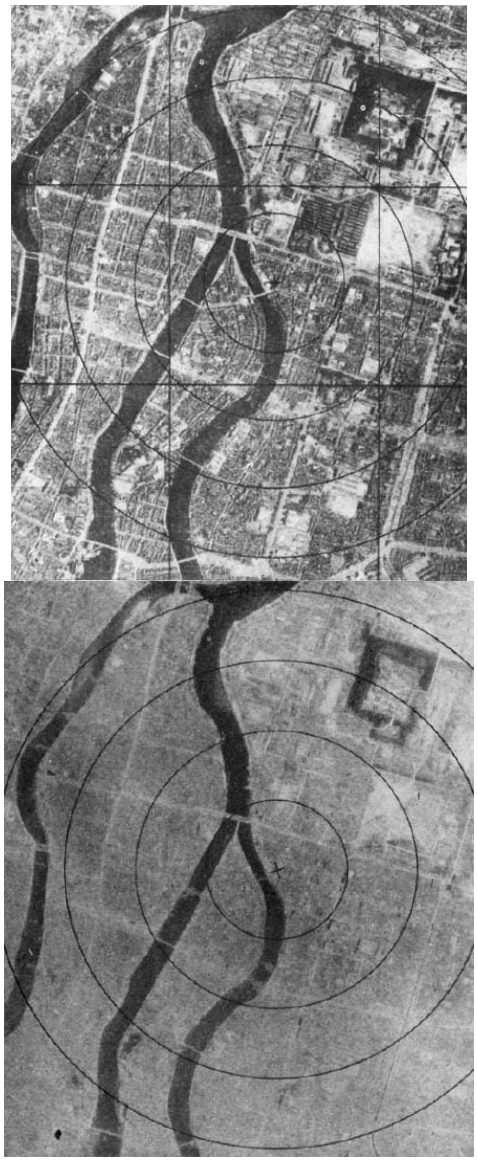


Figure 6.5.4:Hiroshima before bombing. Area around ground zero. 1,000 foot circles.Hiroshima after the bombing and firestorm

Within a few microseconds of detonation, gamma, ultraviolet and visible radiation were emitted to the surrounding areas. Shortly thereafter, a large ball (several hundred feet in diameter) was produced. Intense temperatures((in excess of a million degrees Celsius) and changes in pressures (19 tons per square meter) ultimately produced a growing shockwave. With the formation of a mushroom could, particle (mainly alpha and neutron) and wave (gamma, ultraviolet, and visible) are emitted a second time.

93% of residents who were up to 1000 feet of the epicenter died immediately. Deaths were attributed to thermal burns (from wave radiation), vaporization, flying debris, and extreme radiation exposure. Residents who were about 5,000 feet (almost a mile) from the epicenter had a 49% mortality rate. A distance of 10,000 feet (almost two miles) from the epicenter ensured safety. Individuals who were inside a structure were more likely to survive the detonation. Obviously, reinforced concreted structures provided more protection that wooden or bricked facilities. At Hiroshima, around 60% of the mortalities were attributed to severe burns. Falling debris killed around 30% of deaths. Lastly, pressure changes and radiation exposure killed the remaining 10%. In total, 60,000 - 70,000 people died immediately after the detonation of Little Boy. Within three months of this date, the total would increase to over 130,000 people.

Nagasaki (August 9, 1945)

Three days after the bombing of Hiroshima, the Bockscar with its crew flew towards Kokura. Piloted by Charles Sweeney, this plane carried the plutonium weapon, Fat Man. As with the Enola Gay, the Great Artiste accompanied the overloaded B-29. On that

morning, the Bockscar passed over Kokura three times attempting to drop its nuclear bomb. Cloudy weather prevented the crew from arming the weapon and the target was abandoned. Determined to still use the plutonium weapon, the Bockscar flew towards Nagasaki. At 11:02 a.m., Fat Man detonated 1700 feet above the city. Releasing over 20 kilotons of energy, the bomb exploded between Mitsubishi Steel and Arms(weapons facility) and Mitsibushi-Urakami Works (Torpedo factory). The plane immediately rerouted to Iwo Jima due to a fuel shortage. Hours later, the Bockscar returned to Tinian Island. Like the Enola Gay, both sets of crews were not harmed on either mission.

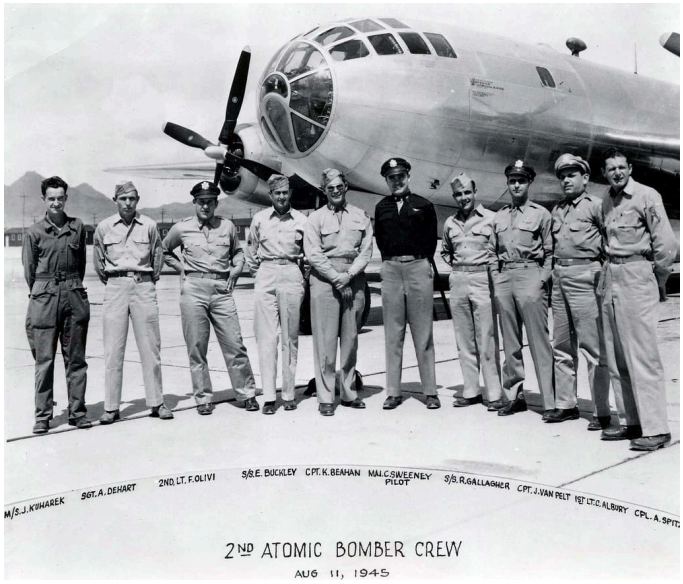


Figure 6.5.5:Crew of the Bockscar. ***Figure*** 6.5.6*: Bomb Technicians adjusting LIttle Boy on Tinian Island before being loaded into the Bockscar.* Images taken from: https://c1.staticflickr.com/4/3247/2...5a4e8794_b.jpg and https://upload.wikimedia.org/wikiped..._77-BT-186.jpg

The pre-bomb population of Nagasaki was approximately 195,000 people. Nestled between mountains, this city was contained in a valley. Differences in topography and population densities resulted in less casualties than the Hiroshima bomb. The larger plutonium weapon instantly killed between 35,000 to 39,000 residents. Within a year's time, a total of 65,000 people would perish.

Being a larger weapon, Fatman had a larger destruction radius when compared to Little Boy. Over 30% of Nagasaki was completely destroyed. Steel framed buildings were severely damaged within 6000 feet of the epicenter. Multi-stored buildings sustained significant destruction 6500 feet from epicenter. At Nagasaki, damage of Japanese homes extended to over 10,500 feet from the epicenter (compared to the distance of 8000 feet at Hiroshima). Nagasaki had fewer building constructed with reinforced concrete. Wood-framed buildings provided little to no protection for this city's residents.

95% of the deaths at Nagasaki were due to burns (gamma and ultraviolet). The remaining percent could be attributed to falling debris, glass, and radiation illnesses. Within 4900 feet of the epicenter, approximately 67% of the residents died. The epicenter distance of 6500-9900 feet (1.2-1.9 miles) resulted in approximately 14% casualty rate. Differences in population density and topography of area resulted in fewer deaths in Nagasaki than Hiroshima.

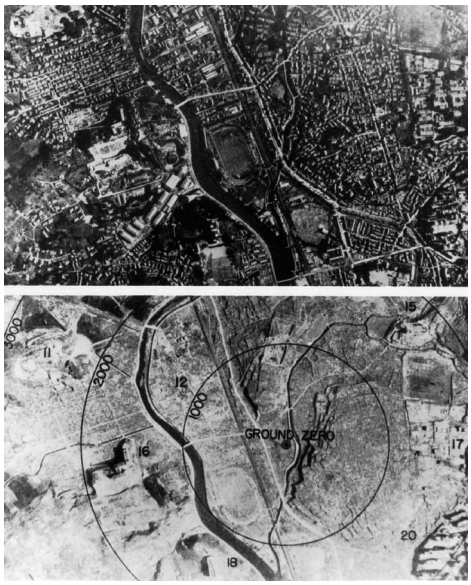


Figure 6.5.7 Pre and post bombing pictures of Nagasaki. Image taken from: https://upload.wikimedia.org/wikiped..._and_after.jpg

On August 15, 1945, Emperor Hirohito publicly announced the surrender of Imperial Japan. Aboard the USS Missouri battleship, members of the Japanese government signed the Japanese Instrument of Surrender on September 2, 1945.

Example 6.5.1:

Watch the installment four of the History Channel's Manhattan Project and 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 powerful was the plutonium test bomb (in tons). What did it do to the steel tower and the surrounding sand?
2. Were the scientists confident about the combat bombs? Which bomb would be dropped first?
3. What was the purpose of the petition at Oak Ridge? Did the other labs (Los Alamos, Hanford, and Chicago) get to see the petition?
4. Give some reasons as to why America still chose to drop the bombs.
5. Why were certain cities in Japan not bombed during 1945?
6. When and where was Little Boy dropped? Give time of day as well. How powerful (TNT) was it.
7. How did Paul Tibbets' describe Hiroshima?
8. What was the death toll at Hiroshima?
9. Where was Fatman dropped?
10. When did Japan surrender?
11. When did Russia detonate their first nuclear weapon?
12. What did Edward Teller design?
13. What was Oppenheimer's attitude toward furthering nuclear bomb 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

Sources

1) http://avalon.law.yale.edu/subject_menus/mpmenu.asp

This page titled 6.5: Hiroshima and Nagasaki is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.6: Aftermath of Bombings

Hiroshima

Sixteen hours after Little Boy was dropped, President Harry S. Truman announced publicly that a nuclear weapon had been used against Japan. Up until that time, Japanese scientists were unsure as to what had happened to their community. There was so discussion as to whether chemical or biological warfare had occurred. Fear spread that disease could be spread by being in contact with a resident of Hiroshima. Healthcare professionals were warned to not enter this area due to fear of contraction of an infectious disease.



Figure 6.6.1: Arial view of Hiroshima shown at the Hiroshima Peace Museum. Photo Courtesy of Shane Patel (Furman University, Class of 2016).

Individuals working at the Shima Hospital noted how all of the x-ray films had mysteriously been exposed. From this observation, resident healthcare workers dismissed this weapon to be chemical or biological in nature. Knowing this, the Japanese had to determine how to combat the effects of this new weapon. Fires burned upon Hiroshima for long periods of time. Water was scarce and contaminated. Families were separated and locating individuals was near to impossible. Communication had been severed from neighboring communities. Moreover, there were no medications designed to fight radiation sickness.

Exploring 24 hours after Hiroshima Video

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

Note: Please be aware of graphic images shown in times 1:22 -1:35 and 3:20-3:44 of this video.

Watch the video and answer the questions below:

1. What distance (in feet) from the epicenter resulted in almost 100% mortality for the residents of Hiroshima?
2. What are atomic shadows? How did scientists use these images to assist in understanding the atomic bomb detonation?
3. Richard Rhodes (atomic bomb book author) describes the melting of what types of materials at 3200°F?
4. Describe the lunch box that was found 1/3 mile from ground zero. Did the owner of this object survive the bombing?
5. How long does it take for the shock wave to flatten Hiroshima? What was the approximate number of buildings destroyed by shock wave?
6. When the shock wave occurred, how far away were the crew of the Enola Gay?
7. What was the radius of heavy damage from the shockwave?
8. Did any people survive from being inside the A-bomb dome? This is the building that survived the atomic blast at Hiroshima.
9. Where was Takashi (the school boy) when the atomic blast occurred? How did he survive?
10. How long did the fires burn after the bomb was dropped. How many square miles of the city were burned?

Typically, individuals who were outside and close (within 0.2 miles) to the epicenter were vaporized or severely burned. Some survivors as close as .4 miles lived if there were inside a concrete based building. Other types of structures offered little to no protection. In many cases, people inside of a structure were crushed to death. People outside during the explosion were exposed to flying debris and extreme pressure changes.

Remembering this event, it is important to people of all backgrounds and cultures. By viewing artifacts and listening to survivor testimonies, one can gain a small amount of insight as to what the residents experienced on August 6, 1945. In addition, listening to Threnody to the Victims of Hiroshima by Penderecki can provide an auditory experience of the sheer terror of nuclear warfare.



Figure 6.6.2: Kunishige Masahiro, a survivor of the Hiroshima bombing, shares his story with students (photo courtesy of Grace Cohen, Furman University, Class of 2016).

Today, Hiroshima has many museums and monuments dedicated to the survivors of Little Boy. One of the most notable sites is the A-bomb dome. This structure was approximately 600 meters from the hypocenter (ground zero) and 160 meters from the epicenter (air detonation). Unfortunately, none of the individuals inside this structure survived due to their close proximity to the explosion.



.



Figure 6.6.3: Picture of the A-bomb Dome in 1945. ***Figure*** 6.6.4**:** Image of the A-bomb Dome in 2016. Images taken from: https://upload.wikimedia.org/wikiped..._Dome_1945.gif and Shane Patel (Furman University, Class of 2016)

Over the years, the A-bomb Dome has undergone many restorations. Other monuments have been constructed and dedicated as well. Click here to see a comprehensive link to these sites. For people who cannot travel to Hiroshima, The Hiroshima Peace Museum offers a virtual tour. Clicking on this link, one can access testimonies, images of Hiroshima and its inhabitants, and messages of peace.

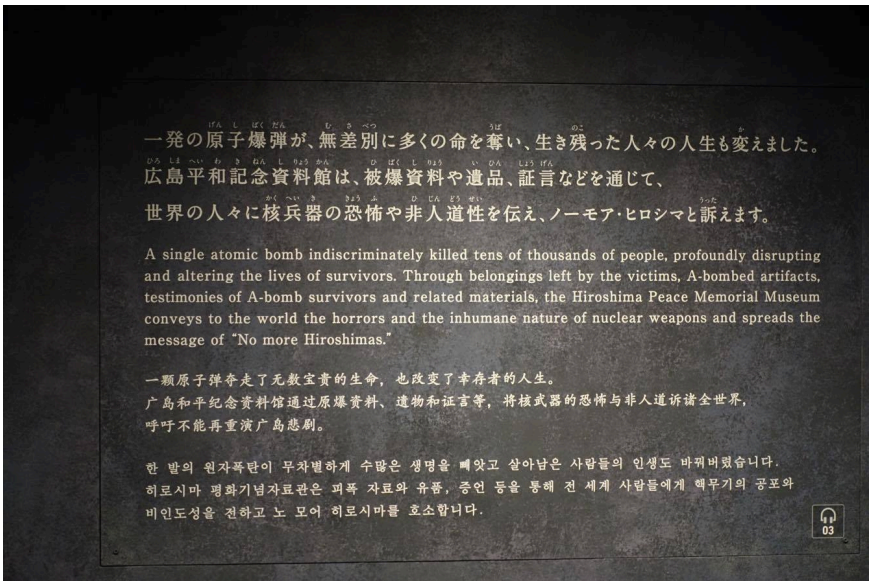


Figure 6.6.5: Plaque at the Hiroshima Peace Museum. Photo courtesy of Shane Patel (Furman University, Class of 2016).

Nagasaki

Unlike the flat terrain of Hiroshima, Nagasaki was located in a valley between mountains. The 20 kiloton blast was contained in this region affecting a ship building center and two Mitsubishi industrial parks. The population of this city was more dispersed than that of Hiroshima. Residents of Nagasaki lived in the valley as well in mountainous regions of this area. Unlike Hiroshima, most of the buildings were composed of wood and not concrete or brick. This material did not adequately protect residents and collapsed easily. Moreover, wooden structures became deadly projectiles of shrapnel during the detonation of Fat Man.

The most famous structure that survived the bombing of Nagasaki is the one-legged entrance to the Sanno Shrine. This building was approximately 800 meters from the bomb's hypocenter. Today, it has been incorporated into the newly constructed buildings of the city.





Figure 6.6.5: Image of the Sanno Shrine immediately after the bombing of Nagasaki. .**Figure** 6.6.6**:** A 2007 picture of the same shrine. Images taken from: https://upload.wikimedia.org/wikiped...-afterbomb.jpg and https://upload.wikimedia.org/wikiped...orii_C1946.jpg

Like Hiroshima, Nagasaki has numerous statues and a peace museum memorizing August 9, 1945. A tall tower has been erected at the hypocenter distance of 500 meters. At the base of this tower, a black stone describes the nuclear bombing and its effects. Each year, a memorial service is held at this site to celebrate survivors, remember victims, and promote peace. Just down from this tower, a large bronze statue (9.7 meters in height and 30 tons in weight) sits with its arms outstretched to the sky. Designed by a local sculptor, Seibo Kitamura, this Peace Statue symbolizes different aspects of nuclear weapons and the bombing of Nagasaki. The vertically raised arm represents the threat of nuclear bombing. The left hand with palm held flat symbolizes peace for all the world's inhabitants. The crossed leg illustrates quite meditation while the left leg shows an intense readiness to rescue and provide assistance to those affected by the destruction. Lastly, the face is not limited to one culture with eyes closed to offer prayers for those who perished.



Figure 6.6.7: The hypocenter of Fat Man and **Figure** 6.6.8**:** the Nagasaki Peace Statue. Both images courtesy of Shane Patel (Furman University, Class of 2016)

For a quick overview of some of the memorials, click on the video underneath this paragraph. The videographer provides interesting footage of the inside of the Nagasaki Peace Museum. You will see preserved artifacts and a three dimensional image of the blast radius. Once outside, he shows the hypocenter tower, Peace Statue, gardens, and various other memorials. For an explanation to the other memorials shown in his video, click on this link.

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

Sources

1. https://www.atomicheritage.org/histo...-nagasaki-1945

This page titled 6.6: Aftermath of Bombings is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

6.7: Fusion

Mike -The First Fusion Device

The first fusion device was tested by the United States on November 1st, 1952. On Elugelab Island located in the Marshall Islands, this nondeliverable weapon measured 20 feet in height and weighed over 140,000 pounds. During the eighth round of United States nuclear testing (named Operation Ivy), Mike (fusion weapon) vaporized the entire island where it was detonated. The 10.4 megaton blast was 500 times more powerful than the Nagasaki weapon and left a crater in the ocean that was 6,240 feet wide and 164 feet deep.

Example 6.7.1:

From video and classroom discussion, 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

A scene from our PBS documentary "Dr. Teller's Very Large Bomb" (with perhaps a bit too much information on how to build your own hydrogen bomb.) (Complete hour-long documentary available at www.foolishearthling.com)

1. Could the United States transported Mike by airplane? Was this bomb deliverable or nondeliverable?
2. What nuclear isotopes did this device contain?
3. What two types of nuclear reactions were used in Mike?
4. Is fission or fusion more technologically complex?
5. What was Teller's response after Mike exploded?
6. How far away were the observers of Mike?

Edward Teller, one of the Manhattan Project scientists, is credited designing the fusion bomb. Unlike J. Robert Oppenheimer, Teller felt it was important and necessary for the United States to pursue new nuclear weapons. After the dropping of Little Boy and Fat Man, Edward Teller testified against Oppenheimer in order to get his security clearance revoked at Los Alamos National Laboratory. He was successful in his testimony and Oppenheimer left Los Alamos for Princeton University. In 1963, Oppenheimer was presented with the Enrico Fermi Award. This accolade recognized international achievement in the field of energy. Only three years later, Oppenheimer died of throat cancer in Princeton, New Jersey.

Figure 6.7.1: Edward *Teller, the physicist, who invented the fusion bomb.* Image taken from: https://upload.wikimedia.org/wikiped...NL-cropped.png

Teller was not present for the testing of Mike. For personal and professional reasons, he left the south Pacific and head back to Berkley, California. At the university, he waited for Mike to affect the university's seismograph machine. This instrument detects earthquake activity due to natural or artificial causes. Before the bomb detonation, Teller had calculated the exact time the vibration would reach the University of California, Berkley.

The Mechanics of Fusion

The process of converting very light nuclei into heavier nuclei is also accompanied by the conversion of mass into large amounts

of energy, a process called fusion. The principal source of energy in the sun is a net fusion reaction in which four hydrogen nuclei fuse and produce one helium nucleus and two positrons. This is a net reaction of a more complicated series of events:

$$4\,{}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{1}_{0}\mathrm{n} \tag{6.7.1}$$

A helium nucleus has a mass that is 0.7% less than that of four hydrogen nuclei; this lost mass is converted into energy during the fusion. This reaction produces about 3.6×10^{11} kJ of energy per mole of ${}^{4}_{2}\mathrm{He}$ produced. This is somewhat larger than the energy produced by the nuclear fission of one mole of U-235 (1.8×10^{10} kJ), and over 3 million times larger than the energy produced by the (chemical) combustion of one mole of octane (5471 kJ).

It has been determined that the nuclei of the heavy isotopes of hydrogen, a deuteron, ${}^{2}_{1}H$ and a triton, ${}^{3}_{1}H$, undergo fusion at extremely high temperatures (thermonuclear fusion). They form a helium nucleus and a neutron:

$${}^{2}_{1}\mathrm{H} + {}^{3}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{1}_{0}\mathrm{n} \tag{6.7.2}$$

This change proceeds with a mass loss of 0.0188 amu, corresponding to the release of 1.69×10^{9} kilojoules per mole of ${}^{4}_{2}\mathrm{He}$ formed. The very high temperature is necessary to give the nuclei enough kinetic energy to overcome the very strong repulsive forces resulting from the positive charges on their nuclei so they can collide.

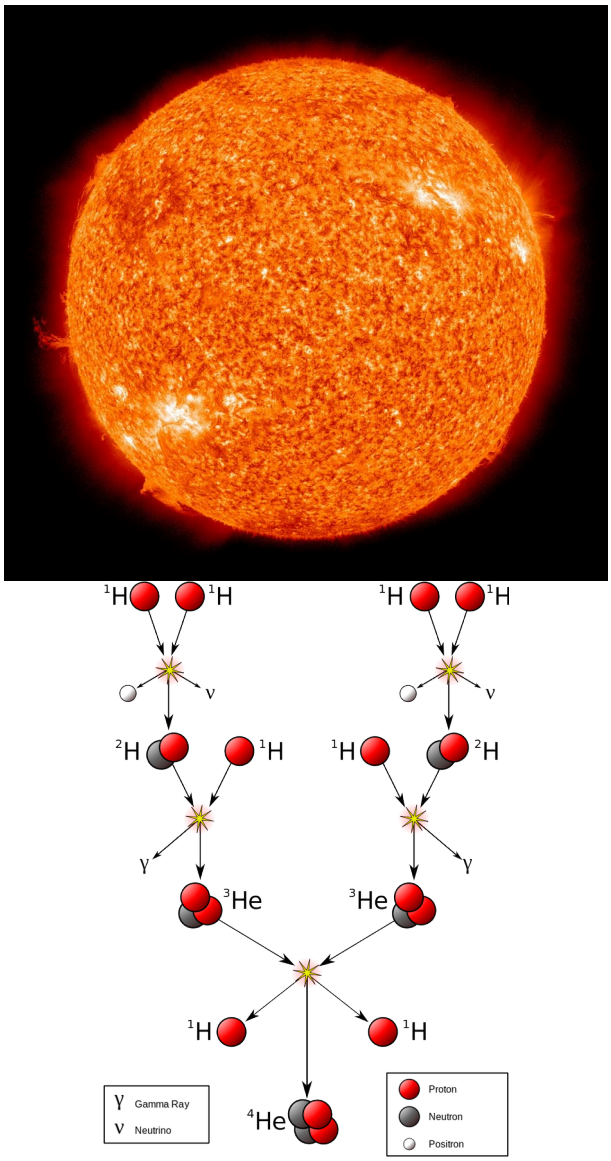


Figure 6.7.2: (left) The Sun is a main-sequence star, and thus generates its energy by nuclear fusion of hydrogen nuclei into helium. In its core, the Sun fuses 620 million metric tons of hydrogen each second. (right) The proton-proton chain dominates in stars the size of the Sun or smaller.

The most important fusion process in nature is the one that powers stars. In the 20th century, it was realized that the energy released from nuclear fusion reactions accounted for the longevity of the Sun and other stars as a source of heat and light. The fusion of nuclei in a star, starting from its initial hydrogen and helium abundance, provides that energy and synthesizes new nuclei as a byproduct of that fusion process. The prime energy producer in the Sun is the fusion of hydrogen to form helium, which occurs at a solar-core temperature of 14 million kelvin. The net result is the fusion of four protons into one alpha particle, with the release of two positrons, two neutrinos (which changes two of the protons into neutrons), and energy (Figure 6.7.2).

Overview of Ivy Mike's Mechanism

Fusion requires extreme temperatures. In order to achieve this, all fusion weapons contain a primary fission component. U-235 or Pu-239 are detonated first and the resulting heat fuses the deuterium and tritium components of a thermonuclear weapon.

Initially, the fission reaction of the primary releases gamma and x-ray, neutrons, and heat to produce the high temperatures required for following the fusion reactions. The fuel for the fusion is lithium deuteride, $^6_3\mathrm{Li}^2_1\mathrm{H}$. The neutrons react with lithium to produce tritium, $^3_1\mathrm{H}$, and alpha radiation.

$$^{6}_{3}\mathrm{Li} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{3}_{1}\mathrm{H} + {}^{4}_{2}\mathrm{He} \tag{6.7.3}$$

At the high temperature created by fission of the primary, the tritium then undergoes fusion with the deuterium in the lithium deuteride fuel.

$$^{3}_{1}\mathrm{H} + {}^{2}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{1}_{0}\mathrm{n} + \mathrm{energy} \tag{6.7.4}$$

The fusion reaction liberates a huge amount of energy and creates large numbers of high energy neutrons, which causes the additional fission of the uranium surrounding the secondary.

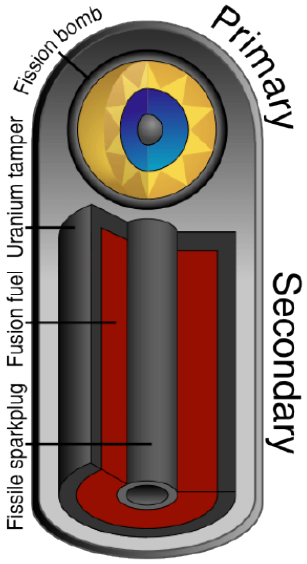


Figure 6.7.3: A Thermonuclear Weapon. The basics of the Teller–Ulam design for a thermonuclear weapon. Radiation from a primary fission bomb compresses a secondary section containing both fission and fusion fuel. The compressed secondary is heated from within by a second fission explosion. Image used with permission (Public Domain).

More recent fusion weapons utilize different mechanisms. Today's, warheads use a series of fission-fusion-fission chain reactions to produce yields several orders of magnitude greater than Little Boy or Fat Man. At this time, fusion weapons still require:

1. knowledge and construction of fission
2. isolation of H-2 and H-3 isotopes, and
3. extreme temperatures.

The five permanent members of the United Nations Security Council all possess the knowledge and the technology of fusion. These countries with their fusion device detonation dates are: United States (1952), the former Soviet Union (1953), the United Kingdom (1957), China (1967) and France (1968). Note that these dates refer to device detonations and not deliverable nuclear weapons. As of November 2017, it is still unclear if North Korea has detonated a fusion device.

In March of 1953, the United States dropped its largest fusion bomb over the Bikini Atoll in the Marshall Islands. Castle Bravo produced a punch equivalent to 15 megatons of TNT. The fallout isotopes affected over 11,000 square kilometers of area. Unfortunately, winds carried the lingering radiation to neighboring islands that were inhabited. Only 145 kilometers away from the epicenter, a Japanese fishing vessel (Lucky Dragon #5) was fishing with a crew of 24 people. Castle Bravo killed one person on the Lucky Dragon #5 and exposed the other crew members to extremely high dosages of radioactive isotopes (fallout).





Figure 6.7.4: (left) The *explosion of Castle Bravo in 1954. (right)* ***Figure*** 6.7.5*: The family of Kuboyama Aikichi, the Japanese fisherman who perished after Castle Braro, processing to his funeral. Images taken from: https://upload.wikimedia.org/wikiped..._Bravo_007.jpg and https://c2.staticflickr.com/8/7456/1...70054a7b_b.jpg*

Tsar Bomba-The Largest Nuclear Weapon

The former Soviet Union is credited with dropping the largest fusion bomb in history. On October 30, 1961, a fifty megaton three staged fusion bomb was dropped over the Arctic Circle. This weapon was calculated to be 1570 times more powerful than Little Boy and Fat Man combined. Tsar Bomba weighed 27 metric tons and was 26 feet long with a 7 foot diameter. The resulting fireball reached a maximum height of 1000 km (620 miles). The mushroom cloud produced was 64 km (40 miles) in height with a maximum width of 95 km (59 miles). Lastly, the flash from the fireball could be seen from 270 km (170 miles) away and burns were experience up to 100 km (62 miles) from epicenter.

Discovery Channel - Ultimates - Explosions - Tsar bomb segment

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 Development of Nuclear Weapons after World War II

The use of nuclear weapons by the United States to hasten the end of the war initiated a nuclear arms race between the United States and the Soviet Union that lasted until the end of the Cold War in 1992. During this period, scientists in both countries developed powerful thermonuclear fusion weapons (hydrogen bombs) and more efficient fission bombs. Table 6.7.2 compares the explosive power of current nuclear warheads. A nuclear weapon includes the warhead encased in a bomb or missile.

Table 6.7.1***:*** *Yields (Equivalent Tons of TNT) of Nuclear Warheads*

Terrorist weapon (improvised nuclear device)	< 3 kilotons
Bunker buster or tactical nuclear weapon	3 to 5 kilotons
Fission weapon (Pu-239 or U-235)	15 to 50 kilotons
Boosted fission weapon	100 to 500 kilotons
Thermonuclear weapon	1 to 57 megatons

Improvised nuclear devices (IND) are low yield weapons that might be assembled by non-state groups or “rogue” nations for a terrorist attack. The most likely scenario would be the acquisition of a critical mass of HEU and construction of a simple gun-type assembly mechanism.

Tactical nuclear warheads are placed in a variety of delivery systems such as artillery shells, torpedoes, cruise missiles, and bombs. These weapons may be defined by their range and types of targets. Tactical nuclear weapons are short-range weapons that are to be used on a battlefield or a a military target. On the other hand, strategic nuclear weapons are longer-range weapons that would disrupt the infrastructure of a country or a nation. Bunker busters are designed to deliver nuclear warheads to hardened, underground targets.

Nuclear weapons with yields greater than 5 kilotons of TNT are known as strategic weapons. They are designed to be used on targets such as missile launch sites, command and control centers, large cities, or industrial sites. The two types of early fission weapons have been discussed previously. Boosted warheads contain a small amount of deuterium and tritium (isotopes of hydrogen) gases, which undergo fusion with the initiation of the fission reaction.

$$ {}_{1}^{2}\mathrm{H} + {}_{1}^{3}\mathrm{H} \longrightarrow {}_{2}^{4}\mathrm{He} + {}_{1}^{1}\mathrm{n} + \text{energy} \tag{6.7.5} $$

The fusion releases an intense burst of high energy neutrons which amplifies the fission chain reaction, along with a small amount of energy This boosting allows increased yields from smaller, lighter warheads. Most current fission weapons contain boosted warheads. The 12 year half-life of tritium requires that these warheads be replenished at regular intervals.

Finally thermonuclear warheads, developed during the Cold War, use a series of fission-fusion-fission reactions to produce yields several orders of magnitude greater than those of fission devices . Weapons designers have also reduced the size of thermonuclear warheads so that multiple warheads could be carried by a single missile. Today, a ballistic missile submarine can carry 24 missiles each with 6 thermonuclear warheads.

A nuclear weapon can be modified from being tactical to strategic. In 1968, Los Alamos National Laboratory build the B61 thermonuclear bomb. This weapon was only twelve feet long and thirteen inches wide. The power range could be easily adjusted from .3 to 340 kilotons. B61 was a two stage thermonuclear weapon and over 3100 were constructed in the United States.



Figure 6.7.6: Technicians assembling a B61 nuclear weapon. Image take from: https://upload.wikimedia.org/wikiped...sGulf_B-61.jpg

Some countries possess ICBMs which are intercontinental ballistic missiles. These weapons can travel a minimum of 3400 miles (5500 kilometers). ICBMs can contain chemical or nuclear weapons. Noted countries that have or have had ICBMs are: Russia, United States, India, China, Israel, France, and North Korea. In 2017, China presented the longest reaching ICBM to the world. This weapon can reach a destination of over 6000 miles.

World Arsenals

According to the Arms Control Association October 2017's report, the countries contained in the table below possess nuclear warheads. This type of weapon would have a fission or fusion portion fastened to the front of a guided missile, rocket, or torpedo. In the world, there are approximately 15,000 nuclear warheads. Of this number, ninety percent are in military service while the remaining percent await dismantlement.

***Table** 6.7.2**:** Stats for countries that possess nuclear weapons.*

Country	Amount of Nuclear Warheads
Russia	7000
United States	6800
France	300
China	270
United Kingdom	215
Pakistan	140
India	130
Israel	80
North Korea	10

Template:HideTOC

At one time, former Soviet Union countries like Belarus, Kazakhstan, and the Ukraine stored nuclear weapons. Now, they do not possess any nuclear warheads. South Africa, Brazil, Iraq, Libya, Argentina, South Korea, and Taiwan have abandoned their nuclear weapon programs as well.

In August of 1949, the American monopoly on nuclear weapons technology ended. The USSR ignited a 20 kiloton Pu-239 weapon that they named First Lightning. American nuclear scientists code named this detonation to be Joe-1. The weapon was duplicate of the Trinity device that was tested by the Manhattan Project scientists in July of 1945. A spy ring of scientists and civilians passed implosion mechanics and hydrogen bomb technology to KGB agents. The most notable of these people were Klaus Fuchs

and Julius and Ethel Rosenberg. During the Manhattan Project, Fuchs worked at the gaseous diffusion facility (University of Columbia) and Pu-239 implosion laboratory (Los Alamos). In addition, Fuchs assisted Edward Teller with hydrogen bomb research. Fuchs passed knowledge of implosion and fusion directly to KGB agents and through civilians, Ethel and Julius Rosenberg.



Figure \({6{.}7{.}\eatSpaces{7\): *Klaus Fuchs and* ***Figure*** Extra close brace or missing open brace *: Ethel and Julius Rosenberg*
Images taken from: https://upload.wikimedia.org/wikiped...A_-_278757.jpg and https://upload.wikimedia.org/wikiped...berg_NYWTS.jpg

The USSR had started fission research in the mid to late 1930's. They primarily focused on uranium enrichment and determination of critical mass. After the bombings of Hiroshima and Nagasaki, Stalin pushed USSR scientists harder to obtain nuclear weapons technology. During the Manhattan Project, Stalin was aware that the United States was pursuing the development of fission weapons. Espionage provided USSR scientists with blueprints of the Trinity device and information regarding fusion. This data expedited the USSR to detonate their first fission device in 1949. Less than a year after the American testing of Mike (first fusion device), USSR exploded a 400 kiloton, H-bomb device named Joe-4 . Hans Bethe, an American Manhattan Project scientists, debated the nature of this explosion. Fragments and resulting isotopes did not indicate that a fusion device had exploded. In 1955, the USSR dropped their first fusion weapon which yielded energy within the megaton range.

Individuals who participated in the espionage of nuclear bomb secrets received different types of punishments. Fuchs, a British citizen who had immigrated from Nazi Germany, served nine years in prison and then was released to East Germany. The Rosenburgs were executed at Sing Sing Prison in New York State. Ethel Rosenberg's brother, David Greenglass, had testified against the couple in order to save his own wife from prison. Greenglass, an army sergeant at Los Alamos, had obtained bomb secrets and had his wife type documents for couriers. For his participation, Greenglass served a decade in prison. His wife was

not tried or punished for her involvement. To read more about this fascinating event in history, click on this link.

Since the late 1960's, the United States and the former Soviet Union/Russia have signed several treaties to downside their nuclear arsenals. The START (Strategic Nuclear Arms Control Agreements) program has greatly reduced the numbers of nuclear weapons for both countries. For more information regarding START, click on this link.

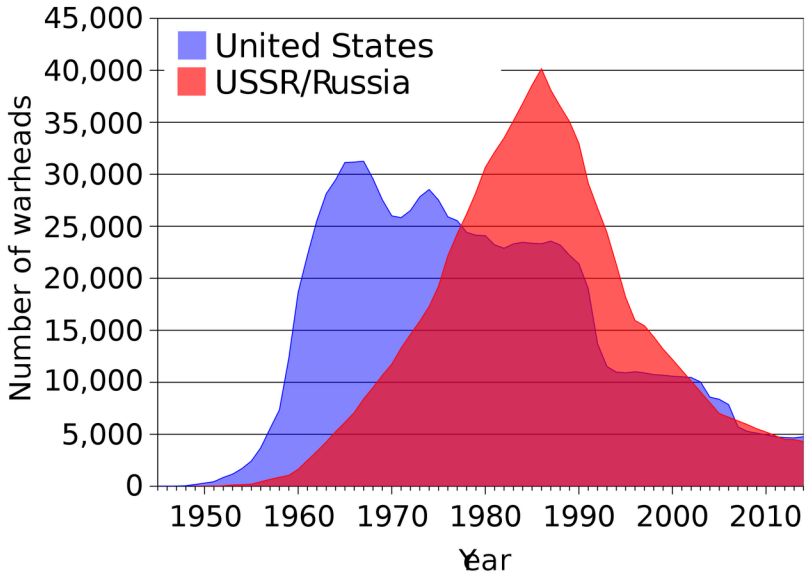


***Figure** 6.7.8:Nuclear Arsenals of the United States and USSR/Russia over the past sixty years. Image taken from: https://upload.wikimedia.org/wikiped...kpiles.svg.png*

Requirements for Producing Nuclear Weapons

A nation seeking to produce nuclear weapons must complete the following basic steps:

- Develop a weapon design or obtain one from an external source.
- Produce highly enriched uranium or plutonium for the core of the device or obtain this material from an external source.
- Fabricate this material into the fissile component of the weapon.
- Fabricate or obtain from external sources the non-nuclear components of the weapon. These include high explosives and a triggering mechanism to detonate the nuclear core.
- Verify the reliability of all components individually and as a system.
- Assemble the components into a deliverable weapon.
- For fusion weapons, a through knowledge of fission is required. A fusion bomb must contain a fission counterpart.

Nuclear Terrorism

Most recently, attention has centered on the possibility of “rogue” nations or terrorist groups obtaining a fission device. Two possibilities exist: the acquisition of an intact weapon or the construction of an improvised nuclear device (IND) device from components either stolen or acquired illegally from a rogue state. In the first case, an organized group must obtain the financial resources to initiate this act of extreme violence. The group must then acquire an intact weapon through purchase, theft, diversion, or as a gift. It then has to ascertain how to bypass or disable safeguards incorporated in the weapon to prevent its unauthorized use. Lastly, the weapon must be transported to a high value target and detonated. The challenges involved with the successful completion of each step are formidable.

The second case would also involve sizeable obstacles. As in the first case, an organization and funding would be required. The group would have to acquire sufficient fissile material, fabricate it into a fissile component, and assemble an IND. This would require significant technical expertise. A device similar to a uranium gun-type bomb (Figure 2) would be easier to build than a plutonium implosion weapon. Transport and detonation of an IND would involve greater risks than those associated with the

intact weapon.

The lack of complete control of fissile materials and the economic conditions in the former Soviet Union have increased the possibility that HEU and plutonium from weapons programs might be available on the international black market. Several instances involving plutonium and enriched uranium smuggled from the former Soviet Union have been observed since 1994. The United States and Russia have had a “blend down” agreement since 1993 that renders excess HEU incapable of being fabricated into nuclear weapons. Under the terms of this agreement, the United States purchases from Russia uranium of which the concentration of U-235 has been reduced by mixing it with U-238 to that suitable for fueling nuclear power reactors.

Other nations also have quantities of HEU and plutonium from their weapon programs and the production of civilian nuclear power. Figures ***Figure*** 6.7.9 and ***Figure*** 6.7.10 show the global disposition and current status of HEU and plutonium.

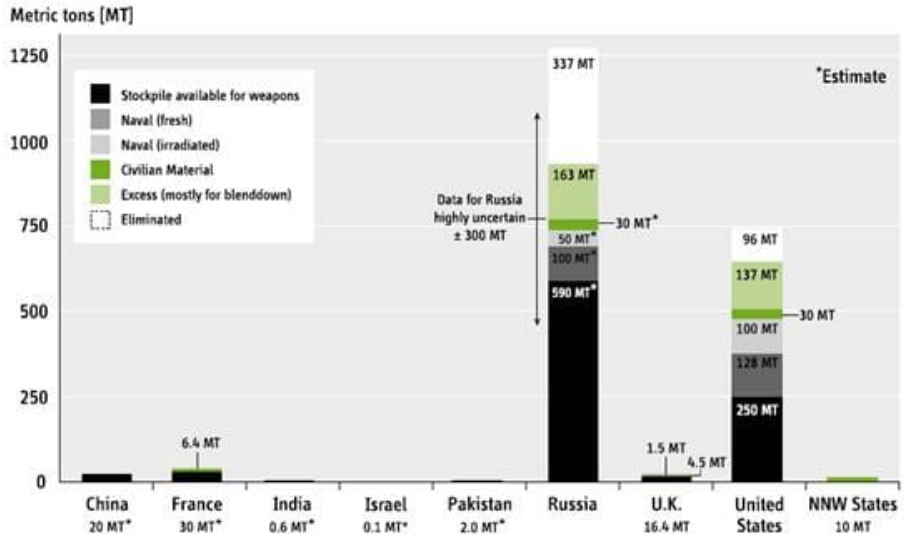


Figure 6.7.9: - Global Stockpiles of HEU in 2008. From the Global Fissile Material Report 2008, p. 11

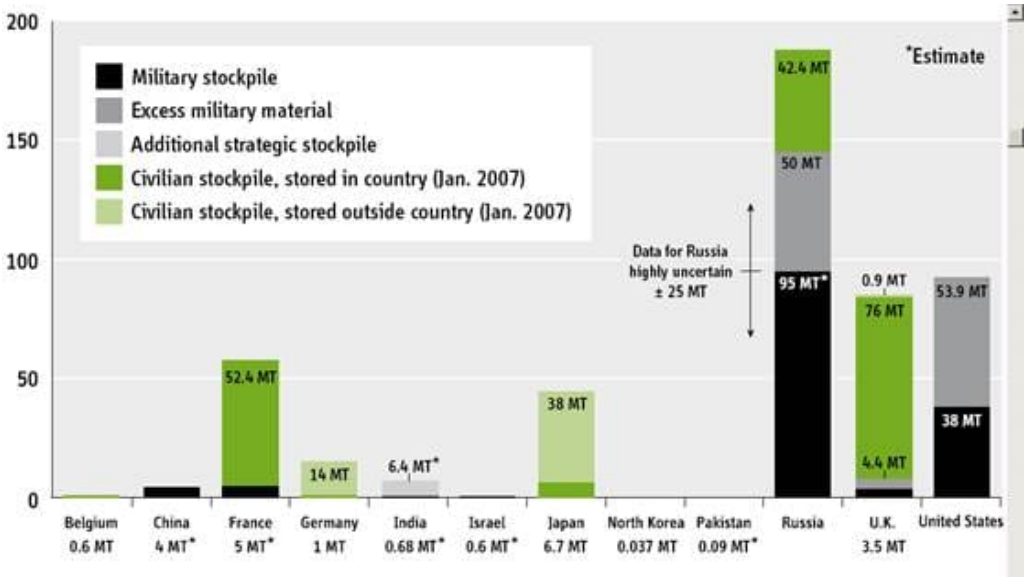


Figure 6.7.10: - Global Stockpiles of Plutonium in 2007. From the Global Fissile Material Report 2008, p. 16

References

1. http://www.atomicarchive.com/History.../page_13.shtml
2. https://www.ctbto.org/specials/testi...-castle-bravo/

3. https://www.armscontrol.org/factshee...ponswhohaswhat

Contributors

- Frank A. Settle (Washington and Lee University)

This page titled 6.7: Fusion is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.8: Fallout and Weapons Testing

Fallout

Fallout refers to the radioactive material that "falls out" of the atmosphere after a nuclear explosion (reactor or bomb). It consists of dust and radioactive particles that can contaminate an area with radioactivity and pose a huge health hazard to biological organisms. It can contaminate the animal food chain which can have drastic effects on the affected region. Weather has a huge impact on fallout, wind currents can spread radioactive fallout either over a large area, such as in the case of *Castle Bravo*, or not.

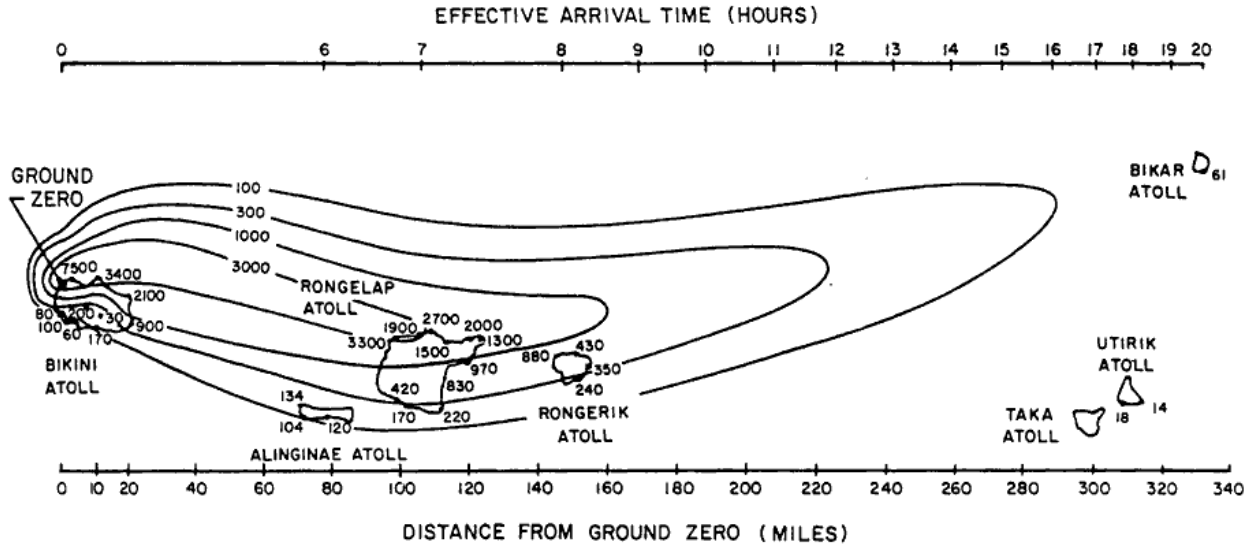


Figure 6.8.1:The 450 km (280 mi) fallout plume from 15 Mt shot Castle Bravo, 1954. Image taken from: https://upload.wikimedia.org/wikiped...o_fallout2.png

In regards to nuclear reactions, fission produces more fallout products that fusion. Intense heat of the latter type of reactions reduces the amount and type of radioisotopes produced. Although more powerful, fusion is a "cleaner" type of nuclear reaction. By increasing the fission component of a fusion weapon, bombs can be designed to generate more fallout products.Of the different types of isotopes that fission produces, the most worrisome are **Sr-90**, **Cs-137**,and **I-131**. The first of these isotopes has a half-life of 28.8 years and decays by gamma and beta emission. Once ingested or inhaled, Sr-90 deposits in teeth, bone,and bone marrow. Sr-90 exposure can lead to bone cancer and/or leukemia. Cs-137 ($t_{1/2}$ = 30.17 years) emits beta and gamma radiation. This ionizing radiation deposits in soft tissues of the body. In addition, this particular radioisotope is water soluble and can easily contaminant food and water supplies. I-131 had a half-life over 8.1 days and undergoes gamma/beta decay. Once this fallout product enters the body, it attacks the thyroid gland. Extensive damage could affect heart rate, blood pressure, body temperature, and childhood growth. In addition, I-131 exposure could result in thyroid cancer.

Example 6.8.1:

Watch the fallout video 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. Describe the physical appearance of fallout.
2. How does fallout spread?
3. What was the baby tooth study?
4. What particular isotopes does the speaker mention in this video?
5. What is the 7-10 rule?
6. What mistake does the speaker say when he is discussing different isotopes?
7. Why do Americans who lived in country from 1945-1963 have more I-131 in their bodies?

Antidotes

Today, there are medications that can be dispensed to help with fallout exposure. On the (Center for Disease Control)'s link, there are four different substances described. Research one of these substances by evaluating how they protect the body and what particular isotope the antidote can counteract. In addition, be aware of one group of individuals who cannot take this medication. **This information is required information for your next test.**



Figure 6.8.2: Dutch Artist, Vincent Van Gogh (1853-1890), used Prussian *Blue as one the signature blues in his Starry Night (1889). Prussian Blue is also uses as an antidote radioactive thallium and cesium* Image taken from: https://upload.wikimedia.org/wikiped...arry_Night.jpg

Limiting Nuclear Weapon Testing

On August 5, 1963, the leaders of the United States (President John F. Kennedy), the Soviet Union, and the United Kingdom signed the Nuclear Test Ban Treaty. This document would limit where nuclear bombs/devices could be tested. These three countries would no longer perform nuclear trials under water, in the atmosphere, or in outer space. This treaty would allow countries to remain testing underground with certain limits. Also, these countries would attempt to move towards disarmament. Reducing nuclear weapons and testing would allow the environment to recover from decades of contamination.



Figure 6.8.3: President John F. Kennedy signing the Nuclear Test Ban Treaty in 1963. Image taken from: https://upload.wikimedia.org/wikiped...tober_1963.jpg

During the early ninety's, President George H.W. Bush announced that the United States would no longer perform full scale nuclear tests. Four years later, the the United Nations General Assembly adopted the Comprehensive Test Ban Treaty. This agreement (known as the CTBT) would halt nuclear bomb/device testing entirely. To go in effect, the United Nations required forty-four of the current nuclear states (those with nuclear capabilities) sign and ratify the CTBT. In 1996, the United States (President Bill Clinton) quickly signed the CTBT. However, the United State Senate refused to provide support and the United States has not ratified it. There are several countries that have not signed and refused to ratify the CTBT. These countries include North Korea, India, and Pakistan. Others have signed but not ratified the treaty: China, Israel, Iran, Egypt, and the United States. For a comprehensive list of countries, click on this link.

From 1945 to 1992, the United States has tested approximately 1,054 nuclear bombs and devices. Of this amount, 216 of these tests were performed in the atmosphere, under water, and in outer space. The majority of these tests detonated in the Nevada Test Site, New Mexico Test Site, and the Marshall Islands (South Pacific).

Other countries have done significant nuclear testing as well. Looking at the figure below, once can visualize how many nuclear bombs/devices have been detonated over a fifty year period. This link shows all (not just the United States') nuclear test sites.

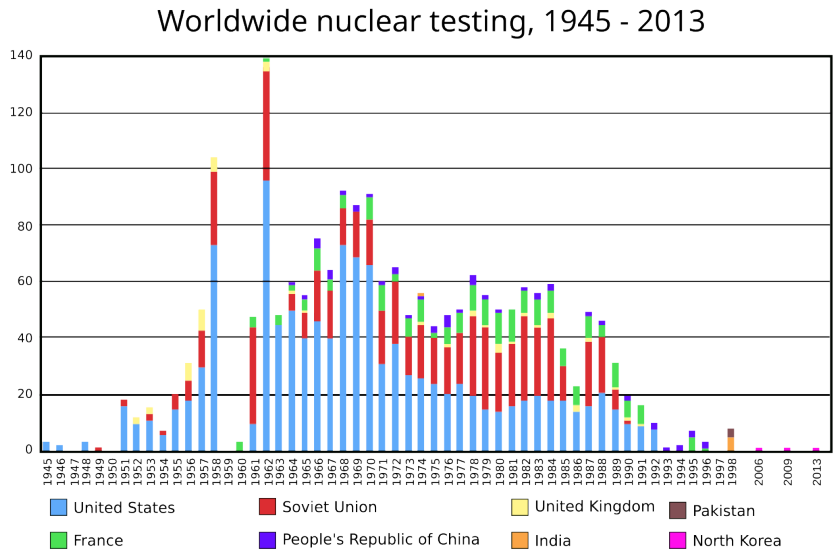


Figure 6.8.4: Image taken from: https://upload.wikimedia.org/wikiped...esting.svg.png

Contributors

- Frank A. Settle (Washington and Lee University)

Complete Bibliography on Radiation from the Alsos Digital Library for Nuclear Issues

This page titled 6.8: Fallout and Weapons Testing is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.9: Additional Readings and Media Fun

Books

Pulitzer Prize winner, Richard Rhodes, has written many nonfiction accounts of the development of nuclear weapons. Rhodes, an American journalist and historian, published *The Making of the Atomic Bomb* (1986) and *Dark Sun* (1996). The first of these books describes accounts of Manhattan Project employees, technical aspects of Little Boy and Fat Man, testing of the Trinity device, and the dropping of the weapons over Japan. He goes into great detail in regards to the construction of the atomic weapons. In his second installment, Rhodes chronicles the development of Edward Teller's fusion bomb. He discusses the trial of J. Robert Oppenheimer, who opposed the construction of new nuclear weapons. With court testimony and images, Rhodes show how Teller and Oppenheimer verbally opposed one another. Lastly, Dark Sun provides evidence as to the USSR espionage during the Manhattan Project. Stolen data and schematics provided information for a Soviet Union Pu-239 fission bomb (1949) and Hydrogen device (1953).

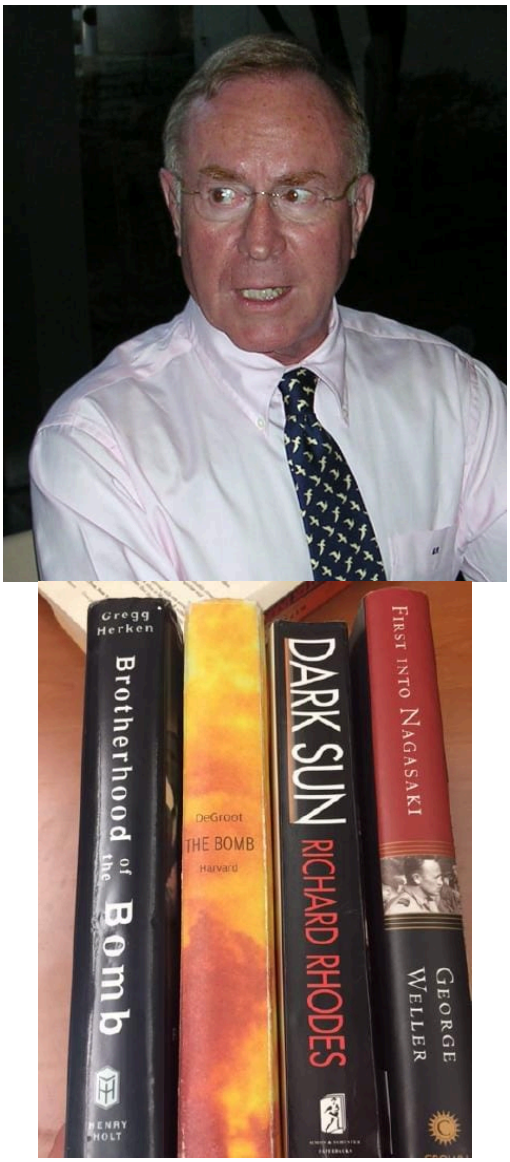


Figure 6.9.1 (left) Richard Rhodes. (right) Technical based books describing the Manhattan or fusion Projects. Image taken from: https://upload.wikimedia.org/wikiped...ard_Rhodes.jpg and the author's private library.

Other authors have recorded atomic bomb history. Greg Herken's *Brotherhood of the Bomb* focuses on the men of the Manhattan Project. Personal and academic lives of these individuals are explored. Candid and serious pictures of Oppenheimer, Lawrence, Fermi and even Teller are displayed in the pages of this book. In addition, Herken includes the blueprints of Little Boy and Fat Man.

The Bomb by Degroot is less technical read that starts with the Manhattan Project and extends throughout the Cold War. There are several interesting songs, poems, pictures, and accounts by scientists and people who lived during this era. Degroot goes on to include the Cold War and how nuclear weapons have affected a person's day to day living.

George Weller's *First into Nagasaki* is one of the few accounts of individuals who witnessed what happened after Fat Man was dropped. Weller was a American novelist, playwright, and war correspondent who illegally entered Nagasaki after the bomb was dropped. While in Nagasaki, Weller was able to meet and interview several members of Japanese POW camps. He was able to record their experiences and stories of survival. Weller never published his work while he was alive. Upon his death, his son discovered these written chapters. Weller's son edited and add a third section to this book. *First Into Nagasaki* was finally published in 2006.

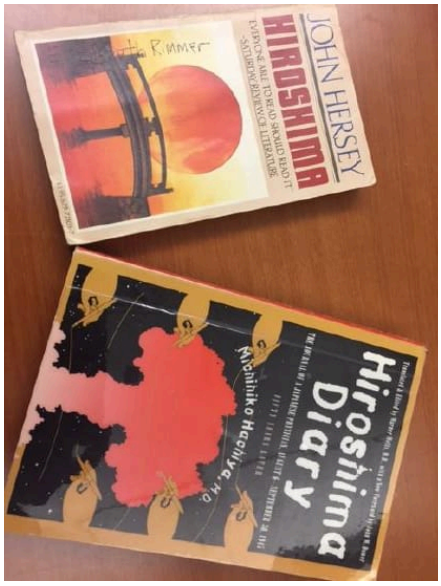


Figure 6.9.2: Diaries accounting the dropping of Little Boy

Two books that read more like diaries are shown above. Hershey's Pulitizer Prize winning account follows the lives of six Hiroshima residents before and up to one year after the dropping of Little Boy. Unlike Weller, Hershey was allowed to publish his experiences as a World War II war correspondent. Michihiko Hachiya, a Japanese medical practitioner, recorded his observations of Hiroshima from August 6th through September 30, 1945. He explains the state of chaos involved in managing a nonfunctioning hospital. With humor and grace, Dr. Hachiya describes his circumstances of surviving f nuclear bomb detonation and fallout

A brief overview of German atomic bomb efforts is described in "Hitler's Scientists." Although the Germans first explained fission, they focused money and research effort on rockets. After the German surrender in May 1945, their scientists were contained and monitored in a farmhouse. Allied forces had hoped to learn secrets of their atomic bomb efforts and capture their reactions to the American bombings of Japan. This investigation was part of the Aslos Project that led allied forces to determine the Germans had not gotten very far with their nuclear endeavors

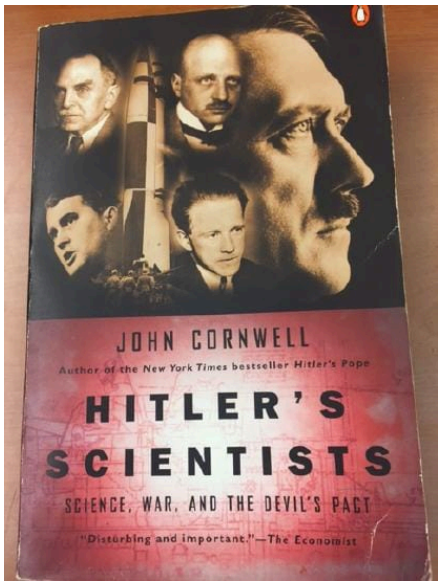


Figure 6.9.3: German Bomb Effort

Several individuals aboard the Enola Gay and Bocks Car have written eyewitness accounts of the dropping of Little Boy or Fat Man. Both the pilot and the navigator of the Enola Gay have published books about these days. Charles Sweeney , pilot of the Bocks car, and Abe Spitzer, radio man on the Great Artiste, also have their recollections published as well.

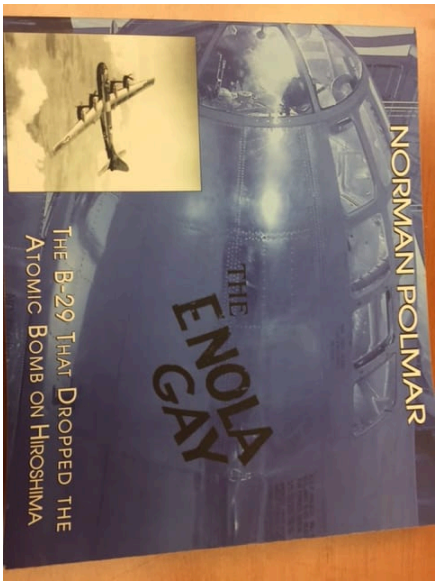


Figure 6.9.5: Enola Gay

The planes that carried Little Boy and Fat Man have been preserved are are on permanent display. Enola Gay is housed at the Smithsonian National Air and Space Museum in Chantilly, Virginia. The Bockscar is located at the National Museum of the Air Force in Dayton, Ohio. The other planes that flew on both missions were scrapped for their metal and are not longer in existence.

Movies

Example 6.9.1: godzilla

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 first three minutes of this video to get an idea of the origin of Godzilla.

1. What year was the first Godzilla movie produced?
2. What country is created the Godzilla character?
3. What American character rivals Godzilla?
4. What two types of animals/creatures compose Godzilla?
5. What does Godzilla represent to the Japanese people? Describe some of his powers.

By the late 1940's, the USSR had constructed a detonated its first nuclear device. In response to this event, the United States Civil Defense Ministry recorded the Duck and Cover. This 1951 film was to be shown in American classroom in order to prepare school children on how to respond to a nuclear attack. It provided kids and adults with tips on how to avoid immediate radiation and lingering fallout after an attack.

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

Over the years, the movie industry has included nuclear weapons in quite a few of their action films. For example,many of the James Bond or 007 movies incorporate the stealing of nuclear material or weapons. Movies made during the Cold War usually involved nuclear espionage or accidental detonations. Nuclear exposure has also been used to explain why zombies exist. Several horror movies focus on survivors of nuclear fallout that morph into altered beings due to radiation.

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

Atomic Music

Songs has been written to highlight atomic themes as well. Many of these became popular during the Cold War. Click on a few of these to determine the atomic component.

- Threnody to the Victims of Hiroshima (Penderecki): https://www.youtube.com/watch?v=Pu371CDZ0ws
- Radioactive (Imagine Dragons) https://www.youtube.com/watch?v=ktvTqknDobU
- Political Science (Randy Newman) https://www.youtube.com/watch?v=Kg_LDeUEiWY
- Russians (Sting) https://www.youtube.com/watch?v=wHylQRVN2Qs
- Ninety Nine Red Balloons (Nena): https://www.youtube.com/watch?v=La4Dcd1aUcE
- Manhattan Project (Rush): https://www.youtube.com/watch?v=ktvTqknDobU
- Melt With You (Modern English): https://www.youtube.com/watch?v=LuN6gs0AJls
- We Didn't Start the Fire (Billy Joel):
- Breathe (Kate Bush):
- Two Minutes to Midnight (Iron Maiden):

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 6.9: Additional Readings and Media Fun is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

6.E: Nuclear Weapons- Fission and Fusion (Exercises)

1) Translate the following artificial transmutation reactions into nuclear symbols and balance:

a. Be-9 is hit with an alpha particle to produce C-12 and ?
b. Li-6 is bombarded with deuterium to form ? and Li-7
c. A neutron hits N-14 to produce an alpha particle and ?
d. Na-23 combines a neutron to form Na-24 and ?

2) What is the difference between emission/decay and transmutation?

3) Which of the following instruments can artificially transmute elements: cyclotron, electromagnetic separation, gaseous diffusion, and/or linear accelerator?

4) Which country split the atom first?

5) What is the significance of $E = mc^2$?

6) Review the Manhattan questions from the Modern Marvel movie.

7) Compare/contrast the bombings of Hiroshima and Nagasaki.

8) Did the same plane drop both of the bombs over Japan? Did the crew of the airplane(s) perish during the bombing?

9) Review basic bomb drawings illustrated on 6.4.3

10) Review questions from video in section 6.6.1 (Exploring 24 hours after Hiroshima).

11) Who was the father of the atomic bomb? Who was the father of the H-bomb?

12) What are the basic materials that are used in a fusion weapon? Who dropped the largest H-bomb? Was this bomb a gadget/device?

13) What is the destruction tonnage range of a fission bomb versus a thermonuclear bomb?

14) What is the difference between a tactical and strategic nuclear weapon? Was the Trinity bomb a strategic or tactical weapon?

15) What is an ICBM? Do all ICBM's contain a nuclear component?

16) Name the countries that have fusion technology.

17) At this time, which country has the most nuclear warheads? Does Japan and Germany possess any nuclear weapons?

18) What types of nuclear secrets were passed by the Rosenbergs and Fuchs?

19) What is START and who does this treaty involve?

20) How does a blend down of HEU reduce the chances of a rogue nation producting IND?

21) Which type of weapon produces more fallout: fission bomb or a H-bomb?

22) Name the three fallout isotopes and how they impact the human body.

23) Review the questions regarding the fallout video in section 6.8.

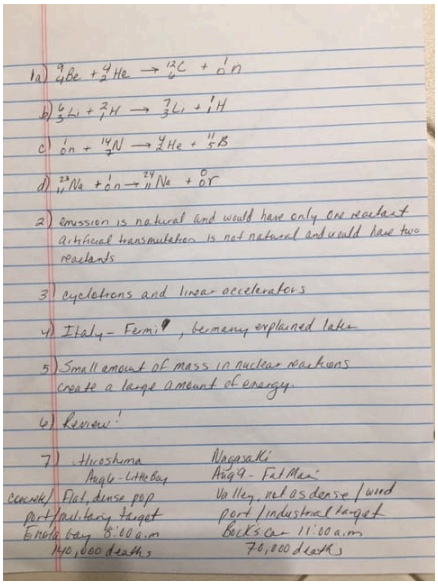
24) Know one antidote to radiation poisoning (see highlighted region in 6.8)

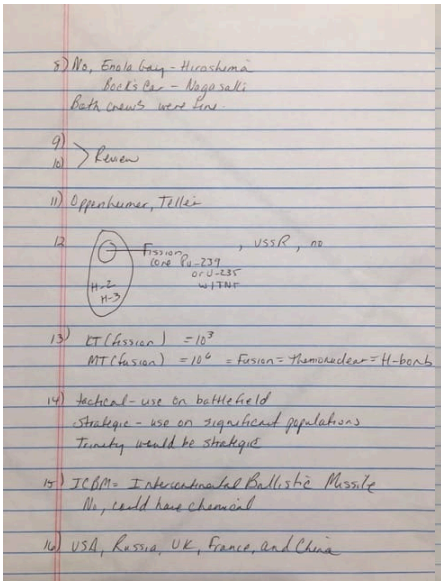
25) What is the difference between the Nuclear Test Ban Treaty (1963) and Comprehensive Test Ban Treaty (1996).

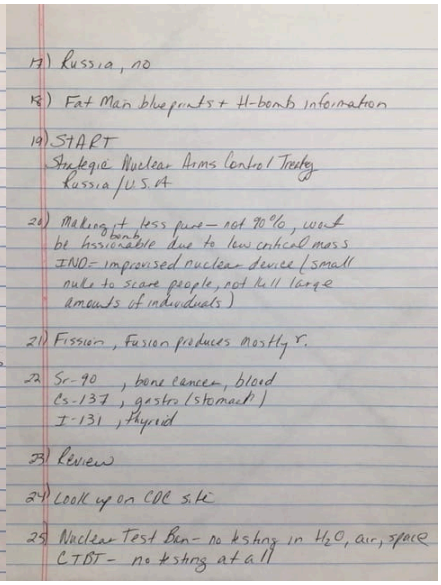
26) Name some of the countries that have not ratified the CTBT. Why do you think they have not done this?

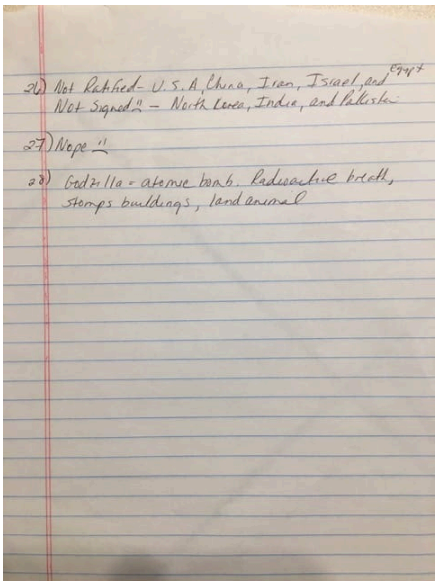
27) Does Duck and Cover help to protect you in any way from nuclear annihilation?

28) What does Godzilla represent in Japanese culture?









This page titled 6.E: Nuclear Weapons- Fission and Fusion (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7: NUCLEAR ENERGY

Even before the atomic bomb had been produced, scientists and engineers had begun to think about the possibility of using the energy released by the fission process for the production of electrical energy. In the United States only 19.7 percent of the electrical energy in 2016 was produced by this nuclear energy. The proportion is a higher in some other countries, notably France and Slovaka. Unfortunately, producing power from atomic fission has turned out to be much more expensive than was previously expected. Even in these days of high prices for the fossil fuels it is still only barely competitive.

7.1: Sharing and Monitoring Nuclear Technology

7.2: Generating Electricity

7.3: Powering the United States

7.4: Generating Electricity by Using Fission

7.5: United States Commercial Nuclear Power Reactors

7.6: South Carolina Energy

7.7: Three Mile Island Nuclear Accident

7.8: Chernobyl Nuclear Disaster

7.9: Fukushima Nuclear Disaster

7.10: Storing and Transporting Nuclear Waste Materials

7.11: New Generation Nuclear Reactors

7.E: Nuclear Energy (Exercises)

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 7: Nuclear Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.1: Sharing and Monitoring Nuclear Technology

Atoms For Peace Speech

On December 8, 1953, President Dwight D. Eisenhower addressed the United Nations with a twenty minute speech entitled, " Atoms for Peace." In this event, Eisenhower attempted to diminish the nuclear arms race between the United States and the Soviet Union. Eisenhower believed that peaceful uses of nuclear technology should be available to all countries. Moreover, he encouraged non-nuclear states (countries) to obtain information regarding advances in nuclear medicine and nuclear reactors. The United States and Soviet Union would no longer have a monopoly on nuclear applications. President Eisenhower wanted a central agency that would collect, store, and distribute radioactive materials to all countries. This "uranium bank" would allow non-nuclear states to obtain fissionable material for medicine and energy. Watch the video below to gain more insight into the "Atoms for Peace"speech.

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

IAEA

In response to Eisenhower's speech, the United Nations established the IAEA (International Atomic Energy Agency) in 1957. This group serves as the world's nuclear watchdog and has three goals to oversee:

- Assist all member countries with peaceful uses of nuclear technology
- Inspect nuclear facilities to ensure weapons are not being constructed
- Provide information regarding safety of nuclear materials.

The IAEA is housed in Vienna, Austria. It has two regional offices in Toronto, Canada and Tokyo, Japan. There are liaison offices for this organization in New York City and Geneva, Switzerland. Research facilities for the IAEA are located in Austria and Monaco. As of November 2017, there were 168 member countries of the IAEA. A listing of these countries with their dates of membership can be found by clicking on this link. On this list, North Korea has been removed from membership. This can be attributed to their lack of cooperation regarding reactor inspections.



Image taken from: https://upload.wikimedia.org/wikiped...AEA_emblem.jpg

NPT

The IAEA has also been charged with enforcing the NPT (Nuclear **Non**-Proliferation Treaty). Entered into force in March of 1970, this involves nuclear and non-nuclear states. Nuclear states are the five countries that exploded a nuclear device before 1967. This group includes the United States, Russia, United Kingdom, France, and China. All other countries besides these are labeled as being non-nuclear states. The five nuclear states and the IAEA are responsible for enforcing the NPT. This treaty expects the following to occur:

- States that do not have nuclear weapons will not build or buy them.
- States that possess nuclear weapons will work towards disarmament.
- All states will be provided information regarding peaceful applications of nuclear technology.

As of November 2017, there were over 190 countries who had signed the NPT. This link shows countries and their NPT statues. Many countries have ratified or consent to abide by this document. Others have assessed the treaty which means they have accepted NPT terms after the treaty has been entered into force. Lastly, some countries have succeeded or revoked membership from the NPT. For a basic overview of the NPT, click on 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

References

- https://www.chemheritage.org/distill...nuclear-gambit
- https://www.iaea.org/about/overview/history

Contributors

- Elizabeth R. Gordon (Furman University)
- Caralyn Zehnder et al. at Georgia College and State University

This page titled 7.1: Sharing and Monitoring Nuclear Technology is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.2: Generating Electricity

Learning Objectives

- Know the basic governmental agencies that have and now regulate with nuclear facilities and materials.
- Understand the basics of how electrical energy is produced.
- Calculate the cost of running an appliance for a period of time (typically for a month).
- Have knowledge of "vampire" devices.

AEC/NRC History

During the Manhattan Project, Enrico Fermi and Leo Szliard constructed the world's first nuclear reactor in the basement of the University of Chicago. It consisted of layers of graphite (carbon) embedded with pieces of uranium. The sustained chain reaction was measured via a Geiger counter. After World War II, Manhattan Project scientists knew that fission could be used for power as well. In 1946, Harry S. Truman signed the McMahon/Atomic Energy Act. This transferred atomic energy control from the military to civilian power. The Atomic Energy Commission (AEC) was established to develop nuclear energy for peaceful purposes. This group would study how fission could be used to power homes, industries, submarines, and battleships. In addition, the AEC would investigate how nuclear isotopes could be used in the fields of medicine and agriculture.

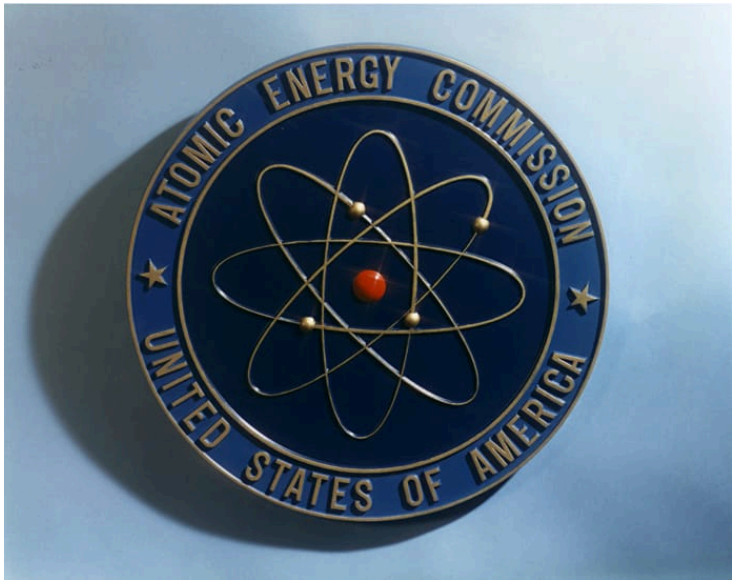


Figure 7.2.1: Image taken from: https://upload.wikimedia.org/wikiped...ssion_logo.jpg

The AEC constructed the first experimental breeder nuclear (this type of equipment uses uranium to produce weapons grade plutonium) reactor in Arco, Idaho. On December 20, 1951, the reactor went online and produced enough energy to power four light bulbs. In 1957, the first full scale American commercial nuclear reactor was completed in Shippingport, Pennsylvania. Unit One reactor was operated by Duquesne Light Company and cost approximately 73 million dollars. The electrical power this reactor produced varied from 60 to 100 MW. Over its twenty-five year lifetime, this reactor produced over 7.4 billion kilowatt hours of energy. In 1989, Shippingport Reactor 1 was shut down and decommissioned.

Figure 7.2.2: *Installation of Shippingport Reactor 1 in Pennsylvania. Image is taken from https://upload.wikimedia.org/wikiped...C_135430pu.jpg*

From its conception, the AEC appeared to take a greater interest in nuclear weapons security than nuclear reactor safety. For this reason, Congress dismantled the AEC in 1974. The existing Department of Energy (DOE) would focus on nuclear security (weapons and materials) and nonproliferation (stopping construction of nuclear weapons). The newly formed Nuclear Regulatory Commission (NRC) would concentrate on ensuring the safe use of radioactive materials. In other words, radioisotopes used in

medical, agricultural, and industrial applications would be watched to minimize contamination of people and the environment. Additionally, NRC would regulate all commercial reactors. These energy sources would be licensed and inspected for safety purposes. If a reactor failed inspection, then the NRC had the power to enforce fines and penalties on that particular reactor.

General Energy Production

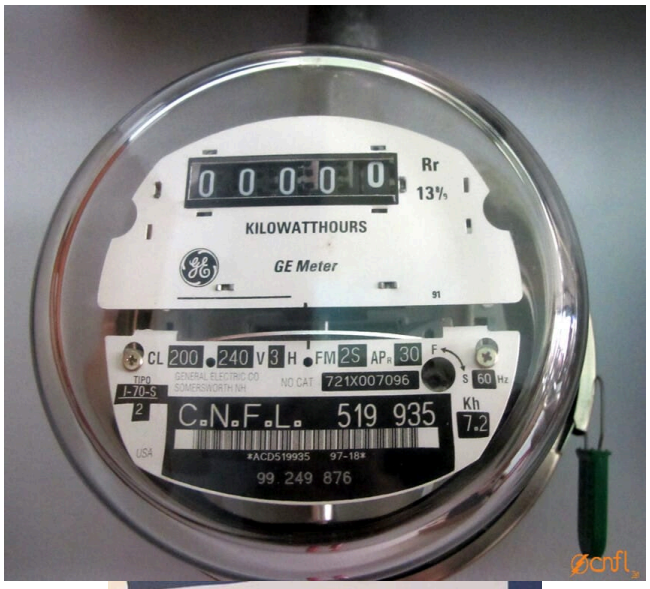
The great majority of all electrical generating systems (whether coal-burning power plants, hydroelectric plants, or nuclear power plants) all follow a reasonably simple design. The electricity is produced by spinning a coil of wire inside a magnetic field. When a fluid (air, steam, water) is forced through the pipe, it spins the fan blades which in turn spin the axle. To generate electricity, the axle of a turbine is attached to the loop of wire in a generator. When a fluid is forced through the turbine, the fan blades turn, the turbine axle turns, and the loop of wire inside the generator turns, thus producing electricity.

Figure 7.2.3: Image taken from: https://upload.wikimedia.org/wikiped...iagram.svg.png

The essential difference in various kinds of electrical generating systems is the method used to spin the turbine. For a wind generator, the turbine is a windmill. In a geothermal generator, steam from a geyser is forced through the turbine. In hydroelectric generating plants, water falling over a dam passes through the turbine and spins it. In fossil fuel (coal, oil, natural gas) generating plants, the fossil fuel is burned and the heat is used to boil water into steam and then the steam passes through the turbine and makes it spin. In a fission reactor generating plant, a fission reaction is used to boil the water into steam and the steam passes through the turbine to make it spin.

Basic Energy Units

When understanding electrical energy consumption, the kilowatt (kW) is the most commonly used unit in the United States. This unit is a measure of power. A more specific unit, the kilowatt-hour (kWh) is utilized to denote how much fuel is used to generate electrical energy over a specific amount of time. Power meters are installed in consumer and industrial facilities to measure kilowatt-hour usage. Power companies charge various rates around the country for electrical consumption. The pictures below show a traditional power meter and a typical residential power bill. Watch this video to get an understanding of the kilowatt hour and how different appliance use different amounts of kilowatts.



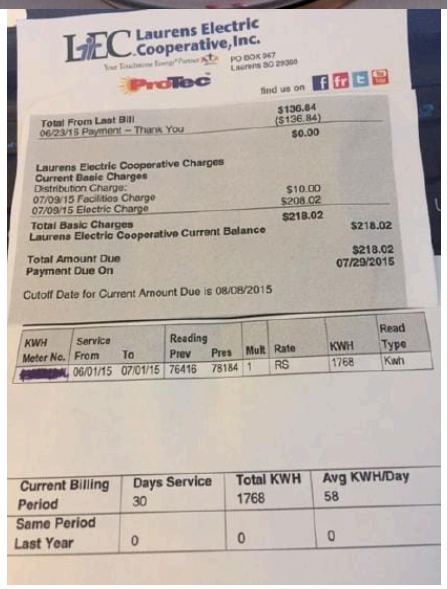


Figure 7.2.4: Meter image taken from: https://c1.staticflickr.com/7/6013/5...116c6fa6_b.jpg and (right) A power bill showing kilowatt hours.

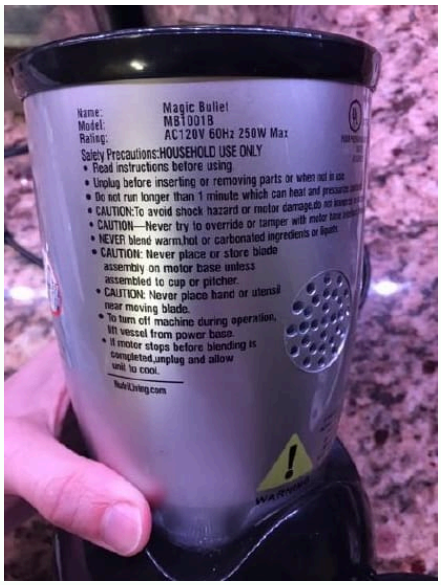
Wattages for common household appliances can be obtained by looking at labels or using the manufacturer's website. Typically, power units for these devices are listed in watts. Calculating a price for the monthly use can be done by following the rules listed in the next box.

Be aware of "vampire" devices. Even when these items are turned off, they still drain energy from your home. Conserve energy by unplugging these apparatuses. Examples of these include: desktop computers/displays, faxes, printers, TVs, microwaves, garage door openers, video game consoles, coffee makers, electric toothbrushes, cable boxes, and device charges. If a consumer is unwilling to unplug these devices, then they should invest in power cords with surge protection. Turning off power strips after use will allow the consumer to still conserve energy.

Calculating Appliance Energy Usage

When purchasing appliances for the home or business, be sure to note the wattage of each product. Energy efficient devices will save money a consumer money on their monthly electric bill.





To calculate monthly electrical bill for an appliance, use the following steps listed below:

1) Locate the wattage on a given appliance.

2) Convert this value to kilowatts using dimensional analysis (recall 1000 watts = 1 kW)

3) Multiply the calculated kW by the hours of operation for a month.

4) Research the energy value your power company charges per month (Furman U uses Duke Energy and current values can be located using this link: https://www.duke-energy.com/home/billing/rates, October 2019 value is around .11).

5) Multiple the the kWh by the monthly rate. The final answer will be a dollar value.

Sources

- https://en.wikipedia.org/wiki/Shippi..._Power_Station
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 7.2: Generating Electricity is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.3: Powering the United States

Learning Objectives

- Determine if an energy source is renewable or nonrenewable.
- Realize what sources provide the most power for the United States.
- Compare and contrast the different types of energy sources.
- Understand the environmental impacts of energy sources.
- Compare and calculate the amounts of energy produced from each source (using kilo, mega, and giga units)

Source of Energy

Production of electricity can involve either one of two primary energy sources. For example, a country employs nonrenewable sources like natural gas, petroleum, coal, or nuclear energy to supply power to homes and industry. In addition, an area might choose to incorporate renewable energy sources to meet their consumers' needs. Examples of these would include biomass, hydropower, geothermal, wind, and solar energies.

Figure 7.3.5: Chart illustrating electrical power generation in the United States. Image taken from: https://upload.wikimedia.org/wikiped..._source_v2.png

As of 2017, the current United States population is approximately 326 million people. Supplying electricity to a population of this size requires many different types of sources. At this time, the majority of electrical output is the result of fossil fuel combustion.

***Table** 7.3.2**:** Energy Conversion: Typical Heat Values of Various Fuels (units are Megajoule/kilogram)*

Firewood (dry)	16 MJ/kg
Brown coal (lignite)	10 MJ/kg
Black coal (low quality)	13-23 MJ/kg
Black coal (hard)	24-30 MJ/kg
Natural Gas	38 MJ/m^3
Crude Oil	45-46 MJ/kg
Uranium - in typical reactor	500,000 MJ/kg (of natural U)

Coal

In the United States, coal is the cheapest option as an electrical source. The coal industry estimates that there is approximately 260 billion tons of coal yet to be mined in this country. Assuming the daily rate of coal consumption remained constant, the United States should have enough coal reserves to last over 230 years. Annually, The United States produces over one billion tons of coal. This value equates to approximately 12% of the world's supply.

Figure 7.3.6: Image taken from: https://upload.wikimedia.org/wikiped...the_US.svg.png

Heavy equipment is utilized to extract coal from the surface of the earth. Surface mining can lead to residual chemicals from the coal entering the water supply and surrounding soils. Many states, namely West Virginia and Pennsylvania, still use coal miners to dig coal from underground reserves. Although regulated, this occupation can be particularly risky with exposure to particulate chemicals that can affect the lung and hearts of workers. In addition, gaseous chemicals inside the mine can combust easily if a spark is present. This could result in trapped workers and fatalities at the site.

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

There are 600 coal-burning electric plants in the US delivering 44.9% of American electricity and producing 2 billion tons of CO_2 annually, accounting for 40% of U.S. CO_2 emissions and 10% of global emissions. Coal-burning facilities also produce 64% of the sulfur dioxide emissions, 26% of the nitrous oxide emissions, and 33% of mercury emissions. Sulfur dioxide and nitrous oxide combustion are precursors to acid rain and industrial and photochemical smog. Mercury, a heavy metal toxin, affects the neurological system and can cause behavior and intellectual disorders. Incomplete combustion of carbon can produce particulate matter. These carbon-based solids gather in the atmosphere and make visibility difficult.

Trace amounts of radio nuclides can be detected in coal. Uranium and thorium isotopes that are present will eventually decay to other species. When burned, these substances become airborne and release their decay products. All chemicals and matter resulting from the combustion process can cause heart, lung, and immunity disorders. In addition, high combustion areas can affect the average lifespan of individuals and species living in these areas.

Figure 7.3.7: Two pictures of Beijing in 2005. The picture on the left shows the city after two days of rain. The picture on the right shows smog gathering in the atmosphere. The image is taken from https://upload.wikimedia.org/wikiped...ugust_2005.png

Coal can be cleaned to remove some impurities (like sulfur-based species), but there is no such thing as "clean" coal. Industries that burn coal can install scrubbers in their smokestacks to grab many pollutants before their enter the atmosphere. In the United States, the Environmental Protection Agency (EPA) regulates all pollutants resulting from the combustion process except for carbon dioxide.

Natural Gas

Natural gas combustion does not produce as many types of byproducts. It is a much cleaner fuel to burn and there would be less environmental and health problems related to this fuel choice. Greenhouse gases are still produced when this fuel is burned. Unfortunately, natural gas exists deep within the ground and can be difficult to extract.

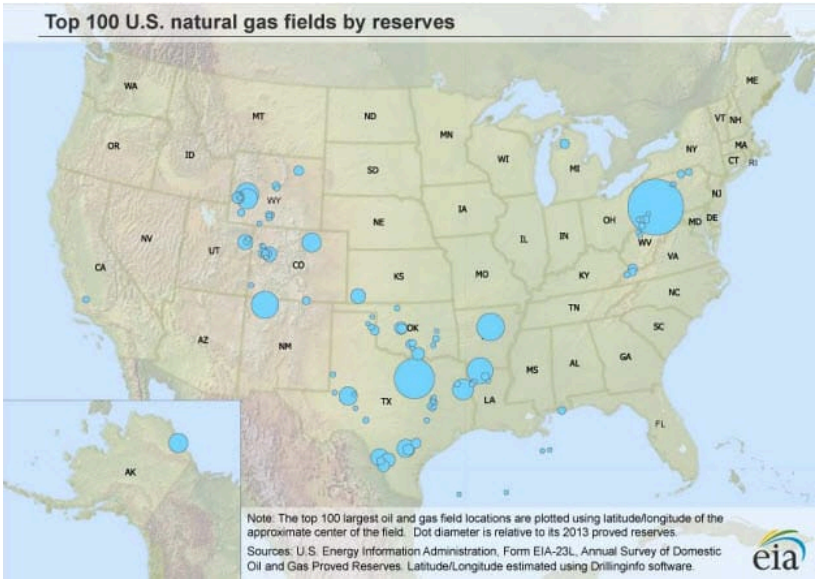


Figure 7.3.8: Natural gas reserves in the United States. Image is taken from https://c1.staticflickr.com/9/8743/1...08652510_z.jpg

Hydroelectric

Hydroelectric sources provide almost 7% of the United State's electricity. To use this type of energy, a large body of water needs to be present within the area. A river is intentionally dammed to create a large water supply. With monitored releases, the dammed water is allowed to fall to a lower altitude. This motion will turn a turbine which will produce energy. Unlike fossil fuel combustion, hydroelectric sources do not release toxins into the atmosphere. As a result, this method of energy production does not contribute to global warming, acid rain, or smog production. Some disadvantages of hydroelectric power include disruption of an aquatic ecosystem. Construction of hydroelectric facilities requires large areas of land that could involve uprooting and affecting human and animal populations. human population. Lastly, altering land geology can cause severe flooding.

The largest hydroelectric facility in the United State is located in Washington State along the Columbia River. Grand Coulee Dam produces over 6800 Megawatts of electrical energy for this region of the country. At this time, there are over 1600 hydroelectric plants in the United States.

Figure 7.3.9: Arial view of Grand Coulee Dam. The image is taken from https://upload.wikimedia.org/wikiped...Dam_aerial.jpg

Solar and Wind Power

In recent years, the use of wind and solar sources has increased in this country. Both of these types of energy require large plots of land for their structures. In addition, both the sun and wind need to be present to generate energy. Wind and solar energy can be stored for days for days in which the climate does not produce adequate amounts of either. With wind energy, there is some concern about how the structures appear and affect wildlife. These devices can also be very noisy. At this time, Texas leads the country in the production of wind energy.

Research has focused on making components of solar panels less toxic and more affordable for the average consumer. Many utility companies offer incentives for residents to install solar panels. In addition, some utility companies will buy back excess solar wattage from its own consumers. The state of California leads the country in the production of solar energy. The Ivanapah Solar Farm in the Mohave Desert generates over 390 megawatts of energy per year.

Figure 7.3.10: Ivanapah is the largest solar farm in the United States in California.The imagee is taken from https://upload.wikimedia.org/wikiped...ity_Online.jpg

Recently, Furman University installed its own solar farm. Click on this link to read more about how this university is using renewable energy to cut down on carbon emissions.

Nuclear Reactors

There are over thirty countries that use nuclear power to generate electricity. Commercial nuclear reactors can be found in North and South America, Europe, Africa, and Asia. The United States has the most reactors of any other countries. There are around ninety-nine reactors in the United States that provide around twenty percent of the electrical energy in the United States. Other countries, like France, employ around sixty nuclear reactors to produce 80% of their electrical power.

***Table** 7.3.1**:** Nuclear Share of Electricity Generation in 2016. Data from International Atomic Energy Agency*

Country	Number of Operated Reactors	Total Net Electrical Capacity	[MW]Nuclear Electricity Supplied	[GW.h]Nuclear Share
FRANCE	58	63130	386452.88	72.3
SLOVAKIA	4	1814	13733.35	54.1
UKRAINE	15	13107	76077.79	52.3
BELGIUM	7	5913	41430.45	51.7
HUNGARY	4	1889	15183.01	51.3
SWEDEN	10	9740	60647.40	40.0
SLOVENIA	1	688	5431.27	35.2

Country	Number of Operated Reactors	Total Net Electrical Capacity	[MW]Nuclear Electricity Supplied	[GW.h]Nuclear Share
BULGARIA	2	1926	15083.45	35.0
SWITZERLAND	5	3333	20303.12	34.4
FINLAND	4	2764	22280.10	33.7
ARMENIA	1	375	2194.85	31.4
KOREA, REPUBLIC OF	25	23077	154306.65	30.3
CZECH REPUBLIC	6	3930	22729.87	29.4
SPAIN	7	7121	56102.44	21.4
UNITED KINGDOM	15	8918	65148.98	20.4
UNITED STATES OF AMERICA	100	100351	804872.94	19.7
ROMANIA	2	1300	10388.20	17.1
RUSSIA	36	26496	184054.09	17.1
CANADA	19	13554	95650.19	15.6
GERMANY	8	10799	80069.61	13.1
SOUTH AFRICA	2	1860	15209.47	6.6
MEXICO	2	1552	10272.29	6.2
ARGENTINA	3	1632	7677.36	5.6
PAKISTAN	4	1005	5438.90	4.4
CHINA	36	31384	197829.04	3.6
INDIA	22	6240	35006.83	3.4
NETHERLANDS	1	482	3749.81	3.4
BRAZIL	2	1884	14970.46	2.9
JAPAN	43	40290	17537.14	2.2
IRAN	1	915	5923.97	2.1

Today's nuclear power plants use fission reactions to generate energy for electrical production. One kilogram of U-235 can supply over 2.4 x 10^7 kilowatt-hours (kWh) of energy. In contrast, one kilogram of coal can only produce 8 kilowatt-hours of energy. Also, nuclear does not contribute to the problems of acid rain, smog, or greenhouse gas production.

Nuclear and coal power plants can appear similar to most people. Both of these types of facilities have large cooling towers that release excess steam energy into the atmosphere. Nuclear power plants will have small, round, and dome-like structures to house the nuclear reactors. In contrast, coal-fired plants will have slender chimney stacks that tower above or at the rounder cooling towers.



Figure 7.3.11: (left) View of Vogtle Nuclear Power Plant (Units 1 and 2) in Augusta, Georgia produces approximately 2.3 gigawatts of power a year. (right) The *Bruce Mansfield Coal Fire Plant in Shipping Port, Pennsylvania produces 2.4 gigawatts of energy a year. Images* are taken from https://en.wikipedia.org/wiki/Vogtle...Vogtle_NPP.jpg and https://3.bp.blogspot.com/-WycgHj-jt...eMansfield.jpg

Sources

- https://en.wikipedia.org/wiki/Shippi..._Power_Station
- https://www.census.gov/popclock/
- http://www.hydro.org/policy/faq/#494
- https://www.euronuclear.org/info/enc...comparison.htm
- https://pubs.usgs.gov/fs/1997/fs163-97/FS-163-97.html

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

This page titled 7.3: Powering the United States is shared under a license and was authored, remixed, and/or curated by LibreTexts.

7.4: Generating Electricity by Using Fission

How a Nuclear Reactor Works

Electrical Power can be generated by using a variety of methods

YouTube Video

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

View this YouTube Video online

Video 7.4.1*: Nuclear Energy Explained: Risk or Opportunity. Anything with the word nuclear next to it usually comes with a fair bit of misunderstanding. Hopefully, this video demystifies the process of how nuclear fuels are turned into electricity and how we can use them in combination with renewables to reduce greenhouse gas emissions and the effects on the climate that comes with high levels of them.*

Components of a Reactor

Chain reactions of fissionable materials can be controlled and sustained without an explosion in a nuclear reactor. Any nuclear reactor that produces power via the fission of uranium (U-235) or plutonium (Pu-239) by bombardment with neutrons must have at least five components: nuclear fuel consisting of fissionable material, a nuclear moderator, reactor coolant, control rods, and a shield / containment system.

The reactor works by separating the fissionable nuclear material such that a critical mass cannot be formed, controlling both the flux and absorption of neutrons to allow shutting down the fission reactions. In a nuclear reactor used for the production of electricity, the energy released by fission reactions is trapped as thermal energy and used to boil water and produce steam. The steam is used to turn a turbine, which powers a generator for the production of electricity.

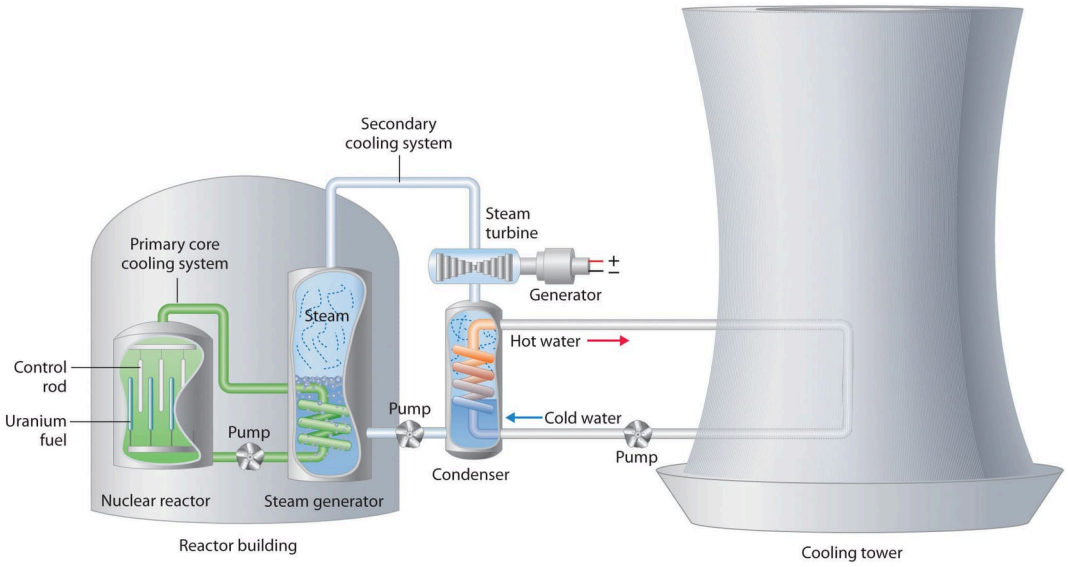


Figure 7.4.2: A Light-Water Nuclear Fission Reactor for the Production of Electric Power. The fuel rods are made of a corrosion-resistant alloy that encases the partially enriched uranium fuel; controlled fission of ^{235}U in the fuel produces heat. Water surrounds the fuel rods and moderates the kinetic energy of the neutrons, slowing them to increase the probability that they will induce fission. Control rods that contain elements such as boron, cadmium, or hafnium, which are very effective at absorbing neutrons, are used to control the rate of the fission reaction. A heat exchanger is used to boil water in a secondary cooling system, creating steam to drive the turbine and produce electricity. The sizeable hyperbolic cooling tower, which is the most visible portion of the facility, condenses the steam in the secondary cooling circuit; it is often located at some distance from the actual reactor.

Nuclear Fuels

Naturally occurring uranium is composed almost totally of two uranium isotopes. It contains more than 99% uranium-238 and less than 1% uranium-235. It is the uranium-235. However, that is **fissionable** (will undergo fission). For uranium to be used as fuel in a fission reactor, the percentage of uranium-235 must be increased, usually to about 3%. (Uranium in which the U-235 content is more than 1% is called **enriched uranium**.)The enriched UF_6 gas is collected, cooled until it solidifies, and then taken to a fabrication facility where it is made into fuel assemblies. Each fuel assembly consists of fuel rods that contain many thimble-sized, ceramic-encased, enriched uranium (usually UO_2) fuel pellets. Modern nuclear reactors may contain as many as 10 million fuel pellets. The amount of energy in each of these pellets is equal to that in almost a ton of coal or 150 gallons of oil. Once the supply of U-235 is acquired, it is placed in a series of long cylindrical tubes called fuel rods. These fuel cylinders are bundled together with **control rods** made of neutron-absorbing material. The amount of U-235 in all the fuel rods taken together is adequate to carry on a chain reaction but is less than the critical mass.

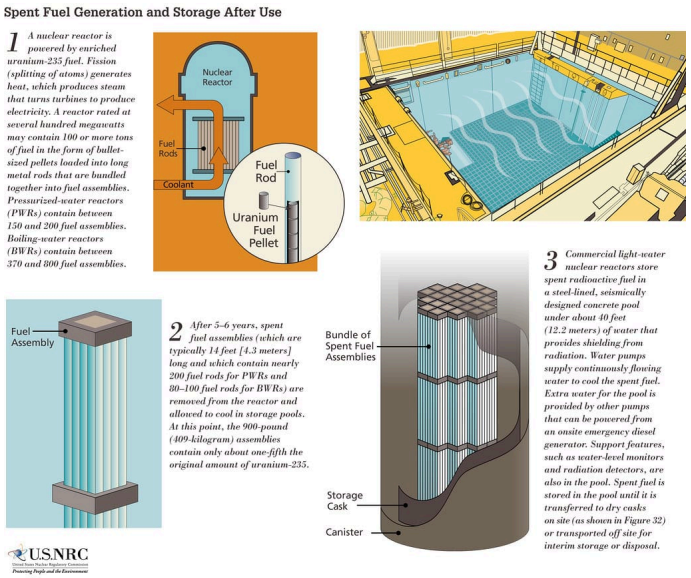


Figure 7.4.3: Fuel Assemblies in nuclear reactors. Image taken: https://c1.staticflickr.com/9/8858/2...7481725c_b.jpg

Nuclear Moderators

Neutrons produced by nuclear reactions move too fast to cause fission (Figure 7.4.4). For efficient collisions to occur, a moderator must be used to slow the the speeds of the neutrons. Moderators can be composed of many different types of chemical substances. The first experimental nuclear reactors used high-purity graphite (or carbon) as a moderator. Today, many countries use light water (LW) as moderators. Otherwise known as H_2O, this substance needs to in large supply. Most LW reactors are built near lakes or other fresh water sources. Depending of the type of fuel, a country would select a corresponding moderator for determining wattage output. Other types of moderators that are used today include: heavy water (HW which is D_2O), carbon dioxide, beryllium, or graphite.

Figure 7.4.4: The Magnox reactor is utilized by Russia and the United Kingdom. wikipedia/commons/thumb/1/10/Magnox_reactor_schematic.svg/2000px-Magnox_reactor_schematic.svg.png

Reactor Coolants

A nuclear reactor coolant is used to carry the heat produced by the fission reaction to an external boiler and turbine, where it is transformed into electricity. Two overlapping coolant loops are often used; this counteracts the transfer of radioactivity from the reactor to the primary coolant loop. All nuclear power plants in the US use light water as a coolant. Other coolants include molten sodium, lead, a lead-bismuth mixture, or molten salts.

Control Rods

Nuclear reactors use control rods (Figure 7.4.4) to control the fission rate of the nuclear fuel by adjusting the number of slow neutrons present to keep the rate of the chain reaction at a safe level. Control rods are made of boron, cadmium, hafnium, or other elements that are able to absorb neutrons. Boron-10, for example, absorbs neutrons by a reaction that produces lithium-7 and alpha particles:

$$^{10}_{5}\mathrm{B} + ^{1}_{0}\mathrm{n} \longrightarrow ^{7}_{3}\mathrm{Li} + ^{4}_{2}\mathrm{He} \tag{7.4.1}$$

When control rod assemblies are inserted into the fuel element in the reactor core, they absorb a larger fraction of the slow neutrons, thereby slowing the rate of the fission reaction and decreasing the power produced. Conversely, if the control rods are removed, fewer neutrons are absorbed, and the fission rate and energy production increase. In an emergency, the chain reaction can be shut down by fully inserting all of the control rods into the nuclear core between the fuel rods.

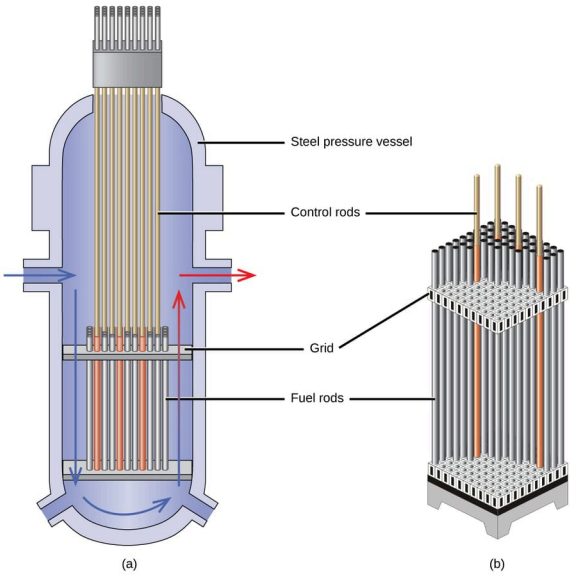


Figure 7.4.5: The nuclear reactor core shown in (a) contains the fuel and control rod assembly shown in (b). (credit: modification of work by E. Generalic, http://glossary.periodni.com/glossar...en=control+rod)

Shield and Containment System

During its operation, a nuclear reactor produces neutrons and other types radiation. Even when shut down, the decay products remain radioactive. In addition, an operating reactor is thermally very hot, and high pressures result from the circulation of water or another coolant through it. Thus, a reactor must withstand high temperatures and pressures, and must protect operating personnel from the radiation. Reactors are equipped with a containment system (or shield) that consists of three parts:

1. The reactor vessel, a steel shell that is 3–20-centimeters thick and, with the moderator, absorbs much of the radiation produced by the reactor
2. A main shield of 1–3 meters of high-density concrete
3. A personnel shield of lighter materials that protects operators from γ rays and X-rays

In addition, reactors are often covered with a steel or concrete dome that is designed to contain any radioactive materials might be released by a reactor accident.

Figure 7.4.6***:*** *Construction on Millstone Nuclear Power Plant in Connecticut. Image taken from: https://upload.wikimedia.org/wikiped..._structure.jpg*

Nuclear power plants are designed in such a way that they cannot form a supercritical mass of fissionable material and therefore cannot create a nuclear explosion. But as history has shown, failures of systems and safeguards can cause catastrophic accidents, including chemical explosions and nuclear meltdowns (damage to the reactor core from overheating).

Reviewing Reactor components

Read over this particular section of the OER and then watch both embedded videos. Test your knowledge of nuclear reactors

by answering 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. What particle bombards U-235 to produce fission?

2. How does the fission reaction in a nuclear weapon differ than than of a reactor?

3. What is the purpose of a control rod in a nuclear reactor?

4. In moving the control rods, must electricity always be applied?

5. What is the purpose of the moderator? What is the chemical formula of the moderator described in this video?

6. What is the purpose of a coolant? What is the chemical formula of the coolant described in this video?

7. If coolant leaks out of the reactor, then what could happen?

8. If a nuclear reactor experiences meltdown, how has this affected the structure of the nuclear reactor?

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 7.4: Generating Electricity by Using Fission is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.5: United States Commercial Nuclear Power Reactors

Location Commercial Reactors

In 2017, there were ninety-nine working nuclear reactors in the United States. The majority of these reactors are on the eastern side of the country. Each reactor site houses between one and three reactors. The average age of these nuclear reactors is approximately thirty-six years. Each year, reactor units are rigorously inspected for safety. Also, reactors that cannot meet safety measures or are not able to be upgraded will be placed on shut-down. These reactors will eventually be closed forever or decommissioned.

Figure 7.5.1: Map of working United States commercial reactors. Image taken from https://upload.wikimedia.org/wikiped...actors_map.gif

Two Types of United States Commercial Reactors

In the United States, commercial reactors can be one of two types. Sixty-five reactors are pressurized water reactors (or PWRs). This type of reactor uses light water as a moderator and a coolant. The water in the reactor is kept under high pressures. This type of configuration does not allow water to boil. The water inside the reactor transfers its heat to the water in a steam generator. Then, steam is produced to aid in the production of electrical energy.

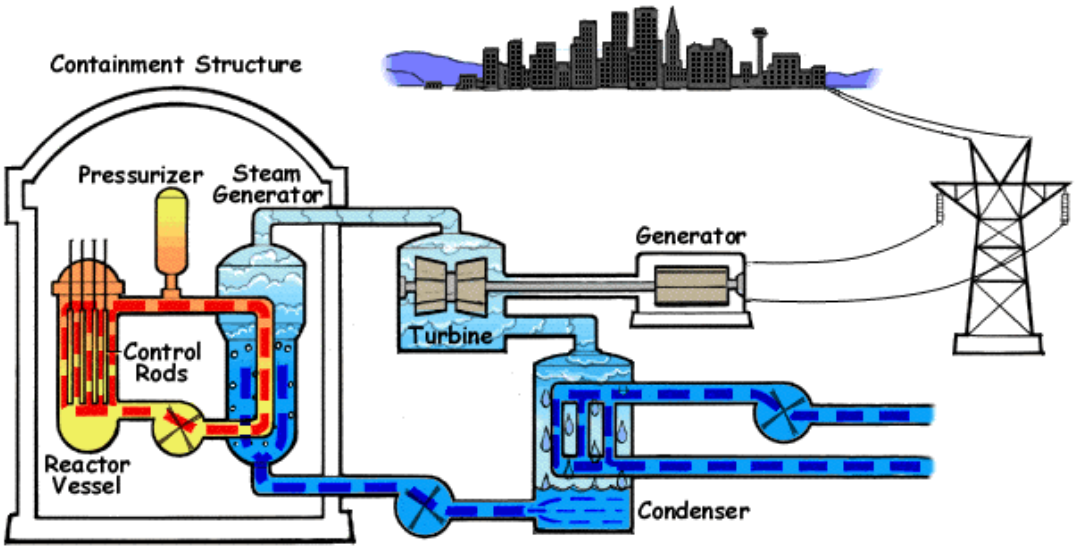


Figure 7.5.7: Pressurized Water Reactor. Image is taken from https://upload.wikimedia.org/wikiped...terReactor.gif

The other thirty-four commercial reactors are boiling water reactors (BWRs). These reactors also use light water as their moderator and coolant. Here, the water in the reactor is allowed to boil. This heated water goes directly to the turbine house which is separated from a generator. The radioactive water that goes through the turbine is contained inside the nuclear facility. Cooling this water with a condenser allows it to be returned into the reactor and used for coolant once again.

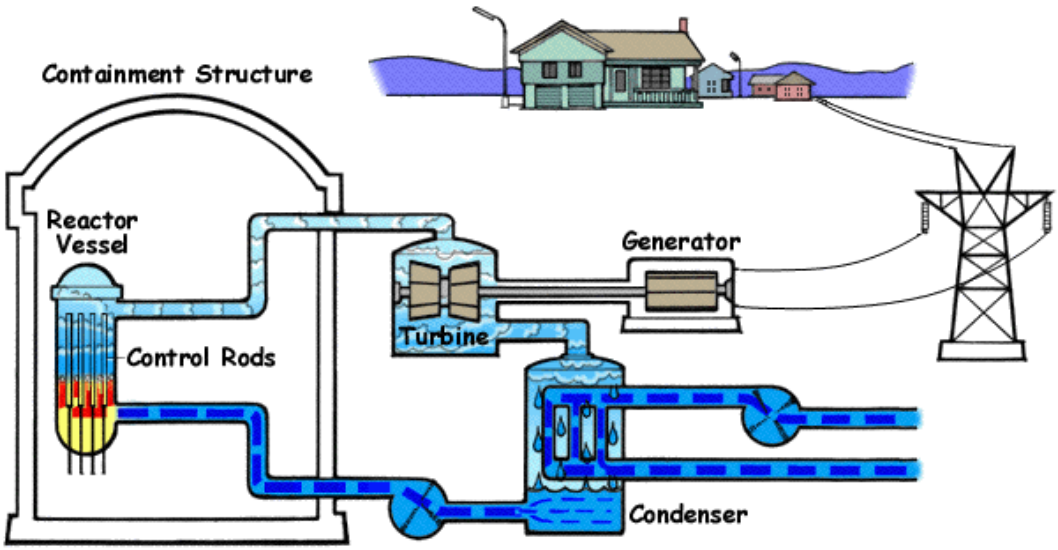


Figure 7.5.8: Boiling Water Reactor. Image taken from https://upload.wikimedia.org/wikiped...terReactor.gif

References

1. https://www.eia.gov/tools/faqs/faq.php?id=228&t=21

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 7.5: United States Commercial Nuclear Power Reactors is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.6: South Carolina Energy

Energy Sources

According to the Department of Energy, the state of South Carolina generates 96.8Twh (tera watt hours= 10^{12} watt hours) of electrical energy a year. This accounts for only 2% of the country's electrical power. The sources of South Carolina energy production are shown in the table below.

Source	Twh (Terawatt Hours of Energy)	Percentage of Electrical Power
Coal	28.4	29%
Petroleum	0.1	<1%
Natural Gas	14.3	15%
Nuclear	51.1	53%
Hydroelectric	0.5	<1%
Other Renewable	0	<1%

Nuclear Power in South Carolina

Oconee Nuclear Power Plant is located in Seneca, South Carolina. This facility is owned and operated by Duke Energy. The three pressurized water reactors (PWR) were constructed and then commissioned in the early 1970's. The approximate cost for the building of these three units was 1.96 billion dollars (2007). Each of the three units produces 846 megawatts (10^6 watts) of energy per year. In total, Oconee Nuclear Station has a maximum output capacity of 2500 megawatts of energy per year. In terms of hours, Oconee Nuclear provides 20,565 Gigawatt (10^9 watts) hours of power.

Like all US commercial reactors, Oconee Nuclear Station uses light water as a moderator and coolant. PWR contain between 150-200 fuel rod assemblies. Both BWRs and PWR plants in the United States use low enriched uranium (LEU) embedded in each fuel rod as an energy source.

Figure 7.6.1: Image taken from: https://upload.wikimedia.org/wikiped...nt_diagram.svg

For backup power, Oconee units 1-3 rely on hydroelectric sources provided by Lake Keowee. If this system were to fail, a nearby combustion facility will provide energy to cool the core uncase of emergencies.

The Fate of V.C. Summer Units 2 and 3.

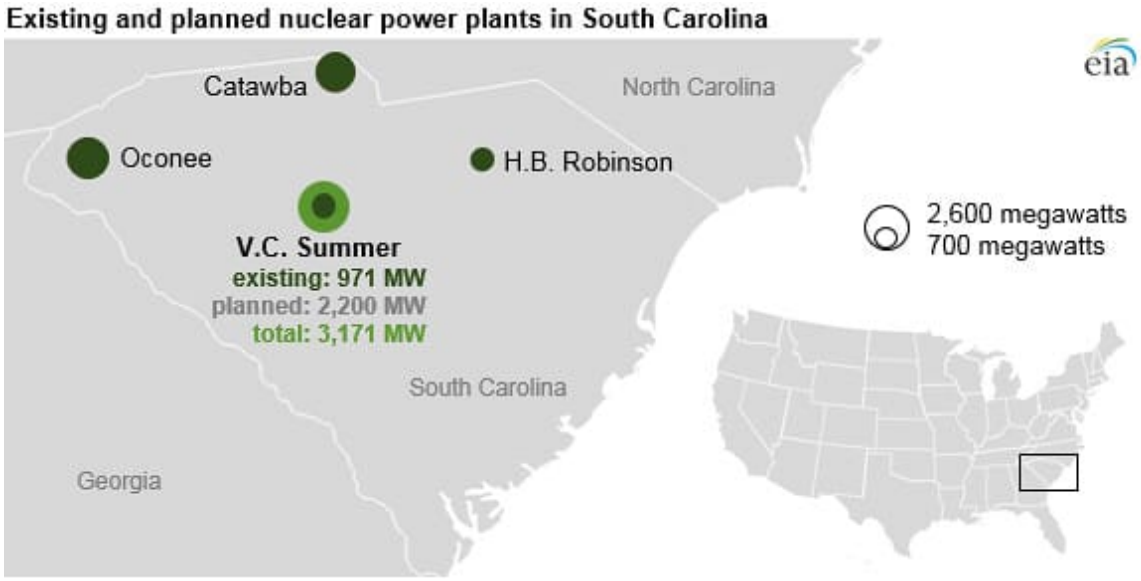


Figure 7.6.2: Oconee reactor image taken from: https://upload.wikimedia.org/wikiped...arolina%29.JPG **Figure** 7.6.3**:**Nuclear power plants in South Carolina. Image taken from:https://c1.staticflickr.com/5/4410/3...258ed1cd_z.jpg

Blah blah vc summer drama/power bills. cancelled construction

Sources

1. https://www.energy.gov/sites/prod/fi...%20Profile.pdf

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 7.6: South Carolina Energy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.7: Three Mile Island Nuclear Accident

The importance of cooling and containment are amply illustrated by three major accidents that occurred with the nuclear reactors at nuclear power generating stations in the United States (Three Mile Island), the former Soviet Union (Chernobyl), and Japan (Fukushima). Other nuclear incidents have occurred around the world that have involved the storage or transportation of nuclear materials.

Three Mile Island (Harrisburg, Pennsylvania, March 28, 1979)

The two reactors housed at this facility are located in Middletown, Pennsylvania. This site is approximately 10 miles southeast of Harrisburg, Pennsylvania. Both units are PWRs (pressurized water reactors) are were commissioned in the mid to late 1970's. Before the accident, each of the Three Mile Island (TMI) reactors generated around 800 MW of electrical power. The smaller cylindrical buildings in the photograph below are the unit 1 and 2 reactors at this facility. The four larger structures in this picture are the cooling towers that allow excess heat to be transferred to the atmosphere.



Figure 7.7.1: Three Mile Island Nuclear Facility. Image taken from: https://upload.wikimedia.org/wikiped...MileIsland.jpg

On March 28, 1979, Three Mile Island's Unit Two reactor lost cooling water in the secondary loop (note the arrow pointing down in the schematic below). This event occurred around 4:00 a.m. and alarms sounded in the control room immediately. Within one minute, the reactor automatically shutdown. Meanwhile, water in the primary loop (yellow tube) overheated and the pressure of the reactor increased. To relief this excess pressure inside the reactor, the pressure relief valve (top yellow/red tube) opened up. Backup pumps restarted the water in the secondary loop to assist cooling efforts.

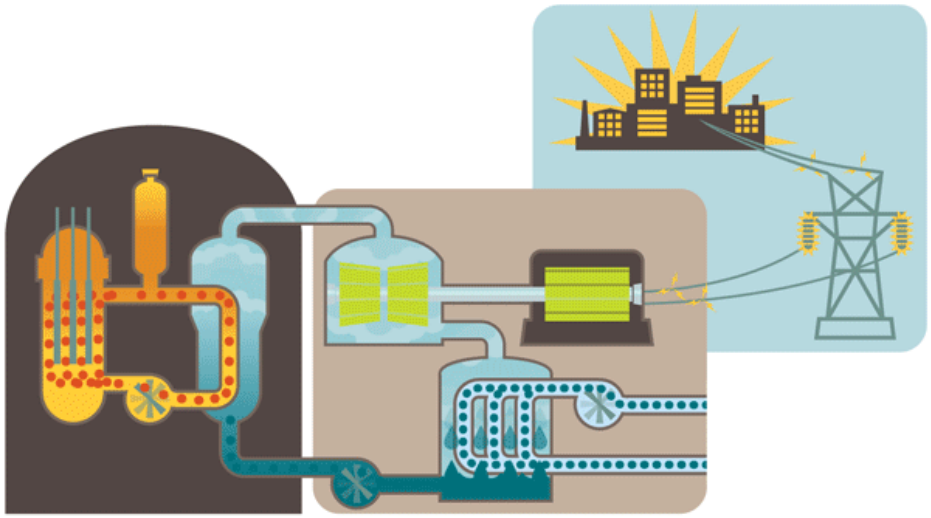


Figure 7.7.2: The following animated diagram graphically depicts the sequence of events associated with the accident at TMI-2. Image used with permission (United State Nuclear Regulatory Commission; Public Domain).

In the control room, operators noted the decrease in pressure of the reactor and a reduction of cooling water in the primary loop. Within fifteen minutes of the original malfunction, approximately 3000 gallons of cooling water had escaped the reactor but remained in the containment building. The reactor components (fuel and control rods) became exposed with lack of water. These materials were melted to various degrees. Back in the control room, instrumentation noted that the pressure valve remained closed. Operators attempted to cool down the control and fuel rods by pumping more water into the primary loop of the reactor. Unfortunately, the pressure value was still open and most of the water converted to steam. This made temperatures increase greatly in the reactor core.

Figure 7.7.3: Image taken from: https://upload.wikimedia.org/wikiped...ntrol_room.png

By 6:22 a.m., the operators closed a valve before the pressure release valve. This allowed water to remain in the primary loop. Unfortunately, most of the water in the core was in the form of steam and did not circulate efficiently throughout the core. In addition, the steam reacted with metals in the control and fuel rods. These reactions produced combustible hydrogen gas which rose to the top of the reactor core. By 7:50 a.m., all cooling water was restored to the core successively, but a large hydrogen bubble hovered over the top of the reactor. At this time, operators were fearful the bubble might ignite and then explode the core to the outside environment. To prevent this explosion from happening, operators and engineers vented the hydrogen gas outside the nuclear reactor into the atmosphere. This venting occurred from March 30 until April 1, 1978.

opy and Paste Caption here

Residents in nearby communities were informed of a nuclear issue around 7:24 a.m. on March 28, 1979. Radio and television broadcasts started within the next hour. By 12:45 p.m., representatives of the Department of Energy (DOE) arrived to evaluate the current status of the reactor. DOE agents put together a plan of action to keep the reactor under control. Numerous state governmental officials met with the DOE and devised safety protocols (remain indoors) and emergency evacuation plans if the hydrogen bubble were to burst. To see a comprehensive timeline of these events, please click here.

In the venting process, minimal amount of radioisotopes were released to the areas surrounding TMI nuclear power station. Several independent research facilities have quantified the radiation that was released during the accident These studies state that residents within 10 miles of TMI were exposed to radiation within the range of 0.08 mSv to 1 mSv. Referring back to Section 5.8, this level of radiation is comparable to having a chest x-ray exam (Table 7.7.2).

***Table** 7.7.1**:** Fission products vented to the atmosphere from TMI 2*

Isotope	Emits	Half-life ($t_{1/2}$)
Kr-85	beta particles, gamma rays	10.8 years
Cs-137	beta particles, gamma rays	30.2 years

Isotope	Emits	Half-life ($t_{1/2}$)
I-131	beta particles, gamma rays	8.0 days
Xe-133	beta particles, gamma rays	5.3 days

In 1984, nuclear scientists were able to survey the destruction of the reactor core. They concluded that approximately 1/3 of the fuel rods melted inside the reactor. This accident was labeled a partial meltdown (full meltdown would indicate the majority of fuel rods were melted). The contaminant structure remained in tact and protected the citizens of the area.

The clean-up process for TMI unit 2 started in August of 1979. It was a twelve year project that employed over 1000 people. All radioactive fuel and water were shipped to existing nuclear waste storage facilities in the United States (Washington State and Idaho). The containment building for this reactor still remains and houses waste for the TMI unit 1. The approximate clean-up price for this accident was 973 million dollars.

The accident at Three Mile Island halted all construction of new nuclear power plants in the United States. Many citizens became fearful of nuclear power in general and became supportive of other types of energy sources. Films like the China Syndrome catalyzed this feeling for decades after this catastrophe.

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

At the time of the partial meltdown of TMI Unit 2, reactor 1 was shutdown for refueling. Despite much opposition, Unit was restarted in the fall of 1985 has been providing power to the residents of this area since then. The NRC has extended the license of this reactor through the spring of 2034.

References

1. Harwood, William, Herring, Geoffrey, Madura, Jeffry, and Petrucci, Ralph, General Chemistry: Principles and Modern Applications, Ninth Edition, Upper Saddle River,New Jersey, Pearson Prentice Hall, 2007.
2. Larabee, Ann. Decade of Disaster. Chicago: Board of Trustees of the University of Chicago, 2000.
3. http://www.world-nuclear.org/informa...-accident.aspx

Contributors

- Abheetinder Brar (UCD)

This page titled 7.7: Three Mile Island Nuclear Accident is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.8: Chernobyl Nuclear Disaster

Chernobyl (Pripyat, Ukraine, April 26, 1986)

On April 26, 1986, a test was scheduled at the Chernobyl Nuclear Power Plant to evaluate residual cooling capacity in the event of a power grid failure. At the time of the accident the Chernobyl facility used four RBMK (reaktor bolshoy moshchnosty kanalny or high power channel) reactors to produce a total of 4000 MW of energy. Two more reactors were being constructed to produce additional wattage. RBMKs not only produce large amount of energy, but also produce weapons grade Pu-239. The excessive power, design, and materials of these reactors have discouraged many countries (the United States) from building them.

Figure 7.8.1: Schematic of a RBMK reactor. Image taken from: https://upload.wikimedia.org/wikiped..._schematic.svg

When comparing the differences of RBMK reactors to those of typically used in most other countries, view the table below:

Table 7.8.1: Types of Nuclear Reactor

Type of Reactor	Containment	Fuel	Coolant	Moderator
PWR	repetitive layers of lead/ concrete and top is closed inside reactor building	slightly enriched U-235 (2-4%)	LW = light water	LW = light water
BWR	repetitive layers of lead/ concrete and top is closed inside reactor core building	slightly enriched U-235 (2-4%l	LW = light water	LW = light water
RBMK	layers of lead/concrete and top is open inside reactor core building to allow extraction of rods while reactor is producing energy	low enriched U-235 (<2%)	LW = light water	Graphite (carbon)

Unlike light water reactors, RBMK reactors use graphite moderators to slow the neutrons of the fission process. This type material can combust easily. The less enriched U-235 fuel of this reactor produces Pu-239 to be produced. Unfortunately, the Chernobyl reactor technicians felt that removing the new Pu-239 fuel would be too difficult with a full containment covering of the unit 4. The new Pu-239 material would extracted while the reactor continued to produce energy. The high power wattage, combustible moderator, and open containment convinced many other countries to steer clear of building RBMKs.

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

Meltdown occurred at 1:23 AM, starting a fire that dispersed large quantities of radioactive materials into the atmosphere. The amount of radioactive material released was 400 times more than the amount the atomic bombing of Hiroshima released. The fallout would be detected in almost all parts of Europe.

photo of the Chernobyl Unit 4 reactor

Chernobyl's Unit 4 Nuclear Reactor after full meltdown. Image taken from: https://upload.wikimedia.org/wikiped...yl_Reactor.jpg

Before the accident

Nuclear reactors require active cooling in order to remove the heat generated by radioactive decay. Even when not generating power, reactors still generate some heat, which must be removed in order to prevent damage to the reactor core. Cooling is usually accomplished through fluid flow, water in Chernobyl's case.

The problem at the Chernobyl plant was that following an emergency shutdown of all power, diesel generators were needed to run the cooling pumps. These generators took about a minute to attain full speed, which was deemed an unacceptably long time for the reactor to be without cooling. It was suggested that the rotational momentum of the winding down steam turbine be used to power the pumps in the time between shutdown and the generators being ready. A test was devised to test this method in 1982, but the turbine did not prove to be successful in providing the required voltage as it spooled down. Two more tests would be conducted in the following years, but would also be unsuccessful. The fourth test was scheduled to be run on April 25, 1986.

The experiment was devised in such a way that if it had gone as planned, the disruption and danger to the plant would be very minimal. First, the reactors would be brought down to low power, between 700 and 800 megawatts. Then the steam turbine would be run up to full speed and then turned off. The power generated by the winding down generators would then be measured to determine if it was sufficient to power the cooling pumps in the time before the diesel generators got up to full speed.

By 1986, the plant had been running for two years without the implementation of a method to keep the cooling pumps running continuously following an emergency shutdown. This was an important safety measure that the plant was lacking, which presumably gave the plant managers a considerable amount of urgency in completing another test.

The experiment

Preparing for the experiment

The experiment was scheduled to run during the day shift of 1985, while the night shift would only have to maintain cooling of the radioactive decay in the shut-down plant. However, another power generator nearby unexpectedly shut down, necessitating the need for the Chernobyl plant to delay the test and continue producing power. The experiment would be resumed at 11:04 PM, by which time the day shift had departed and the evening shift was about to leave. This meant that the experiment would be conducted in the middle of two shifts, leaving very little time for the night shift employees to be briefed about the experiment and told what to do.

The power reduction of reactor 4 to 700 MW was accomplished at 00:05 AM on the 26th of April. However, the natural production of a neutrino absorber, Xenon-135, led to a further decrease in power. When the power dropped to about 500 MW, the night shift operator committed an error and inserted the reactor control rods too far. This caused the reactor to go into a near-shutdown state, dropping power output to around 30 MW.

Since this was too low for the test, it was decided to restore power by extracting the control rods. Power would eventually rise and stabilize at around 200 MW.

The operation of the reactor at such a low power level would lead to unstable temperature and flow. Numerous alarms and warnings were recorded regarding emergency measures taken to keep the reactor stable. In the time between 0:35 and 0:45 AM, alarm signals regarding thermal-hydraulic parameters were ignored in order to preserve the reactor's power level.

The test continued, and at 1:05 AM extra water pumps were activated in order to increase the water flow. The increased coolant flow rate led to an increase of the coolant temperature in the core, reducing the safety margin. The extra water flow also led to a decrease in the core's temperature and increased the neutron absorption rate, decreasing the reactor's power output. Operators removed the manual control rods in order to maintain power.

All these actions led to the reactor being in an unstable state that was clearly outside safe operation protocol. Almost all the control rods had been removed, which reduced the effectiveness of inserting safety rods in an emergency shutdown. The water was very close to boiling, which meant that any power increase would cause it to boil. If it started boiling, it would be less effective at absorbing neutrons, further increasing the reactor's power output.

Conducting the experiment

The experiment was started at 1:23:04 AM. The steam to the turbines was shut off, causing the turbines to start spooling down. Four of the eight cooling pumps were also shut down. The diesel generator was started and began powering the cooling pumps after at 1:23:43. Between this time, the four pumps were powered by the slowing steam turbines. As the turbines slowed down, their power output decreased, slowing the cooling pumps. This lead to increased formation of steam voids in the core, reducing the ability of the cooling water to absorb neutrons. This increased the power output of the reactor, which caused more water to boil into steam, further increasing the reactor's power. However, during this time the automatic control system was successful in limiting power increase through the insertion of control rods.

At 1:23:40, a button was pressed that initiated the emergency shutdown of the reactor and the insertion of all control rods. It is believed that this was done as a routine method to shut down the reactor to conclude the experiment and not as an emergency measure.

The process of inserting the control rods was initiated, but it took about 20 seconds for the rods to be completely inserted. A flawed design in the graphite-tip control rod meant that coolant was displaced before the neutron absorbing material could be fully inserted and slow down the reaction. This meant that the process of inserting the control rods actually increased the reaction rate in the lower half of the core.

A massive power spike occurred, causing the core to overheat. Some of the fuel rods fractured, causing the control rods to become stuck before they were fully inserted. Within three seconds the core's power output rose to above 500 MW. According to simulation, it is estimated that power output then rose to 30 GW, ten times the normal power output. This was caused by the rising power output causing massive steam buildup, which destroyed fuel elements and ruptured their channels.

It is not possible to know precisely what sequence of events led to the destruction of the reactor. It is believed that the steam buildup entered the reactor's inner structure and lifted the 2000 ton upper plate. This steam explosion further ruptured fuel channels, resulting in more coolant turning into steam and leaving the reactor core. This loss of coolant further increased the reactor's power. A nuclear excursion (an increasing nuclear chain reaction) caused a second, even more powerful explosion.

The explosion destroyed the core and scattered its contents in the surrounding area, igniting the red-hot graphite blocks. Against safety regulations, a flammable material, bitumen, had been used in roof of the reactor. When this was ignited and scattered into the surrounding area, it started several fires on reactor 3. Those working there were not aware of the damage that had been done and continued running the reactor until it was shut down at 5:00 AM.

Crisis Management

Radiation Levels

In the worst-hit parts of the reactor building, radiation levels were high enough to cause fatal doses in a matter of minutes. However, all dosimeters available to the workers did not have the ability to read radiation levels so high and thus read "off scale." Thus, the crew did not know exactly how much radiation they were being exposed to. It was assumed that radiation levels were much lower than they actually were, leading to the reactor crew chief to believe that the reactor was still intact. He and his crew would try to pump water into the reactor for several hours, causing most of them to receive fatal doses of radiation.

Containment

Fire crew were called in to protect the remaining buildings from catching fire and to extinguish the still burning reactor 4. While some firefighters were not aware of the harmful doses of radiation they were receiving and had assumed it to be a simple electrical fire, others knew that they would probably receive fatal doses of radiation. However, their heroic efforts were necessary in order to try to contain the large amounts of radiation being released into the atmosphere. The fires in the surrounding buildings were extinguished by 5:00 AM, but it would take firefighters until May 10 before they could fully extinguish reactor 4.

In order to prevent a steam explosion from occurring, volunteers were needed to swim through radioactive water and drain a pool of water under the reactor core. While they were successful, they would later succumb to the high doses of radiation that they had received. The worst of the radioactive debris was shoveled back into the reactor by crew wearing heavy protective

gear. In total, 600,000 people worked in the cleanup, about 250,000 of which received their lifetimes' limit of radiation. It is estimated that over 10,000 eventually died from the radiation.

By December, a concrete sarcophagus had been completed that sealed off the reactor. This was never meant to be a completely permanent solution, however, and is now in danger of collapsing. A collapse could cause a large amount of radioactive material to once again be released and spread around the world. This is why it is necessary that a new structure be constructed to contain reactor 4.

Evacuation of Pripyat

Pripyat, a city nearby the power plant, was not immediately evacuated. At first, the government denied that the reactor had exploded and insisted that it was only a small accident. By April 27, though, investigators were forced to acknowledge that the reactor had exploded and ordered Pripyat to be immediately evacuated.

Effects

400 times more radiation was released by the disaster than had been by the atomic bombing of Hiroshima. The radiation would later be detected in almost all parts of Europe. Over one million people could have been adversely affected by the radiation. The radiation would cause numerous problems, including Down's Syndrome, chromosomal aberrations, mutations, leukemia, thyroid cancer, and birth defects.

The radiation would affect all parts of the environment surrounding the plant, killing plants and animals and infecting the soil and groundwater. Life has returned to the area and seems to be flourishing, possibly due to the lack of human intrusion. Remarkably, numerous species have been reported to have adapted to their environment and have developed increased tolerance of radiation, making it possible for them to live with the radiation that is still prevalent in the soil and plants around the plant. It has even been reported that radiotrophic fungi have been growing on the walls of reactor 4.

Today, radiation levels are still higher than normal in the areas surrounding the plant, but have dropped considerably from the levels that they were at twenty years ago. It is now considered safe to visit the areas immediately surrounding the plant for short periods of time. However, it is estimated that it will take 20,000 years for reactor 4's core to be completely safe.

Another major nuclear accident involving a reactor occurred in April 1986, at the Chernobyl Nuclear Power Plant in Ukraine, which was still a part of the former Soviet Union. While operating at low power during an unauthorized experiment with some of its safety devices shut off, one of the reactors at the plant became unstable. Its chain reaction became uncontrollable and increased to a level far beyond what the reactor was designed for. The steam pressure in the reactor rose to between 100 and 500 times the full power pressure and ruptured the reactor. Because the reactor was not enclosed in a containment building, a large amount of radioactive material spewed out, and additional fission products were released, as the graphite (carbon) moderator of the core ignited and burned. The fire was controlled, but over 200 plant workers and firefighters developed acute radiation sickness and at least 32 soon died from the effects of the radiation. It is predicted that about 4000 more deaths will occur among emergency workers and former Chernobyl residents from radiation-induced cancer and leukemia. The reactor has since been encapsulated in steel and concrete, a now-decaying structure known as the sarcophagus. Almost 30 years later, significant radiation problems still persist in the area, and Chernobyl largely remains a wasteland.

Outside Links

- http://www.nucleartourist.com/events/chernobl.htm
- http://en.wikipedia.org/wiki/Chernobyl_disaster

This page titled 7.8: Chernobyl Nuclear Disaster is shared under a CC BY-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.9: Fukushima Nuclear Disaster

In 2011, the Fukushima Daiichi Nuclear Power Plant in Japan was badly damaged by a 9.0-magnitude earthquake and resulting tsunami. Three reactors up and running at the time were shut down automatically, and emergency generators came online to power electronics and coolant systems. However, the tsunami quickly flooded the emergency generators and cut power to the pumps that circulated coolant water through the reactors. High-temperature steam in the reactors reacted with zirconium alloy to produce hydrogen gas. The gas escaped into the containment building, and the mixture of hydrogen and air exploded. Radioactive material was released from the containment vessels as the result of deliberate venting to reduce the hydrogen pressure, deliberate discharge of coolant water into the sea, and accidental or uncontrolled events.

An evacuation zone around the damaged plant extended over 12.4 miles away, and an estimated 200,000 people were evacuated from the area. All 48 of Japan's nuclear power plants were subsequently shut down, remaining shuttered as of December 2014. Since the disaster, public opinion has shifted from largely favoring to largely opposing increasing the use of nuclear power plants, and a restart of Japan's atomic energy program is still stalled (Figure 7.9.4).



Figure 7.9.1: (a) After the accident, contaminated waste had to be removed, and (b) an evacuation zone was set up around the plant in areas that received heavy doses of radioactive fallout. (credit a: modification of work by "Live Action Hero"/Flickr).

The energy produced by a reactor fueled with enriched uranium results from the fission of uranium as well as from the fission of plutonium produced as the reactor operates. As discussed previously, the plutonium forms from the combination of neutrons and the uranium in the fuel. In any nuclear reactor, only about 0.1% of the mass of the fuel is converted into energy. The other 99.9% remains in the fuel rods as fission products and unused fuel. All of the fission products absorb neutrons, and after a period of several months to a few years, depending on the reactor, the fission products must be removed by changing the fuel rods. Otherwise, the concentration of these fission products would increase and absorb more neutrons until the reactor could no longer operate.

Spent fuel rods contain a variety of products, consisting of unstable nuclei ranging in atomic number from 25 to 60, some transuranium elements, including plutonium and americium, and unreacted uranium isotopes. The unstable nuclei and the transuranium isotopes give the spent fuel a dangerously high level of radioactivity. The long-lived isotopes require thousands of years to decay to a safe level. The ultimate fate of the nuclear reactor as a significant source of energy in the United States probably rests on whether or not a politically and scientifically satisfactory technique for processing and storing the components of spent fuel rods can be developed.

This page titled 7.9: Fukushima Nuclear Disaster is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

7.10: Storing and Transporting Nuclear Waste Materials

Waste from Uranium Production

The tailings or waste produced by the extraction or concentration of uranium from its ore contain radioactive isotopes of uranium, thorium, and radium as well as significant concentrations of heavy metal including chromium, lead, molybdenum, and vanadium. More than 200 pounds of tailings are produced for each pound of uranium. This sandy waste material must be contained in carefully monitored sites known as tailings piles (Figure 7.10.1).

Figure 7.10.1: Moab uranium tailings site is located in Utah. Image taken from: https://upload.wikimedia.org/wikiped...lings_pile.jpg

Uranium production processes do not affect the level of radioactivity and do not produce significant chemical waste. An enrichment process for one ton of uranium hexafluoride produces 130 kg of UF_6 (3.5% U-235) and 870 kg of depleted UF_6 containing U-238. Depleted uranium has few applications. However, its high density 18.7 g/cm^3 makes it useful in armor plating and radiation shielding. In addition, depleted uranium is a potential energy source for fast breeder reactors.

Waste from Nuclear Power Generation

In the open fuel cycle, the spent fuel rods remain in the pools under at least 20 feet of water. This aqueous layer protects the surroundings from radiation. Currently, most reactor sites in the United States utilize this type of storage. The Nuclear Regulatory Commission (NRC) requires all facilities to store spend fuel rods at least three years under water. Most nuclear plants will keep rods in water up to around ten years and then transfer waste to a dry system (cask).

Figure 7.10.2: Spent fuel pond at San Onofre Nuclear Generating Station in San Clemente, California. Image taken from: https://upload.wikimedia.org/wikiped...ol%2C_2014.jpg

Once fuel rods are removed from the pool, they are placed in a cylinder in a chemically inert atmosphere of helium gas (Figure 7.10.3). The cylinder is then sealed and encased in steel and concrete to contain the radiation and enhance security for storage or transportation to a permanent repository. All nuclear power plants are allowed to use dry cask storage on site. If a facility is full, then dry fuel rods can be transferred to a decommissioned nuclear power plant. Rods stored on or off site can be transported to a temporary consolidated storage site.

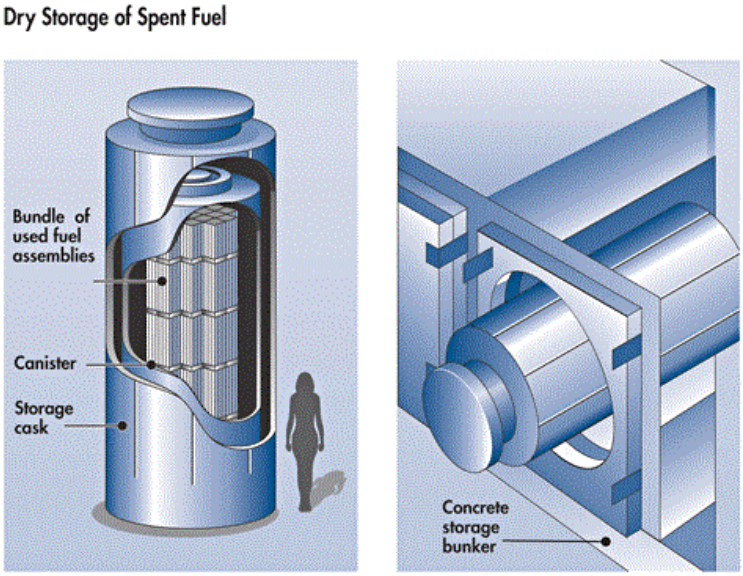


Figure 7.10.3: Dry Cast Storage. Courtesy of the U. S. Nuclear Regulatory Commission

No Term range plan

At this time (2017), there is **no** permanent location in the United States that will store nuclear power waste.

In 2009, the total amount of nuclear waste generated by reactors was 62,683 metric tons. Seventy eight percent of this waste is in wet storage. The remaining twenty two percent is contained in dry cask. Annually, United States' reactors produce between 2000-2400 tons (1 ton = 2000 pounds) of nuclear waste. To understand this amount of weight, view the vehicle below. Built by Ford Motor Company, the 2017 version of the all wheel drive Edge weighs 4,060 pounds (Figure 7.10.4). This amount is relative in size to the total amount of nuclear waste generated by all United States reactors in one year.



Figure 7.10.4: Ford Motor Company's mid-sized SUV named the Edge. Image taken from: https://upload.wikimedia.org/wikiped...Y7A3333%29.jpg

Nuclear waste is radioactive waste, meaning that it spontaneously emits radiation. It usually originates from the by-products of nuclear reactions in applications such as medicine and research. Radioactive waste degrades with time, releasing alpha, beta, and gamma radiation that pose many health risks to the environment and most organisms, including humans. Due to the harmful nature of nuclear waste, there is strict government regulation on the safe disposal of it. There are several types of waste, and as such, many different ways of discarding it.

Closed or Open Fuel Cycle?

Spent fuel in the open and closed fuel cycles generates radioactive waste. The components of spent reactor fuel can either be treated as waste (in the open fuel cycle) or reprocessed (in the closed fuel cycle). At this time, the United States has a open nuclear fuel cycle. This means that once nuclear fuel has been used it will be stored in ponds or casks. Other countries close their fuel cycles. . In other words, the waste from fuel rods is extracted and reprocessed to be reused again for fuel or for nuclear weapons.

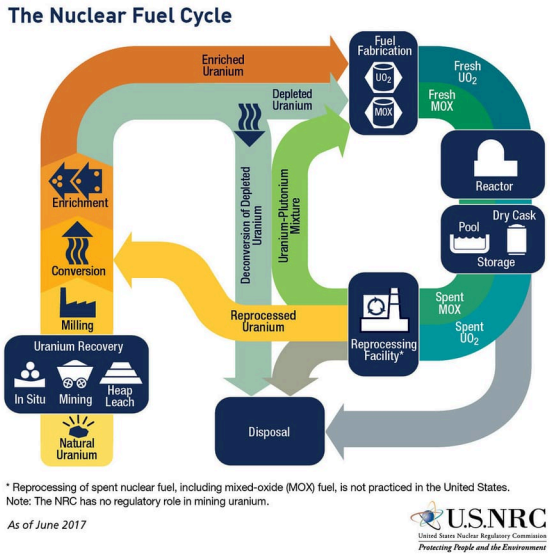


Figure 7.10.5: Image taken from: https://c1.staticflickr.com/5/4372/3...e2d0031c_b.jpg

Importance of Nuclear Waste

The controversy behind nuclear technology is due to the radioactive waste it creates. Some elements used in nuclear reactors have extremely long half-lives and must be shielded from humans and the environment for thousands of years. For example, plutonium-239, an isotope used in the production of nuclear weapons, has a half-life of 24,200 years while uranium-235 Hiroshima has a half-life of 700 million years. These elements emit large quantities of radioactivity that is extremely dangerous. Too much exposure can be followed by Acute Radiation Syndrome (ARS), which includes skin burns, nausea, vomiting, and eventually death within days if the exposure and dosage of radiation is high.

All Sources of Nuclear Waste

There are currently a number of nuclear productions that result in radioactive waste.

1. Nuclear Weapons: From the weapons to the tools and machinery used in its production, proportional amounts of radioactivity can be found in all of them. After their use, these contaminated items must be disposed of while the radioactivity slowly degrades.
2. Medicine and Research Applications: X-rays and other disease detecting technology in the medical field also consist of radiation albeit in a less harmful amount. For example, technetium-99m is an isomer that can be consumed to allow doctors to take images of the body's process. However, even the syringes contribute to the problem of nuclear waste
3. Nuclear Power: Most radioactive waste comes from the nuclear power plants situated around the world. There is a 20-30 ton waste that comes from each nuclear reactor every month it is in use.
4. Agriculture: Nuclear power is also used in eliminating bacteria through the disruption of their genetic structure. This insures that they can no longer proliferate and grow in the food.

Classification of Nuclear Waste

There are three different types of radioactive waste that result from nuclear applications in the United States:

1. *Low Level Waste* (LLW) – waste that is usually results from medicine or other industrial uses such as tools, rags, medical tubes, protective clothing, and others. There are three facilities in the U.S. that handle low-level waste by land disposal. This waste is shipped to four compact states in the country.



Figure 7.10.6: Locations that receive low level nuclear waste. These include: **Richland, WA, Clive, UT, Barnwell, SC, and Andrews TX.** Image taken from: https://www.nrc.gov/waste/llw-dispos...locations.html

2. High Level Waste (HLW) – is the highly radioactive waste resulting from spent nuclear fuel from production or power reactors, as well as from the chemical processing of spent nuclear fuel and irradiated target assemblies Nuclear weapon components consisting of Pu-239 or U-235 are also classified as being high level waste. Because of the level of contamination, the waste must stored for long periods of time at some sort of geological repository. It is not uncommon for spent nuclear fuel to be stored underwater. **HLW is currently being stored at the ninety-nine reactors sites, Hanford, WA, and Savannah River Site (Aiken, SC)**

3. Uranium Mill Tailings – this category of waste comes from remains after extracting uranium from its natural ore. Large amounts are currently left out in the open of abandoned mining sites. However, if it remains uncovered and not disposed of correctly, the waste can mix with the sand and travel to water sources, polluting the environment to a great extent.

Weapons Grade Plutonium Disposal

The highly radioactive liquid waste from reactors used to produce plutonium for the nuclear weapons of the United States is stored in tanks at the Hanford, Washington, and Savannah River, South Carolina. The Hanford site manages the largest volume of high-level waste, but the Savannah River site contains more total radioactivity. At Hanford, high-level waste alkaline liquid, salt cake, and sludge are stored in 149 single-shell and 28 double-shell underground tanks, while Savannah River has 51 tanks. These tanks contain approximately 88 million gallons of liquid, which is not only radioactive but also chemically toxic. The composition of the liquid varies from tank to tank. These facilities produced a combined total of 120 tons of plutonium for 20,000 nuclear warheads.

Management of Nuclear Waste

1. Deep Geological Repository: This method is currently still being tested and facilities are being built. The idea is to seal the radioactive waste into special casks and to deposit them hundreds meters deep into a geographically stable area to allow it to decay. There is currently only one facility in the U.S. but a few more located around the globe.
2. Transmutation: at the moment researchers are searching for a method to transmute the dangerous material into something less harmful, making it easier to dispose.
3. Re-use: Currently, there is a Canadian patent to filter the waste and reuse whatever material that has not been altered. However, this practice requires a substantial amount of energy and is still in the very early stages of development.
4. Space disposal: This has the best alternative to geological repositories because the waste can no longer harm the environment on earth. Of course, there are significant consequences of shooting radioactive waste into space such as damaging space shuttles and pollution. At the same time, space disposal would have to be agreed upon internationally.

Possible HLW Storage in the United States

In 1982, the Nuclear Waste Policy Act required the Department of Energy (DOE) to remove spent fuel from all United States' reactor sites by 1988. Unfortunately, this has not occurred yet. As a result, nuclear facilities have been awarded a total of five billion dollars (2015) to keep their waste fuel on campuses. These monies originated from a taxpayer fund.

In 1987, the DOE was encouraged to investigate an area that could house significant amounts of HLW radioactive waste. By the year 2001, the Department of Energy had spent nine billion federal dollars on their research. They selected Yucca Mountain, located about 100 miles northwest of Las Vegas, Nevada. This chosen site is arid, geographically stable, and remote. In addition, the waste would be store 1000 feet below dry rock and 1000 feet above the water table. It was hoped that this national geological repository would receive both high-level spent nuclear fuel from civilian power plants and defense-related activities (Figure 7.10.8). T

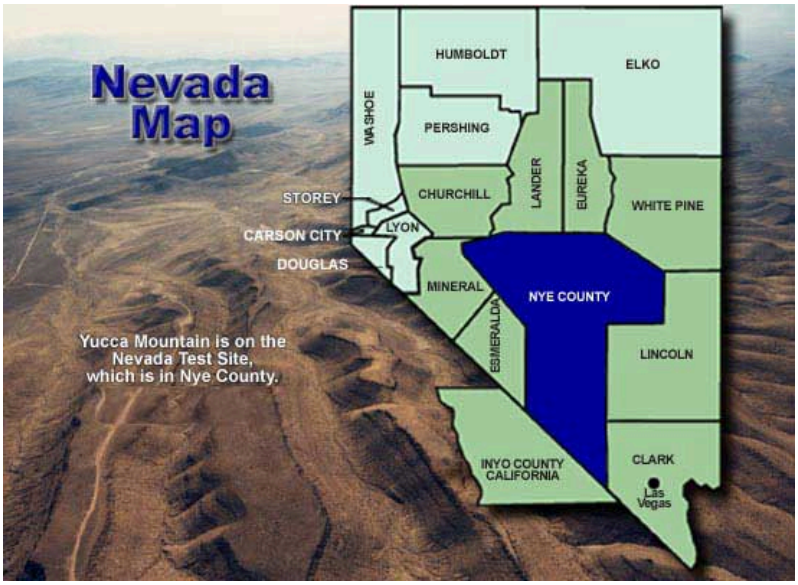


Figure 7.10.7: Yucca Mountain Site in Nevada. Courtesy of the U.S. Department of Defense

As of 2017, all high level waste is being stored at 131 different locations. Thirty-nine states and sixty-six reactor sites utilize above and underground storage. Already, more than 3000 shipments of used reactor fuel has crossed over 1.7 million miles nationwide by tractor trailer or railcar.

Funding for the construction of Yucca Mountain was awarded by 2002. From 1983-2006, **13.54 billion dollars** was spent on the study and construction of Yucca Mountain. In 2007, the NRC estimated the lifetime cost (this would include construction,transportation, and ultimate decommissioning of site) of Yucca Mountain would be **96.18 billion** dollars. Federal funding for Yucca Mountain ended in 2011 for various reasons.

If Yucca were to start receiving waste, a railroad would need to be completed near the repository. The DOE estimates that an average of 175 rail and truck shipments annually could transport the 65,000 metric tons of waste. At this rate, this process would take twenty-four years.

Waste from Reprocessing Spent Fuel – The Closed Fuel Cycle

Reducing weapons grade Pu-239 has been a goal of both the United States and Russia. In 2000, both countries signed an agreement entailing that each country would dispose of 34 metric tons (1 metric ton = 2,205 pounds) of weapons grade plutonium. One approach to this would be to incorporate Pu-239 into reactor fuel. Current United States reactors would need to be altered for this alternative fuel. Pure weapon's grade Pu-239 would not be placed in a nuclear reactor. Instead, the material would undergo the MOX (mixed oxide) process. Excess plutonium (from weapons) can be combined with LEU (in oxide form) and placed in an altered reactor. This would decrease plutonium supplies as well as reduce required U-235 enrichment.

The MOX process can be applied to used U-235 fuel rods. The fission process process transmutates U-238 into Pu-239. If fuel

rods are extracted safely from reactor, the plutonium can be isolated and combined into MOX fuel. Depending upon reactor design, MOX fuel can contain between 5 to 10% weapons grade plutonium Purified U-235 levels of MOX fuel would remain under 5%.

From 2005-2008, Duke Energy Corporation ran MOX in Catawba Nuclear Unit One reactor. This particular unit is located in York, South Caroline (18 miles south of Charlotte, NC). Protestors marched in opposition of MOX being used. In addition, the United States did not have a industrial plant to produce MOX. This fuel was fabricated by the company, AERVA MOX Services which is located in France. The United States would ship weapons grade plutonium out of Charleston Harbor across the ocean to this country. AERVA MOX Services would ship the fuel back to Charleston Harbor and tractor trailers would transport MOX to Catawba.

In 2007, the United States began construction on MOX facility at Savannah River Site (Aiken, South Carolina). Shaw AREVA MOX services consortium will build a Mixed Oxide Fuel Fabrication Facility (MFFF) to turn 3.5 tons/year of weapons grade plutonium into 150 MOX fuel assemblies. This company was awarded $2.7 billion for construction costs. At this time (2017), the United States' MOX facility is still being built.

Figure 7.10.8: A 2010 image of the United States' MFFF at Savannah River Site. As of this time, 62% of the facility was completed. Image taken from: https://c1.staticflickr.com/5/4131/5...33110d2c_b.jpg

MOX has been used by many countries. Most of the fuel is produced by France and the United Kingdom. Commercial (LW) reactors in France, Germany, Belgium, Switzerland, and the United Kingdom have used MOX for several years.

Study Questions 7.10.1

1. What are the two ways nuclear reactor facilities store used fuel rods?
2. What is HLW and LLW? How would radioactive medical waste, nuclear fuel waste, and nuclear bomb waste be classified?
3. Can nuclear waste be chemically altered to reduce half-life?
4. At this time, does the United States have an open or closed nuclear fuel cycle? Explain your answer.
5. In the United States, where is the LLW stored?
6. In the United States, where is reactor waste stored. Where is bomb waste stored?
7. What is MOX? What is the purpose of using MOX?
8. Does the United States use MOX fuel?
9. How did the United States obtain MOX fuel?
10. Where is MOX facility being built in the United States?
11. What is Yucca Mountain? Where is located? Why was it selected for nuclear waste storage? Is it currently receiving nuclear waste? How would the waste be transported to this destination if it were to open?

Answer

1. Nuclear fuel rods are first stored in a cooling pool above ground. After a couple of years, the cooled rods can be transferred to a dry, cask system.

2. HLW = high level nuclear waste and LLW = low level nuclear waste. Medical waste would be classified as being low level. Bomb and reactor waste are both high level. High level usually contains plutonium and/or uranium.

3. No, you cannot chemically alter a nucleus. Half-life cannot be changed by chemical means. One would have to transmute the waste to alter it.

4. The United States does not reprocess used fuel rods. This would be mean the fuel cycle is still open and rods must be stored. Other countries extract plutonium from fuel rods and use it for fuel. These countries would have a closed nuclear fuel cycle.

5. LLW is stored at four compact states: South Carolina, Texas, Utah, and Washington State.

6. Reactor waste is stored at commercial reactor sites (66 areas), Hanford, WA, and Aiken, SC (Savannah River Site). Bomb waste is stored at Hanford and Savannah River Site (SRS).

7. MOX stands for mixed oxide fuel. LEU is mixed with oxygen and plutonium to make a new type of fissionable

fuel for commercial reactors. This helps reduce plutonium supplies and conserves uranium.

8. The United States did use MOX in the Catawba reactors. At this time, the United States does not use MOX fuel (we can't make it).

9. The United States shipped Pu-239 out of Charleston Harbor to France. AERVA corporation moxed fuel for us and shipped it back to Charleston. MOX was transported to the Catawba reactors (York, SC) by tractor trailer or railroad car.

10. A MOX facility is being built at SRS (Savannah River Site) in Aiken, SC.

11. Yucca is a hollowed out mountain that was designed to receive all HLW. It was selected due to climate, geography, and distance from people. It is not currently receiving waste due to governmental lawsuits. The waste would be transported from all reactors sites, Hanford, WA, and Aiken, SC by tractor trailer or railroad car (no planes!!!).

References

1. Ojovan, M.I. *An introduction to nuclear waste immobilisation*. 1st. United Kingdom: Elsevier, 2005. Print.
2. Scarborough, Kate. *Nuclear Waste*. United Kingdom: Franklin Watts, 2003. Print.
3. http://www.yuccamountain.org/pdf/ocrwm_tslc-2008.pdf

Contributors

- Frank A. Settle (Washington and Lee University)
- **Complete Bibliographies on Nuclear Waste from the Alsos Digital Library for Nuclear Issues**
- Karen Chan (University of California, Davis)

This page titled 7.10: Storing and Transporting Nuclear Waste Materials is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.11: New Generation Nuclear Reactors

Categorizing Nuclear Reactors

As of 2017, the United States only used light water (LW) moderator thermal reactors. Fission powered both boiling water (BWR) and pressurized water (PWR) reactors. For research purposes, the United States does utilize heavy water, graphite moderator, and fast neutron reactors. The United States does not employ graphite moderated or liquid metal cooled reactors for commercial purposes.

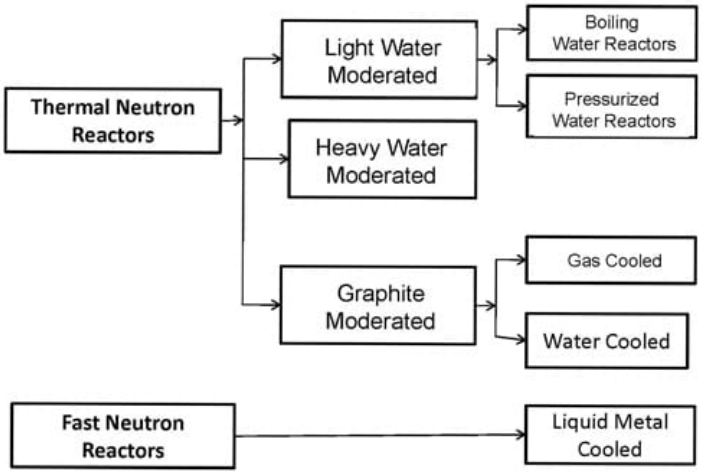


Figure 7.11.1: Types of Nuclear Reactors

Breeder Reactors

Because $^{235}_{92}\mathrm{U}$ is only 0.7 percent of naturally occurring uranium, its supply is fairly limited and could well only last for about 50 years of full-scale use. The other 99 percent of the uranium can also be utilized if it is first converted into plutonium by neutron bombardment:

$$^{238}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{239}_{94}\mathrm{Pu} + 2\,{}^{0}_{-1}\mathrm{e} \tag{7.11.1}$$

$^{239}_{94}\mathrm{Pu}$ is also fissionable, and so it could be used in a nuclear reactor as well as $^{235}_{92}\mathrm{U}$.

The production of plutonium can be carried out in a breeder reactor which not only produces energy like other reactors but is designed to allow some of the fast neutrons to bombard the $^{235}_{92}\mathrm{U}$, producing plutonium at the same time. More fuel is then produced than is consumed.

Breeder reactors present additional safety hazards to those already outlined. They operate at higher temperatures and use very reactive liquid metals such as sodium in their cooling systems, and so the possibility of a serious accident is higher. In addition the large quantities of plutonium which would be produced in a breeder economy would have to be carefully safeguarded. Plutonium is an α emitter and is very dangerous if taken internally. Its half-life is 24,000 years, and so it will remain in the environment for a long time if dispersed. Moreover, $^{239}_{94}\mathrm{Pu}$ can be separated chemically (not by the much more expensive gaseous diffusion used to concentrate $^{235}_{92}\mathrm{U}$ from fission products and used to make bombs. Such a material will obviously be attractive to terrorist groups, as well as to countries which are not currently capable of producing their own atomic weapons.

Today many nations are considering an expanded role for nuclear power in their energy portfolios. This expansion is driven by concerns about global warming, growth in energy demand, and relative costs of alternative energy sources. In 2008, 435 nuclear

reactors in 30 countries provided 16% of the world's electricity. In January 2009, 43 reactors were under construction in 11 countries, with several hundred more projected to come on line globally by 2030.



Figure 7.11.2: A Nuclear Power Plant. Dukovany Nuclear Power Station

Nuclear reactors produce energy through a controlled fission chain reaction. While most reactors generate electric power, some can also produce plutonium for weapons and reactor fuel. Power reactors use the heat from fission to produce steam, which turns turbines to generate electricity. In this respect they are similar to plants fueled by coal and natural gas. The components common to all nuclear reactors include a fuel assembly, control rods, a coolant, a pressure vessel, a containment structure, and an external cooling facility.

The speed of the neutrons in the chain reaction determines the reactor type (Figure 7.11.1). Thermal reactors use slow neutrons to maintain the reaction. These reactors require a moderator to reduce the speed of neutrons produced by fission. Fast neutron reactors, also known as fast breeder reactors (FBR), use high speed, unmoderated neutrons to sustain the chain reaction.

***Figure** 7.11.3**:** Types of Nuclear Reactors*

Thermal reactors operate on the principle that uranium-235 undergoes fission more readily with slow neutrons than with fast ones. Light water (H2O), heavy water (D2O), and carbon in the form of graphite are the most common **moderators**. Since slow neutron reactors are highly efficient in producing fission in uranium-235, they use **fuel assemblies** containing either natural uranium (0.7% U-235) or **slightly enriched uranium** (0.9 to 2.0% U-235) fuel. Rods composed of neutron-absorbing material such as cadmium or boron are inserted into the fuel assembly. The position of these **control rods** in the reactor core determines the rate of the fission chain reaction. The **coolant** is a liquid or gas that removes the heat from the core and produces steam to drive the turbines. In reactors using either light water or heavy water, the coolant also serves as the **moderator**. Reactors employing gaseous coolants (CO_2 or He) use graphite as the moderator. The **pressure vessel**, made of heavy-duty steel, holds the reactor core containing the fuel assembly, control rods, moderator, and coolant. The **containment structure,** composed of thick concrete and steel, inhibits the release of radiation in case of an accident and also secures components of the reactor from potential intruders. Finally, the most obvious components of many nuclear power plants are the **cooling towers**, the external components, which provide cool water for condensing the steam to water for recycling into the containment structure. Cooling towers are also employed with coal and natural gas plants.

Reactor Fundamentals

It is important to realize that while the U-235 in the fuel assembly of a thermal reactor is undergoing fission, some of the **fertile** U-238 present in the assembly is also absorbing neutrons to produce **fissile** Pu-239. Approximately one third of the energy produced by a thermal power reactor comes from fission of this plutonium. Power reactors and those used to produce plutonium for weapons operate in different ways to achieve their goals. Production reactors produce less energy and thus consume less fuel than power reactors. The removal of fuel assemblies from a production reactor is timed to maximize the amount of plutonium in the spent fuel (Figure 7.11.3). Fuel rods are removed from production reactors after only several months in order to recover the

maximum amount of plutonium-239. The fuel assemblies remain in the core of a power reactors for up to three years to maximize the energy produced. However it is possible to recover some plutonium from the spent fuel assemblies of a power reactor.

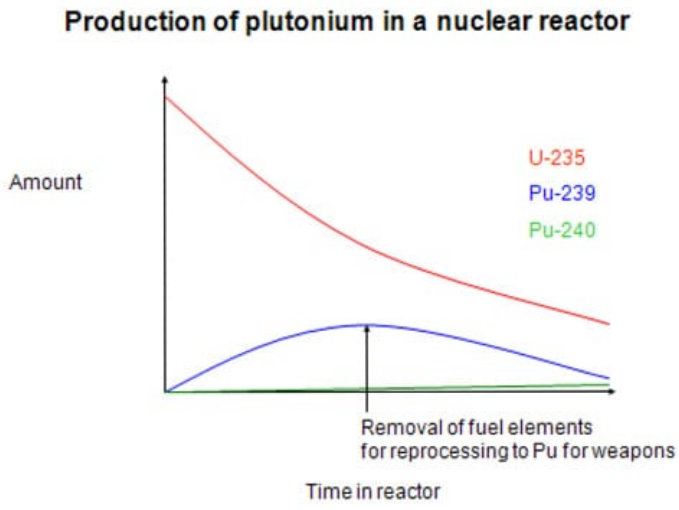


Figure 7.11.3: The Fate of Plutonium in a Thermal Reactor

The power output or capacity of a reactor used to generate electricity is measured in megawatts of electricity, MW(e). However, due to the inefficiency of converting heat into electricity, this represents only about one third of the total thermal energy, MW(t), produced by the reactor. Plutonium production is related to MW(t). A production reactor operating at 100 MW(t) can produce 100 grams of plutonium per day or enough for one weapon every two months.

Another important property of a reactor is its capacity factor. This is the ratio of its actual output of electricity for a period of time to its output if it had been operated at its full capacity. The capacity factor is affected by the time required for maintenance and repair and for the removal and replacement of fuel assemblies. The average capacity factor for U.S. reactors has increased from 50% in the early 1970s to over 90% today. This increase in production from existing reactors has kept electricity affordable.

Thermal Reactors

Currently the majority of nuclear power plants in the world are water-moderated, thermal reactors. They are categorized as either light water or heavy water reactors. Light water reactors use purified natural water (H_2O) as the coolant/moderator, while heavy water reactors employ heavy water, deuterium oxide (D_2O). In light water reactors, the water is either pressured to keep it in superheated form (in a pressurized water reactor, PWR) or allowed to vaporize, forming a mixture of water and steam (in a boiling water reactor, BWR). In a PWR (Figure 7.11.4), superheated water flowing through tubes in the reactor core transfers the heat generated by fission to a heat exchanger, which produces steam in a secondary loop to generate electricity. None of the water flowing through the reactor core leaves the containment structure.

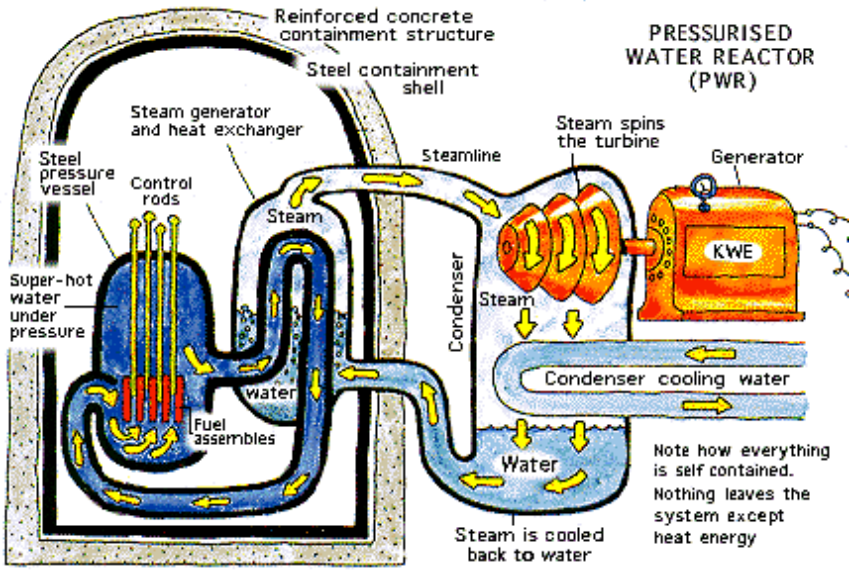


Figure 7.11.4*: Pressurized Water Reactor. Courtesy of the Uranium Information Centre*

In a BWR (Figure 7.11.5), the water flowing through the core is converted to directly to steam and leaves the containment structure to drive the turbines. Light water reactors use low enriched uranium as fuel. Enriched fuel is required because natural water absorbs some of the neutrons, reducing the number of nuclear fissions. All of the 103 nuclear power plants in the United States are light water reactors; 69 are PWRs and 34 are BWRs.

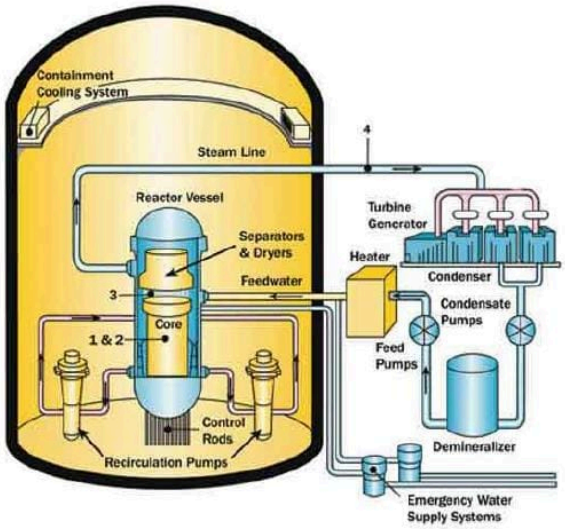


Figure 7.11.5: Boiling Water Reactor. (1) the core inside the reactor vessel creates heat, (2) a steam-water mixture is produced when very pure water (reactor coolant) moves upward through the core, absorbing heat, (3) the steam-water mixture leaves the top of the core and enters the two stages of moisture separation where water droplets are removed before the steam is allowed to enter the steam line, and (4) the steam line directs the steam to the main turbine, causing it to turn the turbine generator, which produces electricity. Courtesy of the U.S. Nuclear Regulatory Commission

Heavy water reactors use D_2O as the coolant/moderator, allowing natural, unenriched uranium to be used as the fuel. This is possible because D_2O absorbs fewer neutrons than H_2O. The heat transfer system is similar to that of the PWR, with the steam generator located within the containment structure. The heavy water reactor shown in Figure 7.11.6, known as CANDU, was developed in Canada and sold globally. The cost advantage in fuel is offset by the expense of producing D_2O by a chemical exchange process or electrolysis. The individual fuel assemblies of a heavy water reactor can be replaced without shutting down the reactor, thus eliminating the down time involved with refueling a light water reactor. However, spent fuel produced by a heavy water reactor contains more plutonium and tritium than that from light water reactors. This, coupled with the difficulty in monitoring a continuously fueled reactor, causes concerns about the proliferation of nuclear weapons. Thus, heavy water

is classified as a "sensitive material" because a nation possessing it can produce plutonium directly from natural uranium, eliminating the need for uranium enrichment.

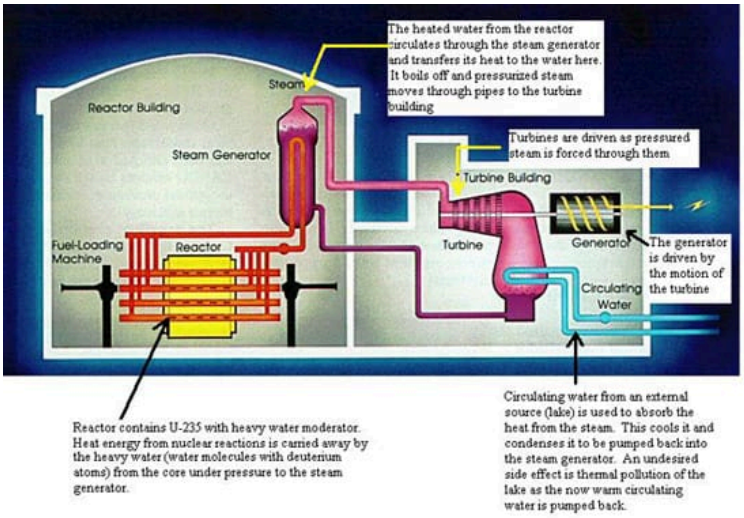


Figure 7.11.6: Heavy Water Reactor. Courtesy of Atomic Energy of Canada Limited

Another type of thermal reactor is graphite moderated and gas cooled. Twenty-six Magnox reactors (Figure 7.11.7), employing pressurized carbon dioxide as the coolant, were built in the United Kingdom but are currently being phased out of service. These gas cooled reactors have the same advantages as the heavy water reactors in that they can use natural uranium fuel and be fueled continuously.

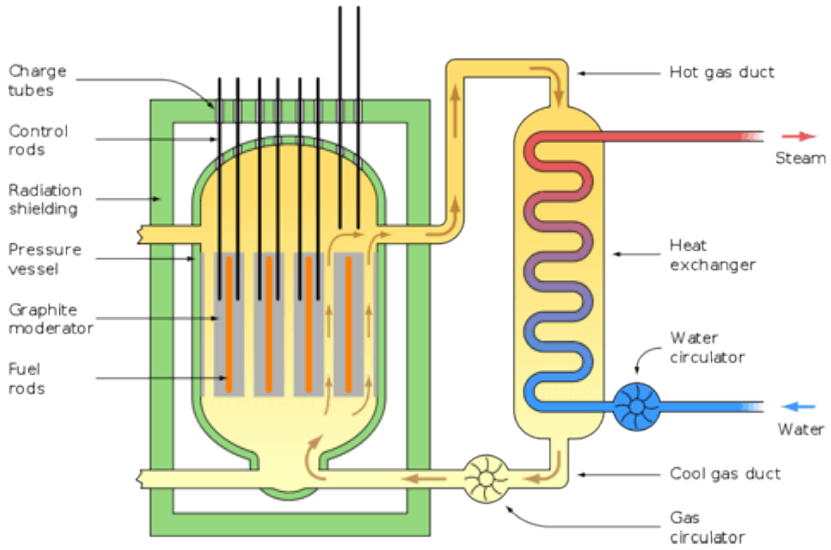


Figure 7.11.7: Gas Cooled Reactor. (Courtesy Osterreichisches Ökologie-Institut)

The pebble bed represents a new design for a gas cooled reactor (Figure 7.11.8). It uses helium as the coolant and a fuel consisting of low enriched uranium dioxide coated with silicon carbide and pyrolitic carbon encased in small graphite spheres. The graphite serves as the moderator. Fresh fuel is added at the top of the reactor and the spent fuel removed from the bottom, allowing continuous operation. The helium coolant is maintained at high temperatures and pressures, which increases the efficiency of heat transfer and power production. South Africa is leading the effort to develop this reactor technology.

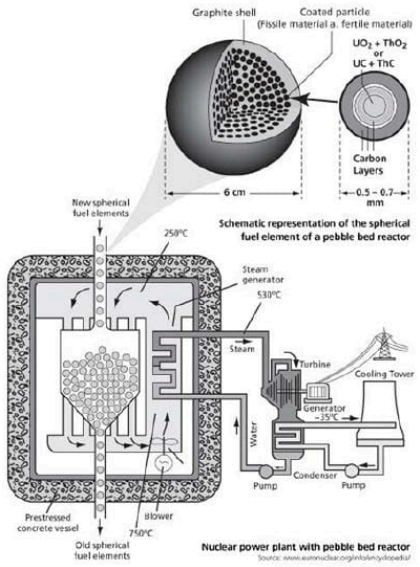


Figure 7.11.8: Pebble Bed Reactor. Courtesy of the European Nuclear Society

Fast Neutron Reactors

In contrast to thermal reactors, the neutrons in a fast neutron reactor (or fast breeder reactor, FBR) are not slowed by the presence of a moderator (Figure 7.11.9). The coolant, usually a liquid sodium or lead, is a substance that does not slow or absorb neutrons. It also has excellent heat transfer properties, which allow the reactor to be operated at lower pressures and higher temperatures than thermal reactors.

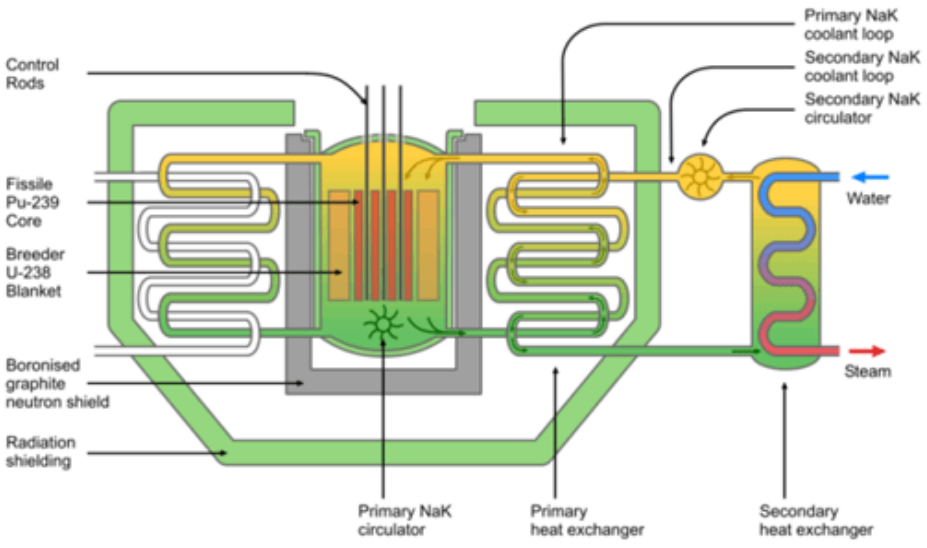


Figure 7.11.9: Fast Neutron Breeder Reactor. Courtesy of Nationmaster

An FBR is configured and operated to produce more fuel than it consumes. Fast neutrons are readily absorbed by fertile uranium-238, which then can undergo successive beta emissions to become fissile Pu-239. Thorium-232 is another fertile isotope that can absorb neutrons and produce fissile uranium-233 by beta emissions. These fissile isotopes can be reprocessed for nuclear reactor fuel or weapons. Because fast neutrons are not as efficient in producing fission as slow ones, FBRs require uranium oxide containing 20% U-235, plutonium oxide, or a mixture of these oxides, known as MOX, as fuel.

Originally FBRs were thought to be a means of extending global uranium resources by producing fissile Pu-239 or U-233 as reactor fuel. However, problems with reactor operations and material components combined with the discovery of new uranium deposits mean that FRBs are not economically competitive with existing thermal reactors. FBR research has produced technical

advances but the limiting factor continues to be the price of FBR-produced reactor fuel versus the cost of uranium fuel. FBRs are more complex than other types of reactors and also raise concerns about the proliferation of plutonium for use in nuclear weapons.

Summary of Reactor Types

Table 1 summarizes the characteristics of the reactors discussed above.

Table 7.11.1*: Reactor Characteristics*

Reactor Type	Function	Coolant	Moderator	Chemical Form of Fuel	Fuel Enrichment Level*
Thermal					
Boiling Water	electricity	light water	light water	uranium dioxide	low enriched uranium
Pressurized Water	electricity, nautical power	light water	light water	uranium dioxide	low enriched uranium
Heavy Water	electricity, plutonium production	heavy water	heavy water	uranium dioxide or uranium metal	natural, unenriched uranium
Gas Cooled Graphite Moderated	electricity, plutonium production	carbon dioxide or helium	graphite	uranium dicarbide or uranium metal	slightly enriched or natural uranium
Water Cooled Graphite Moderated	electricity, plutonium production	light water	graphite	uranium dicarbide or uranium metal	slightly enriched uranium
Pebble Bed Gas Cooled Graphite Moderated**	electricity	pressurized helium	graphite and silicon carbide	uranium dioxide or thorium dioxide	low enriched uranium
Fast Neutron					
Fast Neutron Breeder	electricity, plutonium production	molten sodium or lead	none required	various mixtures of plutonium dioxide and uranium dioxide	various mixtures of plutonium dioxide and uranium dioxide

*Percentage of U-235 isotope in the fuel compared to U-238 isotope. Natural uranium contains 0.7% U-235, slightly enriched uranium from 0.8 to 3.0% U-235, and low enriched uranium from 3.0 to 5.0% U-235.
** Currently under development.

Reactor Safety Concerns

Reactor safety has been called into question by two major accidents, one in 1979 at the Three Mile Island (TMI) nuclear power plant in Pennsylvania and the other in 1986 at Chernobyl RBMK (Reactor Bolshoy Moshchnosty Kanalnyin) in the former Soviet Union (Figure 7.11.10). The latter was the worst reactor accident in history, with 31 people dying of direct radiation poisoning and thousands more exposed to high doses of radiation over a period of nearly 20 years. Disputes still fester regarding total damage of the Chernobyl disaster, although an extensive study reported by a UN sponsored group of scientists indicates that the damage was somewhat less than earlier feared. The Chernobyl reactor used a graphite moderator and water coolant without a containment structure. Although some of the design flaws were corrected after this accident, RMBKs are considered to be the most dangerous reactors.



Figure 7.11.10***:*** *Chernobyl Nuclear Power Plant Following the 1986 Accident. Courtesy of the U.S. Department of Energy*

Although the TMI accident was much smaller in scope and resulted in no deaths, it had a profound psychological impact on the public's view of commercial nuclear power. The TMI plant was a pressurized light-water reactor, whereas Chernobyl had a reactor design similar to the one used for the reactors at Hanford, Washington, which produced plutonium for the United States nuclear weapons. Both accidents involved human error overriding safety features built into the reactor systems. Features of the new generations of reactors including passive safety systems and fewer pipes and valves are designed to minimize the potential for accidents and to limit the damaging effects should such an accident occur.

A Brief History of Nuclear Power Reactors

The first nuclear reactors were used to produce plutonium for nuclear weapons. These water-cooled, graphite reactors operated in the United States from 1944 to 1982 and were also used in the Soviet Union during the Cold War. Nuclear reactors were first used to power a submarine, the USS *Nautilus*, in 1954. That same year the Obninsk 5 megawatt nuclear power plant in the Soviet Union became the first reactor to be linked to an electrical grid. The first commercial nuclear power plants went online at Calder Hall in the United Kingdom in 1956 and Shippingport, Pennsylvania in 1957. The number of nuclear power reactors grew at a rapid rate before leveling off in the late 1980s.

In addition to the reactors for generating electricity mentioned above, there are also 220 reactors powering ships and submarines. Another 284 reactors operating in 56 nations are used for research in a variety of areas. Approximately 20% of the electricity used in the United States is generated by 103 nuclear reactors, although no new reactors have been placed in operation since the 1970s. France leads the world in the generation of electricity with more than 70% coming from nuclear sources. France, Japan, Russia, and Great Britain reprocess spent reactor fuel from commercial reactors.

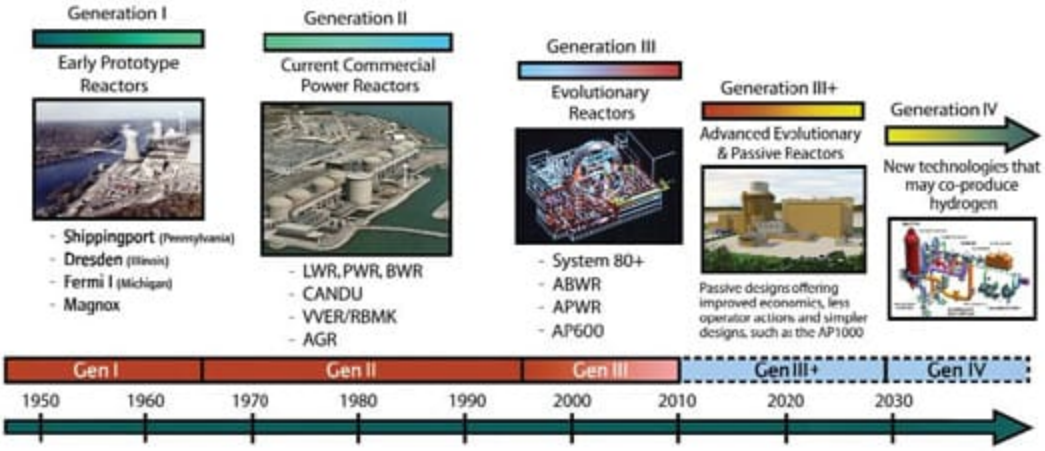


Figure 7.11.11: Evolution of Nuclear Reactors. Courtesy of the U.S. Energy Information Administration

Interest in nuclear power has increased driven in large part by concerns over energy supplies and climate charge. Reactor designs have evolved from the first generation of reactors cited above to the generation II systems that included the PWR, BWR, and Magnox reactors of the 1970s and 1980s (Figure 7.11.11). Table 7.11.2 includes some of the designs for generation III and III+ reactors. These systems are touted to be more economical to build and operate, as well as containing improved passive safety features. Figures 11 and 12 show some of the advances in reactor design for the European Pressured Water Reactor and the Westinghouse AP1000 PWR.

***Table** 7.11.1**:** Designs for Generation II and III+ Reactors*

Design	Manufacturer	Approximate Capacity (MWe)	Reactor type	Status (2008)
AP 600	Westinghouse	650	PWR	Certified
AP1000	Westinghouse	1117	PWR	Certified
ABWR	GE and Tosuiba	1371	BWR	Certified
System 80+	Westinghouse	1300	PWR	Certified
ESBWR	GE	1550	BWR	Undergoing Certification
EPR	AREVA NP	1600	PWR	Pre-certification
Pebble Bed Modular	Westinghouse, Eskom	180	High Temperature Graphite	Pre-certification

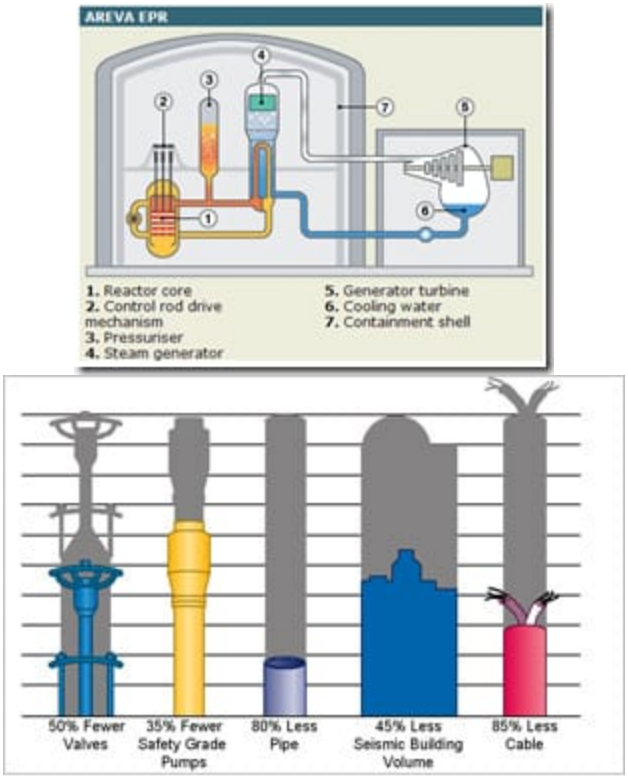


Figure 7.11.12: European Pressured Water Reactor Courtesy of Areva Figure 12 – Design Features of the Westinghouse AP1000, Courtesy of Westinghouse

Finally, designs for Generation IV reactors are being developed for deployment by 2030. These reactor designs will continue to address concerns associated with economics, safety, proliferation, waste, and reactor security.

Useful fusion reactions require very high temperatures for their initiation—about 15,000,000 K or more. At these temperatures, all molecules dissociate into atoms, and the atoms ionize, forming plasma. These conditions occur in an extremely large number of locations throughout the universe—stars are powered by fusion. Humans have already figured out how to create temperatures high enough to achieve fusion on a large scale in thermonuclear weapons. A thermonuclear weapon such as a hydrogen bomb contains a nuclear fission bomb that, when exploded, gives off enough energy to produce the extremely high temperatures necessary for fusion to occur.

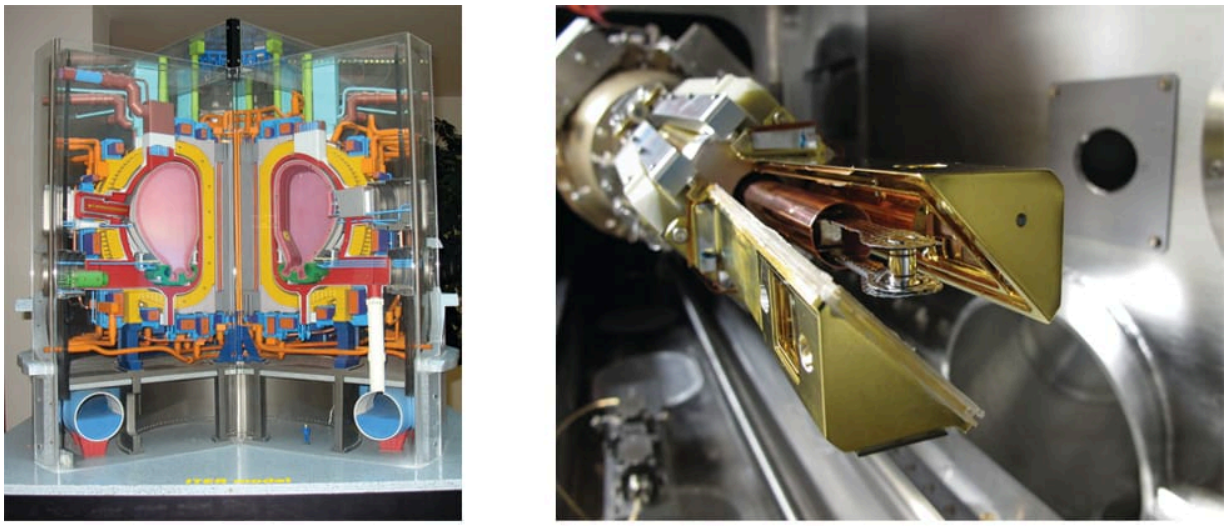


Figure 7.11.13: (a) This model is of the International Thermonuclear Experimental Reactor (ITER) reactor. Currently under construction in the south of France with an expected completion date of 2027, the ITER will be the world's largest experimental Tokamak nuclear fusion reactor with a goal of achieving large-scale sustained energy production. (b) In 2012, the National Ignition Facility at Lawrence Livermore National Laboratory briefly produced over 500,000,000,000 watts (500 terawatts, or 500 TW) of peak power and delivered 1,850,000 joules (1.85 MJ) of energy, the largest laser energy ever produced and 1000 times the power usage of the entire United States in any given moment. Although lasting only a few billionths of a second, the 192 lasers attained the conditions needed for nuclear fusion ignition. This image shows the target prior to the laser shot. (credit a: modification of work by Stephan Mosel)

Another much more beneficial way to create fusion reactions is in a fusion reactor, a nuclear reactor in which fusion reactions of light nuclei are controlled. Because no solid materials are stable at such high temperatures, mechanical devices cannot contain the plasma in which fusion reactions occur. Two techniques to contain plasma at the density and temperature necessary for a fusion reaction are currently the focus of intensive research efforts: containment by a magnetic field and by the use of focused laser beams (Figure 7.11.13). A number of large projects are working to attain one of the biggest goals in science: getting hydrogen fuel to ignite and produce more energy than the amount supplied to achieve the extremely high temperatures and pressures that are required for fusion. At the time of this writing, there are no self-sustaining fusion reactors operating in the world, although small-scale controlled fusion reactions have been run for very brief periods.Contributors

Complete Bibliography on Nuclear Reactors from Alsos Digital Library for Nuclear Issues

Contributors

- Frank A. Settle (Washington and Lee University)
- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 7.11: New Generation Nuclear Reactors is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

7.E: Nuclear Energy (Exercises)

This page titled 7.E: Nuclear Energy (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8: WATER CHEMISTRY

8.1: Water Cycle

8.2: EPA's Primary and Secondary Drinking Water Standards

8.3: Biological Contamination of Water

8.4: Chemical Contamination of Water

8.5: Water Concentration Units

8.6: Sanitation of Drinking Water

8.7: Fluoridation of Water

8.8: Other Water treatment processes

8.9: Sanitizing water while traveling

8.10: Bottled Water Basics

8.11: Love Canal Case Study

8.12: Flint River Case Study

This page titled 8: Water chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.1: Water Cycle

Learning Objectives

- Recall the percentages of fresh and salt water sources on Earth.
- Know the processes and phases of the water (hydrologic) cycle.
- Understand how air pollution affects water quality
- Recall the large consumers of water in the United States and the world.

Water merely is two atoms of hydrogen and one atom of oxygen bonded together. Despite its simplicity, water has remarkable properties. It expands when it freezes, has high surface tension (because of the polar nature of the molecules, they tend to stick together), and has a high boiling point for a liquid. Without water, life might not be able to exist on Earth, and it certainly would not have the tremendous complexity and diversity that we see.



Figure 8.1.1: This Florentine Bell Tower is one of Furman University most beloved landmarks. The original Furman Bell Tower was built in 1854 on what used to be the men's campus in Downtown Greenville. During that time,the bells rang to signal the start of classes and to celebrate victories for the athletic teams. Image courtesy of Elizabeth R. Gordon.

Distribution of Earth's Water

Earth’s oceans contain 97% of the planet’s water, so just 3% is fresh water, water with low concentrations of salts (Figure 8.1.2). The majority of the Earth's water can be classified as being **saline** (or salt containing). Most freshwater is trapped as ice in the vast glaciers and ice sheets of Greenland. A storage location for water such as an ocean, glacier, pond, or even the atmosphere is known as a **reservoir**. A water molecule may pass through a reservoir very quickly or may remain for much longer. The amount of time a molecule stays in a reservoir is known as its **residence time**.

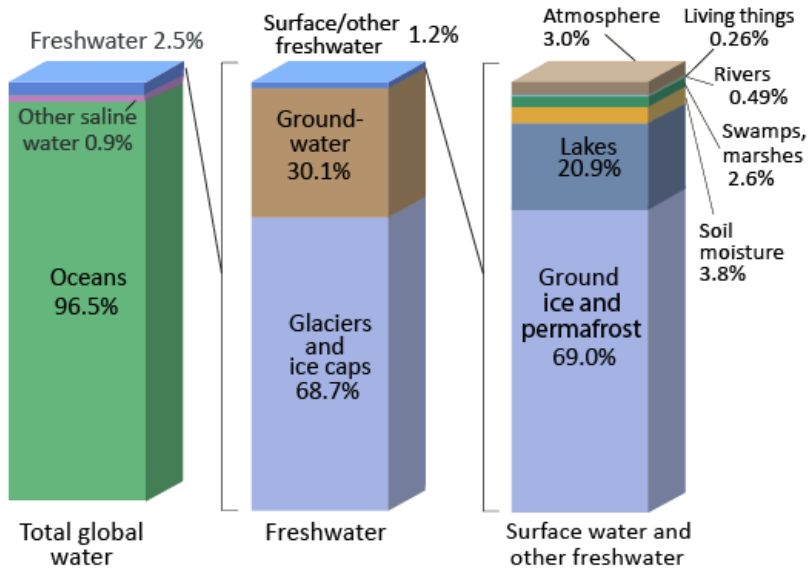


Figure 8.1.2***:*** *The distribution of Earth's water. In the first bar, notice how only 2.5% of Earth's water is freshwater - the amount needed for life to survive. The middle bar shows the breakdown of freshwater. Almost all of it is locked up in ice and in the ground. Only a little more than 1.2% of all freshwater is surface water, which serves most of life's needs.The right bar shows the breakdown of surface freshwater. Source: Igor Shiklomanov's chapter "World Freshwater resources" in Peter H. Gleick (editor), 1993 Water in Crises: A guide to the World's Freshwater Resources. The numbers are rounded.*

The Hydrologic Cycle

Because Earth's water is present in all three states, it can get into a variety of environments around the planet. The movement of water around Earth's surface is the **hydrologic (water) cycle** (Figure 8.1.3).

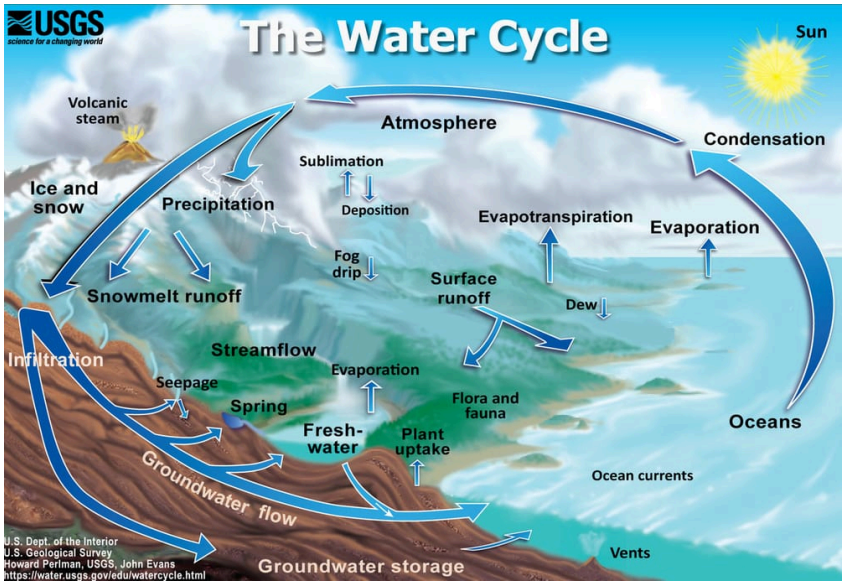


Figure 8.1.3: Because it is a cycle, the water cycle has no beginning and no end. Earth's water is always in movement, and the natural water cycle, also known as the hydrologic cycle, describes the continuous movement of water on, above, and below the surface of the Earth. Water is always changing states between liquid, vapor, and ice, with these processes happening in the blink of an eye and over millions of years.

The Sun, many millions of kilometers away, provides the energy that drives the water cycle. Our nearest star directly impacts the water cycle by supplying the energy needed for evaporation. Most of Earth's water is stored in the oceans where it can remain for hundreds or thousands of years. Water changes from a liquid to a gas by **evaporation** to become water vapor. The Sun's energy can evaporate water from the ocean surface or from lakes, streams, or puddles on land. Only the water molecules evaporate; the salts remain in the ocean or a freshwater reservoir.

The water vapor remains in the atmosphere until it undergoes **condensation** to become tiny droplets of liquid. The droplets gather in clouds, which are blown around the globe by the wind. As the water droplets in the clouds collide and grow, they fall from the sky as precipitation. **Precipitation** can be rain, sleet, hail, or snow. Sometimes precipitation falls back into the ocean and sometimes it falls onto the land surface.



Figure 8.1.4: Snow is one form of precipitation and a key transport mechanism in the water cycle. Image courtesy of Elizabeth R. Gordon

When water falls from the sky as rain it may enter streams and rivers that flow downward to oceans and lakes. Water that falls as snow may sit on a mountain for several months. Snow may become part of the ice in a glacier, where it may remain for hundreds or thousands of years. Snow and ice may go directly back into the air by **sublimation**, the process in which solid changes directly into a gas without first becoming a liquid. Although you probably have not seen water vapor sublime from a glacier, you may have seen dry ice sublime in the air.

Snow and ice slowly melt over time to become liquid water, which provides a steady flow of fresh water to streams, rivers, and lakes below. A water droplet falling as rain could also become part of a stream or a lake. At the surface, the water may eventually evaporate and reenter the atmosphere. A significant amount of water infiltrates into the ground. Soil moisture is an important reservoir for water (Figure 8.1.5). Water may seep through dirt and rock below the soil through pores infiltrating the ground to go into Earth's groundwater system. Groundwater enters **aquifers** that may store fresh water for centuries. Alternatively, the water may come to the surface through springs or find its way back to the oceans.

Water trapped in soil is important for plants to grow. Plants and animals depend on water to live and they also play a role in the water cycle. Plants take up water from the soil and release large amounts of water vapor into the air through their leaves, a process is known as *transpiration*.

It is important to note that water molecules cycle around. If climate cools and glaciers and ice caps grow, there is less water for the oceans and sea level will fall. The reverse can also happen.

Components of the Hydrologic Cycle

Most precipitation falls in the form of rain but there are other forms such as snow, hail, and sleet. Once it runs sufficiently, surface water runoff is generated when the ground is saturated or impervious. **Surface water** is a major component of the hydrological cycle and one that we interact with very regularly. It includes lakes, wetlands, stormwater **runoff** (overland flow), ponds, potholes, rivers and streams.

Streams and Rivers

A river forms from water moving from a higher altitude to lower altitude, under the force of gravity. When rain falls on the land, it either evaporates, seeps into the ground or becomes runoff (water running on the surface). When water runs on the land surface it

usually converges as it moves towards lower elevation. The converging runoff can concentrate into single channels of conveyance called creeks, stream, or rivers. Usually these start as small rill and rivulets that would join up downhill into larger streams and creeks which can also join up downstream to form even bigger rivers. The streams and rivers that join up to form a larger river are called **tributaries**, Figure 8.1.5. The land area drained by a river and all its tributaries is called a **watershed** or catchment or river basin.

The area adjacent to a river that floods frequently is a called a floodplain. **Floodplains** are areas that rivers use to temporarily store excess water during storm events and frequently contain very fertile soils. This has historically encouraged humans to move into floodplains and use them for agriculture, resulting in a reduction in the capacity of the floodplain to act as temporally storage for excess water during storm events, causing increased damaging flooding downstream. Properly functioning floodplains reduce the negative impacts of floods (by reducing severity of flood), and they assist in filtering stormwater and protecting the water quality of rivers. They also act as areas of recharge for groundwater.

Figure 8.1.5: *River systems. (A) A river with a small tributary (B) A meandering river with a mature floodplain (C) A satellite image of river system with multiple tributaries (Source USGS)*

In South Carolina, the Santee River system flows throughout the state (Figure 8.1.6). From North Carolina, the Saluda River runs through the South Carolina upstate region. The Saluda River meets up with the Broad River south of Columbia, South Carolina. The combination of these two waterways form the Congaree River. Just north of Rock Hill, South Carolina, the Catawba River enters into

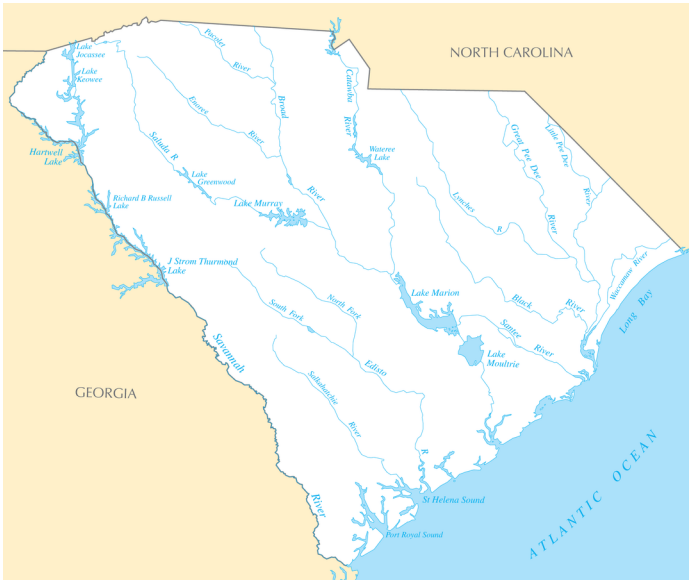


Figure 8.1.6: Major watershed of Georgia representing the main rivers in the state. (CC BY-SA 1.0; http://mapsof.net)

These rivers are very important for supplying water to the cities and populations of the states. The rivers also contain important biological communities and provide opportunities for recreation such as swimming, fishing, and white water rafting. Rivers are so important and largely control settlement patterns all over the world. Major cities, communities, factories, industries, and power stations are located along rivers. It is, therefore, very important to protect the quality and integrity of rivers all over the world.

Unfortunately, most of the rivers in the world are too polluted to support certain human activities, especially swimming, fishing, and drinking. Close to half of the rivers in the US have been deemed too polluted to support swimming and fishing. A lot of the rivers have also been channelized, dredged, or impounded by dams which have ruined their ability to support a lot of human and biological activities. It is estimated that over 600, 000 river miles have been dammed in the US. Benefits of dams to humans include providing a source of water (reservoirs and farms ponds), recreation waters and reducing local flooding. On the other side, dams can also have negative impacts on people and the environment. They can lead to increased severe flooding downstream of the dam, especially during high rain events.

The impoundments can trap stream sediments resulting in reduced sediment supply downstream as well as increased deposition behind the dam. This shift in sediments flow can disrupt and damage aquatic habitats and can increase downstream stream erosion due to lack of sediment supply. The impoundments can also prevent certain aquatic organisms from migrating either upstream or downstream, therefore reducing their range and abilities to survive environmental changes as well cutting them off

from spawning areas. Construction of dams can also result in displacement of the local people and loss of traditional lands and cultural history. Reservoirs and ponds usually form behind these impoundments.

Lakes, Reservoirs and ponds

If water flows to a place that is surrounded by higher land on all sides, a lake will form (Figure 8.1.7). A lake, pond or reservoir is a body of standing water on the land surface. When people build dams to stop rivers from flowing, the lakes that form are called reservoirs. It is estimated that over 300 million water bodies in the world are lakes, reservoirs, and ponds. Most of the Earth's lakes (about 60%) are found in Canada. Even though lakes and rivers contain less than 1% of the Earths water, the US gets over two thirds (70%) of its water (for drinking, industry, irrigation, and hydroelectric power generation) from lakes and reservoirs. Lakes are also the cornerstone of the US's freshwater fishing industry and are the backbone of the nation's state tourism industries and inland water recreational activities. (http://water.epa.gov/type/lakes/)

Figure 8.1.7: *Lake Sinclair in Baldwin and Putnam counties (Photo Credits: GCSU Hydrology Research lab)*

Wetland

A **wetland** is an area which is home to standing water for notable parts of the year, has saturated soils for a large part of the year and has plants that require large amounts of water to survive. Wetlands include swamps, marshes, and bogs. Wetlands are identified using three characteristics: soils (water-saturated soils are present), hydrology (shallow water table) and vegetation (wetland plants that are adapted to areas that are saturated with water for long periods of time). Wetlands are very important areas of biological diversity and productivity. These are also important areas where geochemical and biological cycles/ processes are consistently taking place. For instance, wetlands are considered as areas of significant carbon sequestration (storage), which impacts global climate change. They also act as filters for storm-water runoff before it enters rivers and lakes.

Oceans

As you have probably already guessed, oceans are an important component of the hydrologic cycle because they store majority of all water on Earth (about 95%). Most of the major rivers drain into them. The five oceans covering the surface of the Earth are the Atlantic, Indian, Pacific, Arctic and the Southern Ocean (Figure 8.1.8).

Figure 8.1.8: *The five oceans found on planet Earth. The Pacific Ocean is the largest. Source: https://woodshole.er.usgs.gov/projec...es/primer.html*

Approximately 90 % of the water that is evaporated into the hydrologic cycle comes from the ocean. Oceans are an important and large part of the hydrologic cycle, with lots biological diversity and many landforms. D*id you know that the average depth of the oceans is about 3.6 km with a maximum depth that can exceed 10 kilometers in areas known as ocean trenches?* The ocean is also home to many forms of life uniquely adapted to survive in this habitat. Unfortunately, humans have degraded the oceans and their life through pollution, overfishing, carbon dioxide acidification and resource exploitation. Figure 8.1.9 shows a couple of examples of human impacts on the ocean environment.

Figure 8.1.9: *Trash washed up on the beach (A) and seal tangles up and being struggled by plastic trash in the ocean (B).*

Storage and Flow

Almost 99% of the available fresh water is found below the surface as groundwater. Groundwater is not created by some mysterious processes below ground, but is part of the recycled water in the hydrologic cycle. When precipitation falls, some of the water runs off on the surface while some infiltrates into the ground. Groundwater is replenished when water moves from the surface, through unsaturated rocks or sediment (**unsaturated**), all the way down the **saturated parts** (saturated zone) in a process called infiltration and becomes groundwater (Figure 8.1.10). The top of the saturated portion is called the **water table**, which is the boundary between saturated and unsaturated zone.

Groundwater is found in **aquifers**, which are bodies of rock or sediment that store (and yield) large amounts of usable water in their pores. Aquifer productivity is controlled by **porosity** and **permeability**. Porosity is the percentage of open space in a rock or sediment body. Permeability is the ability of subsurface material to transmit fluids. Groundwater is found in the saturated zone of a rock body where all pores are filled with water. An important concept is that surface water always moves from higher elevation

to lower elevation while groundwater always moves from higher energy (hydraulic head) to lower energy.

Figure 8.1.10: *Model of groundwater system showing the different components of an unconfined groundwater system: http://water.usgs.gov/edu/earthgwaquifer.html*

Groundwater will continue to flow until it emerges as a spring, or discharges into surface water bodies on the land or in the ocean. To utilize groundwater, we drill holes (wells) into the ground and pump the water out.

Water Consumption

***Table** 8.1.2**: Water Use in the United States and Globally*

Use	United States	Global
Agriculture	34%	70%
Domestic (drinking, bathing)	12%	10%
Industry	5%	20%
Power plant cooling	49%	small

Contributors

- Template:ContribCandelaGeo
- Elizabeth R. Gordon (Furman University)

This page titled 8.1: Water Cycle is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

8.2: EPA's Primary and Secondary Drinking Water Standards

The 1974 Act

Up until 1974, public drinking water supplies in the United States were monitored and regulated by state and local authorities. Lists of contaminants with their various concentrations could vary from state to state. As the chemical industry grew, these same state agencies noted the presence of more and new organic chemicals in public water systems. In order to standardize drinking water across the country, the Environmental Protection Agency (EPA) enacted the Safe Water Drinking Act of 1974.



Figure 8.2.1: AFGE members and environmental activists participated in an event at the Washington, D.C. EPA headquarters to thank EPA workers for their work. Image used with permission (CC BY 2.0; AFGE).

The 1974 act enabled the EPA to monitor and regulate public water systems that serve over 25 people. Implementation and enforcement of drinking water standards were still performed by each state. Regarding drinking water sources (surface and ground), the EPA and state agencies protect and monitor these as well. The EPA does not regulate private wells or bottled water in the United States. Well water monitoring is the responsibility of the owner. As for bottled water, the Food and Drug Administration (FDA) oversees this commodity by using EPA drinking water standards.

Insert bottled water picture here with brief explanation of tap vs spring

The first set of drinking water standards included only 22 chemicals and/or pathogens. EPA established to major types of contaminants: primary and secondary. The first of these types (primary) of contaminants are substances (examples could include Hg, As, and U) that can be toxic in small amounts. On the other hand, secondary contaminants are less toxic species (Fe and Zn) and would include cosmetic issues (color, taste, and odor) of drinking water.

All primary contaminants have enforceable concentration values. For the majority of these pollutants, EPA lists specific limits by using the term **Maximum Contaminant Level (MCL)**. If a water supplier exceeds a given MCL for a toxin, then fines and penalties could by imposed by the EPA. A few pathogens (Giardia Lamblia and Legionella) use Treatment Technique (TT) notation rather than numeric MCL concentrations. Water that contains any amount of these pathogens must be sanitized immediately with a standardized EPA procedure.

Table 8.2.1: Exert of the EPA's Primary Drinking Water Contaminant Listing

Primary Contaminant	Classification	MCLG[1] (mg/L)	MCL or TT[2] (mg/L)	Potential Health Effects from Long-Term Exposure Above the MCL (unless specified as short-term)	Sources of Contaminant in Drinking Water
Giardia lamblia	Microorganism	zero	TT[3]	Gastrointestinal illness (such as diarrhea, vomiting, and cramps)	Human and animal fecal waste
Thallium	Inorganic	0.0005	0.002	Hair loss; changes in blood; kidney, intestine, or liver problems	Leaching from ore-processing sites; discharge from electronics, glass, and drug factories

In Table 8.2.1, a **Maximum Contaminant Level Goal (MCLG)** column is displayed for each of the two drinking water impurties. MCLG values are not enforced by the EPA. Instead, MCLG are suggested values that water suppliers should strive to meet. Every six years, the EPA reviews each primary contaminant with its MCL standard. During these times, they analyze data in regards to health risk assessment. If they choose to lower a MCL (smaller value) of a contaminant, then less health issues should occur. Unfortunately, reducing a concentration requires more technology which will cost the supplier and the consumer more money. By providing MCLG limits, the EPA encourages a water company to gradually work towards lowering a toxin's concentration.

Primary Drinking Water Contaminants

These types of toxins are classified into one of the six EPA classifications: microorganisms, disinfectants, disinfectant byproducts, inorganics, organics, or radionuclides. Primary contaminants are regulated because the have the capacity to do great harm to humans, plants, and animals. If a water distributor (must serve at least 200 homes) exceeds one of these mandated standards, then the EPA may impose a fine on the company.



Figure 8.2.2: Pretty picture. Image used with permission (CC BY; Elizabeth Gordon).

Typically, concentrations of these contaminants are listed in parts per million. Metric amounts of a part per million are milligrams of the toxin per liter of sample.

$$\text{ppm} = \frac{mg}{Liters} \tag{8.2.1}$$

For extremely toxic substance, smaller units like parts per billion (ppb) might be used to express concentration. Metric amounts of a part per billion are micrograms per liter of sample. (or micrograms of contaminant per liter of water).

$$\text{ppm} = \frac{mg}{Liters} \tag{8.2.2}$$

mention differences between mcl and mclg. weird radionuclide units and asbestos.....

Amendment SDWA of 1986

Amendment of SDWA of 1996

This page titled 8.2: EPA's Primary and Secondary Drinking Water Standards is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.3: Biological Contamination of Water

Learning Objectives

- Identify the three categories of pathogens that produce water-borne diseases.
- Provide one example of each pathogen, typical sources, and health problems.
- Summarize how water-borne diseases can be prevented and treated.
- Recall areas of the world where water is not properly sanitized.
- List foods and beverages that need to be avoided when using water that is not sanitized.
- Know how to sanitize water when traveling or camping.
- Know the limitations of the LifeStraw and Steripen.
- Recall how long to boil water if under an advisory.
- Realize that boiling does not remove inorganics, metals, or organics.
- Provide examples of chemicals used to sanitize water (refer to CDC chart below).
- Know the limitations of chemical sanitation (cryptosporidium is hard to kill this way).
- Besides boiling, discuss the chemical sanitation process of how to sanitize water at home (EPA method below).

Water Borne Diseases

If water is not cleaned properly, residents of a community can contract various illnesses. Viruses, bacteria, and parasites can enter a water supply unknowingly. All three of these species are pathogenic, or disease-causing. In the United States, the Environmental Protection Agency (EPA) quantifies and monitors these pathogens in water systems that serve at least 200 consumers. The EPA classifies these contaminants as being microorganisms (even though a biologist might dispute this terminology).

YouTube Video

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

View this YouTube Video online

Video 8.3.1***:*** *The Coalition for Global Community Health is working within existing social structures in Belén, Iquitos, Peru to uphold the human rights of the community members. We speak directly with community members in an open forum to learn about their needs, desires, and ideas for creating an opportunity to change their communities for the better*

The picture below shows how a person might contract Giardiasis from *giardia,* a parasite. This particular pathogen can live in a body up to six months. Once detected through a stool sample, a patient can be prescribed specific antibiotics like Flagyl to treat the infection. Unfortunately, there is no vaccine for preventing Giardiasis.

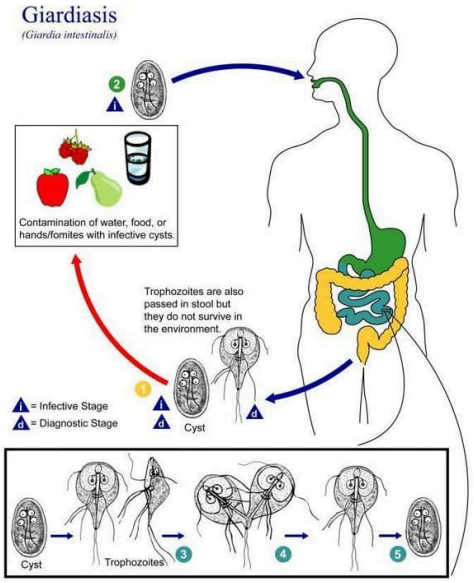


Figure 8.3.1: The life cycle of Giardia.

The table below shows water-borne diseases that can result from viruses, bacteria, and parasites. In some cases, vaccines are available. When eating, drinking, or swimming, it is important to be aware of how you could be affected by these pathogens. Sanitation of drinking water with chlorine-based compounds reduces the power of these pathogens. In addition, proper handling of foods and beverages could reduce your risk of developing one or more of the following health problems.

Table 8.3.1: Pathogens that cause waterborne illnesses.

Pathogen Name	Pathogen Type	Source	Health problem	Prevention/Treatment
Giardia	Parasite	Fecal contamination and uncooked food	Vomiting, diarrhea, and cramps	Medication afterward
Cryptosporidium	Parasite	Fecal contamination	Vomiting, diarrhea, fever, and cramps	Medication afterward
Typhoid	Bacteria	Fecal contamination	High fever, stomach pains, headache, and rash	Vaccination/Antibiotics
E. coli	Bacteria	Fecal contamination	Diarrhea and cramps	Fluids
Legionella	Bacteria	Found naturally in heated water	Causes Legionnaires (a type of pneumonia)	Medications afterward
Cholera	Bacteria	Related to fecal contamination or undercooked or raw shellfish	Diarrhea	Vaccine/Rehydration, antibiotics, and Zinc
Hepatitis A	Virus	Contaminated food and water	Vomiting, dark urine, and yellowing of the eyes.	Vaccination/Fluids
Polio	Virus	Fecal contamination	Flu symptoms, paralysis	Vaccination

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

Municipal water systems in the United States use chlorine compounds, ozonation or ultraviolet irradiation to sanitize the drinking water. When camping or hiking, it is important to know that river, stream, and lake water is not sanitized. Water that appears clear looking could contain any of the above-mentioned pathogens. When traveling or enjoying the outdoors, you might want to invest in a portable water purification system. Two of the most common of these are the LifeStraw and Steripen. LifeStraws uses physical filtration with no chemical treatment. This device is effective in removing bacteria and parasites. A personal device can filter around 1000 liters of water. A LifeStraw family unit can filter up to 18,000 liters of water. Unfortunately, most LifeStraws cannot remove metals, salt, viruses, or organic chemicals. The average price for a personal unit is $25-$40.

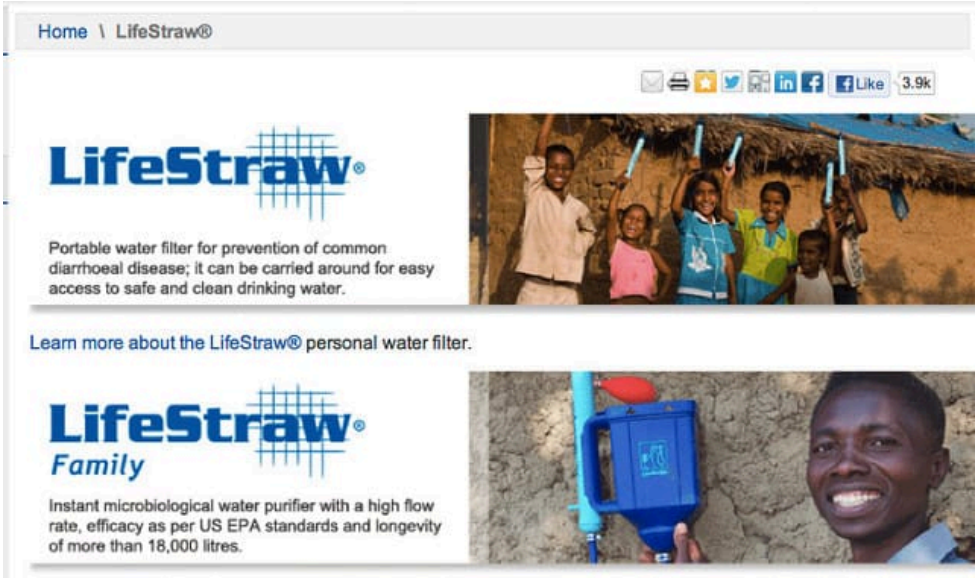


Figure 8.3.2: Images taken from: https://www.flickr.com/photos/caseorganic/8010028560

SteriPens utilize UVC radiation to kill parasites, viruses, and bacteria. Cost of a personal SteriPen device range from $80 to $100. SteriPens can treat up to 8000 liters of water. Like the LifeStraw, these apparatus cannot remove organic, inorganic, or metallic contaminants.

YouTube Video

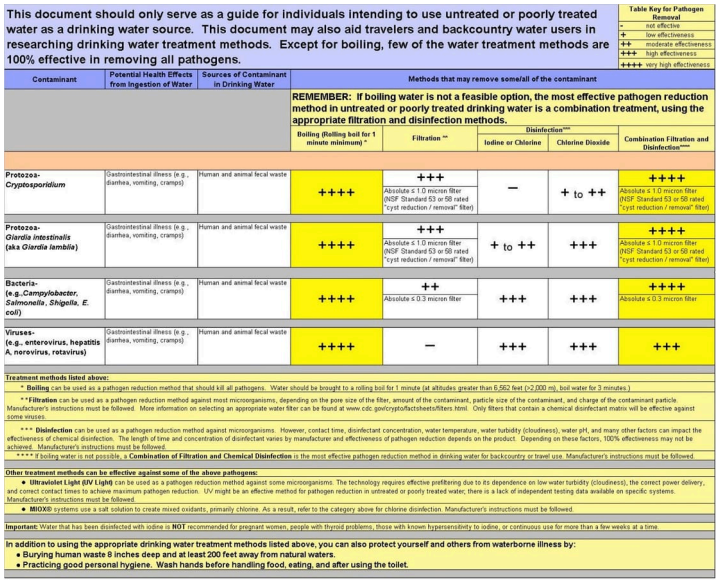
This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

***Video** 8.3.2:*

Other Methods of Sanitizing Drinking Water

Water advisories may be declared if sanitation is compromised. In these situations, boiling, halogen disinfection, or filtration could be utilized to make water **potable** (drinkable). The Centers for Disease Control has developed a document to compare these different methods of sanitation. Look below to see the most effective ways to make water safe.



This document should only serve as a guide for individuals intending to use untreated or poorly treated water as a drinking water source. This document may also aid travelers and backcountry water users in researching drinking water treatment methods. Except for boiling, few of the water treatment methods are 100% effective in removing all pathogens.

Table Key for Pathogen Removal

-	not effective
+	low effectiveness
++	moderate effectiveness
+++	high effectiveness
++++	very high effectiveness

Contaminant	Potential Health Effects from Ingestion of Water	Sources of Contaminant in Drinking Water	Methods that may remove some/all of the contaminant				
			REMEMBER: If boiling water is not a feasible option, the most effective pathogen reduction method in untreated or poorly treated drinking water is a combination treatment, using the appropriate filtration and disinfection methods.				
			Boiling (Rolling boil for 1 minute minimum) *	Filtration **	Disinfection*** Iodine or Chlorine	Disinfection*** Chlorine Dioxide	Combination Filtration and Disinfection****
Protozoa- *Cryptosporidium*	Gastrointestinal illness (e.g., diarrhea, vomiting, cramps)	Human and animal fecal waste	++++	+++ Absolute ≤ 1.0 micron filter (NSF Standard 53 or 58 rated "cyst reduction / removal" filter)	–	+ to ++	++++ Absolute ≤ 1.0 micron filter (NSF Standard 53 or 58 rated "cyst reduction / removal" filter)
Protozoa- *Giardia intestinalis* (aka *Giardia lamblia*)	Gastrointestinal illness (e.g., diarrhea, vomiting, cramps)	Human and animal fecal waste	++++	+++ Absolute ≤ 1.0 micron filter (NSF Standard 53 or 58 rated "cyst reduction / removal" filter)	+ to ++	+++	++++ Absolute ≤ 1.0 micron filter (NSF Standard 53 or 58 rated "cyst reduction / removal" filter)
Bacteria- (e.g.,*Campylobacter, Salmonella , Shigella, E. coli*)	Gastrointestinal illness (e.g., diarrhea, vomiting, cramps)	Human and animal fecal waste	++++	++ Absolute ≤ 0.3 micron filter	+++	+++	++++ Absolute ≤ 0.3 micron filter
Viruses- (e.g., enterovirus, hepatitis A, norovirus, rotavirus)	Gastrointestinal illness (e.g., diarrhea, vomiting, cramps)	Human and animal fecal waste	++++	–	+++	+++	+++

Treatment methods listed above:

* **Boiling** can be used as a pathogen reduction method that should kill all pathogens. Water should be brought to a rolling boil for 1 minute (at altitudes greater than 6,562 feet (>2,000 m), boil water for 3 minutes.)

** **Filtration** can be used as a pathogen reduction method against most microorganisms, depending on the pore size of the filter, amount of the contaminant, particle size of the contaminant, and charge of the contaminant particle. Manufacturer's instructions must be followed. More information on selecting an appropriate water filter can be found at www.cdc.gov/crypto/factsheets/filters.html. Only filters that contain a chemical disinfectant matrix will be effective against some viruses.

*** **Disinfection** can be used as a pathogen reduction method against microorganisms. However, contact time, disinfectant concentration, water temperature, water turbidity (cloudiness), water pH, and many other factors can impact the effectiveness of chemical disinfection. The length of time and concentration of disinfectant varies by manufacturer and effectiveness of pathogen reduction depends on the product. Depending on these factors, 100% effectiveness may not be achieved. Manufacturer's instructions must be followed.

**** If boiling water is not possible, a **Combination of Filtration and Chemical Disinfection** is the most effective pathogen reduction method in drinking water for backcountry or travel use. Manufacturer's instructions must be followed.

Other treatment methods can be effective against some of the above pathogens:

- **Ultraviolet Light (UV Light)** can be used as a pathogen reduction method against some microorganisms. The technology requires effective prefiltering due to its dependence on low water turbidity (cloudiness), the correct power delivery, and correct contact times to achieve maximum pathogen reduction. UV might be an effective method for pathogen reduction in untreated or poorly treated water, there is a lack of independent testing data available on specific systems. Manufacturer's instructions must be followed.
- **MIOX®** systems use a salt solution to create mixed oxidants, primarily chlorine. As a result, refer to the category above for chlorine disinfection. Manufacturer's instructions must be followed.

Important: Water that has been disinfected with iodine is **NOT** recommended for pregnant women, people with thyroid problems, those with known hypersensitivity to iodine, or continuous use for more than a few weeks at a time.

In addition to using the appropriate drinking water treatment methods listed above, you can also protect yourself and others from waterborne illness by:

- **Burying human waste 8 inches deep and at least 200 feet away from natural waters.**
- **Practicing good personal hygiene. Wash hands before handling food, eating, and after using the toilet.**

Halogen based solids can be purchased at most home improvement stores, pool stores, Walmart, and Target. These chemicals will alter taste and smell of water. These chemicals are also needed when properly sanitizing a pool or a hot tub. Halogen tablets are quite cheap and an effective way for sanitizing any type of water.

EPA: EMERGENCY DISINFECTION OF DRINKING WATER

In an emergency situation where regular water service has been interrupted – like a hurricane, flood, or water pipe breakage – local authorities may recommend using only bottled water, boiled water, or disinfected water until regular water service is restored. The instructions below show you how to boil and disinfect water to kill most disease-causing microorganisms that may be present in the water. However, boiling or disinfection will not destroy other contaminants, such as heavy metals, salts, and most other chemicals.

> ONLY USE WATER THAT HAS BEEN PROPERLY DISINFECTED FOR DRINKING, COOKING, MAKING ANY PREPARED DRINK, WASHING DISHES, AND FOR BRUSHING TEETH.

Use **bottled wate**r or water you have properly prepared and stored as an emergency water supply.

Boil water, if you do not have bottled water. Boiling is sufficient to kill pathogenic bacteria, viruses and protozoa (WHO, 2015).

- If water is cloudy, let it settle and filter it through a clean cloth, paper towel, or coffee filter.
- Bring water to a rolling boil for at least one minute. At altitudes above 5,000 feet (1,000 meters), boil water for three minutes.
- Let the water cool naturally and store it in clean containers with covers.
- To improve the flat taste of boiled water, add one pinch of salt to each quart or liter of water, or pour the water from one clean container to another several times.

Disinfect water using household bleach, if you can't boil water. Only use regular, unscented chlorine bleach products that are suitable for disinfection and sanitization as indicated on the label. The label may say that the active ingredient contains 6 or 8.25% of sodium hypochlorite. Do not use scented, color safe, or bleaches with added cleaners.

- If water is cloudy, let it settle and filter it through a clean cloth, paper towel, or coffee filter.
- Locate a clean dropper from your medicine cabinet or emergency supply kit.
- Locate a fresh liquid chlorine bleach or liquid chlorine bleach that is stored at room temperatures for less than one year.
- Use the table on the next page as a guide to decide how much bleach you should add to the water, for example, add 8 drops of 6 % bleach or 6 drops of 8.25% bleach to each gallon of water. Double the amount of bleach if the water is cloudy, colored, or very cold.
- Stir and let stand for 30 minutes. The water should have a slight chlorine odor. If it doesn't, repeat the dosage and let

stand for another 15 minutes before use.
- If the chlorine taste is too strong, pour the water from one clean container to another and let it stand for a few hours before use.

Volume of Water	Amount of 6% Bleach to Add†	Amount of 8.25% Bleach to Add†
1 quart/liter	2 drops	2 drops
1 gallon	8 drops	6 drops
2 gallons	16 drops (1/4 tsp.)	12 drops (1/4 tsp.)
4 gallons	1/3 tsp.	1/4 tsp.
8 gallons	2/3 tsp.	1/2 tsp.
†Bleach may contain 6 or 8.25% sodium hypochlorite		

ADDITIONAL WATER GUIDANCE FOR EMERGENCIES

Prepare and store an emergency water supply. Visit the Federal Emergency Management Agency (FEMA) website for additional guidance on preparing and storing an emergency water supply.

Look for other sources of water in and around your home. Although bottled water is your best choice, you may be able to find other sources of water by melting ice cubes or draining your hot water tank or pipes. You can also use river or lake water. It is generally better to use flowing water than still, stagnant water. However, do not use water with floating material in it or water that has a dark color or questionable odor. Regardless of the source, treat the water by following the instructions above. If you have a well on your property that has been flooded, make sure to disinfect and test the well water after the flood. Contact your state or local health department for advice or go here.

Consider how the water looks and how to filter it if needed. Disinfection does not work as well when water is cloudy or colored. If water is cloudy, let it settle. Then filter the water through a clean cloth, paper towel, or coffee filter. Store the settled and filtered water in clean containers with covers

OTHER DISINFECTION METHODS

If you don't have liquid bleach, you can use one of the other disinfection methods described below.

- *Granular calcium hypochlorite*. The first step is to make a chlorine solution that you will use to disinfect your water. For your safety, do it in a ventilated area and wear eye protection. Add one heaping teaspoon (approximately ¼ ounce) of high-test granular calcium hypochlorite (HTH) to two gallons of water and stir until the particles have dissolved. The mixture will produce a chlorine solution of approximately 5000 milligrams per liter. To disinfect water, add one part of the chlorine solution to every 100 parts of water you are treating. This is about the same as adding 1 pint (16 ounces) of the chlorine solution to 12.5 gallons of water. If the chlorine taste is too strong, pour the water from one clean container to another and let it stand for a few hours before use. CAUTION: HTH is a very powerful oxidant. Follow the instructions on the label for safe handling and storage of this chemical.
- *Common household iodine* (or "tincture of iodine"). You may have iodine in your medicine cabinet or first aid kit. Add five drops of 2% tincture of iodine to each quart or liter of water that you are disinfecting. If the water is cloudy or colored, add 10 drops of iodine. Stir and let the water stand for at least 30 minutes before use.
- *Water disinfection tablets*. You can disinfect water with tablets that contain chlorine, iodine, chlorine dioxide, or other disinfecting agents. These tablets are available online or at pharmacies and sporting goods stores. Follow the instructions on the product label as each product may have a different strength.

Safe Drinking Water Hotline: 1-800-426-4791 water.epa.gov/drink/hotline

Example 8.3.1:

1. Do all waterborne pathogens cause stomach illnesses?
2. If your community is under a water advisory, then how long should you boil tap water to kill pathogens?

3. Can Steripens use ultraviolet light to remove metals from drinking water?
4. Vaccinations and preventive medications can protect you from which pathogens?
5. When properly diluted, which common household cleaner can sanitize tap water?
6. What is the specific type of ultraviolet radiation that Steripens used to kill bacteria, viruses, and parasites? Where else does this type of radiation occur?
7. The residents of Flint, Michigan experienced heavy metal and pathogenic contamination in their tap water. If the state of Michigan provided individuals with Lifestraws, then would they have clean water for drinking and cooking purposes?

Solutions

1. Legionella bacteria causes respiratory problems. Also, hepatitis and polio viruses do not produce stomach illnesses.
2. Water should be boiled rigorously for at least one minute to kill most pathogens.
3. Metals cannot be removed by irradiating water.
4. Vaccines can be used to prevent: hepatitis, typhoid, cholera, and polio. In the United States, most children have been vaccinated from hepatitis and polio. When traveling to destinations that lack sanitized water, it would be beneficial to request prescription medications from your health care professional. Refer to the chart above regarding which diseases can be treated with medications.
5. A diluted (watered down) form of bleach (sodium hypochlorite) can be used to sanitize water. Please note the different types of bleach that can be purchased. In addition, do not use color safe or scented bleach.
6. Steripens use UVC radiation to kill water pathogens. The sun generated UVC radiation. Unlike UVA and UVB radiation, UVC does not reach the troposphere (where you live) on the earth.
7. The lifestraws mentioned on this page do not remove toxic metals from drinking water. These types of

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 8.3: Biological Contamination of Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.4: Chemical Contamination of Water

The US Safe Drinking Water Act defines the term "contaminant" as meaning any physical, chemical, biological, or radiological substance or matter in water. Therefore, the "contaminant" definition very broadly applies as being anything other than water molecules. Drinking water may reasonably be expected to contain at least small amounts of some contaminants. Some drinking water contaminants may be harmful if consumed at certain levels in drinking water while others may be harmless. The presence of contaminants does not necessarily indicate that the water poses a health risk.

Air Pollutants

Water can become contaminated at any part of the water cycle. Air pollution can affect water vapor and water liquid. Combustion sources like vehicles and power plants generate compounds like NO_xs, SO_xs, CO, CO_2, and other various inorganic and organic volatile organic compounds (VOC) species. Some of these compounds can become soluble in water. This could affect pH (or acidity) level of water surface water. Normally, the pH of water is a neutral value (or pH= 7). When NO_xs, SO_xs, CO, CO_2 enter the water cycle, then the pH level is lowered below 7.0. If these gases are absorbed in rain clouds, then acid rain results. Specific acids involved in acid rain are **sulfuric, nitric,** and **carbonic**. This environmental problem affects living organisms and building materials. Acid solutions can corrode metals and make them soluble as well.

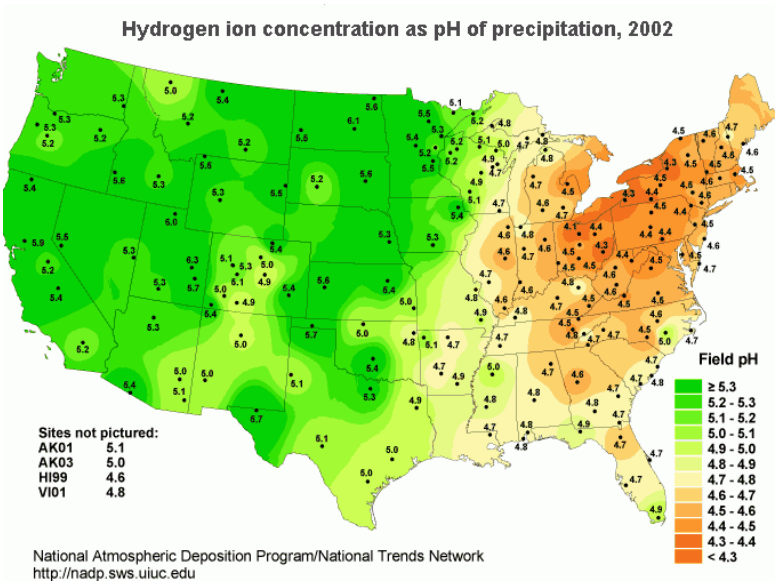


Figure $8.4.6$: pH of rainwater over the continental United States.

In the United States, the northeast has the most problems with acid rain. Concentrated populations that use electrical energy and vehicles contribute greatly to the pH reduction of rainwater. Reducing gaseous output requires capping combustion sources (vehicles and power plants). Acidity in rain is measured by collecting samples of rain and measuring its pH. To find the distribution of rain acidity, weather conditions are monitored and rain samples are collected at sites all over the country (Figure $8.4.6$). The areas of greatest acidity (lowest pH values) are located in the Northeastern United States. This pattern of high acidity is caused by a large number of cities, the dense population, and the concentration of power and industrial plants in the Northeast. In addition, the prevailing wind direction brings storms and pollution to the Northeast from the Midwest, and dust from the soil and rocks in the Northeastern United States is less likely to neutralize acidity in the rain.

Acid rain from sulfuric acid

However, the majority of acid rain is accounted for by the presence of sulfuric acid ($\mathrm{H_2SO_4}$).

$$\mathrm{SO_2\,(g) + O_2\,(g) \longrightarrow SO_3\,(g)} \tag{8.4.1}$$

Sulfur dioxide reacts with water to form sulfuric acid

$$\mathrm{SO_3\,(g) + H_2O \longrightarrow H_2SO_4} \tag{8.4.2}$$

Although sulfuric acid may be produced naturally in small quantities from biological decay and volcanic activity, it is produced almost entirely by human activity, especially the combustion of sulfur-containing fossil fuels in power plants (Figure 8.4.6).

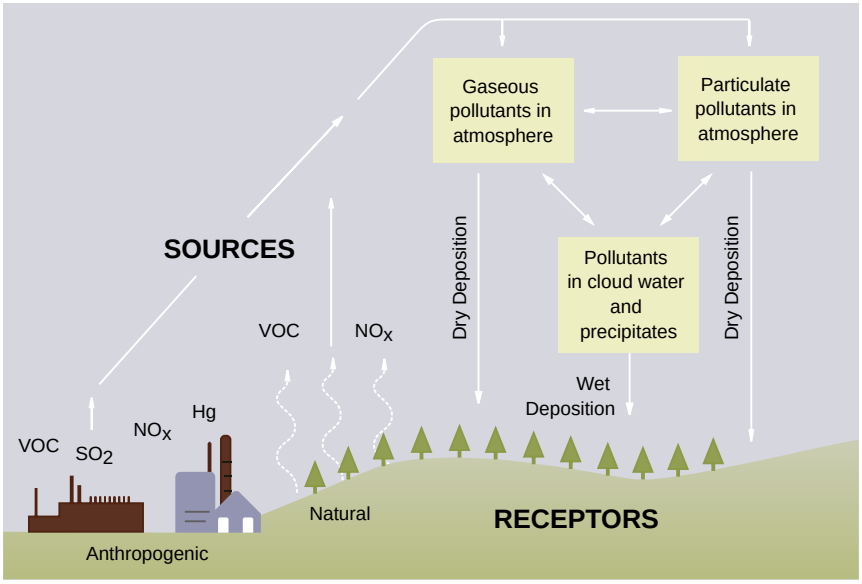


Figure 8.4.6: Processes involved in acid deposition (note that only $\mathrm{SO_2}$ and NOx play a significant role in acid rain). Image used with permission (public domain; EPA)

Acid rain from nitric acid

Some of acid rain is accounted for by nitric acid ($\mathrm{HNO_3}$) that originates from natural processes, but especially from high-temperature air combustion, such as occurs in car engines and power plants, that produces large amounts of NO gas. This gas then forms nitric acid via several steps. First nitric oxide (NO) is formed during lightning storms by the reaction of nitrogen and oxygen

$$\mathrm{N_2\,(g) + O_2\,(g) \longrightarrow 2\,NO\,(g)} \tag{8.4.3}$$

and NO is oxidized to nitrogen dioxide ($\mathrm{NO_2}$)

$$\mathrm{NO\,(g) + \tfrac{1}{2}\,O_2\,(g) \longrightarrow NO_2\,(g)} \tag{8.4.4}$$

$\mathrm{NO_2\,(g)}$ then reacts with water droplets to generate nitric acid ($\mathrm{HNO_3}$)

$$\mathrm{3\,NO_2\,(g) + H_2O\,(l) \longrightarrow 2\,HNO_3\,(aq) + NO\,(g)} \tag{8.4.5}$$

Table 8.4.2*: Source of Water contamination in the United States*

Combustion products	Sources
CO_2 and CO	Combustion of any material (any fuel or tree)
NO_X (NO_2 and NO^3)	High-temperature combustion of any fuel (gas, diesel, or coal), a product of lightning
SO_x (SO_2 and SO_3)	Combustion of sulfur-based fuels (diesel and coal), volcanic release
VOC (volatile organic compound)	Combustion of any carbon-based fuel (gas, diesel, or coal), fumes from paints or solvents

Inorganic Contaminants

Solubility of these species can be determined by consulting a solubility table (e.g., Table E3). These chemicals might be naturally present in a water source. Industry and agriculture also use and dispose of these types of chemicals as well. In determining the source of any pollutant, it is helpful to know a region's geology and proximity to chemical consumers. In this section, we will focus on a few inorganic species that can affect drinking water quality.

Hard Water

When water is described as being hard, then this means it has a large concentration of minerals. The **ions** that contribute to this problem are **calcium, magnesium, carbonate**, and **iron**. Hard water is not toxic but can affect industrial processes. These ions will bond together to form insoluble salts that are called precipitates. Building up of these newly formed solids is called scale. Precipitates can collect in boilers, cooling towers, and any equipment that employs water. This buildup can affect how the equipment works by damaging it internally.

Video

Watch this video below to get an idea of what hard water is.

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

Please be sure to note that the speaker is incorrect when saying that calcium carbonate can be boiled off. Ionic compounds **cannot** be removed by boiling. The speaker should mention distillation when discussing the separation of calcium carbonate from the water.

Soft Water

Soft water does not contain an excessive amount of minerals. Instead, potassium and sodium ions are present. Soft water does not form scale but does produce excessive lathers with soap. Although industry and consumers prefer soft water, the sodium content can affect health if ingested at a high concentration.

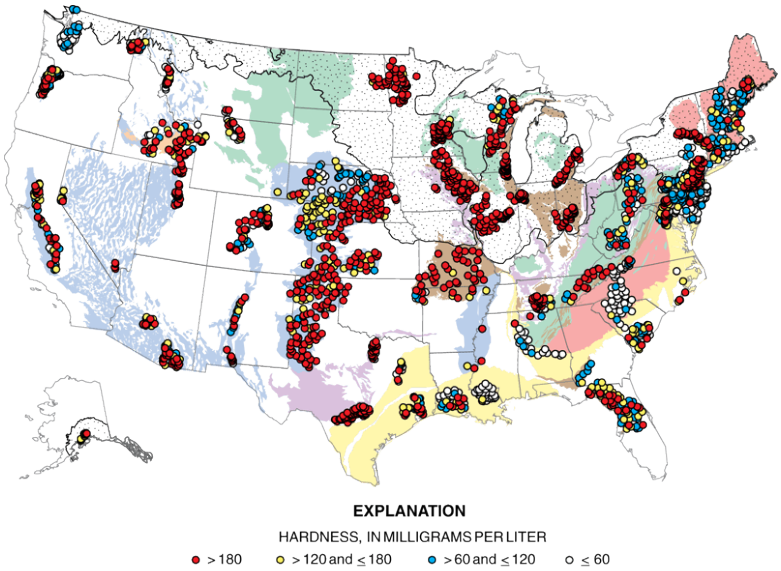


Figure 8.4.1: Hardness of groundwater from domestic wells, a USGS study. A study from the National Water-Quality Assessment (NAWQA) Program assessed water-quality conditions for about 2,100 domestic wells across the United States. Image used with permission (Public Domain; USGS).

Eutrophication

When pollutants enter a water system they originate from either point or non-point sources. If a single identifiable source can be located, then the pollutant originated from a point source. In contrast, nonpoint sources indirectly contaminant water, air, or soil. Examples of non-point sources would be run-off and acid rain. Obviously, non-point pollutants are more difficult to regulate.

In agriculture, nitrates are used as fertilizers. Farmers can choose to use synthetic or natural (manure) nitrates. Regulating concentrations of these species could be difficult if all sources are not investigated and monitored. An excessive amount of these substances leads to eutrophication. Many nitrate and phosphate containing compounds are soluble in water. Once dissolved, these substances can travel to lakes and ponds. Here, these chemicals activate plants and algae to grow. Overgrowth of plant-life in water can inhibit oxygen from entering the water. As a result, aquatic animals are more likely to die in this environment. In addition, decomposition of organic material (once living) will be slow and smelly.

picture6 

Image taken from: https://theskyguys.ca/2017/02/water-...e-great-lakes/ (I took this pic from a site that I don't think is cc). and Image used with permission (CC BY-SA 2.0; N Chadwick).

Eutrophication can be halted in phosphate contamination is controlled. This involves limiting the use of some fertilizers and removing phosphates from detergents. Reducing nitrates can be a more challenging task. All compounds containing this polyatomic ion are extremely soluble. In addition, controlling nitrate concentration from fecal material can be difficult.

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

Toxic Metals

Toxic metals, including "heavy metals," are individual metals and metal compounds that negatively affect people's health. Some toxic, semi-metallic elements, including arsenic and selenium, are discussed in this page. In very small amounts, many of these metals are necessary to support life. However, in larger amounts, they become toxic. They may build up in biological systems and become a significant health hazard.

Lead (Action Level is 0.015 ppm)

Occupational exposure to lead is one of the most prevalent overexposures. Industries with high potential exposures include construction work, most smelter operations, radiator repair shops, and firing ranges.



https://media.defense.gov/2017/May/2...-PO640-001.JPG (please add this picture beside the house)

Arsenic (0.010 ppm)

Common sources of exposure to higher-than-average levels of arsenic include near or in hazardous waste sites and areas with high levels naturally occurring in soil, rocks, and water. Exposure to high levels of arsenic can cause death.

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

Mercury (0.002 ppm)

Common sources of mercury exposure include mining, production, and transportation of mercury, as well as mining and refining of gold and silver ores. High mercury exposure results in permanent nervous system and kidney damage.

Cadmium (0.005 ppm)

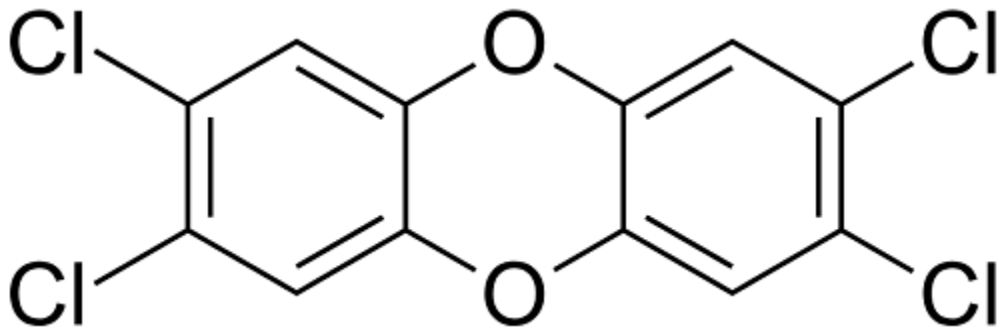
Cadmium is an extremely toxic metal commonly found in industrial workplaces, particularly where any ore is being processed or smelted. Several deaths from acute exposure have occurred among welders who have unsuspectingly welded on cadmium-containing alloys or with silver solders.

Uranium (30 ppb)

Deadly Organics

Dioxin (3.0×10^{-8} ppm)

Dioxins are mainly byproducts of industrial practices. They are produced through a variety of incineration processes, including improper municipal waste incineration and burning of trash, and can be released into the air during natural processes, such as forest fires and volcanoes. Almost every living creature has been exposed to dioxins.



PCB (5.0×10^{-4} ppm)

Love Canal

file:///C:/Users/egordon/Downloads/c...disaster-5.pdf

Contributors

- Elizabeth R. Gordon (Furman University)

This page titled 8.4: Chemical Contamination of Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.5: Water Concentration Units

How to calculate ppm and ppb units..... coding nightmare here. showing metrics and short conversion.

example box with problems. moodle mention of fab video

be able to go between

Love canal image/case study article q/a

This page titled 8.5: Water Concentration Units is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.6: Sanitation of Drinking Water

Water treatment is` a process of making water suitable for its application or returning its natural state. Thus, water treatment required before and after its application. The required treatment depends on the application. For example, treatment of greywater (from bath, dish and wash water) differs from the black water (from flush toilets). Composting toilet is not allowed in urban dwelling. Yet, composting toilets are used in a 30,000-square-foot office complex at the Institute of Asian Research, University of British Columbia.

Water treatment involves science, engineering, business, and art. The treatment may include mechanical, physical, biological, and chemical methods. As with any technology, science is the foundation, and engineering makes sure that the technology works as designed. The appearance and application of water is an art.

In terms of business, RGF Environmental, Water Energy Technologies, Aquasana Store, Vitech, Recalyx Industrial SDN BHD and PACE Chemicals ltd are some of many companies that offer various processes for water treatment. Millipore, a Fisher Scientific partner, offers many lines of products to produce *ultrapure water*, using a combination of active charcoal membranes, and reverse osmosis filter. Internet sites of these companies offer useful information regarding water.

An environmental scientist or consultant matches the service provider, modify if necessary, with the requirement.

- Natural Water includes some discussion on hard and soft water. Softening hard water for boiler, cooler, and domestic application is discussed therein. These treatments prepare water so that it is suitable for the applications.
- Water Biology deals with water and biology. Drinking water is part of making water suitable for living. Thus, this link gives some considerations to drinking water problems.
- There are many different industry types, and waters from various sources are usually treated before and after their applications. Pre-application treatment and wastewater treatment offer a special opportunity or challenge. Only a general consideration will be given to some industrial processes.
- General municipal and domestic wastewater treatment converts used water (waste) into environmentally acceptable water or even drinking water. Every urban centre requires such a facility.

General Wastewater Treatment

Water is a renewable resource. All water treatments involve the removal of solids, bacteria, algae, plants, inorganic compounds, and organic compounds. Removal of solids is usually done by filtration and sediment. Bacteria digestion is an important process to remove harmful *pollutants*. Converting used water into environmentally acceptable water or even drinking water is **wastewater treatment**

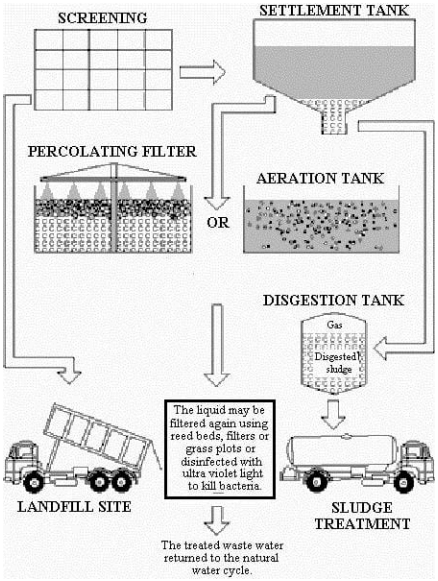
Water in the Great Lakes Region is an organization dealing with the water resources. Ontario Clean Water Agency (OCWA) is a provincial Crown corporation in business to provide environmentally responsible and cost-efficient water and wastewater services. It currently operates more than 400 facilities for 200 municipalities. This web site provides information on water and water treatment.

In April 1993, 403,000 people in Milwaukee were ill as a result of *cryptosporidium* contaimination of water due to spring run off. This outbreak caused the more stringent regulations to be implemented in the public dringking water system. The measures were aimed at removing cryptosporidium.

In May 2000, due to torrential downpour surface water got into shallow wells in a small town Walkerton, Ontario, Canada. On May 17, some residents complained of fever, bloody diarrhea and vomiting. This was know as the Walkerton *E. Coli* Outbreak. Nearly half of the population of the town fell ill, and several people died due to the E. Coli O157:H7 infection. A public inquiry recommended many measures to prevent similar outbreaks. These measures were aimed at eliminating E. Coli.

Sewage Treatment

As a general discussion, let us look at a typical process in sewage treatment. A flow diagram for a general sewage treatment plant from Water Education, Department of Computer Science, University of Exeter, U.K., is shown below:



Sewage is SCREENED to remove large solid chunks, which are disposed in LANDFILL SITE. It flows over to the SETTLEMENT TANK to let the fine particles to settle. The settlement is called the activated SLUDGE. The supernatant is then PERCOLATING FILTERED and/or AERATED. The water can be filtered again, and then disinfected (chlorinated in most cases). When there is no other complication, the water is returned to nature back to the ecological cycle.

The SLUDGE removed from the settlement is composed of living biological material. A portion of it may be returned to the AERATION TANK, but the raw SLUDGE is digested by both microorganism. Anaerobic (without oxygen) and aerobic (with air) bacteria digestions are used. At the digestion stage, carbon dioxide, ammonia, and methane gases are evolved. Volume of the digested sludge is reduced, and it is acceptable as a fertilizer supplement in farming.

Wastewater Treatment

Although the sewage water may be discharged back to the ecological system after AERATED DIGESTION and PERCOLATING FILTRATION, but in some cases, further treatment is required. Some general consideration of water treatment is given below.

A rather recent book, *Chemistry of Water Treatment* by S.D. Faust and O.M. Aly, 2nd Ed. (1998) [TD433 F38 1998], addresses the problem of quality natural and treated water.

The first three chapters discuss the criteria and standards for drinking water quality, organic compounds in waters, taste and order of water. Understandably, the standards change over the years. So are the standards of treated waters. Guidelines are available from government agencies such as Environment Canada which is equivalent to U.S. Public Health Service and the Environment Protection Agency (EPA). We have talked about drinking water in Water Biology.

Next seven chapters deal with the removal of the following:

- organics and inorganics by activated carbon
- particulate matter by coagulation
- particulate matter by filtration and sedimentation
- hardness and other scale-forming substances
- inorganic contaminants
- corrosive substances
- pathogenic (disease producing) bacteria, viruses, and protozoans (microorganisms).

There is a chapter dealing with **aeration**

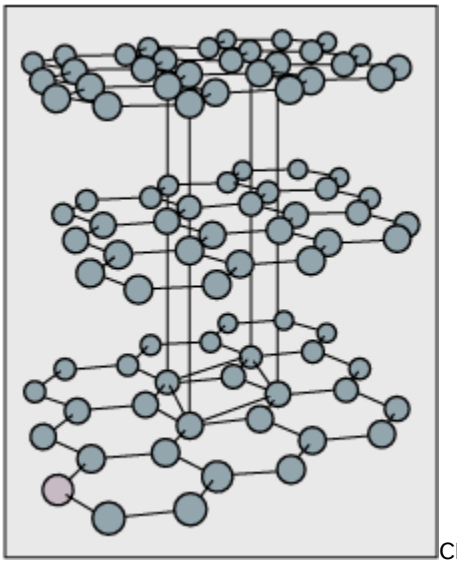
These items cover the chemistry, biology, and physics involved in the treatment of water. Some of these topics have been discussed in chemistry of water, physical properties of water, biology of water, and natural water. Introductions are going to be given to some selected topics below.

Treatment by Activated Carbon

Treatment by activated carbon is mostly due to adsorption or absorption. When a chemical species is adhered to the surface of a solid, it is an **adsorption**. When partial chemical bonds are formed between adsorbed species or when the absorbate got into the channels of the solids, we call it **absorption**. However, these two terms are often used to mean the same, because to distinguish one from type from the other is very difficult.

Application of activated charcoal for the removal of undesirable order and taste in drinking water has been recognized at the dawn of civilization. Using bone char and charred vegetation, gravel, and sand for the filtration of water for domestic application has been practised for thousands of years. Active research and production of activated charcoal was accelerated during the two world wars. The use of poison gas prompted the development of masks. They are still in use today.

Charcoal absorbs many substances, ranging from colored organic particulates to inorganic metal ions. Charcoal has been used to remove the colour of raw sugar from various sources.



Charcoal consists of microcrystallites of graphite. The particles are so small in charcoal that they were considered amorphous. The crystal structure of graphite consists of layers of hexagonal networks, stacked on top of each other. Today, making activated carbon is a new and widely varied industry. Other molecules attach themselves to the porous surface and dangling carbons in these microcrystallites.

Carbon containing substances are charred at less than 900 K to produce carbon in the manufacture of **activated carbon**. However, the carbon is activated at 1200 K using oxidizing agent to selectively oxidize portions of the char to produce pores in the material. Because of the special process to produce used, these materials with high surface to mass ratio, they are called **activated carbon** rather than activated charcoal. Factors affecting the absorption are particle size, surface area, pore structure, acidity (pH), temperature, and the nature of the material to be absorbed. Usually, adsorption (absorption) equilibria and rate of adsorption must be considered for effective applications.

Coagulation, Flucculation and Sedimentation

Natural and wastewater containing small particulates. They are suspended in water forming a **colloid**. These particles carry the same charges, and repulsion prevents them from combining into larger particulates to settle. Thus, some chemical and physical techniques are applied to help them settle. The phenomenon is known as **coagulation**. A well known method is the addition of electrolyte. Charged particulates combine with ions neutralizing the charges. The neutral particulates combine to form larger particles, and finally settle down.

Another method is to use high-molecular-weight material to attract or trap the particulates and settle down together. Such a process is called **flocculation**. Starch and multiply charged ions are often used.

Historically, dirty water is cleaned by treating with alum, $Al_2(SO_4)_3 \cdot 12\ H_2O$, and lime, $Ca(OH)_2$. These electrolytes cause the pH of the water to change due to the following reactions:

Al2(SO4)3.12 H2O, -> Al(aq)3+ + 3 SO4(aq)2- + 12 H2O
SO4(aq)2- + H2O -> HSO4(aq)- + OH- (causing pH change)
Ca(OH)2 -> Ca(aq)2+ + 2 OH- (causing pH change)

The slightly basic water causes $Al(OH)_3$, $Fe(OH)_3$ and $Fe(OH)_2$ to precipitate, bringing the small particulates with them and the water becomes clear. Some records have been found that Egyptians and Romans used these techniques as early as 2000 BC.

Suspension of iron oxide particulates and humic organic matter in water gives water the yellow muddy appearance. Both iron oxide particulates and organic matter can be removed from coagulation and flocculation. The description given here is oversimplified, and many more techniques have been applied in the treatment of water. Coagulation is a major application of lime in the treatment of wastewater.

Other salts such as iron sulfates $Fe_2(SO_4)_3$ and $FeSO_4$, chromium sulfate $Cr_2(SO_4)_3$, and some special polymers are also useful. Other ions such as sodium, chloride, calcium, magnesium, and potassium also affect the coagulation process. So do temperature, pH, and concentration.

Disposal of coagulation sludge is a concern, however.

Sedimentation let the water sit around to let the floculated or coagulated particles to settle out. It works best with relatively dense particles (e.g. silt and minerals), while flotation works better for lighter particles (e.g. algae, color). A settling tank should be big enough so that it takes a long time (ideally 4 hours +) to get through. Inlets and outlets are designed so the water moves slowly in the tank. Long and narrow channels are installed to let the water to snake its way through the tank. The settled particles, sludge, must occasionally be removed from the tanks. The water is next ready to be filtered. Sedimentation is used in pre-treatment and wastewater treatment.

Filtration

Filtration is the process of removing solids from a fluid by passing it through a porous medium. Coarse, medium, and fine porous media have been used depending on the requirement. The filter media are artificial membranes, nets, sand filter, and high technological filter systems. The choice of filters depends on the required filtering speed and the *cleanness* requirement. The flow required for filtration can be achieved using gravity or pressure. In pressure filtration, one side of the filter medium is at higher pressure than that of the other so that the filter plane has a pressure drop. Some portion of this filter type must be enclosed in a container.

The process of removing the clogged portion of the filter bed by reversing the flow through the bed and washing out the solid is called **back washing**. During this process, the solid must be removed out of the system, but otherwise the filters must be either replaced or taken out of service to be cleaned.



Aqua-Rain manufacture water filters as shown here. This unit consists of four filters. Regarding the filtering system, its technical info gave the following statement.

> *At the heart of the AquaRain™ Water Filtration System are Marathonr State-of-the-art ceramic elements utilizing a long-proven filtration process that is over 100 years old which will safely remove dangerous waterborne pathogens such as cysts (Cryptospordium, Giardia lamblia) and bacteria (E. coli, Samonelli typhus, etc...). These innovative Marathonr ceramic elements are also filled with a high grade silvered granulated activated carbon (GAC). The GAC reduces pesticides, chemicals, chlorine, tastes & odors, while leaving the naturally occurring minerals found in the water unaffected.*

The units are designed for emergency and perhaps undeveloped countries.

AquaSelect of Mississauga has a pitcher water filter system, and *its cartridge contains hundreds of high efficiency activated carbon and ion exchange beads*, its web site claims. Brita filters is very popular.

Aeration

Bringing air into intimate contact with water for the purpose of exchanging certain components between the two phases is called **aeration**. Oxygenation is one of the purposes of aeration. Others are removal of volatile organic substances, hydrogen sulfide, ammonia, and volatile organic compounds.

A gas or substance dissolved in water may further react with water. Such a reaction is called **hydration**. Ionic substance dissolve due to hydration, for example:

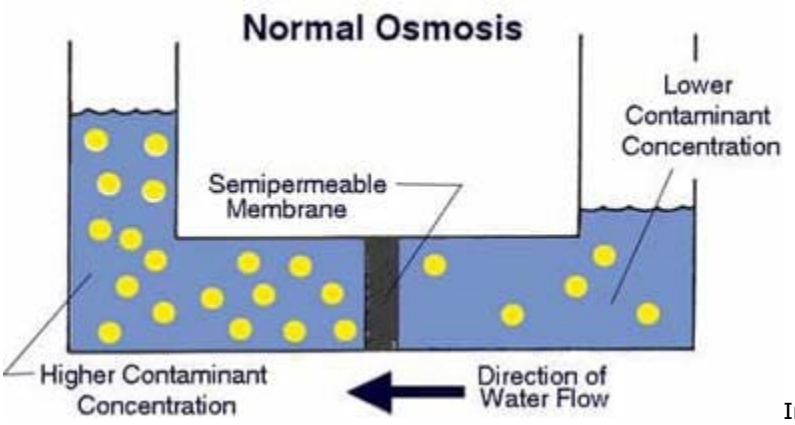
$$\mathrm{HCl\,(g)} + x\,\mathrm{H_2O} \rightleftharpoons \mathrm{H(H_2O)}_x{}^{+} + \mathrm{Cl(aq)}^{-} \tag{8.6.1}$$

$$\mathrm{H_2S} \rightleftharpoons \mathrm{H^{+}\,(aq)} + \mathrm{HS^{-}\,(aq)} \tag{8.6.2}$$

These reactions are reversible, and aeration may also causes dehydration resulting in releasing the gas from water. Henry's law is applicable to this type of equilibrium for consideration. Methods of aeration are

- Diffused aeration - Air bubbles through water.
- Spray aeration - Water is sprayed through air.
- Multiple-tray aeration - Water flows through several trays to mix with air.
- Cascade aeration - Water flows downwards over many steps in the form of thin water falls.
- Air stripping - A combination of multiple tray and cascade technique plus random packed blocks causing water to mix thoroughly with air.

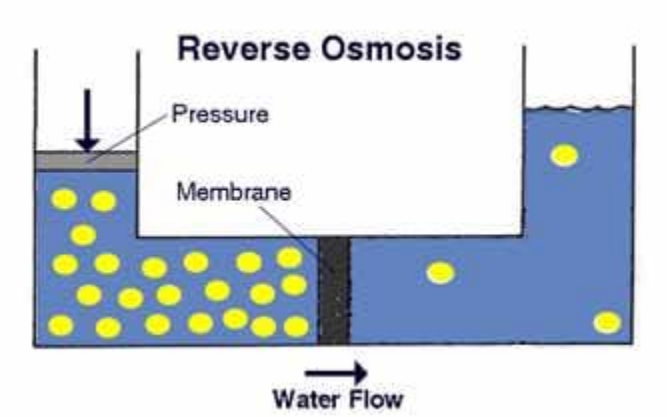
Reverse Osmosis Water Filter System



In the following discussion, a dilute solution and a concentrated solution are considered. The dilute solution can be a clean water whereas the concentrated solution contains undesirable solute (electrolyte or others).

When a compartment containing a dilute solution is connected to another compartment containing a concentrated solution by a semipermeable membrane, water molecules move from the dilute solution to concentrated solution. This phenomenon is called osmosis. Pig bladders are natural semipermeable membranes. As the water molecules migrate through the semipermeable membrane, water level in the solution will increase until the (osmotic) pressure prevents a net migration of water molecules in one direction. A pressure equivalent to the height difference is called the osmotic pressure. The illustration given on the right is from

the PurePro, one of the many companies that manufacture reverse osmosis water filter devices. Millipore also use this technique.



By applying pressure in the higher concentration solution, water molecules migrate from a high concentration solution to a low concentration solution. This method is called **reverse osmosis water filter system**. The concept of reverse osmosis is illustrated in the diagram here from PurePro.

In this technique, the membrane must be able to tolerate the high pressure, and prevent solute molecules to pass through. Regarding membranes, PurePro made the following statement:

> *Semipermeable membranes have come a long way from the natural pig bladders used in the earlier osmosis experiments. Before the 1960's, these membranes were too inefficient, expensive, and unreliable for practical applications outside the laboratory. Modern advances in synthetic materials have generally solved these problems, allowing membranes to become highly efficient at rejecting contaminants, and making them tough enough to withstand the greater pressures necessary for efficient operation.*

This technology certainly works, and it has been used to convert salt (ocean or sea) water into fresh water. With this technique, the water with higher concentration is discharged. Thus, this technology is costly in regions where the water cost is high. Free Drinking Water also uses reverse osmosis filter system for domestic applications.

Industrial Wastewater Treatment

The Environment Canada's (Atlantic Region) Waste Management and Remediation Section contains links to detailed information on the programs and activities relating to Petroleum and Allied Petroleum Storage Tank Systems, Ocean Disposal, Contaminated Sites, and Hazardous Waste Disposal Advice for the provinces of Newfoundland and Labrador, Nova Scotia, New Brunswick and Prince Edward Island. It also contains links to sites within and outside of Environment Canada with related information.

Contributors and Attributions

- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)

This page titled 8.6: Sanitation of Drinking Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.7: Fluoridation of Water

This page titled 8.7: Fluoridation of Water is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.8: Other Water treatment processes

Water treatment is a process of making water suitable for its application or returning its natural state. Thus, water treatment required before and after its application. The required treatment depends on the application. For example, treatment of greywater (from bath, dish and wash water) differs from the black water (from flush toilets). Composting toilet is not allowed in urban dwelling. Yet, composting toilets are used in a 30,000-square-foot office complex at the Institute of Asian Research, University of British Columbia.



Water treatment involves science, engineering, business, and art. The treatment may include mechanical, physical, biological, and chemical methods. As with any technology, science is the foundation, and engineering makes sure that the technology works as designed. The appearance and application of water is an art.

In terms of business, RGF Environmental, Water Energy Technologies, Aquasana Store, Vitech, Recalyx Industrial SDN BHD and PACE Chemicals ltd are some of many companies that offer various processes for water treatment. Millipore, a Fisher Scientific partner, offers many lines of products to produce *ultrapure water*, using a combination of active charcoal membranes, and reverse osmosis filter. Internet sites of these companies offer useful information regarding water.

An environmental scientist or consultant matches the service provider, modify if necessary, with the requirement.

- Natural Water includes some discussion on hard and soft water. Softening hard water for boiler, cooler, and domestic application is discussed therein. These treatments prepare water so that it is suitable for the applications.
- Water Biology deals with water and biology. Drinking water is part of making water suitable for living. Thus, this link gives some considerations to drinking water problems.
- There are many different industry types, and waters from various sources are usually treated before and after their applications. Pre-application treatment and wastewater treatment offer a special opportunity or challenge. Only a general consideration will be given to some industrial processes.
- General municipal and domestic wastewater treatment converts used water (waste) into environmentally acceptable water or even drinking water. Every urban centre requires such a facility.

General Wastewater Treatment

Water is a renewable resource. All water treatments involve the removal of solids, bacteria, algae, plants, inorganic compounds, and organic compounds. Removal of solids is usually done by filtration and sediment. Bacteria digestion is an important process to remove harmful *pollutants*. Converting used water into environmentally acceptable water or even drinking water is **wastewater treatment**

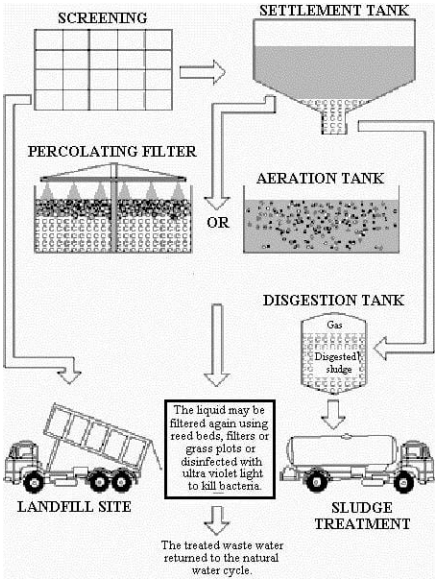
Water in the Great Lakes Region is an organization dealing with the water resources. Ontario Clean Water Agency (OCWA) is a provincial Crown corporation in business to provide environmentally responsible and cost-efficient water and wastewater services. It currently operates more than 400 facilities for 200 municipalities. This web site provides information on water and water treatment.

In April 1993, 403,000 people in Milwaukee were ill as a result of **cryptosporidium** contaimination of water due to spring run off. This outbreak caused the more stringent regulations to be implemented in the public dringking water system. The measures were aimed at removing cryptosporidium.

In May 2000, due to torrential downpour surface water got into shallow wells in a small town Walkerton, Ontario, Canada. On May 17, some residents complained of fever, bloody diarrhea and vomiting. This was know as the Walkerton *E. Coli* Outbreak. Nearly half of the population of the town fell ill, and several people died due to the E. Coli O157:H7 infection. A public inquiry recommended many measures to prevent similar outbreaks. These measures were aimed at eliminating E. Coli.

Sewage Treatment

As a general discussion, let us look at a typical process in sewage treatment. A flow diagram for a general sewage treatment plant from Water Education, Department of Computer Science, University of Exeter, U.K., is shown below:



Sewage is SCREENED to remove large solid chunks, which are disposed in LANDFILL SITE. It flows over to the SETTLEMENT TANK to let the fine particles to settle. The settlement is called the activated SLUDGE. The supernatant is then PERCOLATING FILTERED and/or AERATED. The water can be filtered again, and then disinfected (chlorinated in most cases). When there is no other complication, the water is returned to nature back to the ecological cycle.

The SLUDGE removed from the settlement is composed of living biological material. A portion of it may be returned to the AERATION TANK, but the raw SLUDGE is digested by both microorganism. Anaerobic (without oxygen) and aerobic (with air) bacteria digestions are used. At the digestion stage, carbon dioxide, ammonia, and methane gases are evolved. Volume of the digested sludge is reduced, and it is acceptable as a fertilizer supplement in farming.

Wastewater Treatment

Although the sewage water may be discharged back to the ecological system after AERATED DIGESTION and PERCOLATING FILTRATION, but in some cases, further treatment is required. Some general consideration of water treatment is given below.

A rather recent book, *Chemistry of Water Treatment* by S.D. Faust and O.M. Aly, 2nd Ed. (1998) [TD433 F38 1998], addresses the problem of quality natural and treated water.

The first three chapters discuss the criteria and standards for drinking water quality, organic compounds in waters, taste and order of water. Understandably, the standards change over the years. So are the standards of treated waters. Guidelines are available from government agencies such as Environment Canada which is equivalent to U.S. Public Health Service and the Environment Protection Agency (EPA). We have talked about drinking water in Water Biology.

Next seven chapters deal with the removal of the following:

- organics and inorganics by activated carbon
- particulate matter by coagulation
- particulate matter by filtration and sedimentation
- hardness and other scale-forming substances
- inorganic contaminants
- corrosive substances
- pathogenic (disease producing) bacteria, viruses, and protozoans (microorganisms).

There is a chapter dealing with **aeration**

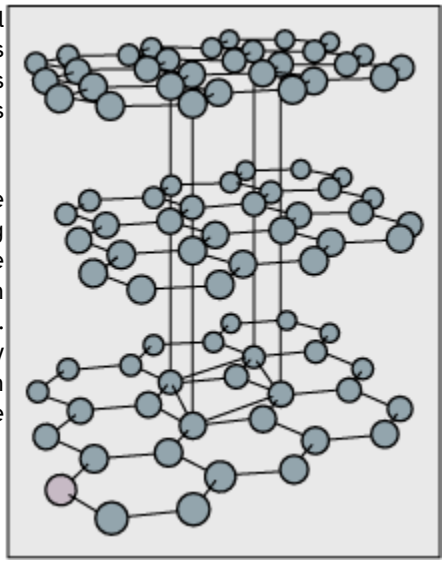
These items cover the chemistry, biology, and physics involved in the treatment of water. Some of these topics have been discussed in chemistry of water, physical properties of water, biology of water, and natural water. Introductions are going to be given to some selected topics below.

Treatment by Activated Carbon

Treatment by activated carbon is mostly due to adsorption or absorption. When a chemical species is adhered to the surface of a solid, it is an **adsorption**. When partial chemical bonds are formed between adsorbed species or when the absorbate got into the channels of the solids, we call it **absorption**. However, these two terms are often used to mean the same, because to distinguish one from type from the other is very difficult.

Application of activated charcoal for the removal of undesirable order and taste in drinking water has been recognized at the dawn of civilization. Using bone char and charred vegetation, gravel, and sand for the filtration of water for domestic application has been practised for thousands of years. Active research and production of activated charcoal was accelerated during the two world wars. The use of poison gas prompted the development of masks. They are still in use today.

Charcoal absorbs many substances, ranging from colored organic particulates to inorganic metal ions. Charcoal has been used to remove the colour of raw sugar from various sources.



Charcoal consists of microcrystallites of graphite. The particles are so small in charcoal that they were considered amorphous. The crystal structure of graphite consists of layers of hexagonal networks, stacked on top of each other. Today, making activated carbon is a new and widely varied industry. Other molecules attach themselves to the porous surface and dangling carbons in these microcrystallites.

Carbon containing substances are charred at less than 900 K to produce carbon in the manufacture of **activated carbon**. However, the carbon is activated at 1200 K using oxidizing agent to selectively oxidize portions of the char to produce pores in the material. Because of the special process to produce used, these materials with high surface to mass ratio, they are called **activated carbon** rather than activated charcoal. Factors affecting the absorption are particle size, surface area, pore structure, acidity (pH), temperature, and the nature of the material to be absorbed. Usually, adsorption (absorption) equilibria and rate of adsorption must be considered for effective applications.

Coagulation, Flucculation and Sedimentation

Natural and wastewater containing small particulates. They are suspended in water forming a **colloid**. These particles carry the same charges, and repulsion prevents them from combining into larger particulates to settle. Thus, some chemical and physical techniques are applied to help them settle. The phenomenon is known as **coagulation**. A well known method is the addition of electrolyte. Charged particulates combine with ions neutralizing the charges. The neutral particulates combine to form larger particles, and finally settle down.

Another method is to use high-molecular-weight material to attract or trap the particulates and settle down together. Such a process is called **flocculation**. Starch and multiply charged ions are often used.

Historically, dirty water is cleaned by treating with alum, $Al_2(SO_4)_3{\cdot}12\ H_2O$, and lime, $Ca(OH)_2$. These electrolytes cause the pH of the water to change due to the following reactions:

Al2(SO4)3.12 H2O, -> Al(aq)3+ + 3 SO4(aq)2- + 12 H2O
SO4(aq)2- + H2O -> HSO4(aq)- + OH- (causing pH change)
Ca(OH)2 -> Ca(aq)2+ + 2 OH- (causing pH change)

The slightly basic water causes $Al(OH)_3$, $Fe(OH)_3$ and $Fe(OH)_2$ to precipitate, bringing the small particulates with them and the water becomes clear. Some records have been found that Egyptians and Romans used these techniques as early as 2000 BC.

Suspension of iron oxide particulates and humic organic matter in water gives water the yellow muddy appearance. Both iron oxide particulates and organic matter can be removed from coagulation and flocculation. The description given here is oversimplified, and many more techniques have been applied in the treatment of water. Coagulation is a major application of lime in the treatment of wastewater.

Other salts such as iron sulfates $Fe_2(SO_4)_3$ and $FeSO_4$, chromium sulfate $Cr_2(SO_4)_3$, and some special polymers are also useful.

Other ions such as sodium, chloride, calcium, magnesium, and potassium also affect the coagulation process. So do temperature, pH, and concentration.

Disposal of coagulation sludge is a concern, however.

Sedimentation let the water sit around to let the floculated or coagulated particles to settle out. It works best with relatively dense particles (e.g. silt and minerals), while flotation works better for lighter particles (e.g. algae, color). A settling tank should be big enough so that it takes a long time (ideally 4 hours +) to get through. Inlets and outlets are designed so the water moves slowly in the tank. Long and narrow channels are installed to let the water to snake its way through the tank. The settled particles, sludge, must occasionally be removed from the tanks. The water is next ready to be filtered. Sedimentation is used in pre-treatment and wastewater treatment.

Filtration

Filtration is the process of removing solids from a fluid by passing it through a porous medium. Coarse, medium, and fine porous media have been used depending on the requirement. The filter media are artificial membranes, nets, sand filter, and high technological filter systems. The choice of filters depends on the required filtering speed and the *cleanness* requirement. The flow required for filtration can be achieved using gravity or pressure. In pressure filtration, one side of the filter medium is at higher pressure than that of the other so that the filter plane has a pressure drop. Some portion of this filter type must be enclosed in a container.

The process of removing the clogged portion of the filter bed by reversing the flow through the bed and washing out the solid is called **back washing**. During this process, the solid must be removed out of the system, but otherwise the filters must be either replaced or taken out of service to be cleaned.

Aqua-Rain manufacture water filters as shown here. This unit consists of four filters. Regarding the filtering system, its techinical info gave the following statement.

> *At the heart of the AquaRain™ Water Filtration System are Marathonr State-of-the-art ceramic elements utilizing a long-proven filtration process that is over 100 years old which will safely remove dangerous waterborne pathogens such as cysts (Cryptospordium, Giardia lamblia) and bacteria (E. coli, Samonelli typhus, etc...). These innovative Marathonr ceramic elements are also filled with a high grade silvered granulated activated carbon (GAC). The GAC reduces pesticides, chemicals, chlorine, tastes & odors, while leaving the naturally occurring minerals found in the water unaffected.*



The units are designed for emergency and perhaps undeveloped countries.

AquaSelect of Mississauga has a pitcher water filter system, and *its cartridge contains hundreds of high efficiency activated carbon and ion exchange beads*, its web site claims. Brita filters is very popular.

Aeration

Bringing air into intimate contact with water for the purpose of exchanging certain components between the two phases is called **aeration**. Oxygenation is one of the purposes of aeration. Others are removal of volatile organic substances, hydrogen sulfide, ammonia, and volatile organic compounds.

A gas or substance dissolved in water may further react with water. Such a reaction is called **hydration**. Ionic substance dissolve due to hydration, for example:

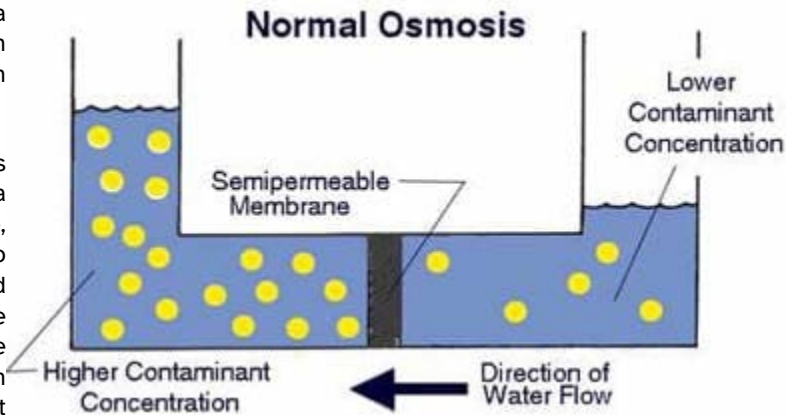
$$HCl\ (g) + x\ H2O = H(H2O)x+ + Cl(aq)-$$
$$H2S = H(aq)+ + HS(aq)-$$

These reactions are reversible, and aeration may also causes dehydration resulting in releasing the gas from water. Henry's law is applicable to this type of equilibrium for consideration. Methods of aeration are

- Diffused aeration - Air bubbles through water.
- Spray aeration - Water is sprayed through air.
- Multiple-tray aeration - Water flows through several trays to mix with air.
- Cascade aeration - Water flows downwards over many steps in the form of thin water falls.
- Air stripping - A combination of multiple tray and cascade technique plus random packed blocks causing water to mix thoroughly with air.

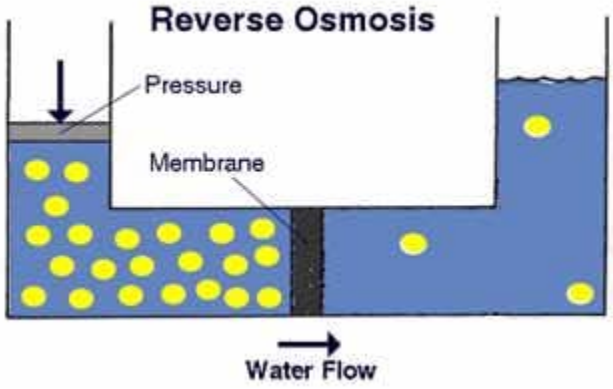
Reverse Osmosis Water Filter System

In the following discussion, a dilute solution and a concentrated solution are considered. The dilute solution can be a clean water whereas the concentrated solution contains undesirable solute (electrolyte or others).



When a compartment containing a dilute solution is connected to another compartment containing a concentrated solution by a semipermeable membrane, water molecules move from the dilute solution to concentrated solution. This phenomenon is called **osmosis**. Pig bladders are natural semipermeable membranes. As the water molecules migrate through the semipermeable membrane, water level in the solution will increase until the (osmotic) pressure prevents a net migration of water molecules in one direction. A pressure equivalent to the height difference is called the **osmotic pressure**. The illustration given on the right is from the PurePro, one of the many companies that manufacture reverse osmosis water filter devices. Millipore also use this technique.

By applying pressure in the higher concentration solution, water molecules migrate from a high concentration solution to a low concentration solution. This method is called **reverse osmosis water filter system**. The concept of reverse osmosis is illustrated in the diagram here from PurePro.



In this technique, the membrane must be able to tolerate the high pressure, and prevent solute molecules to pass through. Regarding membranes, PurePro made the following statement:

> *Semipermeable membranes have come a long way from the natural pig bladders used in the earlier osmosis experiments. Before the 1960's, these membranes were too inefficient, expensive, and unreliable for practical applications outside the laboratory. Modern advances in synthetic materials have generally solved these problems, allowing membranes to become highly efficient at rejecting contaminants, and making them tough enough to withstand the greater pressures necessary for efficient operation.*

This technology certainly works, and it has been used to convert salt (ocean or sea) water into fresh water. With this technique, the water with higher concentration is discharged. Thus, this technology is costly in regions where the water cost is high. Free Drinking Water also uses reverse osmosis filter system for domestic applications.

Industrial Wastewater Treatment

The Environment Canada's (Atlantic Region) Waste Management and Remediation Section contains links to detailed information on the programs and activities relating to Petroleum and Allied Petroleum Storage Tank Systems, Ocean Disposal, Contaminated Sites, and Hazardous Waste Disposal Advice for the provinces of Newfoundland and Labrador, Nova Scotia, New Brunswick and Prince Edward Island. It also contains links to sites within and outside of Environment Canada with related information.

Water and Waste Treatment in a Brewry, Distillery, or Winery

This section is underconstruction. The fermentation industry is interesting, but you may not think it is.



First of all, for the process involved in a brewry, a proposal called Alcohol Fermentation to Produce Beer gives some details about the brewing process. Brewing equipment is available from CDC Inc.

Water Treatment Related Companies

The listing of these companies is not an endorsement of them. They are listed here for information only.

- Cleanwater H_2OAustralia company for water in the environment.
- accepta Water and waste water treatment chemicals, test kits, analysis etc, a U.K. Company
- Environmental Expert provides a large list of companies for water technology.
- Machines for Water Systems - a company with addresses in Niagara Falls, N.Y. and Hamilton in Ontario, Canada.

Contributors

- Chung (Peter) Chieh (Professor Emeritus, Chemistry @ University of Waterloo)

This page titled 8.8: Other Water treatment processes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.9: Sanitizing water while traveling

This page titled 8.9: Sanitizing water while traveling is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.10: Bottled Water Basics

This page titled 8.10: Bottled Water Basics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.11: Love Canal Case Study

This page titled 8.11: Love Canal Case Study is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

8.12: Flint River Case Study

This page titled 8.12: Flint River Case Study is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9: ORGANIC CHEMISTRY

Organic chemistry studies the structure, properties and reactions of organic compounds that contain carbon in covalent bonding.

9.1: Introduction to Organic Chemistry

9.2: Alkanes

9.3: Alkenes

9.4: Alkynes

9.5: Cyclics

9.6: Isomers

9.7: Hydrogenation

9.8: Functional Group introduction

9.9: Aromatics

9.10: Alcohols

9.11: Ethers

9.12: Carbonyl Friends

9.13: Amines and Thiols

9.E: Organic Chemistry (Exercises)

This page titled 9: Organic Chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.1: Introduction to Organic Chemistry

This page titled 9.1: Introduction to Organic Chemistry is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.2: Alkanes

The names of all alkanes end with *-ane*. Whether or not the carbons are linked together end-to-end in a ring (called *cyclic alkanes* or *cycloalkanes*) or whether they contain side chains and branches, the name of every carbon-hydrogen chain that lacks any double bonds or functional groups will end with the suffix *-ane*.

Alkanes with unbranched carbon chains are simply named by the number of carbons in the chain. The first four members of the series (in terms of number of carbon atoms) are named as follows:

1. CH_4 = **methane** = one hydrogen-saturated carbon
2. C_2H_6 = **ethane** = two hydrogen-saturated carbons
3. C_3H_8 = **propane** = three hydrogen-saturated carbons
4. C_4H_{10} = **butane** = four hydrogen-saturated carbons

Alkanes with five or more carbon atoms are named by adding the suffix *-ane* to the appropriate numerical multiplier, except the terminal *-a* is removed from the basic numerical term. Hence, C_5H_{12} is called *pentane*, C_6H_{14} is called *hexane*, C_7H_{16} is called *heptane* and so forth.

Straight-chain alkanes are sometimes indicated by the prefix *n-* (for normal) to distinguish them from branched-chain alkanes having the same number of carbon atoms. Although this is not strictly necessary, the usage is still common in cases where there is an important difference in properties between the straight-chain and branched-chain isomers: e.g. *n-hexane* is a neurotoxin while its branched-chain isomers are not.

IUPAC nomenclature

The IUPAC nomenclature is a system on which most organic chemists have agreed to provide guidelines to allow them to learn from each others' works. Nomenclature, in other words, provides a foundation of language for organic chemistry.

Number of Hydrogen to Carbons

This equation describes the relationship between the number of hydrogen and carbon atoms in alkanes:

$$\mathbf{H = 2C + 2}$$

where "C" and "H" are used to represent the number of carbon and hydrogen atoms present in one molecule. If C = 2, then H = 6.

Many textbooks put this in the following format:

$$\mathbf{C_nH_{2n+2}}$$

where "C_n" and "H_{2n+2}" represent the number of carbon and hydrogen atoms present in one molecule. If $C_n = 3$, then $H_{2n+2} = 2(3) + 2 = 8$. (For this formula look to the "n" for the number, the "C" and the "H" letters themselves do not change.)

Progressively longer hydrocarbon chains can be made and are named systematically, depending on the number of carbons in the longest chain.

The following table contains the systematic names for the first twenty straight chain alkanes. It will be important to familiarize yourself with these names because they will be the basis for naming many other organic molecules throughout your course of study.

Drawing Hydrocarbons

Recall that when carbon makes four bonds, it adopts the tetrahedral geometry. In the tetrahedral geometry, only two bonds can

occupy a plane simultaneously. The other two bonds point in back or in front of this plane. In order to represent the tetrahedral geometry in two dimensions, solid wedges are used to represent bonds pointing out of the plane of the drawing toward the viewer, and dashed wedges are used to represent bonds pointing out of the plane of the drawing away from the viewer. Consider the following representation of the molecule methane:

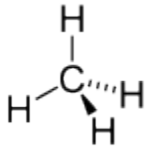


Figure 1: Two dimensional representation of methane

In the above drawing, the two hydrogens connected by solid lines, as well as the carbon in the center of the molecule, exist in a plane (specifically, the plane of the computer monitor / piece of paper, etc.). The hydrogen connected by a solid wedge points out of this plane toward the viewer, and the hydrogen connected by the dashed wedge points behind this plane and away from the viewer.

In drawing hydrocarbons, it can be time-consuming to write out each atom and bond individually. In organic chemistry, hydrocarbons can be represented in a shorthand notation called a **skeletal structure**. In a skeletal structure, only the bonds between carbon atoms are represented. Individual carbon and hydrogen atoms are not drawn, and bonds to hydrogen are not drawn. In the case that the molecule contains just single bonds (sp^3 bonds), these bonds are drawn in a "zig-zag" fashion. This is because in the tetrahedral geometry all bonds point as far away from each other as possible, and the structure is not linear. Consider the following representations of the molecule propane:

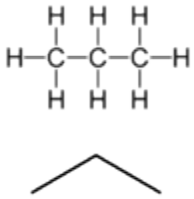
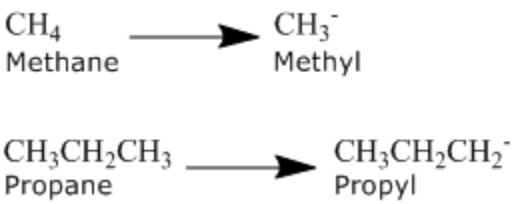


Figure 2: Full structure of propane and skeletal structure of propane

Only the bonds between carbons have been drawn, and these have been drawn in a "zig-zag" manner. Note that there is no representation of hydrogens in a skeletal structure. Since, in the absence of double or triple bonds, carbon makes four bonds total, the presence of hydrogens is implicit. Whenever an insufficient number of bonds to a carbon atom are specified in the structure, it is assumed that the rest of the bonds are made to hydrogens. For example, if the carbon atom makes only one explicit bond, there are three hydrogens implicitly attached to it. If it makes two explicit bonds, there are two hydrogens implicitly attached, etc. Note also that two lines are sufficient to represent three carbon atoms. It is the bonds only that are being drawn out, and it is understood that there are carbon atoms (with three hydrogens attached!) at the terminal ends of the structure.

Alkyl Groups

Alkanes can be described by the general formula C_nH_{2n+2}. An alkyl group is formed by removing one hydrogen from the alkane chain and is described by the formula C_nH_{2n+1}. The removal of this hydrogen results in a stem change from **-ane** to **-yl**. Take a look at the following examples.



The same concept can be applied to any of the straight chain alkane names provided in the table above.

Name	Molecular Formula	Condensed Structural Formula
Methane	CH_4	CH_4
Ethane	C_2H_6	CH_3CH_3
Propane	C_3H_8	$CH_3CH_2CH_3$
Butane	C_4H_{10}	$CH_3(CH_2)_2CH_3$
Pentane	C_5H_{12}	$CH_3(CH_2)_3CH_3$
Hexane	C_6H_{14}	$CH_3(CH_2)_4CH_3$
Heptane	C_7H_{16}	$CH_3(CH_2)_5CH_3$
Octane	C_8H_{18}	$CH_3(CH_2)_6CH_3$
Nonane	C_9H_{20}	$CH_3(CH_2)_7CH_3$
Decane	$C_{10}H_{22}$	$CH_3(CH_2)_8CH_3$
Undecane	$C_{11}H_{24}$	$CH_3(CH_2)_9CH_3$
Dodecane	$C_{12}H_{26}$	$CH_3(CH_2)_{10}CH_3$
Tridecane	$C_{13}H_{28}$	$CH_3(CH_2)_{11}CH_3$
Tetradecane	$C_{14}H_{30}$	$CH_3(CH_2)_{12}CH_3$
Pentadecane	$C_{15}H_{32}$	$CH_3(CH_2)_{13}CH_3$
Hexadecane	$C_{16}H_{34}$	$CH_3(CH_2)_{14}CH_3$
Heptadecane	$C_{17}H_{36}$	$CH_3(CH_2)_{15}CH_3$
Octadecane	$C_{18}H_{38}$	$CH_3(CH_2)_{16}CH_3$
Nonadecane	$C_{19}H_{40}$	$CH_3(CH_2)_{17}CH_3$
Eicosane	$C_{20}H_{42}$	$CH_3(CH_2)_{18}CH_3$

Using Common Names with Branched Alkanes

Certain branched alkanes have common names that are still widely used today. These common names make use of prefixes, such as ***iso-***, ***sec-***, ***tert-***, and ***neo-***. The prefix ***iso-***, which stands for isomer, is commonly given to 2-methyl alkanes. In other words, if there is methyl group located on the second carbon of a carbon chain, we can use the prefix ***iso-***. The prefix will be placed in front of the alkane name that indicates the *total* number of carbons. Examples:

- isopentane which is the same as 2-methylbutane
- isobutane which is the same as 2-methylpropane

To assign the prefixes ***sec-***, which stands for secondary, and ***tert-***, for tertiary, it is important that we first learn how to classify carbon molecules. If a carbon is attached to only one other carbon, it is called a **primary** carbon. If a carbon is attached to two other carbons, it is called a **seconday** carbon. A **tertiary** carbon is attached to three other carbons and last, a **quaternary** carbon is attached to four other carbons. Examples:

- 4-*sec*-butylheptane (30g)
- 4-*tert*-butyl-5-isopropylhexane (30d); if using this example, may want to move sec/tert after iso disc

The prefix ***neo-*** refers to a substituent whose second-to-last carbon of the chain is trisubstituted (has three methyl groups attached to it). A neo-pentyl has five carbons total. Examples:

- neopentane
- neoheptane

Alkoxy Groups

Alkoxides consist of an organic group bonded to a negatively charged oxygen atom. In the general form, alkoxides are written as RO^-, where R represents the organic substituent. Similar to the alkyl groups above, the concept of naming alkoxides can be applied to any of the straight chain alkanes provided in the table above.

RO^-
Alkoxy

CH_3O^-
Methoxy

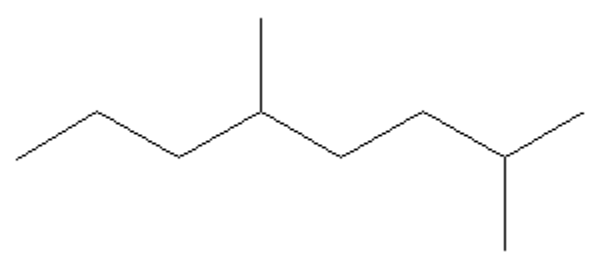
$CH_3CH_2O^-$
Ethoxy

Three Principles of Naming

1. Choose the longest, most substituted carbon chain containing a functional group.
2. A carbon bonded to a functional group must have the lowest possible carbon number. If there are no functional groups, then any substitute present must have the lowest possible number.
3. Take the alphabetical order into consideration; that is, after applying the first two rules given above, make sure that your substitutes and/or functional groups are written in alphabetical order.

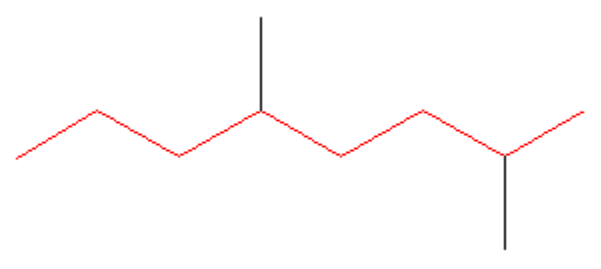
Example 1

What is the name of the follow molecule?



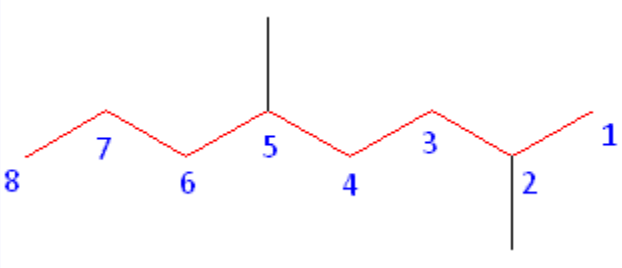
SOLUTION

Rule #1: Choose the longest, most substituted carbon chain containing a functional group. This example does not contain any functional groups, so we only need to be concerned with choosing the longest, most substituted carbon chain. The longest carbon chain has been highlighted in red and consists of eight carbons.



Rule #2: Carbons bonded to a functional group must have the lowest possible carbon number. If there are no functional groups, then any substitute present must have the lowest possible number. Because this example does not contain any functional groups, we only need to be concerned with the two substitutes present, that is, the two methyl groups. If we begin

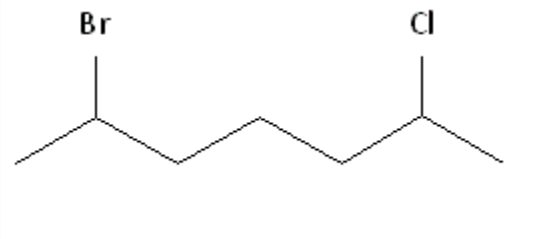
numbering the chain from the left, the methyls would be assigned the numbers 4 and 7, respectively. If we begin numbering the chain from the right, the methyls would be assigned the numbers 2 and 5. Therefore, to satisfy the second rule, numbering begins on the right side of the carbon chain as shown below. This gives the methyl groups the lowest possible numbering.



Rule 3: In this example, there is no need to utilize the third rule. Because the two substitutes are identical, neither takes alphabetical precedence with respect to numbering the carbons. This concept will become clearer in the following examples.

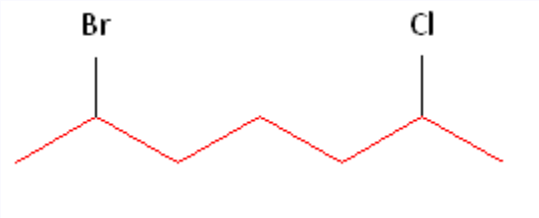
Example 2

What is the name of the follow molecule?

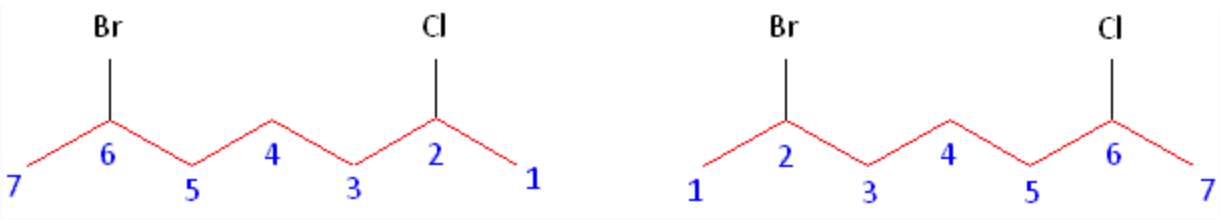


SOLUTION

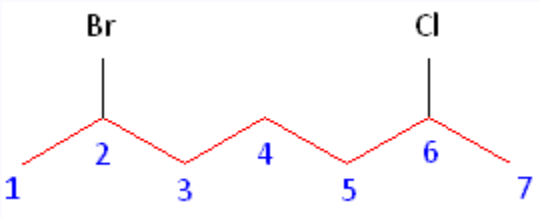
Rule #1: Choose the longest, most substituted carbon chain containing a functional group. This example contains two functional groups, bromine and chlorine. The longest carbon chain has been highlighted in red and consists of seven carbons.



Rule #2: Carbons bonded to a functional group must have the lowest possible carbon number. If there are no functional groups, then any substitute present must have the lowest possible number. In this example, numbering the chain from the left or the right would satisfy this rule. If we number the chain from the left, bromine and chlorine would be assigned the second and sixth carbon positions, respectively. If we number the chain from the right, chlorine would be assigned the second position and bromine would be assigned the sixth position. In other words, whether we choose to number from the left or right, the functional groups occupy the second and sixth positions in the chain. To select the correct numbering scheme, we need to utilize the third rule.

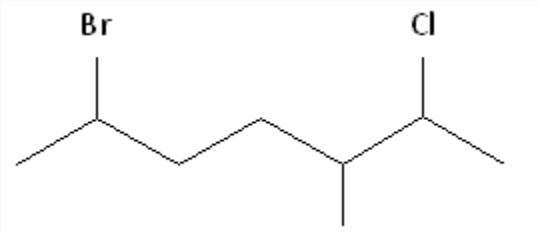


Rule #3: After applying the first two rules, take the alphabetical order into consideration. Alphabetically, bromine comes before chlorine. Therefore, bromine is assigned the second carbon position, and chlorine is assigned the sixth carbon position.



Example 3

What is the name of the follow molecule?

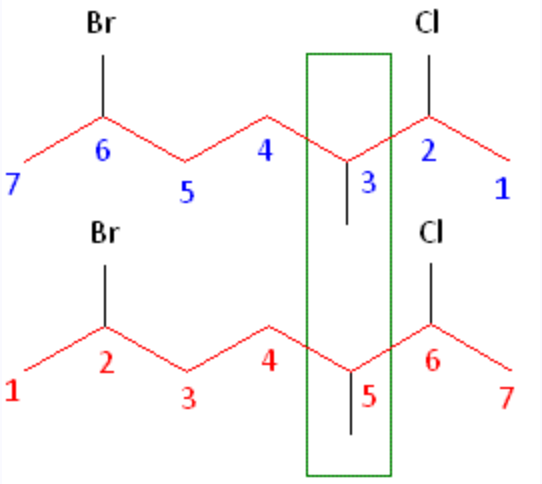


SOLUTION

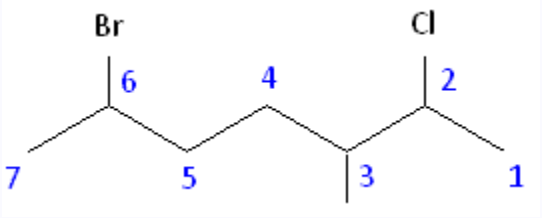
Rule #1: Choose the longest, most substituted carbon chain containing a functional group. This example contains two functional groups, bromine and chlorine, and one substitute, the methyl group. The longest carbon chain has been highlighted in red and consists of seven carbons.



Rule #2: Carbons bonded to a functional group must have the lowest possible carbon number. After taking functional groups into consideration, any substitutes present must have the lowest possible carbon number. This particular example illustrates the **point of difference principle**. If we number the chain from the left, bromine, the methyl group and chlorine would occupy the second, fifth and sixth positions, respectively. This concept is illustrated in the second drawing below. If we number the chain from the right, chlorine, the methyl group and bromine would occupy the second, third and sixth positions, respectively, which is illustrated in the first drawing below. The position of the methyl, therefore, becomes a **point of difference**. In the first drawing, the methyl occupies the third position. In the second drawing, the methyl occupies the fifth position. To satisfy the second rule, we want to choose the numbering scheme that provides the lowest possible numbering of this substitute. Therefore, the first of the two carbon chains shown below is correct.



Therefore, the first numbering scheme is the appropriate one to use.



Once you have determined the correct numbering of the carbons, it is often useful to make a list, including the functional groups, substitutes, and the name of the parent chain.

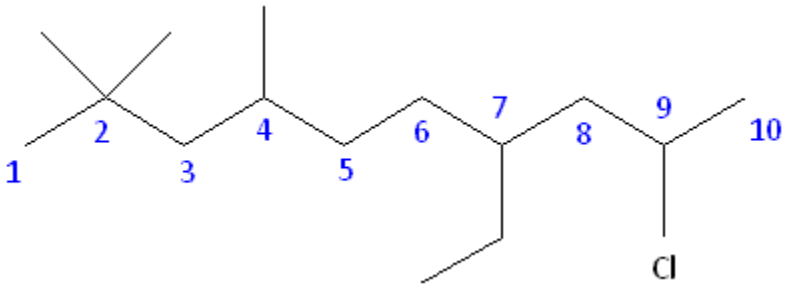
Rule #3: After applying the first two rules, take the alphabetical order into consideration. Alphabetically, bromine comes before chlorine. Therefore, bromine is assigned the second carbon position, and chlorine is assigned the sixth carbon position.

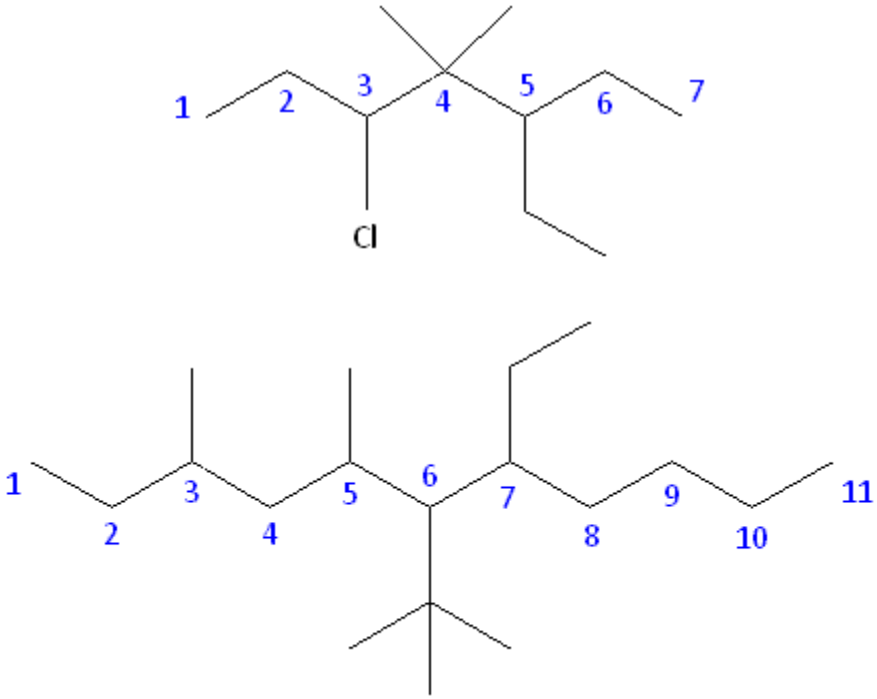
Parent chain: heptane 2-Chloro 3-Methyl 6-Bromo

6-bromo-2-chloro-3-methylheptane

Problems

What is the name of the follow molecules?





Contributors

- Jonathan Mooney (McGill University)
- Template:ContribWikiOrgo

This page titled 9.2: Alkanes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.3: Alkenes

Alkenes contain carbon-carbon double bonds and are unsaturated hydrocarbons with the molecular formula is C_nH_{2n}. This is also the same molecular formula as cycloalkanes. Alkenes are named using the same general naming rules for alkanes, except that the suffix is now -ene. There are a few other small differences:

- The main chain of carbon atoms must contain both carbons in the double bond. The main chain is numbered so that the double bond gets the smallest number.
- Before the root name, the number of the carbon atom at which the double bond starts (the smaller number) is written.
- If more than one double bond is present, the prefixes di-, tri-, tetra-, etc. are used before the -ene, and (strangely) the letter "a" is added after the prefix for the number of carbon atoms.

Introduction

Alkenes and alkynes are hydrocarbons which respectively have **carbon-carbon double bond** and **carbon-carbon triple bond** functional groups. The molecular formulas of these **unsaturated** hydrocarbons reflect the multiple bonding of the functional groups:

Alkenes are named as if they were alkanes, but the "-ane" suffix is changed to "-ene". If the alkene contains only one double bond and that double bond is terminal (the double bond is at one end of the molecule or another) then it is not necessary to place any number in front of the name.

> **butane:** C_4H_{10} ($CH_3CH_2CH_2CH_3$)
> **butene:** C_4H_8 (CH_2=$CHCH_2CH_3$)

If the double bond is not terminal (if it is on a carbon somewhere in the center of the chain) then the carbons should be numbered in such a way as to give the first of the two double-bonded carbons the lowest possible number, and that number should precede the "ene" suffix with a dash, as shown below.

> **correct:** pent-2-ene (CH_3CH=$CHCH_2CH_3$)
> **incorrect:** pent-3-ene (CH_3CH_2CH=$CHCH_3$)
> *The second one is incorrect because flipping the formula horizontally results in a lower number for the alkene.*

If there is more than one double bond in an alkene, all of the bonds should be numbered in the name of the molecule - even terminal double bonds. The numbers should go from lowest to highest, and be separated from one another by a comma. The IUPAC numerical prefixes are used to indicate the number of double bonds.

> **octa-2,4-diene:** CH_3CH=$CHCH$=$CHCH_2CH_2CH_3$
> **deca-1,5-diene:** CH_2=$CHCH_2CH_2CH$=$CHCH_2CH_2CH_2CH_3$

Note that the numbering of "2-4" above yields a molecule with two double bonds separated by just one single bond. Double bonds in such a condition are called "conjugated", and they represent an enhanced stability of conformation, so they are energetically favored as reactants in many situations and combinations.

Alkane	C_nH_{2n+2}	This is the maximum H/C ratio for a given number of carbon atoms.
Alkene	C_nH_{2n}	Each double bond reduces the number of hydrogen atoms by 2.

The parent structure is the longest chain containing both carbon atoms of the double bond. The two carbon atoms of a double bond and the four atoms attached to them lie in a plane, with bond angles of approximately 120° A double bond consists of one sigma bond formed by overlap of sp^2 hybrid orbitals and one pi bond formed by overlap of parallel 2 p orbitals

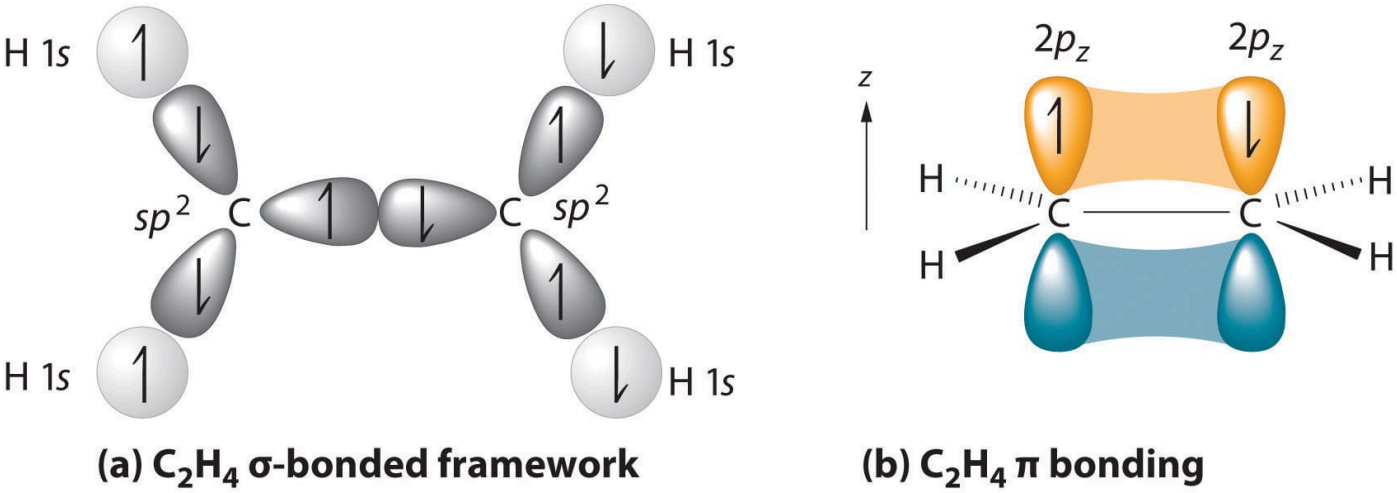
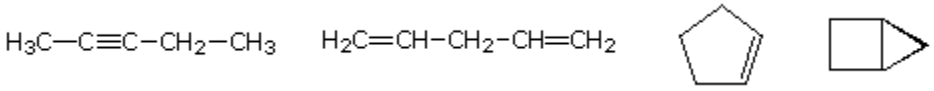


Figure 1: (a) The σ-bonded framework is formed by the overlap of two sets of singly occupied carbon sp^2 hybrid orbitals and four singly occupied hydrogen 1s orbitals to form electron-pair bonds. This uses 10 of the 12 valence electrons to form a total of five σ bonds (four C–H bonds and one C–C bond). (b) One singly occupied unhybridized $2p_z$ orbital remains on each carbon atom to form a carbon–carbon π bond. (Note: by convention, in planar molecules the axis perpendicular to the molecular plane is the *z*-axis.)

The molecular formula of a hydrocarbon provides information about the possible structural types it may represent. For example, consider compounds having the formula C_5H_8. The formula of the five-carbon alkane pentane is C_5H_{12} so the difference in hydrogen content is 4. This difference suggests such compounds may have a triple bond, two double bonds, a ring plus a double bond, or two rings. Some examples are shown here, and there are at least fourteen others!

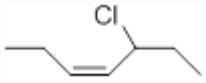


IUPAC Rules for Alkene and Cycloalkene Nomenclature

1. The **ene** suffix (ending) indicates an alkene or cycloalkene.
2. The longest chain chosen for the root name must include both carbon atoms of the double bond.
3. The root chain must be numbered from the end nearest a double bond carbon atom. If the double bond is in the center of the chain, the nearest substituent rule is used to determine the end where numbering starts.
4. The smaller of the two numbers designating the carbon atoms of the double bond is used as the double bond locator. If more than one double bond is present the compound is named as a diene, triene or equivalent prefix indicating the number of double bonds, and each double bond is assigned a locator number.
5. In cycloalkenes the double bond carbons are assigned ring locations #1 and #2. Which of the two is #1 may be determined by the nearest substituent rule.
6. Substituent groups containing double bonds are:

H_2C=CH– Vinyl group
H_2C=CH–CH_2– Allyl group

Example 1



In this diagram this is a cis conformation. It has both the substituents going upward. This molecule would be called (cis) 5-chloro-3-heptene.)

Trans would look like this

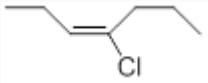
v. On the other hand if there are 3 or 4 non-hydrogen different atoms attached to the alkene then use the E, Z system.

E (entgegen) means the higher priority groups are opposite one another relative to the double bond.

Z (zusammen) means the higher priority groups are on the same side relative to the double bond.

(You could think of Z as Zame Zide to help memorize it.)

Example 2



Solution

In this example it is E-4-chloro-3-heptene. It is E because the Chlorine and the CH_2CH_3 are the two higher priorities and they are on opposite sides.

vi. A hydroxyl group gets precedence over th double bond. Therefore alkenes containing alchol groups are called alkenols. And the prefix becomes --enol. And this means that now the alcohol gets lowest priority over the alkene.

vii. Lastly remember that alkene substituents are called alkenyl. Suffix --enyl.

Here is a chart containing the systemic name for the first twenty straight chain alkenes.

Name	Molecular formula
Ethene	C_2H_4
Propene	C_3H_6
Butene	C_4H_8
Pentene	C_5H_{10}
Hexene	C_6H_{12}
Heptene	C_7H_{14}
Octene	C_8H_{16}
Nonene	C_9H_{18}
Decene	$C_{10}H_{20}$
Undecene	$C_{11}H_{22}$
Dodecene	$C_{12}H_{24}$
Tridecene	$C_{13}H_{26}$
Tetradecene	$C_{14}H_{28}$
Pentadecene	$C_{15}H_{30}$
Hexadecene	$C_{16}H_{32}$
Heptadecene	$C_{17}H_{34}$
Octadecene	$C_{18}H_{36}$

Nonadecene	$C_{19}H_{38}$
Eicosene	$C_{20}H_{40}$

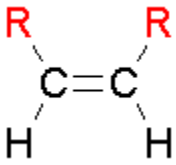
Did you notice how there is no methene? Because it is impossible for a carbon to have a double bond with nothing.

Geometric Isomers

Double bonds can exist as geometric isomers and these isomers are designated by using either the cis / trans designation or the modern E / Z designation.

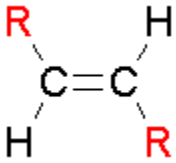
cis Isomers

.The two largest groups are on the same side of the double bond.



trans Isomers

...The two largest groups are on opposite sides of the double bond.



E/Z nomenclature

E = entgegan ("trans") Z = zusamen ("cis")

Priority of groups is based on the atomic mass of attached atoms (not the size of the group). An atom attached by a multiple bond is counted once for each bond.

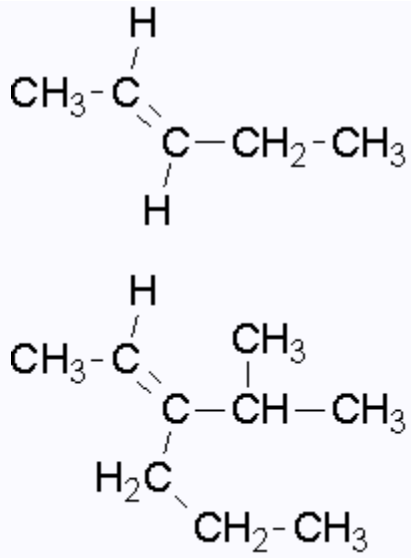
fluorine atom > isopropyl group > n-hexyl group

deuterium atom > hydrogen atom

$-CH_2-CH=CH_2$ > $-CH_2CH_2CH_3$

Example 3

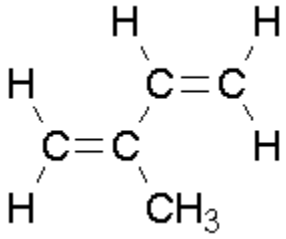
Try to name the following compounds using both conventions...



Common names

Remove the -ane suffix and add -ylene. There are a couple of unique ones like ethenyl's common name is vinyl and 2-propenyl's common name is allyl. That you should know are...

- vinyl substituent $H_2C=CH-$
- allyl substituent $H_2C=CH-CH_2-$
- allene molecule $H_2C=C=CH_2$
- isoprene

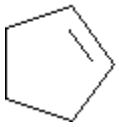


Endocyclic Alkenes

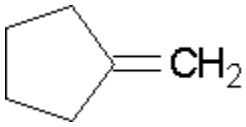
Endocyclic double bonds have both carbons in the ring and exocyclic double bonds have only one carbon as part of the ring.

Cycloalkane: types of alkane which have one or more rings of carbon atoms in the chemical structure of their molecules.

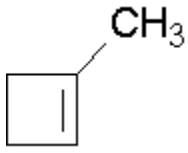
Cyclopentene is an example of an endocyclic double bond.



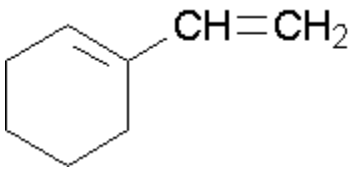
Methylenecylopentane is an example of an exocyclic double bond.



Name the following compounds...

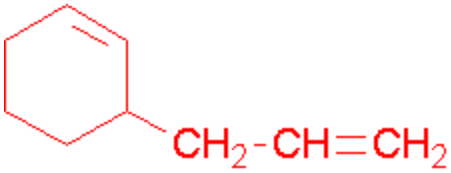


1-methylcyclobutene. The methyl group places the double bond. It is correct to also name this compound as 1-methylcyclobut-1-ene.

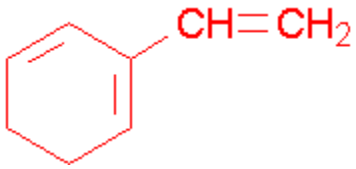


1-ethenylcyclohexene, the methyl group places the double bond. It is correct to also name this compound as 1-ethenylcyclohex-1-ene. A common name would be 1-vinylcyclohexene.

Try to draw structures for the following compounds...



- 2-vinyl-1,3-cyclohexadiene



Outside links

- http://www.vanderbilt.edu/AnS/Chemis...0a/alkenes.pdf

References

1. Vollhardt, Peter, and Neil E. Schore. Organic Chemistry: Structure and Function. 5th Edition. New York: W. H. Freeman & Company, 2007.

Problems

Try to name the following compounds...

1-pentene or pent-1-ene

2-ethyl-1-hexene or 2-ethylhex-1-ene

Try to draw structures for the following compounds...

- 2-pentene

$CH_3–CH=CH–CH_2–CH_3$

- 3-heptene

$CH_3–CH_2–CH=CH–CH_2–CH_2–CH_3$

b. Give the double bond the lowest possible numbers regardless of substituent placement.

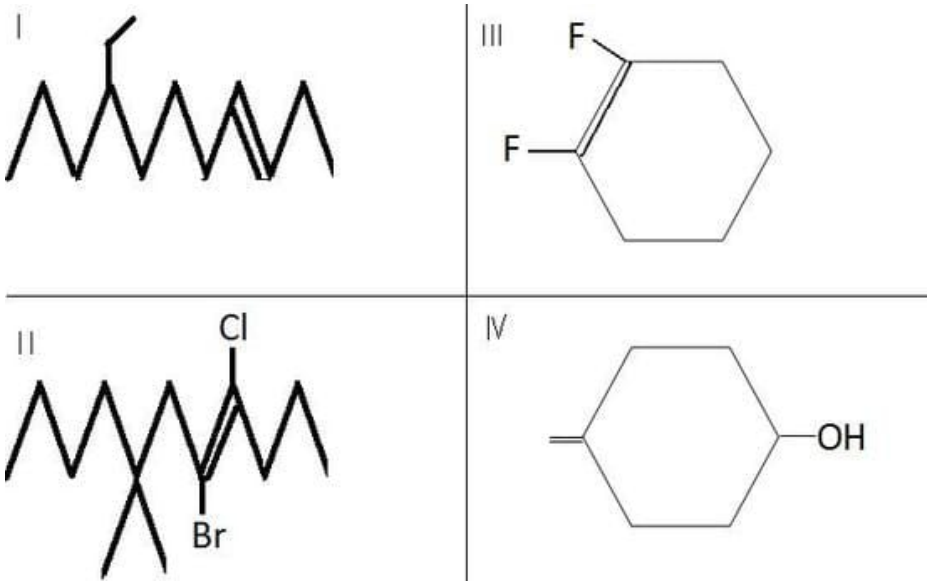
• Try to name the following compound...

J

• Try to draw a structure for the following compound...

4-methyl-2-pentene J

Name the following structures:



v. Draw (Z)-5-Chloro-3-ethly-4-hexen-2-ol.

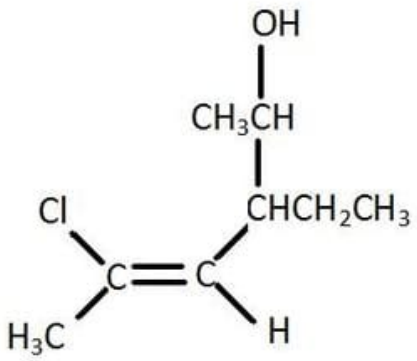
Answers

I. trans-8-ethyl-3-undecene

II. E-5-bromo-4-chloro-7,7-dimethyl-4-undecene

III. Z-1,2-difluoro-cyclohexene

IV. 4-ethenylcyclohexanol.



V.

Contributors

- S. Devarajan (UCD)

Richard Banks (Boise State University)

- William Reusch, Professor Emeritus (Michigan State U.), Virtual Textbook of Organic Chemistry
- Jonathan Mooney (McGill University)

This page titled 9.3: Alkenes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.4: Alkynes

Alkynes are organic molecules made of the functional group carbon-carbon triple bonds and are written in the empirical formula of C_nH_{2n-2}. They are unsaturated hydrocarbons. Like alkenes have the suffix –ene, alkynes use the ending –yne; this suffix is used when there is only one alkyne in the molecule. If a molecule contains both a double and a triple bond, the carbon chain is numbered so that the first multiple bond gets a lower number. If both bonds can be assigned the same number, the double bond takes precedence. The molecule is then named "n-ene-n-yne", with the double bond root name preceding the triple bond root name (e.g. 2-hepten-4-yne).

Introduction



Here are the molecular formulas and names of the first ten carbon straight chain alkynes.

Name	Molecular Formula
Ethyne	C_2H_2
Propyne	C_3H_4
1-Butyne	C_4H_6
1-Pentyne	C_5H_8
1-Hexyne	C_6H_{10}
1-Heptyne	C_7H_{12}
1-Octyne	C_8H_{14}
1-Nonyne	C_9H_{16}
1-Decyne	$C_{10}H_{18}$

The more commonly used name for ethyne is acetylene, which used industrially.

Naming Alkynes

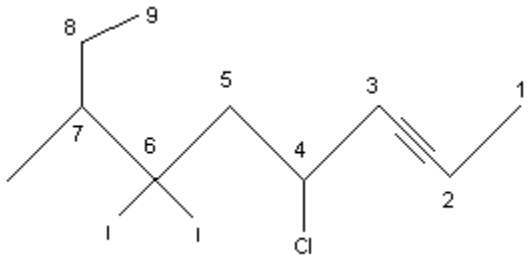
Like previously mentioned, the IUPAC rules are used for the naming of alkynes.

Rule 1

Find the longest carbon chain that includes both carbons of the triple bond.

Rule 2

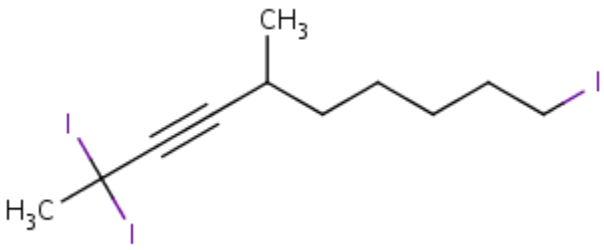
Number the longest chain starting at the end closest to the triple bond. A 1-alkyne is referred to as a terminal alkyne and alkynes at any other position are called internal alkynes. For example:



4-chloro-6-diiodo-7-methyl-2-nonyne

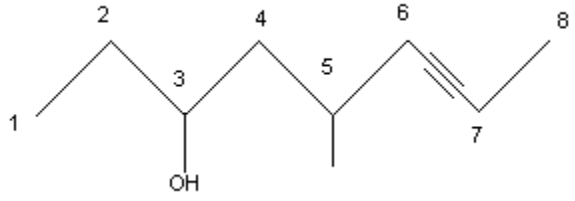
Rule 3

After numbering the longest chain with the lowest number assigned to the alkyne, label each of the substituents at its corresponding carbon. While writing out the name of the molecule, arrange the substituents in alphabetical order. If there are more than one of the same substituent use the prefixes di, tri, and tetra for two, three, and four substituents respectively. These prefixes are not taken into account in the alphabetical order. For example:



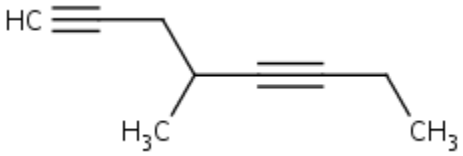
2,2,10-triiodo-5-methyl-3-decyne

If there is an alcohol present in the molecule, number the longest chain starting at the end closest to it, and follow the same rules. However, the suffix would be –ynol, because the alcohol group takes priority over the triple bond.



5- methyl-7-octyn-3-ol

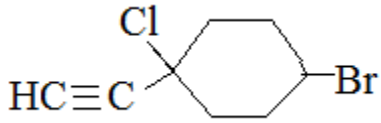
When there are two triple bonds in the molecule, find the longest carbon chain including both the triple bonds. Number the longest chain starting at the end closest to the triple bond that appears first. The suffix that would be used to name this molecule would be –diyne. For example:



4-methyl-1,5-octadiyne

Rule 4

Substituents containing a triple bond are called alkynyl. For example:



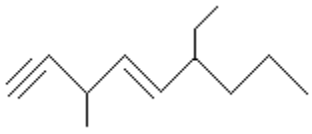
1-chloro-1-ethynyl-4-bromocyclohexane

Here is a table with a few of the alkynyl substituents:

Name	**Molecule**
Ethynyl	-C?CH
2- Propynyl	-CH_2C?CH
2-Butynyl	-CH_3C?CH_2CH_3

Rule 5

A molecule that contains both double and triple bonds is called an alkenyne. The chain can be numbered starting with the end closest to the functional group that appears first. For example:



6-ethyl-3-methyl-1,4-nonenyne

Reference

1. Vollhardt, Peter, and Neil E. Schore. Organic Chemistry: Structure and Function. 5th Edition. New York: W. H. Freeman & Company, 2007.

Problems

Name or draw out the following molecules:

1. 4,4-dimethyl-2-pentyne

2. 4-Penten-1-yne

3. 1-ethyl-3-dimethylnonyne

4.

8.bmp

Contributors

- A. Sheth and S. Sujit (UCD)
- Jonathan Mooney (McGill University)

This page titled 9.4: Alkynes is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.5: Cyclics

Although cyclohexane can be isolated from petroleum products, a major source of this chemical is the hydrogenation of benzene. Much of the cyclohexane produced is used to manufacture intermediates for the production of nylon. Items such as nylon balloons no doubt had their start in a chemical plant where hydrogen gas and benzene were reacted at high temperatures to form cyclohexane. This cycloalkane then undergoes nitration to begin the process of forming the long strands of nylon that can be made into balloons, ropes, clothing, and many other useful products.

Cyclic Hydrocarbons

A **cyclic hydrocarbon** is a hydrocarbon in which the carbon chain joins to itself in a ring. A **cycloalkane** is a cyclic hydrocarbon in which all of the carbon-carbon bonds are single bonds. Like other alkanes, cycloalkanes are saturated compounds. Cycloalkanes have the general formula $\mathrm{C}_n\mathrm{H}_{2n}$. The simplest cycloalkane is cyclopropane, a three-carbon ring.

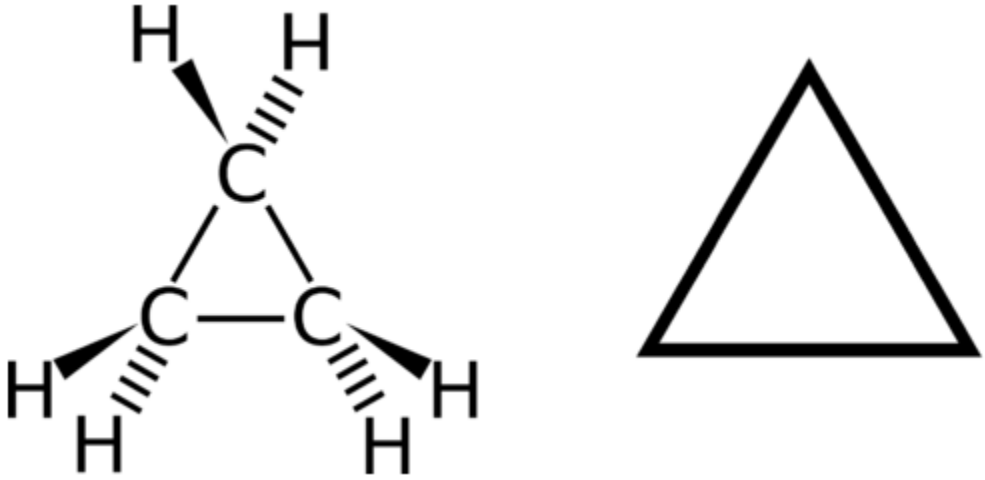
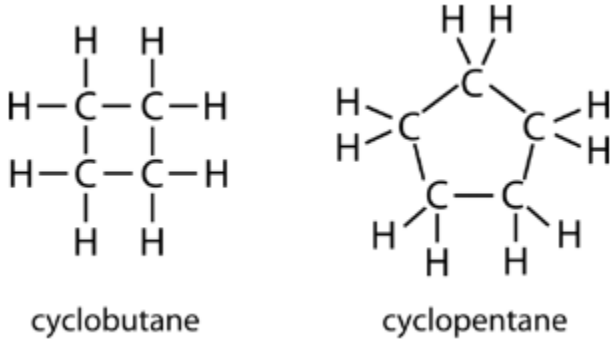


Figure 7.5.1: *Cyclopropane is the simplest cycloalkane. Its highly strained geometry makes it rather unstable and highly reactive.*

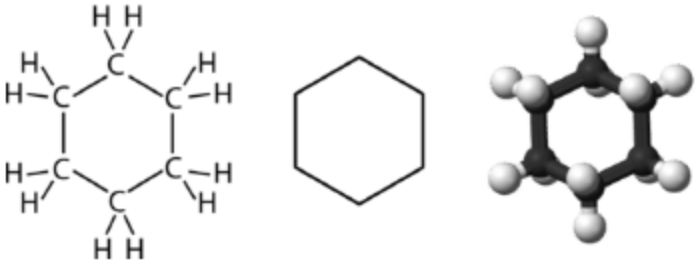
The structural formulas of cyclic hydrocarbons can be represented in multiple ways, two of which are shown above. Each atom can be shown as in the structure on the left from the figure above. A convenient shorthand is to omit the element symbols and only show the shape, as in the triangle on the right. Carbon atoms are understood to be the vertices of the triangle.

The carbon atoms in cycloalkanes are still sp^3 hybridized, with an ideal bond angle of 109.5°. However, an examination of the cyclopropane structure shows that the triangular structure results in a $\mathrm{C{-}C{-}C}$ bond angle of 60°. This deviation from the ideal angle is called ring strain and makes cyclopropane a fairly unstable and reactive molecule. Ring strain is decreased for cyclobutane, with a bond angle of 90°, but is still significant. Cyclopentane has a bond angle of about 108°. This minimal ring strain for cyclopentane makes it a more stable compound.



cyclobutane cyclopentane

Cyclohexane is a six-carbon cycloalkane shown below



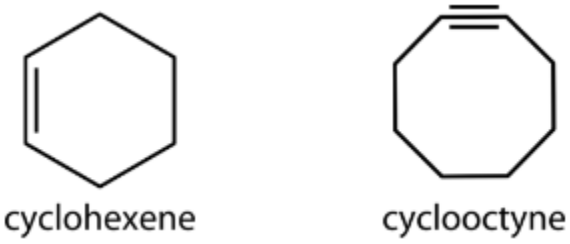
All three of the depictions of cyclohexane are somewhat misleading because the molecule is not planar. In order to reduce the ring strain and attain a bond angle of approximately 109.5°, the molecule is puckered. The puckering of the ring means that every other carbon atom is above and below the plane. The figure below shows two possibilities for the puckered cyclohexane molecule. Each of the structures is called a conformation. The conformation on the right is called the boat conformation, while the one on the left is called the chair conformation.



Figure 7.5.2: *Chair (left) and boat (right) conformations for cyclohexane.*

While both conformations reduce the ring strain compared to a planar molecule, the chair is preferred. This is because the chair conformation results in fewer repulsive interactions between the hydrogen atoms. However, interconversion readily occurs between the two conformations.

Larger cycloalkanes also exist, but are less common. Cyclic hydrocarbons may also be unsaturated. A **cycloalkene** is a cyclic hydrocarbon with at least one carbon-carbon double bond. A **cycloalkyne** is a cyclic hydrocarbon with at least one carbon-carbon triple bond. Shown below are the simplified structural formulas for cyclohexene and cyclooctyne.



Summary

Definition of cyclic hydrocarbon and cyclic alkane are given. Names and structures of typical cyclic hydrocarbons are given.

Contributors

-

This page titled 9.5: Cyclics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.6: Isomers

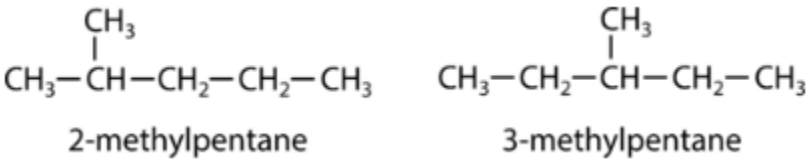
As we get more into the complexities of organic chemistry, we will see how molecular shape affects reactions. One common reaction for alkenes is the addition of hydrogen across the double bond to form the corresponding alkane. Because of the geometry of the reaction, the different 2-butene shapes have different heats of reaction. These differences are important both from a theoretical standpoint as well as from the point of view of industrial applications. Greater energy requirements mean a higher cost and a more expensive product.

Isomers

One of the interesting aspects of organic chemistry is that it is three-dimensional. A molecule can have a shape in space that may contribute to its properties. Molecules can differ in the way the atoms are arranged - the same combination of atoms can be assembled in more than one way. These compounds are known as **isomers**. Isomers are molecules with the same molecular formulas, but different arrangements of atoms. We will look at some isomer possibilities for alkanes and alkenes.

Structural Isomers

A **structural isomer** is one in which two or more organic compounds have the same molecular formulas but different structures. The two pentane molecules below differ only in the location of the methyl group.



Alkenes can also demonstrate structural isomerism. In alkenes, there are multiple structural isomers based on where in the chain the double bond occurs. The condensed structural formulas of 1-butene and 2-butene show this.

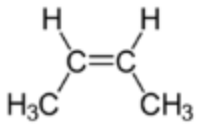


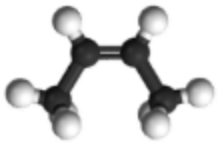
The number in the name of the alkene refers to the lowest numbered carbon in the chain that is part of the double bond.

Geometric Isomers

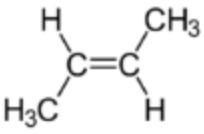
With a molecule such as 2-butene, a different type of isomerism called geometric isomerism can be observed. **Geometric isomers** are isomers in which the order of atom bonding is the same but the arrangement of atoms in space is different. The double bond in an alkene is not free to rotate because of the nature of the pi bond. Therefore, there are two different ways to construct the 2-butene molecule. The image below shows the two geometric isomers, called *cis*-2-butene and *trans*-2-butene.

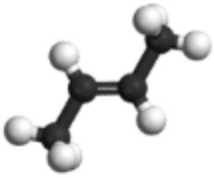
cis-2-butene



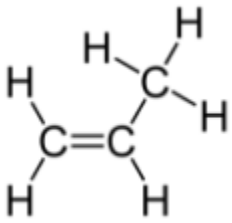


trans-2-butene





The *cis* isomer has the two single hydrogen atoms on the same side of the molecule, while the *trans* isomer has them on opposite sides of the molecule. In both molecules, the bonding order of the atoms is the same. In order for geometric isomers to exist, there must be a rigid structure in the molecule to prevent free rotation around a bond. If the double bond in an alkene was capable of rotating, the two geometric isomers above would not exist. In addition, the two carbon atoms must each have two different groups attached in order for there to be geometric isomers. Propene has no geometric isomers because one of the carbon atoms has two single hydrogens bonded to it.



Physical and chemical properties of geometric isomers are generally different. While *cis*-2-butene is a polar molecule, *trans*-2-butene is nonpolar. Heat or irradiation with light can be used to bring about the conversion of one geometric isomer to another. The input of energy must be large enough to break the pi bond between the two carbon atoms, which is weaker than the sigma bond. At that point,the now single bond is free to rotate and the isomers can interconvert.

As with alkenes, alkynes display structural isomerism beginning with 1-butyne and 2-butyne. However, there are no geometric isomers with alkynes because there is only one other group bonded to the carbon atoms that are involved in the triple bond.

Summary

- Structural and geometric isomers are defined.
- Examples of alkane and alkene isomers are given.

Contributors

-

This page titled 9.6: Isomers is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.7: Hydrogenation

In the late 1970's the lipid hypothesis came in to existences stating that eating saturated fats leads to elevated LDL (Low Density Lipoprotein) which was perceived to be "bad cholesterol." This will result in coronary heart disease which is hardening and narrowing of arteries resulting in heart attack. Fats were eventually classified in to 2 categories: “healthy fats” and “unhealthy fats”. Unhealthy fats where perceived to be *saturated* fats and healthy fats where perceived to be *unsaturated* fats.

A meta-analysis of 72 studies with over 103,052 people have found no validity in the lipid hypothesis. The conclusion of the Meta-Analysis was,“In contrast to current recommendations, this systematic review found no evidence that saturated fat increases the risk of coronary disease, or that polyunsaturated fats have a cardio protective effect.”[1] Dietary fats play a critical role in human health. They help keep cells healthy, help with brain development, help with the use of fat soluble vitamins, and they help cushion organs protecting them against blunt trauma. Fats come in multiple forms, saturated, unsaturated and trans fats just to name a few.

Saturated fats are solid at room temperature due to their molecular shape. The term saturated is in reference to an sp^3 carbon chain that has its remaining sp^3 orbitals bonded with hydrogen atoms. Thus the term “saturated”. It’s “saturated” with hydrogen. Saturated fats have a chain like structure which allows them to stack very well forming a solid at room temperature. Unsaturated fats are not linear due to double bonded carbons which results in a different molecular shape because the sp^2 carbons are trigonal planar, not tetrahedral (sp^3 carbons) as the carbons are in saturated fats. This change in structure will cause the fat molecules to not stack very well resulting in fats that are liquid at room temperature. Butter is mostly saturated fat, that’s why it’s solid at room temperature. Olive Oil is liquid at room temperature, thus it’s an unsaturated fat. An unsaturated fat can be made in to a saturated fat via hydrogenation reactions.

Hydrogenation Reaction

Unsaturated fatty acids may be converted to saturated fatty acids by the relatively simple hydrogenation reaction. Recall that the addition of hydrogen to an alkene (unsaturated) results in an alkane (saturated). A simple hydrogenation reaction is:

$$\mathrm{H_2C{=}CH_2 + H_2 \ \rightarrow CH_3CH_3} \tag{9.7.1}$$

alkene plus hydrogen yields an alkane

Vegetable oils are commonly referred to as "polyunsaturated". This simply means that there are several double bonds present. Vegetable oils may be converted from liquids to solids by the hydrogenation reaction. Margarines and shortenings are "hardened" in this way to make them solid or semi-solids.

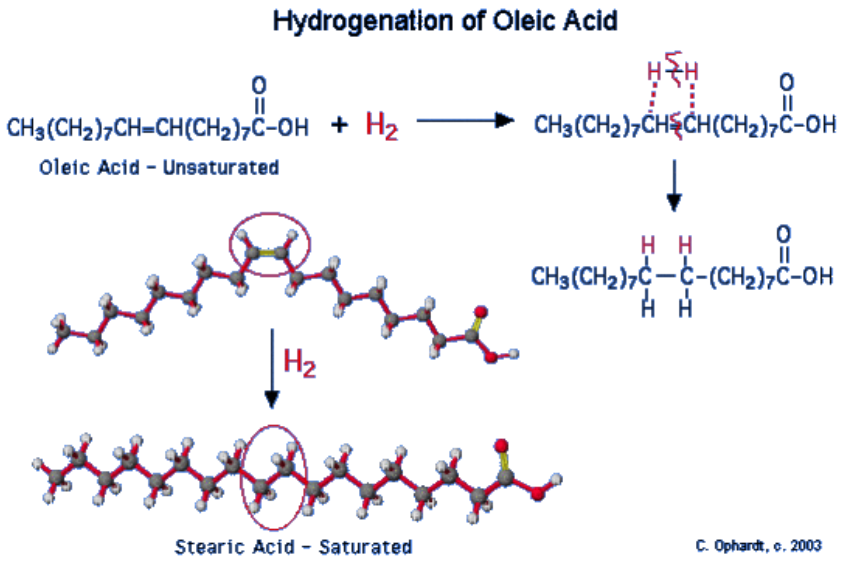
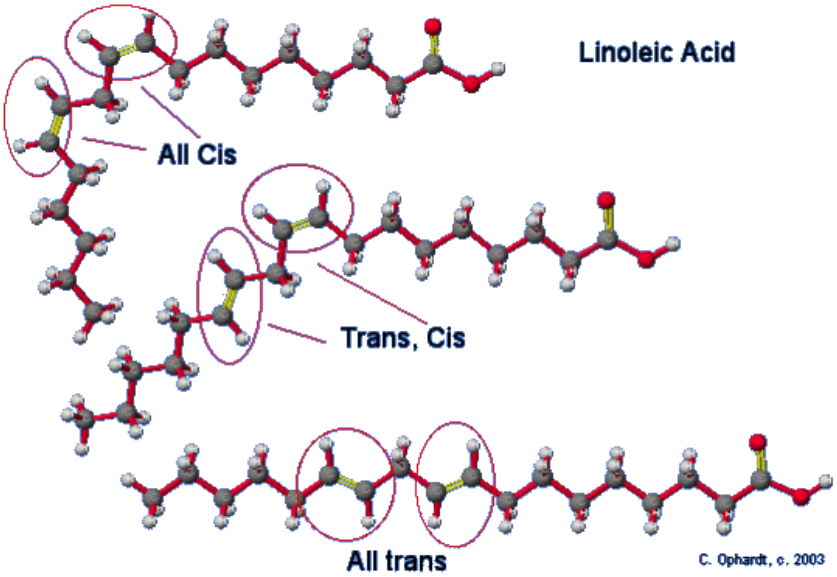


Figure 1: Hydrogenation of a oleic fatty acid

Vegetable oils which have been partially hydrogenated, are now partially saturated so the melting point increases to the point where a solid is present at room temperature. The degree of hydrogenation of unsaturated oils controls the final consistency of the product. What has happened to the healthfulness of the product which has been converted from unsaturated to saturated fats?

Trans Fat

A major health concern during the hydrogenation process is the production of trans fats. Trans fats are the result of a side reaction with the catalyst of the hydrogenation process. This is the result of an unsaturated fat which is normally found as a cis isomer converts to a trans isomer of the unsaturated fat. Isomers are molecules that have the same molecular formula but are bonded together differently. Focusing on the sp^2 double bonded carbons, a cis isomer has the hydrogens on the same side. Due to the added energy from the hydrogenation process, the activation energy is reached to convert the cis isomers of the unsaturated fat to a trans isomer of the unsaturated fat. The effect is putting one of the hydrogens on the opposite side of one of the carbons. This results in a trans configuration of the double bonded carbons. The human body does not recognize trans fats.



Although trans fatty acids are chemically "monounsaturated" or "polyunsaturated," they are considered so different from the cis

monounsaturated or polyunsaturated fatty acids that they can not be legally designated as unsaturated for purposes of labeling. Most of the trans fatty acids (although chemically still unsaturated) produced by the partial hydrogenation process are now classified in the same category as saturated fats.

The major negative is that trans fat tends to raise "bad" LDL- cholesterol and lower "good" HDL-cholesterol, although not as much as saturated fat. Trans fat are found in margarine, baked goods such as doughnuts and Danish pastry, deep-fried foods like fried chicken and French-fried potatoes, snack chips, imitation cheese, and confectionery fats.

Contributors

- Charles Ophardt, Professor Emeritus, Elmhurst College; Virtual Chembook,
- Antonio Rodriguez

This page titled 9.7: Hydrogenation is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

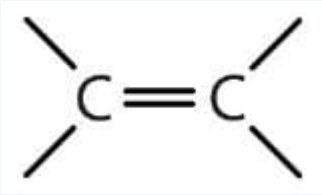
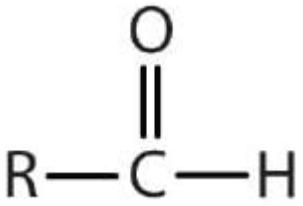
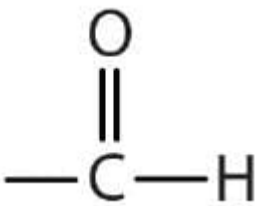
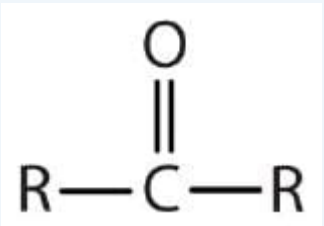
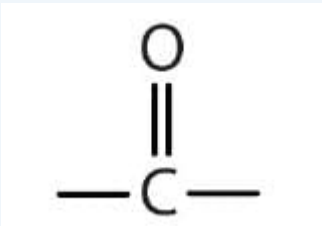
9.8: Functional Group introduction

Learning Objectives

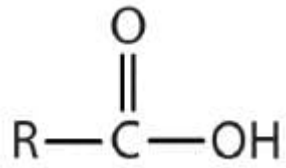
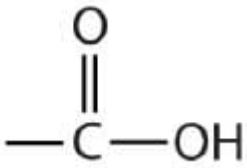
- to describe functional groups and explain why they are useful in the study of organic chemistry.

Previously, we considered several kinds of hydrocarbons. Now we examine some of the many organic compounds that contain functional groups. We first introduced the idea of the functional group, a specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to the molecule. If you understand the behavior of a particular functional group, you will know a great deal about the general properties of that class of compounds. In this chapter, we make a brief yet systematic study of some of organic compound families. Each family is based on a common, simple functional group that contains an oxygen atom or a nitrogen atom. Some common functional groups are listed in Table 7.8.1.

Table 7.8.1 *Selected Organic Functional Groups*

Name of Family	General Formula	Functional Group	Suffix*
alkane	RH	none	*-ane*
alkene	$R_2C{=}CR_2$		*-ene*
alkyne	RC≡CR	–C≡C–	*-yne*
alcohol	ROH	–OH	*-ol*
thiol	RSH	–SH	*-thiol*
ether	ROR	–O–	ether
aldehyde			*-al*
ketone			*-one*

***Ethers do not have a suffix in their common name; all ethers end with the word *ether*.**

Name of Family	General Formula	Functional Group	Suffix*
carboxylic acid			*-oic acid*
***Ethers do not have a suffix in their common name; all ethers end with the word *ether*.**			

Concept Review Exercises

1. What is the functional group of an alkene? An alkyne?
2. Does $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$ have a functional group? Explain.

Answers

1. carbon-to-carbon double bond; carbon-to-carbon triple bond
2. No; it has nothing but carbon and hydrogen atoms and all single bonds.

Key Takeaway

- The functional group, a structural arrangement of atoms and/or bonds, is largely responsible for the properties of organic compound families.

Exercises

1. What is the functional group of 1-butanol ($CH_3CH_2CH_2CH_2OH$)?
2. What is the functional group of butyl bromide, $CH_3CH_2CH_2CH_2Br$?

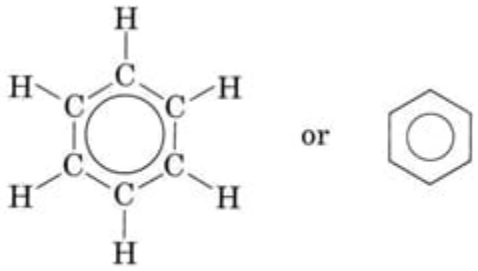
Answer

1. OH

This page titled 9.8: Functional Group introduction is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.9: Aromatics

Aromatic hydrocarbons are one of the three classes of compounds found in petroleum. They are less abundant than the alkanes and cycloalkanes, amounting to only a few percent of the total, but they are quite important commercially. Most aromatic hydrocarbons contain a **benzene ring**. You will recall from the discussion on resonance that benzene, C_6H_6 contains a flat ring of six carbon atoms joined by bonds which are intermediate in character between single and double bonds. The benzene ring is usually indicated by



In the latter structure the lines represent C—C bonds, but carbon and hydrogen atoms, as well as C—H bonds, have been omitted. The benzene ring is very stable, surviving unchanged in most chemical reactions. It is very different in reactivity and shape from the puckered six-membered rings found in cycloalkanes. Below are 3D Jmol models of both cyclohexane and benzene.

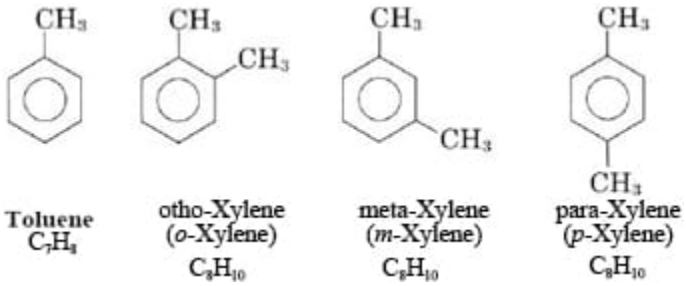
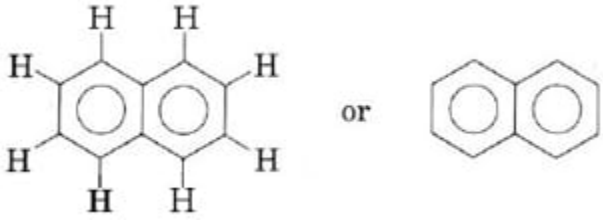


Figure 9.9.1: Examples of aromatic hydrocarbons found in crude oil

Note that the three xylenes are also isomers. Compounds containing two benzene rings joined together, such as naphthalene, are also found in crude oil, though they are much rarer than benzene-related compounds.



Aromatic hydrocarbons are much more common in coal than in petroleum, though in the United States they are mostly manufactured from the latter. In addition to their use in motor fuel, they may be made into dyes, plastics, explosives, detergents, insecticides, medicines, and many other products. In 2000, a total of 6.74×10^{12} liters of benzene were produced in the US, after compensation for exportation and importation.

Some aromatic compounds, benzene among them, are toxic. The compound 1,2-benzopyrene was the cause of the first demonstrated case of occupational disease.



1,2-Benzopyrene

During the eighteenth century chimney sweeps in London were found to have extremely high rates of skin cancer relative to the average person. This was eventually traced to the carcinogenic (cancer-causing) properties of 1,2-benzopyrene in the soot which coated the insides of the chimneys they cleaned. Small quantities of the compound were produced by inefficient combustion of coal in the fireplaces used to heat London houses.

References

1. Price, C. "Summary of available production volumes in millions of gallons per year for selected volatile organic compounds(VOCs) studied by the National Water Quality Assessment(NAWQA) Program." USGS. March 07, 2005. water.usgs.gov/nawqa/vocs/national_assessment/aboutvocs/prod_data.xls

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 9.9: Aromatics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.10: Alcohols

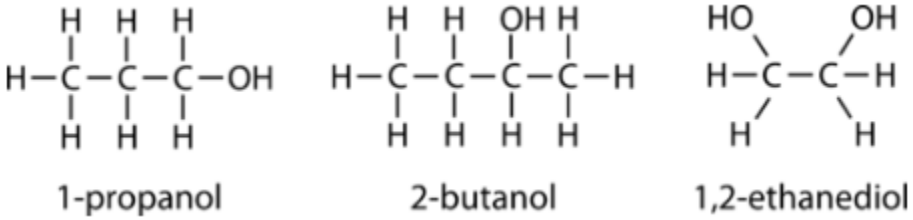
Water freezes at $0^{\circ}\mathrm{C}$ which creates problems for cars in the winter. The water in the engine would freeze and crack the engine block. To prevent this, antifreeze is added to lower the freezing point of the liquid. The most common antifreeze is an alcohol known as propylene glycol. It has largely replaced the much more toxic ethylene glycol. Methanol can also be used as an antifreeze, mainly in windshield wiper fluid.

Alcohols

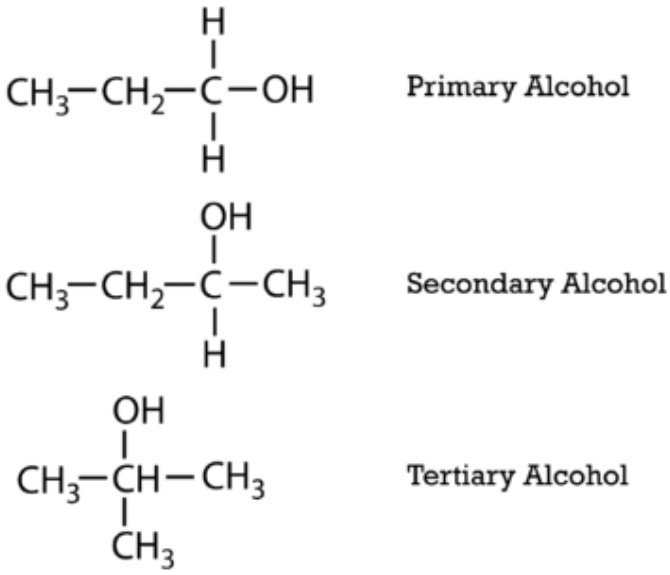
An **alcohol** is an organic compound that contains one or more hydroxyl $(-\mathrm{OH})$ groups. The general formula for alcohols is $\mathrm{R}-\mathrm{OH}$. Do not confuse alcohols with inorganic bases that contain the hydroxide ion (OH^{-}). The $-\mathrm{OH}$ group in an alcohol is covalently bonded to a carbon atom and does not ionize in solution. The steps for naming alcohols are listed below.

1. Name the parent compound by finding the longest continuous carbon atom chain that also contains the hydroxyl group. If there is one hydroxyl group in the molecule, change the final *-e* in the name of the alkane to *-ol*. If there is more than one hydroxyl group, use the full name of the alkane and add a suffix to indicate the number of hydroxyl groups. For example, two hydroxyl groups is *-diol*, three is *-triol*, etc.
2. Number the carbon chain in a way that makes the sum of the hydroxyl numbers as low as possible.
3. Add the numerical prefix into the name before the name of the alcohol.
4. Separate numbers with commas and separate numbers from names or prefixes with a hyphen. There are no spaces in the name.

Following are three examples of alcohols and their IUPAC names.



Aliphatic alcohols can be classified according to the number of R groups attached to the carbon with the hydroxyl group. If one R group is attached to that carbon, the alcohol is a primary alcohol. If two R groups are attached, the alcohol is a secondary alcohol. If three R groups are attached, the alcohol is a tertiary alcohol. Shown below is an example of each. The primary alcohol is 1-propanol, the secondary alcohol is 2-butanol, and the tertiary alcohol is 2-methyl-2-propanol.



Properties of Alcohols

The smallest and lightest alcohols (methanol, ethanol, propanol) are completely soluble in water in all proportions. In a solution, the hydroxyl groups of alcohol molecules and the water molecules form hydrogen bonds with each other, resulting in complete miscibility. However, as the length of the carbon chain increases, the solubility decreases. The solubility of 1-butanol is $7.4\ \text{g}$ per $100\ \text{g}$ of water, while that of 1-pentanol is $2.7\ \text{g}$ per $100\ \text{g}$ of water, and 1-octanol is $0.06\ \text{g}$ per $100\ \text{g}$ of water. The carbon chain portion of the larger alcohol molecule is nonpolar and leads to the decreased solubility of the overall compound.

The presence of hydrogen bonds in alcohols also explains the relatively high boiling points of alcohols compared to the alkanes of similar molar mass (see table below).

***Table 25.9.1**: Boiling Point Comparison of Alkanes and Alcohols*

Compound	Formula	Molar Mass (g/mol)	Boiling Point $(^{\circ}\text{C})$
ethane	CH_3CH_3	32	-88
methanol	CH_3OH	30	64.7
propane	$\text{CH}_3\text{CH}_2\text{CH}_3$	44	-42.1
ethanol	$\text{CH}_3\text{CH}_2\text{OH}$	46	78.3

Only weak London dispersion forces hold molecules of nonpolar alkanes together in the liquid phase. Consequently, less energy is required to break these molecules away from the surface of the liquid and turn them into a vapor. The stronger hydrogen bonding between alcohol molecules means that more energy is required to convert the liquid to vapor, and boiling points are therefore high.

This page briefly discusses some of the important uses of simple alcohols, such as methanol, ethanol and propan-2-ol.

Alcoholic Drinks

The word "alcohol" in alcoholic drinks refers to ethanol ($\text{CH}_3\text{CH}_2\text{OH}$).

Industrial methylated spirits

Ethanol is usually sold as industrial methylated spirits, which is ethanol with a small quantity of methanol and possibly some color added. Because methanol is poisonous, industrial methylated spirits are unfit to drink, allowing purchasers to avoid the high taxes

levied on alcoholic drinks.

Use of ethanol as a fuel

Ethanol burns to produce carbon dioxide and water, as shown in the equation below, and can be used as a fuel in its own right or in mixtures with petrol (gasoline). "Gasohol" is a petrol/ethanol mixture containing approximately 10–20% ethanol. Because ethanol can be produced by fermentation, this is a useful method for countries without an oil industry to reduce the amount of petrol imports.

$$CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O \tag{9.10.1}$$

Ethanol as a solvent

Ethanol is widely used as a solvent. It is relatively safe and can be used to dissolve many organic compounds that are insoluble in water. It is used, for example, in many perfumes and cosmetics.

Methanol as a fuel

Methanol also burns to form carbon dioxide and water:

$$2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O \tag{9.10.2}$$

It can be used a a petrol additive to improve combustion, and its use as a fuel in its own right is under investigation.

Methanol as an industrial feedstock

Most methanol is used to make other compounds, for example, methanal (formaldehyde), ethanoic acid, and methyl esters of various acids. In most cases, these are then converted into further products.

Contributors

-
- Jim Clark (Chemguide.co.uk)

This page titled 9.10: Alcohols is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.11: Ethers

This page titled 9.11: Ethers is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

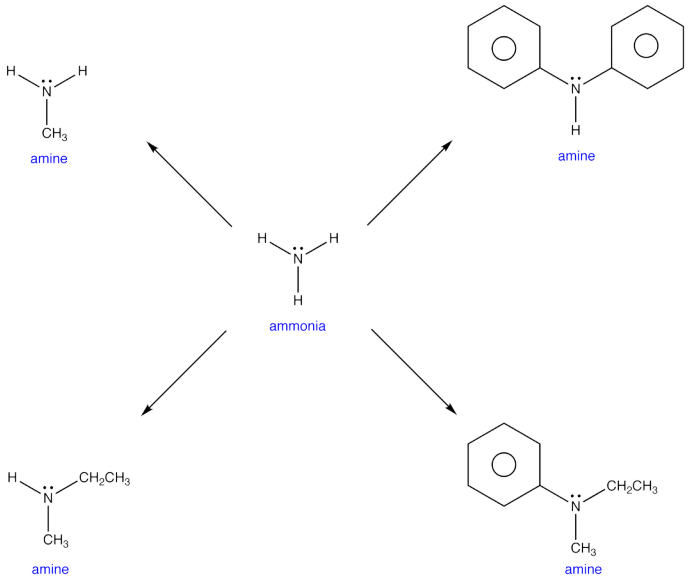
9.12: Carbonyl Friends

This page titled 9.12: Carbonyl Friends is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.13: Amines and Thiols

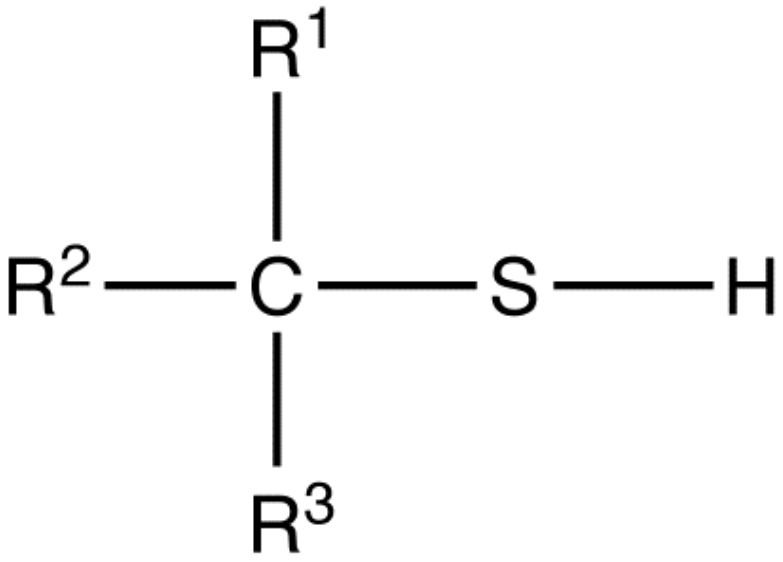
Amine

A compound obtained by replacing one or more hydrogen atoms in the ammonia molecule by alkyl and/or aryl groups is called an amine.



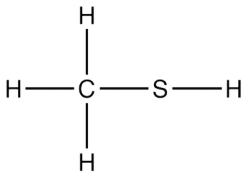
Thiols

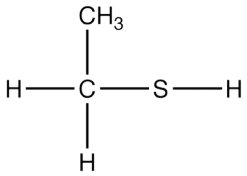
A thiol, thio alcohol, or mercaptan is a compound that has the following general structural formula.

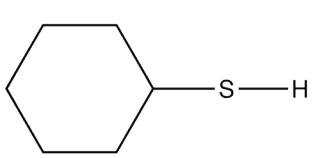


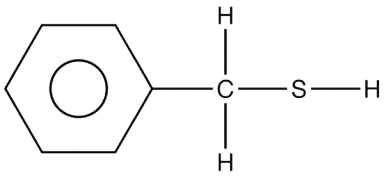
R^1—R^3 could be hydrogen atoms, alkyl groups, aryl groups, or any combination thereof.

eg:









The sulfhydryl group in a thiol is called the thiol group.

Contributors

- Gamini Gunawardena from the OChemPal site (Utah Valley University)

This page titled 9.13: Amines and Thiols is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

9.E: Organic Chemistry (Exercises)

This page titled 9.E: Organic Chemistry (Exercises) is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10: PHARMACEUTICALS

A pharmaceuticals are drugs used to diagnose, cure, treat, or prevent disease. Drug therapy is an important part of the medical field and relies on the science of pharmacology for continual advancement and on pharmacy for appropriate management.A very broad definition of a drug would include "all chemicals other than food that affect living processes." If the affect helps the body, the drug is a medicine. However, if a drug causes a harmful effect on the body, the drug is a poison. The same chemical can be a medicine and a poison depending on conditions of use and the person using it. Another definition would be "medicinal agents used for diagnosis, prevention, treatment of symptoms, and cure of diseases."

10.1: Chemotherapy

10.2: Side Effects and Warnings

10.3: Allergy and Cold Meds

10.4: Combination Medications

10.5: Stomach Medicine

10.6: DEA and Controlled Substances

10.7: Narcotics

10.8: Basics of Neurotransmitters

10.9: Antianxiety Medications

10.10: Amphetamines

10.11: Antipsychotic medication

10.12: Analgesics

10.13: Narcotic Analgesics

10.14: Antidepressant medications

This page titled 10: Pharmaceuticals is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.1: Chemotherapy

A very broad definition of a drug would include "all chemicals other than food that affects living processes." If the effect helps the body, the drug is a medicine. However, if a drug causes a harmful effect on the body, the drug is a poison. The same chemical can be a medicine and a poison depending on conditions of use and the person using it. Another definition would be "medicinal agents used for diagnosis, prevention, treatment of symptoms, and cure of diseases." Contraceptives would be outside of this definition unless the pregnancy was considered a disease.



Image taken from: https://torange.biz/fx/drugs-colored...ct-very-124521

Disease Classification

A disease is a condition of impaired health resulting from a disturbance in the structure or function of the body. Diseases may be classified into the following major categories:

1. Infections caused by viruses, rickettsia, bacteria, fungi, protozoa, and worms
2. Allergic diseases caused by antigens and foreign substances
3. Metabolic disorders caused by defects in the body's ability to carry out normal reactions - these may be hereditary, deficiency, and congenital defects
4. Cancer
5. Toxic diseases caused by poisons
6. Psychosomatic and mental diseases

Chemotherapy, broadly defined, means the treatment of any disease by chemicals including infectious and non-infectious diseases. The original definition applied only to drugs which were used in the treatment of infectious diseases. The proper term for the treatment of non-infectious diseases is pharmacodynamics.

Drug Classification

Drugs can be classified according to various criteria including chemical structure or pharmacological action. The preferred classification is the latter one which may be divided into main groups as follows:

1. Chemotherapeutic agents - used to cure infectious diseases and cancer. (Sulfa drugs, Antibiotics)
2. Pharmacodynamic agents - used in non-infectious diseases (Cholinergic, Adrenergic, Hallucinogenic, Sedatives)
3. Miscellaneous agents (Narcotic Analgesics, Local Anesthetics)

Drug Names

Drugs have three or more names including a: chemical name, brand or trade name, and generic or common name. The chemical name is assigned according to rules of nomenclature of chemical compounds. The brand name is always capitalized and is

selected by the manufacturer. The generic name refers to a common established name irrespective of its manufacturer.

In most cases, a drug bearing a generic name is equivalent to the same drug with a brand name. However, this equivalency is not always true. Although drugs are chemically equivalent, different manufacturing processes may cause differences in pharmacological action. Several differences may be crystal size or form, isomers, crystal hydration, purity-(type and number of impurities), vehicles, binders, coatings, dissolution rate, and storage stability.

Mode of Drug Action

It is important to distinguish between actions of drugs and their effects. Actions of drugs are the biochemical physiological mechanisms by which the chemical produces a response in living organisms. The effect is the observable consequence of a drug action. For example, the action of penicillin is to interfere with cell wall synthesis in bacteria and the effect is the death of the bacteria.

One major problem of pharmacology is that no drug produces a single effect. The primary effect is the desired therapeutic effect. Secondary effects are all other effects beside the desired effect which may be either beneficial or harmful. **Drugs are chosen to exploit differences between normal metabolic processes and any abnormalities which may be present.** Since the differences may not be very great, drugs may be nonspecific in action and alter normal functions as well as the undesirable ones. This leads to undesirable side effects.

The biological effects observed after a drug has been administered are the result of an interaction between that chemical and some part of the organism. Mechanisms of drug action can be viewed from different perspectives, namely, the site of action and the general nature of the drug-cell interaction.

Killing Foreign Organisms

Chemotherapeutic agents act by killing or weakening foreign organisms such as bacteria, worms, viruses. The main principle of action is selective toxicity, i.e. the drug must be more toxic to the parasite than to the host.

Stimulation and Depression

Drugs act by stimulating or depressing normal physiological functions. Stimulation increases the rate of activity while depression reduces the rate of activity.

Sites of Drug Action

Enzyme Inhibition

Drugs act within the cell by modifying normal biochemical reactions. Enzyme inhibition may be reversible or non-reversible; competitive or non-competitive. Antimetabolites may be used which mimic natural metabolites. Gene functions may be suppressed.

Drug-Receptor Interaction

Drugs act on the cell membrane by physical and/or chemical interactions. This is usually through specific drug receptor sites known to be located on the membrane. A receptor is the specific chemical constituents of the cell with which a drug interacts to produce its pharmacological effects. Some receptor sites have been identified with specific parts of proteins and nucleic acids. In most cases, the chemical nature of the receptor site remains obscure.

Non-specific Interactions

Drugs act exclusively by physical means outside of cells. These sites include external surfaces of skin and gastrointestinal tract. Drugs also act outside of cell membranes by chemical interactions. Neutralization of excess stomach acid by antacids is a good example.

Contributors

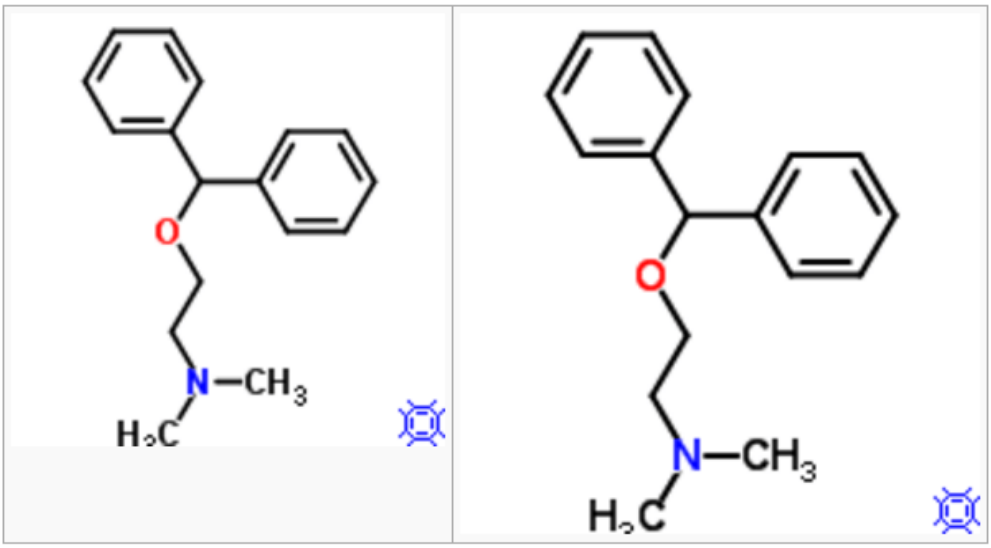
Charles Ophardt (Professor Emeritus, Elmhurst College); Virtual Chembook

This page titled 10.1: Chemotherapy is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.2: Side Effects and Warnings

This page titled 10.2: Side Effects and Warnings is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.3: Allergy and Cold Meds



- **Main ChemSpider page**
- Molecular formula: $C_{17}H_{21}NO$
- Molar mass: 255.355
- CAS Registry Number: 58-73-1
- Appearance: Not available
- Melting point: 169 °C
- Boiling point: Not available
- Solubility: Water, 3060 mg/L (37 deg C)
- Safety sheet: Not available

From Wikipedia

Diphenhydramine is an antihistamine mainly used to treat allergies. Common side effects include sleepiness, poor coordination, and an upset stomach. It is a first generation H1-antihistamine and works by blocking certain effects of histamine. It is available as a generic medication. In the United States, it costs less than for a typical month supply. It is sold under the trade name Benadryl among others.

Sominex is the trademarked name for several over the counter sleep aids.

Two different formulations of Sominex are available, depending upon the market. Both formulations contain a significant dose of a first generation antihistamine with hypnotic properties.

In the United States, Sominex is marketed by Prestige Brands and has as its active ingredient diphenhydramine hydrochloride. Diphenhydramine hydrochloride is also marketed as a sleep aid in Canada under the same name.

In the UK, Actavis UK markets promethazine hydrochloride under the Sominex name. Actavis also markets a product by the name of Sominex Herbal, which is a herbal sleep aid with three natural active ingredients: hops, valerian, and passion flower extracts.

In India, the Sominex name is used by Lunar BioPharma Pvt. Ltd. to market a nutritional supplement intended to provide "long lasting relief from heart burn, acid peptic isorders and indigestion". This formulation of Sominex is not related to the two formulations (containing either diphenhydramine or promethazine) that are marketed as sleep aids by Prestige Brands and Actavis in the United States and the United Kingdom, respectively.

Diphenhydramine and promethazine are both marketed by names other than Sominex for several indications including and besides their use as sleep aids. For example, diphenhydramine is sold under the brand name Benadryl as an anti-histamine for use against allergies. Tylenol PM contains diphenhydramine in addition to acetaminophen for the purpose of inducing sleep and reducing pain and/or fever. Promethazine is marketed by the brand name Phenergan for use against allergy symptoms and motion sickness in the US, where a prescription for promethazine is required. It is also available in the form of a syrup either alone or in combination with one or more of several other active ingredients including codeine and dextromethorphan, for the treatment of allergy symptoms, cough, or cold, among other uses. In Canada, promethazine is available as an over the counter anti-histamine by the name of Histantil.

Other names

2-(Diphenylmethoxy)-N,N-dimethylethanamine (IUPAC Name); 2-Diphenylmethoxy-N,N-dimethylethanamine; alpha-(2-Dimethylaminoethoxy)diphenylmethane; beta-Dimethylaminoethyl Benzhydryl Ether; Dibendrin; Diphenhydramine Base; DPH; 2-(Diphenylmethoxy)-n,n-dimethyl-ethanamine

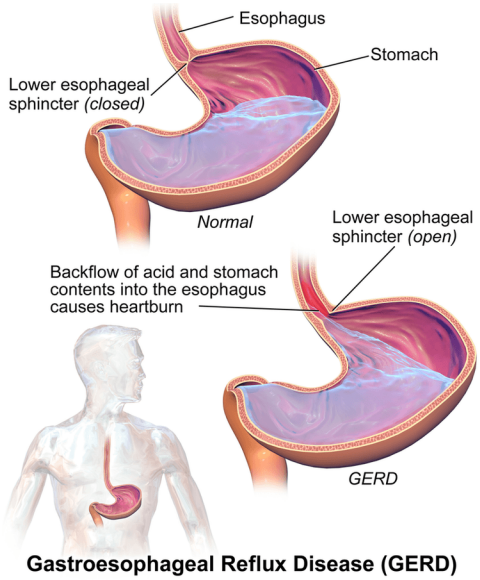
This page titled 10.3: Allergy and Cold Meds is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.4: Combination Medications

This page titled 10.4: Combination Medications is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.5: Stomach Medicine

Heartburn is caused by a buildup of excessive amounts of stomach acid, particularly HCl. This acid is used to digest the food we eat, but it can often back up into the esophagus causing that burning sensation many of us are familiar with. The symptoms of heartburn can be treated with a mild base, which acts to neutralize the excess HCl. For example, Tums is a very commonly used antacid that can be bought over the counter. Other antacids, such as Alka Seltzer, work in similar ways. Below, show the mechanism and products for reaction in which calcium carbonate (the active ingredient in Tums) neutralizes HCl in a simple proton transfer reaction. Following this acid-base reaction, carbonic acid quickly degrades into CO_2 and H_2O.



Gastroesophageal Reflux Disease (GERD)

A comparison of a healthy condition to GERD. Image used with permission (Cc BY-SA 4.0; BruceBlaus .

This page titled 10.5: Stomach Medicine is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.6: DEA and Controlled Substances

This page titled 10.6: DEA and Controlled Substances is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.7: Narcotics

Narcotic agents are potent analgesics which are effective for the relief of severe pain. Analgesics are selective central nervous system depressants used to relieve pain. The term analgesic means "without pain". Even in therapeutic doses, narcotic analgesics can cause respiratory depression, nausea, and drowsiness. Long term administration produces tolerance, psychic, and physical dependence called addiction.

Introduction

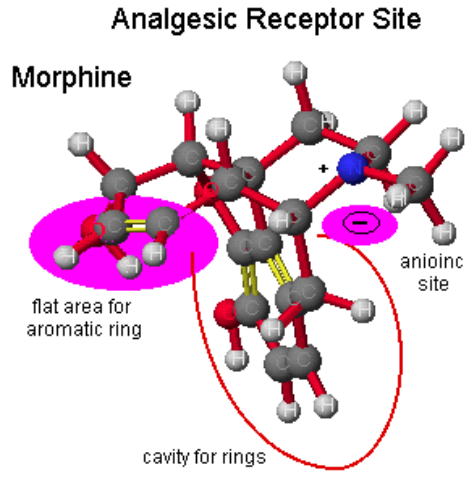
Narcotic agents may be classified into four categories:

1. Morphine and codeine - natural alkaloids of opium.
2. Synthetic derivatives of morphine such as heroin.
3. Synthetic agents which resemble the morphine structure.
4. Narcotic antagonists which are used as antidotes for overdoses of narcotic analgesics.

The main pharmacological action of analgesics is on the cerebrum and medulla of the central nervous system. Another effect is on the smooth muscle and glandular secretions of the respiratory and gastro-intestinal tract. The precise mechanism of action is unknown although the narcotics appear to interact with specific receptor sites to interfere with pain impulses.

Receptor Site

A schematic for an analgesic receptor site may look as shown in the graphic below with morphine. Three areas are needed: a flat areas to accommodate a flat nonpolar aromatic ring, a cavity to accept another series of rings perpendicular, and an anionic site for polar interaction of the amine group.



C. Ophardt, c. 2003

Enkephalins

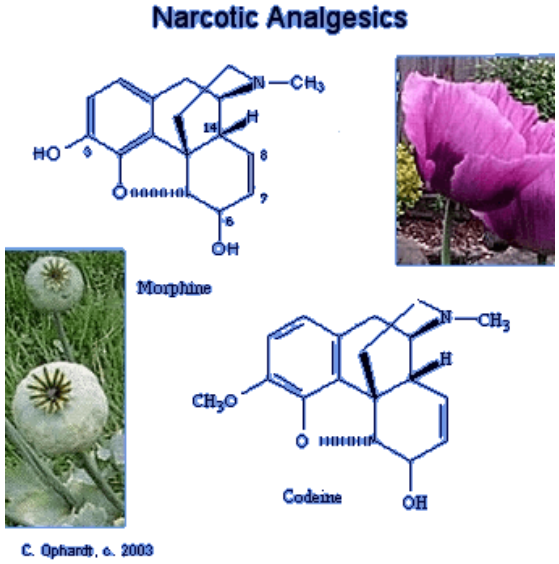
Recently investigators have discovered two compounds in the brain called enkephalins which resemble morphine in structure. Each one is a peptide composed of 5 amino acids and differ only in the last amino acid. The peptide sequences are: tyr-gly-gly-phe-leu and tyr-gly-gly-phe-met. Molecular models show that the structures of the enkephalins has some similarities with morphine. The main feature in common appears to be the aromatic ring with the -OH group attached (tyr). Methadone and other similar analgesics have 2 aromatic rings which would be similar to the enkephalins (tyr and phe).

Analgesics may relieve pain by preventing the release of acetylcholine. Enkephalin molecules are released from a nerve cell and bind to analgesic receptor sites on the nerve cell sending the impulse. The binding of enkephalin or morphine-like drugs changes the shape of the nerve sending the impulse in such a fashion as to prevent the cell from releasing acetylcholine. As a result, the pain impulse cannot be transmitted and the brain does not preceive pain.

Morphine and Codeine

Morphine exerts a narcotic action manifested by analgesia, drowsiness, changes in mood, and mental clouding. The major medical action of morphine sought in the CNS is analgesia. Opiates suppress the "cough center" which is also located in the brain stem, the medulla. Such an action is thought to underlie the use of opiate narcotics as cough suppressants. Codeine appears to be particularly effective in this action and is widely used for this purpose.

Narcotic analgesics cause an addictive physical dependence. If the drug is discontinued, withdrawal symptoms are experienced. Although the reasons for addiction and withdrawal symptoms are not completely known, recent experiments have provided some information. A nucleotide known as cyclicadenosine monophosphate (cAMP) is synthesized with the aid of the enzyme adenylate cyclase. Enkephalin and morphine-like drugs inhibit this enzyme and thus decrease the amount of cAMP in the cells. In order to compensate for the decreased cAMP, the cells synthesize more enzyme in an attempt to produce more cAMP. Since more enzyme has been produced, more morphine is required as an inhibitor to keep the cAMP at a low level. This cycle repeats itself causing an increase in the tolerance level and increasing the amounts of morphine required. If morphine is suddenly withheld, withdrawal symptoms are probably caused by a high concentration of cAMP since the synthesizing enzyme, adenylate cyclase, is no longer being inhibited.



Morphine and codeine are contained in opium from the poppy (*Papaver Somniterum*) plant found in Turkey, Mexico, Southeast Asia, China, and India. This plant is 3-4 feet tall with 5-8 egg shaped capsules on top. Ten days after the poppy blooms in June, incisions are made in the capsules permitting a milky fluid to ooze out. The following day the gummy mass (now brown) is carefully scraped off and pressed into cakes of raw opium to dry.

Opium contains over 20 compounds but only morphine (10%) and codeine (0.5%) are of any importance. Morphine is extracted from the opium and isolated in a relatively pure form. Since codeine is in such low concentration, it is synthesized from morphine by an ether-type methylation of an alcohol group. Codeine has only a fraction of the potency compared to morphine. It is used with aspirin and as a cough suppressant.

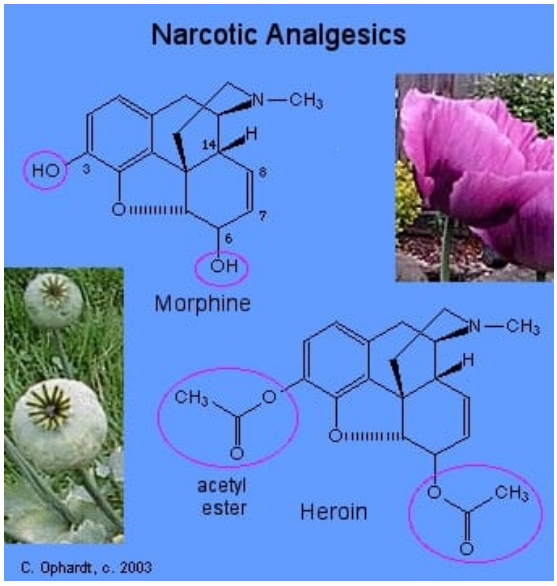
Heroin

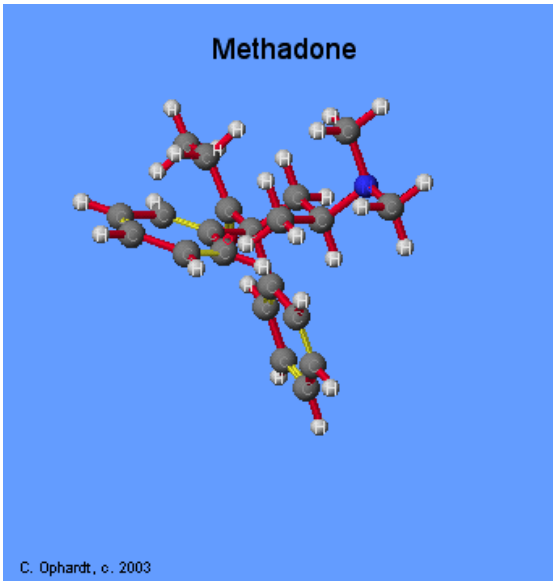
Heroin is synthesized from morphine by a relatively simple esterification reaction of two alcohol (phenol) groups with acetic anhydride (equivalent to acetic acid). Heroin is much more potent than morphine but without the respiratory depression effect. A possible reason may be that heroin passes the blood-brain barrier much more rapidly than morphine. Once in the brain, the

heroin is hydrolyzed to morphine which is responsible for its activity.

Synthetic narcotic analgesics include meperidine and methadone. Meperidine is the most common subsitute for morphine. It exerts several pharmacological effects: analgesic, local anesthetic, and mild antihistamine. This multiple activity may be explained by its structural resemblance to morphine, atropine, and histamine.

Methadone is more active and more toxic than morphine. It can be used for the relief of may types of pain. In addition it is used as a narcotic substitute in addiction treatment because it prevents morphine abstinence syndrome. Methadone was synthesized by German chemists during Wold War II when the United States and our allies cut off their opium supply. And it is difficult to fight a war without analgesics so the Germans went to work and synthesized a number of medications in use today, including demerol and darvon which is structurally simular to methadone. And before we go further lets clear up another myth. Methadone, or dolophine was not named after Adolf Hitler. The "dol" in dolophine comes from the latin root "dolor." The female name Dolores is derived from it and the term dol is used in pain research to measure pain e.g., one dol is 1 unit of pain.

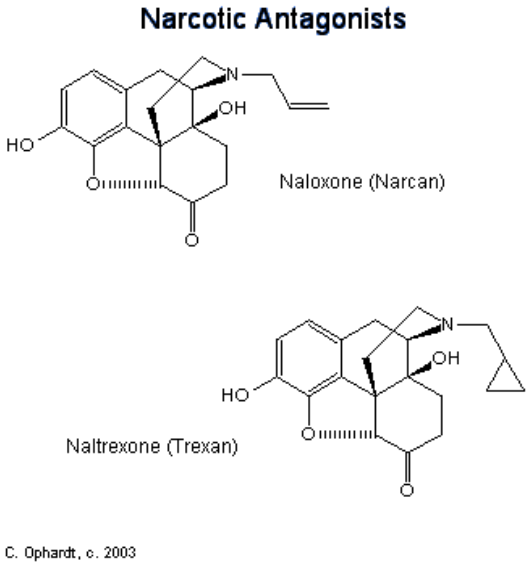




Even methadone, which looks strikingly different from other opioid agonists, has steric forces which produce a configuration that closely resembles that of other opiates. See the graphic on the left and the top graphic on this page. In other words, steric forces bend the molecule of methadone into the correct configuration to fit into the opiate receptor. When you take methadone it first must be metabolized in the liver to a product that your body can use. Excess methadone is also stored in the liver and blood stream and this is how methadone works its 'time release trick' and last for 24 hours or more. Once in the blood stream metabolized methadone is slowly passed to the brain when it is needed to fill opiate receptors. Methadone is the only effective treatment for heroin addiction. It works to smooth the ups and down of heroin craving and allows the person to function normally.

Narcotic Antagonists

Narcotic Antagonists prevent or abolish excessive respiratory depression caused by the administration of morphine or related compounds. They act by competing for the same analgesic receptor sites. They are structurally related to morphine with the exception of the group attached to nitrogen.



Nalorphine precipitates withdrawal symptoms and produces behavioral disturbances in addition to the antogonism action. Naloxane is a pure antagonist with no morphine like effects. It blocks the euphoric effect of heroin when given before heroin. Naltrexone became clinically available in 1985 as a new narcotic antagonist. Its actions resemble those of naloxone, but naltrexone is well is well absorbed orally and is long acting, necessitating only a dose of 50 to 100 mg. Therefore, it is useful in narcotic treatment programs where it is desired to maintain an individual on chronic therapy with a narcotic antagonist. In individuals taking naltrexone, subsequent injection of an opiate will produce little or no effect. Naltrexone appears to be particularly effective for the treatment of narcotic dependence in addicts who have more to gain by being drug-free rather than drug dependant

Contributors and Attributions

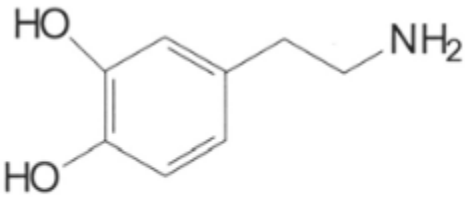
- Charles Ophardt, Professor Emeritus, Elmhurst College; Virtual Chembook
- Poppies image from: leda.lycaeum.org

This page titled 10.7: Narcotics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

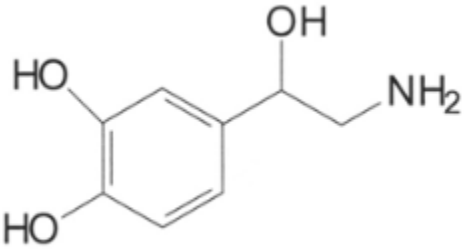
- Narcotic Analgesic Drugs by Charles Ophardt is licensed CC BY-NC-SA 4.0.

10.8: Basics of Neurotransmitters

Dopamine (DA), 4,5-dihydroxyphenethylamine or 4-(2-aminoethyl)1,2-benzenediol, is a known neurotransmitter that is involved in the chemical transmission of nerve impulses in the mammalian brain. It is a member of the catecholamine family and a precursor to epinephrine (adrenaline) and norepinephrine (noradrenaline) in the biosynthetic pathways.



Dopamine



Norepinephrine

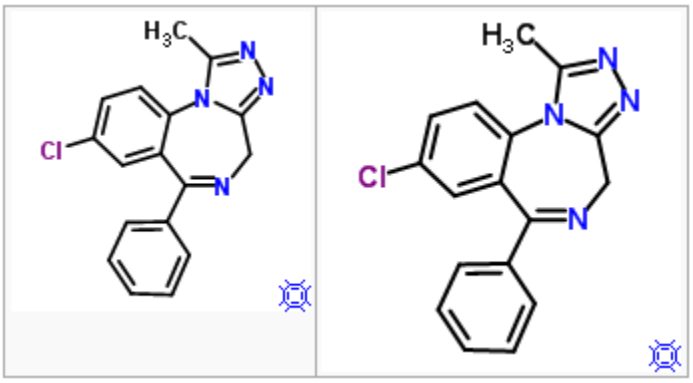
DA has a molecular formula of $C_8H_{11}NO_2$ and a formula weight of 153.18 [ref. 1]. It is a water-soluble hormone released by the hypothalamus. Imbalance in dopamine activity can cause brain dysfunction related to two major disorders, Parkinson's disease and schizophrenia [ref. 2,3]. Researchers are also looking at dopamine neurotransmission in drug abuse ranging from stimulants, such as amphetamines and cocaine, to depressants, such as morphine and other opioids, and alcohol [ref. 3].

Several amine neurotransmitters such as DA, noradrenaline (norepinephrine), adrenaline and serotonin are electroactive so that they can be monitored electrochemically. Most undergo a chemical reaction following the initial electron transfer step, an *ec* mechanism, as evaluated by cyclic voltammetry (CV) in this experiment. In biological fluids, prior separation with HPLC is recommended in conjunction with an electrochemical detector (HPLC-ECD).

Great strides in learning about the role and fate of DA and other neurotransmitters in brains have come about in recent years due to the ability to monitor these compounds in-vivo. The major breakthrough making this possible came when Adams and co-workers [ref. 4] implanted small carbon electrodes (fibers) in rat brain to detect in-vivo catecholamine neurotransmitters. The development and application of the methodology are discussed in *"Probing Brain Chemistry: Voltammetry Comes of Age"* [ref. 5]. This article is available online and is a recommended reading as background to this experiment. Venton and Wightman [ref. 6] in a more recent article propose calling this new subject area "psychoanalytical chemistry" in which sensors, like microelectrodes, can detect neurotransmitter dopamine and determine how its neurochemistry affects and correlates with animal behavior.

This page titled 10.8: Basics of Neurotransmitters is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.9: Antianxiety Medications



- **Main ChemSpider page**
- Molecular formula: $C_{17}H_{13}ClN_4$
- Molar mass: 308.765
- CAS Registry Number: 28981-97-7
- Appearance: Not available
- Melting point: 228 to 229 °C
- Boiling point: Not available
- Solubility: Not available
- Safety sheet: Not available

From Wikipedia

Alprazolam, available under the trade name Xanax, is a potent, short-acting anxiolytic of the benzodiazepine class. It is commonly used for the treatment of anxiety disorders, especially of panic disorder, but also in the treatment of generalized anxiety disorder(GAD) or social anxiety disorder. Alprazolam, like other benzodiazepines, binds to specific sites on the GABAA receptor. It possesses anxiolytic, sedative, hypnotic, skeletal muscle relaxant, anticonvulsant, and amnestic properties. Alprazolam is available for oral administration in compressed tablet (CT) and extended-release capsule (XR) formulations.

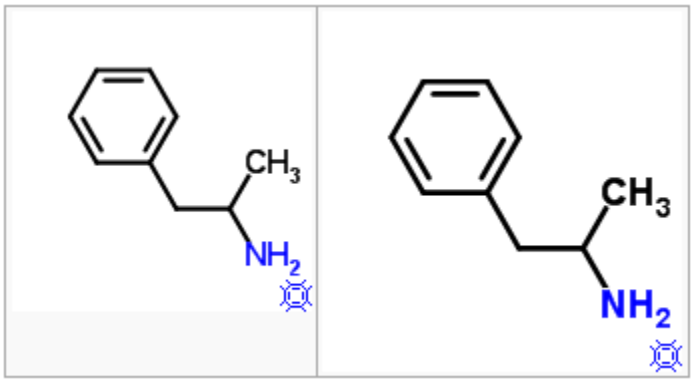
Peak benefits achieved for generalized anxiety disorder (GAD) may take up to a week. Tolerance to the anxiolytic/antipanic effects is controversial with some authoritative sources reporting the development of tolerance, tolerance will, however, develop to the sedative-hypnotic effects within a couple of days. The first approved use was panic disorder and within two years of its original marketing Xanax became a blockbuster drug in the US. Presently, alprazolam is the most prescribed and the most misused benzodiazepine in the US. The potential for misuse among those taking it for medical reasons is controversial with some expert reviews stating that the risk is low and similar to that of other benzodiazepine drugs Compared to the large number of prescriptions, relatively few individuals increase their dose on their own initiative or engage in drug-seeking behavior. Alprazolam is classified as a schedule IV controlled substance by the U.S. Drug Enforcement Administration (DEA).

Other names

Alprax; Anxyl; Panistat; Panix; Xanax; Xanor

This page titled 10.9: Antianxiety Medications is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.10: Amphetamines



- **Main ChemSpider page**
- Molecular formula: $C_9H_{13}N$
- Molar mass: 135.206
- CAS Registry Number: 300-62-9
- Appearance: Not available
- Melting point: Not available
- Boiling point: Not available
- Solubility: Not available
- Safety sheet: Not available

Adderall

Adderall is a combination drug containing salts of the two enantiomers of amphetamine, a central nervous system (CNS) stimulant of the phenethylamine class. Adderall is used in the treatment of attention deficit hyperactivity disorder (ADHD) and narcolepsy. It is also used as an athletic performance enhancer and cognitive enhancer, and recreationally as an aphrodisiac and euphoriant. By salt content, the active ingredients of Adderall are 25% levoamphetamine salts (the levorotary or 'left-handed' enantiomer) and 75% dextroamphetamine salts (the dextrorotary or 'right-handed' enantiomer).

Adderall is generally well-tolerated and effective in treating the symptoms of ADHD and narcolepsy. At therapeutic doses, Adderall causes emotional and cognitive effects such as euphoria, change in desire for sex, increased wakefulness, and improved cognitive control. At these doses, it induces physical effects such as decreased reaction time, fatigue resistance, and increased muscle strength. In contrast, much larger doses of Adderall can impair cognitive control, cause rapid muscle breakdown, or induce a psychosis (e.g., delusions and paranoia). The side effects of Adderall vary widely among individuals, but most commonly include insomnia, dry mouth, and loss of appetite. The risk of developing an addiction is insignificant when Adderall is used as prescribed at fairly low daily doses, such as those used for treating ADHD; however, the routine use of Adderall in larger daily doses poses a significant risk of addiction due to the pronounced reinforcing effects that are present at higher doses. Recreational doses of Adderall are generally much larger than prescribed therapeutic doses, and carry a far greater risk of serious adverse effects.

The two amphetamine enantiomers that compose Adderall (i.e., levoamphetamine and dextroamphetamine) alleviate the symptoms of ADHD and narcolepsy by increasing the activity of the neurotransmitters norepinephrine and dopamine in the brain, which results from their interactions with trace amine associated receptor 1 (TAAR1) and vesicular monoamine transporter 2 (VMAT2) in neurons. Dextroamphetamine is a more potent stimulant than levoamphetamine, but levoamphetamine has slightly stronger cardiovascular and peripheral effects and a longer elimination half-life (i.e., it remains in the body longer) than dextroamphetamine. The levoamphetamine component of Adderall has been reported to improve the treatment response in some individuals relative to dextroamphetamine alone. Adderall's active ingredient, amphetamine, shares many chemical and pharmacological properties with the human trace amines, particularly phenethylamine and , the latter of which is a positional isomer of amphetamine.

Amphetamine

Amphetamine (contracted from) is a potent central nervous system (CNS) stimulant that is used in the treatment of attention deficit hyperactivity disorder (ADHD), narcolepsy, and obesity. Amphetamine was discovered in 1887 and exists as two enantiomers: levoamphetamine and dextroamphetamine. Amphetamine properly refers to a specific chemical, the racemic free base, which is equal parts of the two enantiomers, levoamphetamine and dextroamphetamine, in their pure amine forms. The term is frequently used informally to refer to any combination of the enantiomers, or to either of them alone. Historically, it has been used to treat nasal congestion and depression. Amphetamine is also used as an athletic performance enhancer and cognitive enhancer, and recreationally as an aphrodisiac and euphoriant. It is a prescription drug in many countries, and unauthorized possession and distribution of amphetamine are often tightly controlled due to the significant health risks associated with recreational use.

The first pharmaceutical amphetamine was Benzedrine, a brand which was used to treat a variety of conditions. Currently, pharmaceutical amphetamine is prescribed as racemic amphetamine, Adderall, dextroamphetamine, or the inactive prodrug lisdexamfetamine. Amphetamine, through activation of a trace amine receptor, increases monoamine and excitatory neurotransmitter activity in the brain, with its most pronounced effects targeting the catecholamine neurotransmitters norepinephrine and dopamine.

At therapeutic doses, amphetamine causes emotional and cognitive effects such as euphoria, change in desire for sex, increased wakefulness, and improved cognitive control. It induces physical effects such as decreased reaction time, fatigue resistance, and increased muscle strength. Larger doses of amphetamine may impair cognitive function and induce rapid muscle breakdown. Drug addiction is a serious risk with large recreational doses but is unlikely to arise from typical long-term medical use at therapeutic doses. Very high doses can result in psychosis (e.g., delusions and paranoia) which rarely occurs at therapeutic doses even during long-term use. Recreational doses are generally much larger than prescribed therapeutic doses and carry a far greater risk of serious side effects.

Amphetamine belongs to the phenethylamine class. It is also the parent compound of its own structural class, the substituted amphetamines, which includes prominent substances such as bupropion, cathinone, MDMA, and methamphetamine. As a member of the phenethylamine class, amphetamine is also chemically related to the naturally occurring trace amine neuromodulators, specifically phenethylamine and , both of which are produced within the human body. Phenethylamine is the parent compound of amphetamine, while is a positional isomer of amphetamine that differs only in the placement of the methyl group.

Other names

1-Phenyl-2-propanamine (IUPAC Name); 1-phenyl-2-aminopropane; 1-Phenylpropan-2-amine; adderall; alpha-Methylphenethylamine; alpha-methylphenylethylamine; alpha-Methyl-benzeneethanamine

Contributors

Template:Wikipedia

This page titled 10.10: Amphetamines is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.11: Antipsychotic medication

This page titled 10.11: Antipsychotic medication is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

10.12: Analgesics

A. Aspirin-like Drugs

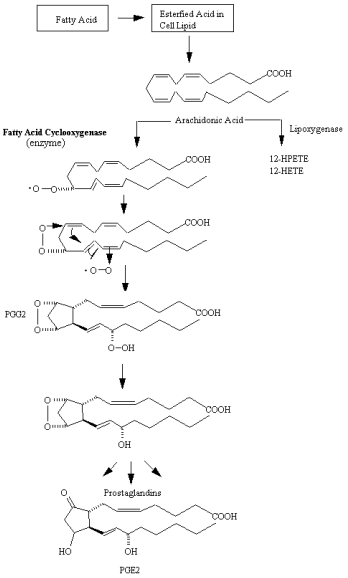
The anti-inflammatory, analgesic, and antipyretic drugs are a heterogeneous group of compounds, often chemically unrelated (although most of them are organic acids), which nevertheless share certain therapeutic actions and side effects. The prototype is aspirin; hence these compounds are often referred to as aspirin-like drugs. All aspirin-like drugs are antipyretic, analgesic, and anti-inflammatory, but there are important differences in their activities. For example, acetaminophen is antipyretic and analgesic but is only weakly anti-inflammatory. The reason for the differences are not clear; variations in the sensitivity of enzymes in the target tissues may be important.

When employed as analgesics, these drugs are usually effective only against pain of low-to-moderate intensity, particularly that associated with inflammation. Aspirin drugs do not change the perception of sensory modalities other than pain. The type of pain is important; chronic postoperative pain or pain arising from inflammation is particularly well controlled by aspirin-like drugs, whereas pain arising from the hollow viscera is usually not relieved.

As antipyretics, aspirin-like drugs reduce the body temperature in feverish states. Although all such drugs are antipyretics and analgesics, some are not suitable for either routine or prolonged use because of toxicity; phenylbutaxone is an example. This class of drugs finds its chief clinical application as anti inflammatory agents in the treatment of musculoskelatal disorders, such as rheumatoid arthritis, osteoarthritis, and ankylosing spondylitis. In general, aspirin-like drugs provide only symptomatic relief from the pain and inflammation associated with the disease and do not arrest the progression of pathological injury.

Chemistry of Action

There has been substantial progress in elucidating the mechanism of action of aspirin-like drugs, and it is now possible to understand why such heterogeneous agents have the same basic therapeutic activities and often the same side effects. Indeed, their therapeutic activity appears to depend to a large extent upon the inhibition of a defined biochemical pathways responsible for the biosynthesis of prostaglandins (see figure below) and related autacoids.



Aspirin-like drugs inhibit the conversion of arachidonic acid to the unstable endoperoxide intermediate, PGG2, which is catalyzed by the cyclooxygenase. Individual agents have differing modes of inhibitory activity on the cyclooxygenase. Aspirin itself acetylates a serine at the active site of the enzyme. Platelets are especially susceptible to this action because (unlike most other cells) they are incapable of regenerating the enzyme, presumably because they have little or no capacity for protein biosynthesis. In practical terms this means that a single dose of aspirin will inhibit the platelet cyclooxygenase for the life of the platelet (8 to

10 days); in man a dose as small as 40 mg per day is sufficient to produce this effect. In contrast to aspirin, salicylic acid has no acetylating capacity and is almost inactive against cyclooxygenase *in vitro*. Nevertheless, it is as active as aspirin in reducing the synthesis of prostaglandins *in vivo*. The basis of this action and, thus, of the anti-inflammatory effect of salicylic acid is not clearly understood. Since aspirin is rapidly hydrolyzed to salicylic acid in vivo (half-life in human plasma, approximately 15 minutes), the acetylated and nonacetylated species probably act as pharmacologically distinct entities.

Most of the other common aspirin-like drugs are irreversible inhibitors of the cyclooxygenase, although there are some exceptions. For indomethacin, the mode of inhibition is particularly complex and probably involves a site on the enzyme different from that which is acetylated by aspirin.

Pain

Prostaglandins are associated particularly with the development of pain that accompanies injury or inflammation. Large doses of PGE2 or PGF2a , given to women by injection to induce abortion, cause intense local pain. Prostaglandins can also cause headache and vascular pain when infused intravenously in man. While the doses of prostaglandins required to elicit pain are high in comparison with the concentrations expected *in vivo*, induction of hyperalgesia occurs when minute amounts of PGE1 are given intradermally to man. Furthermore, in experiments in man where separate infusions of PGE1, bradykinin, or histamine caused no pain, marked pain was experienced when PGE1 was added to bradykinin or histamine. When PGE1 was infused with histamine, itching was also noted.

Fever

The hypothalamus regulates the set point at which body temperature is maintained. In fever, this set point is elevated, and aspirin-like drugs promote its return to normal. These drugs do not influence body temperature when it is elevated by such factors as exercise or increases in the surrounding temperature.

Fever may be a result of infection, tissue damage, inflammation, graft rejection, malignancy, or other disease states. A multitude of microorganisms can cause fever. There is evidence that bacterial endotoxins act by stimulating the biosynthesis and release by neutrophils and other cells of an endogenous pyrogen, a protein with a molecular weight of 10,000 to 20,000. The current view is that the endogenous pyrogen passes from the general circulation into the central nervous system, where it acts upon discrete sites within the brain, especially the preoptic hypothalamic area. There is evidence that the resultant elevation of body temperature is mediated by the release of prostaglandins and that aspirin-like drugs suppress the effects of endogenous pyrogen by inhibiting the synthesis of these substances. The evidence includes the ability of prostaglandins, especially PGE2, to produce fever when infused into the cerebral ventricles or when infected into the hypothalamus. Fever is a frequent side effect of prostaglandins when they are administered to a women as abortifacients. Moreover some studies have demonstrated an increase in prostaglandin-like substances in the cerebrospinal fluid when endogenous pyrogen is injected intravenously. The fever produced by the administration of pyrogen, but not that by prostaglandins, is reduced by aspirin-like drugs.

Side Effects

Aspirin is very useful, but it has many side effects and therefore must be used carefully. Like most powerful drugs, an overdose of aspirin or salicylates can be fatal. If a child or adult takes an overdose of aspirin, induce vomiting to empty the unabsorbed medication from the stomach (if the person is still awake and conscious). Obtain emergency medical care right away.

The most common side effects of aspirin are heartburn and other symptoms of stomach irritation such as indigestion, pain, nausea, and vomiting. The stomach irritation may lead to bleeding from the stomach, which may cause black stools. These symptoms may be reduced by taking aspirin with meals, with an antacid, with a glass of milk, or by taking enteric-coated or timed-release aspirin. Also, it is best not to take aspirin with alcohol or coffee (or other beverages containing caffeine, such as tea or cocoa and many soft drinks). Alcohol and caffeine make the stomach more sensitive to irritation. The non aspirin salicylate preparations sometimes are less irritating to the stomach and may be substituted for aspirin by your doctor.

A few people develop asthma, hay fever, nasal congestion, or hives from aspirin or non-steroid anti-inflammatory drugs (NSAIDs). These people should never take aspirin, nor should people who have active stomach or duodenal ulcers. Anyone who has ever had a peptic ulcer should be very careful about taking aspirin because it can lead to a recurrence.

Aspirin is known to interfere with the action of the platelets. As a result, some people who take a lot of aspirin experience easy

bruising of the skin. Therefore, people who have major bleeding problems should not take aspirin. Also, keep in mind that aspirin should not be taken for 10-14 days before surgery (including surgery in the mouth) to avoid excessive bleeding during or after the operation. These side effects probably depend on aspirin-like drugs' ability to block endogenous prostaglandin biosynthesis. Platelet function appears to be disturbed because aspirin-like drugs prevent the formation by the platelets of thrombozane A2 (TXA2), a potent aggregating agent. This accounts for the tendency of these drugs to increase the bleeding time.

Aspirin increases oxygen consumption by the body, increasing carbon dioxide production-an effect that stimulates respiration. Therefore, overdose with aspirin is often characterized by marked increases in respiratory rate, which cause the overdosed individual to appear to pant. This occurrence results in other, severe, metabolic consequences.

Prolongation of gestation by aspirin-like drugs has been demonstrated in both experimental animals and the human female. Furthermore, prostaglandins of the E and F series are potent uterotropic agents , and their biosynthesis by the uterus increases dramatically in the hours before parturition. It is thus hypothesized that prostaglandins play a major role in the initiation and progression of labor and delivery.

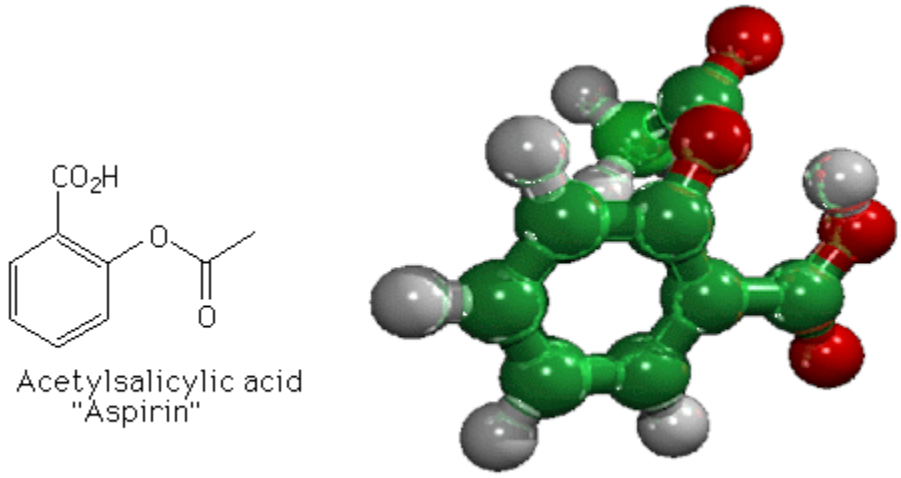
High doses of salicylate may cause ringing in the ears and slight deafness. Sometimes, however, these symptoms indicate mild overdose, which could become more serious.

Aspirin and NSAIDs sometimes affect the normal function of the kidneys and aspirin-like drugs promote the retention of salt and water by reducing the prostaglandin-induced inhibition of both the reabsorption of chloride and the action of antidiuretic hormone. This may cause edema in some patients with arthritis who are treated with an aspirin-like drug.

Recent reports have said there could be a link between the use of aspirin and the development of Reye's syndrome. Reye's syndrome is a rare but possibly fatal disease seen most often in children and teenagers. It usually affects those recovering from chicken pox or a viral illness such as the flu. These reports have raised concern in pediatricians (doctors who specialize in treating children) and parents of children with arthritis who need to take large doses of aspirin to control their disease.

Aspirin

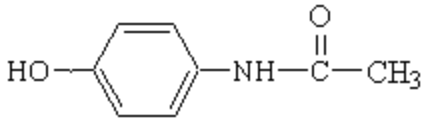
In the U.S., about 10 to 20 thousand tons of aspirin are consumed each year; it is our most popular analgesic. Aspirin is one of the most effective analgesic, antipyretic, and anti-inflammatory agents.



Chemical structure

Acetaminophen (Tylenol)

Acetaminophen is an effective alternative to aspirin as an analgesic and antipyretic agent. However, its anti-inflammatory effect is minor and not clinically useful. It is commonly felt that acetaminophen may have fewer side effects than aspirin, but it should be noted that an acute overdose may produce severe or even fatal liver damage. Acetaminophen does not inhibit platelet aggregation and therefore is not useful for preventing vascular clotting.



Acetaminophen (4'-hydroxyacetanilide)

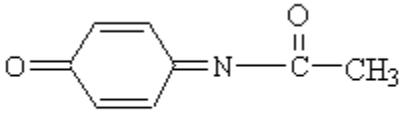
Side effects are usually fewer than those of aspirin; the drug produces less gastric distress and less ringing in the ears. However, as stated previously, overdose can lead to severe damage of the liver.

Acetaminophen has been proved to be a reasonable substitute for aspirin when analgesic or antipyretic effectiveness is desired, especially in patients who cannot tolerate aspirin. This might include patients with peptic ulcer disease of gastric distress or those in whom the anticoagulant action of aspirin might be undesirable.

Aspirin is often combined with acetaminophen in a single tablet for relief of arthritis and other painful conditions. Sometimes other drugs such as caffeine, an antihistamine, nasal drying agents, and sedatives are also added. Although some of these preparations may have special uses for certain acute conditions such as a cold or a headache, they should not be taken for a chronic (long-term) form of arthritis. If a combination is required, each drug should be prescribed separately. The dose of each should be adjusted individually to achieve the greatest benefit with the fewest side effects.

Researchers attribute the pain-relieving activity of acetaminophen to the drug's ability to elevate the pain threshold, although the precise mechanisms involved in this process have not been clearly identified. The antipyretic, or fever-reducing, effect of acetaminophen is far better understood. Research shows that the drug inhibits the action of fever-producing agents on the heat-regulating centers of the brain by blocking the formation and release of prostaglandins in the central nervous system. However, unlike aspirin and other NSAIDs, acetaminophen has no significant effect on the prostaglandins involved in other body processes.

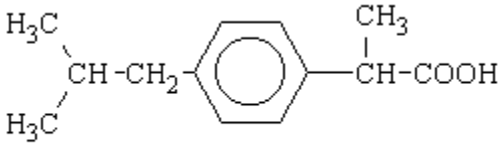
Despite claims to the contrary, stomach upset and hepatic toxicity are statistically as much a problem with acetaminophen as with aspirin-like drugs. Acetaminophen is normally metabolized in the liver and kidney by P450 enzymes. No toxicity is observed with therapeutic doses, however, after ingestion of large quantities (>2,000 mg/kg), a highly reactive metabolite, N-acetyl-p-benzoquinoneimine, is generated (see figure below). This species is electrophilic intermediate which is conjugated with glutathion to a non-toxic compound. Overdosing depletes glutithione and N-acetyl-p-benzoquinone reacts with nucleophilic portions (sulfhdryl groups) of critical liver cell protein. This results in cellular dysfunction and hepatic and renal toxicity. Antidote treatment consists of amino acid supplements to replenish glutathione. The P450 metabolizing enzymes differ somewhat in character between the liver and kidney. Factors that enhance renal toxicity include chronic liver disease, possibly gender, concurrent renal insults, and conditions that alter the activity of P450-metabolizing enzyme systems.



N-acetyl-p-benzoquinoneimine

Other Aspirin-Like Drugs

Other aspirin-like drugs include diflunisal, phenylbutazone, apazone, indomethacine, sulindac, fenamates, tolmetin, ibuprofen (see figure below), and piroxicam.

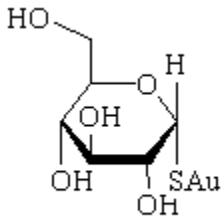


Ibuprofen

Gold

Gold is not of course an aspirin-like drug. However, its end effect is similar to aspirin, so it will be briefly considered here. Gold in elemental form has been employed for centuries as an antipruritic (anti itch medication) to relieve the itching palm. At present, gold treatment includes different forms of gold salts used to treat rheumatoid arthritis and related diseases. In some people, it helps relieve joint pain and stiffness, reduce swelling and bone damage, and reduce the chance of joint deformity and disability.

The significant preparations of gold are all compounds in which the gold is attached to sulfur. The three prominent drugs are aurothioglucose, auranofin, and gold sodium thiomalate.



Aurothioglucose

It takes months for gold compounds to leave the body. This means that side effects to gold therapy may take some time to resolve. Sometimes side effects even appear after the last gold injection. Rash and a metallic taste in the mouth are side effects of gold injections that may not seem serious at first. However, they are early warning signs for more serious reactions. If either of these side effects develop, the health care provider should be contacted promptly.

Some side effects may cause multiple symptoms, not all of which may occur. Side effects with multiple symptoms are:

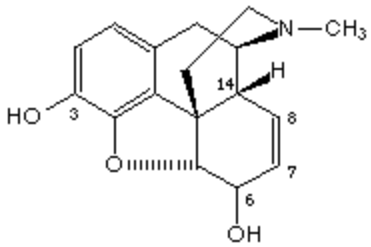
1. Low platelet count (thrombocytopenia). Most often this appears to be an immunological disturbance that results in an accelerated degradation of platelets. Symptoms may include: black, tarry stools, blood in urine, stool, or vomit, small red dots on the skin, nose bleeds, unusual bruising or bleeding.
2. Anaphylaxis. Symptoms include sudden onset of the following soon after a gold injection: trouble swallowing, tightness in the throat, fainting, trouble breathing, wheezing, hives, swelling of the face, usually the lips or around the eyes, liver damage, abdominal pain for more than a few days, light colored stools, albuminuria, yellow eyes or skin, ulcerative colitis, severe abdominal pain or cramps, diarrhea that lasts more than a few days, blood in stool.

B. Opiates, Opiate Antagonists, and Opiate Receptors

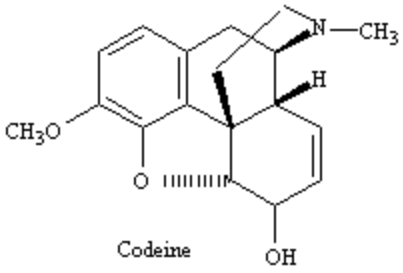
The term opiate refers to any natural or synthetic drug that exerts actions upon the body similar to those induced by morphine, the major pain-relieving agent obtained from the opium poppy (Papaver somniferum). They were so highly regarded in the nineteenth century as remedies for pain, anxiety, cough, and diarrhea that some physicians referred to them as G.O.M.- `God's Own Medicine'. Opiates interact with what appear to be several closely related receptors, and they share some of the properties of certain naturally occurring peptides, the enkephalins, endorphins, and dynorphins.

Opiates

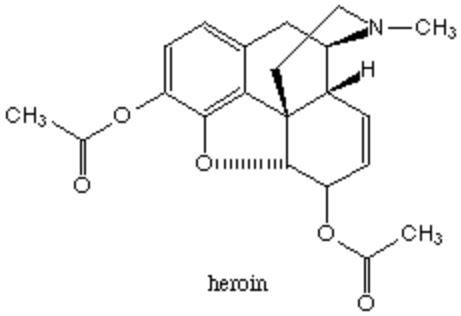
The term opium refers to the crude resinous extract obtained from the opium poppy. Crude opium contains a wide variety of ingredients, including morphine and codeine, both of which are widely used in medicine. The bulk of the ingredients of opium, however, consists of such organic substances as resins, oils, sugars, and proteins that account for more than 75 % of the weight of the opium but exert little pharmacological activity. Morphine is the major pain relieving drug found in opium, being approximately 10% of the crude exudate. Codeine is structurally close to morphine (see figs below), although it is much less potent and amounts to only 0.5% of the opium extract. Heroin does not occur naturally but is a semisynthetic derivative produced by a chemical modification of morphine that increases the potency (see figs. below). It takes only 3 mg. of heroin to produce the same analgesic effect as 10 mg of morphine. However, at these equally effective doses, it may be difficult to distinguish between the effects of the two compounds.



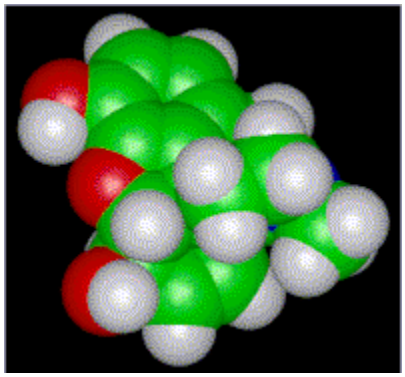
Morphine



Codeine



heroin



filled structure of morphine

Mechanism of Action

Studies of the binding of opioid drugs and peptides to specific sites in brain and other organs have suggested the existence of perhaps as many as eight types of receptors. In the CNS, there is reasonably firm evidence for four major categories of receptors, designated m, k, d , and s . To add confusion, there may well be subtypes of each of these receptors. Although there is considerable variation in binding characteristics and anatomical distribution among different species, inferences have been drawn from data that attempt to relate pharmacological effects to interactions with a particular constellation of receptors. For example, analgesia has been associated with both m and k receptors, while dysphoria or psychotomimetic (alteration of behavior or personality) effects have been ascribed to s receptors; based primarily on their localization in limbic regions of the brain, d receptors are thought to be involved in alterations of affective behavior. The actions of opioid drugs that are currently available have usually been interpreted with respect to the participation of only three types of receptors - m, k, and s ; at each, a given agent may act as an agonist, a partial agonist, or an antagonist (see table). The m receptor is thought to meduate supraspinal analgesia, respiratory depression, euphoria, and physical dependance; the k receptor, spinal analgesia, miosis, and sedation; the s receptor, dysphoria, hallucinations, and respiratory and vasomotor stimulation.

It has been observed that opioids can selectively inhibit certain excitatory inputs to identified neurons. For example, the iontophoretic administration (the induction of an ionized substance through intact skin by the application of a direct current) of morphine into the substantia gelatinosa suppresses the discharge of spinal neurons in lamina IV of the dorsal horn that is evoked by noxious stimuli (e.g. heat) without changing responses to other inputs. While a postsynaptic action at discrete dedritic sites cannot be excluded, these findings suggest that opioids selectively inhibit the release of excitatory transmitters from terminals of nerves carrying pain related stimuli. In other situations, postsynaptic actions of opioids appear to be important. For example, application of opioids to neurons in the locus ceruleus reduces both spontaneous discharge and responses evoked by noxious stimuli. However, excitation of the neurons by antidromic stimulation (i.e. causing the neurons to fire backwards) is also suppressed, and the cells are hyperpolarized by the drugs.

Opioids have been observed to inhibit prostaglandin-induced increases in the accumulation of cyclic AMP in in brain tissue. Of potential relevance to mechanisms that underlie the phenomena of tolerance and withdrawal, the responses to prostaglandins

recover in the continued presence of opioids.

Opioid-induced analgesia is due to actions of several sites within the CNS and involves several systems of neurotransmitters. Although opioids do not alter the threshold or responsivity of afferent nerve endings to noxious stimulation or impair the conduction of the nerve impulses along peripheral nerves, they may decrease conduction of impulses of primary afferent fibers when they enter the spinal cord and decrease activity in other sensory endings. There are opioid binding sites (m receptors) on the terminal axons of primary afferents within laminae I and II (substantia gelatinosa) of the spinal cord and in the spinal nucleus of the trigeminal nerve. Morphine-like drugs acting at this site are thought to decrease the release of neurotransmitters, such as substance P, that mediate transmission of pain impulses.

High doses of opioids can produce muscular rigidity in man, and both opioids and endogenous peptides cause catalepsy, circling, and stereotypical behavior in rats and other animals. These effects are probably related to actions at opioid receptors in the substania nigra and striatum, and involve interactions with both dopaminergic and GABA-ergic neurons.

The mechanism by which opioids produce euphoria, tranquility, and other alterations of mood remains unsettled. Microinjections of opioids into the ventral tegmentum activate dopaminergic neurons that project to the nucleus accumbens. Animals will work to receive such injections, and activation (or disinhibition) of these neurons has been postulated to be a critical element in the reinforcing effects of opioids and opioid-induced euphoria. However, the administration of dopaminergic antagonists does not consistently prevent these reinforcing effects. The neural systems that mediate opioid reinforcement in the ventral tegmentum appear to be distinct from those involved in the classical manifestations of physical dependence and analgesia.

Basic Effects of Morphine

CNS. Morphine exerts a narcotic action manifested by analgesia, drowsiness, changes in mood, and mental clouding. The major medical action of morphine sought in the CNS is analgesia, which may usually be induced by doses below those that cause other effects on the CNS, such as sedation or respiratory depression. The relief of pain by morphine-like opioids is relatively selective, in that other sensory modalities (touch, vibration, vision, hearing, etc.) are not inhibited. Patients frequently report that the pain is still present but that they feel more comfortable. Continuous dull pain is relieved more effectively than sharp intermittent pain, but with sufficient amounts of morphine it is possible to relieve even the severe pain associated with renal or biliary colic. In fact, its analgesic action appears to result not from a decrease of pain impulses into the CNS but from an altered perception of the painful stimuli.

Respiration. A second major action of morphine-like drugs is to depress respiration through interaction with m receptors located in the brainstem. At high doses, respiration may become so slow and irregular that life is threatened. In man, death from morphine poisoning is nearly always due to respiratory arrest. The primary mechanism of respiratory depression by morphine involves a reduction in the responsiveness of the brain stem respiratory centers to increases in carbon dioxide tension (P_{CO2}). High concentrations of opioid receptors, as well, as endogenous peptides, are found in the medullary areas believed to be important in ventilatory control. Respiratory depression is mediated by a subpopulation of m receptors (m 1), distinct from those that are involved in the production of analgesia (m 2). Thus, a 'pure' m 1-opioid agonist could theoretically produce analgesia with little respiratory depression.

Cough. Opiates suppress the "cough center" which is also located in the brainstem, the medulla. Such an action is thought to underlie the use of opiate narcotics as cough suppressants. Codeine appears to be particularly effective in this action and is widely used for this purpose.

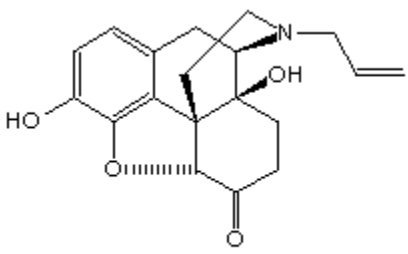
Gastrointestinal Tract. The opiates have been used for centuries for the relief of diarrhea and for the treatment of dysentery, and these uses were developed long before these agents were used as analgesics or euphoiants. Opiates appear to exert their effect on the gastrointestinal tract primarily in the intestine, where peristaltic movements, which normally propel food down the intestine, are markedly diminished. Also, the tone of the intestine is greatly increased to the point where almost complete spastic paralysis of movement occurs. This combination of decreased propulsion and increased tone leads to a marked decrease in the movement of food through the intestine. This stasis is followed by a dehydration of the feces, which hardens the stool and further retards the advance of material. All these effects contribute to the constipating properties of opiates. Indeed, nothing more effective has yet been developed for treating sever diarrhea.

Opiate Antagonists

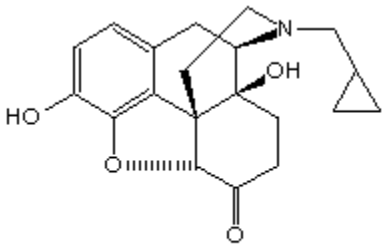
Naloxone, when administered to normal individuals, produces no analgesia, euphoria, or respiratory depression. However, it rapidly precipitates withdrawal in narcotic-dependent individuals. Naloxone antagonizes the actions of morphine at all its

receptors; however its affinity for m receptors is generally more than ten fold higher than for k or d receptors.

The uses of naloxone include the reversal of the respiratory depression that follows acute narcotic intoxication and the reversal of narcotic-induced respiratory depression in newborns of mothers who have received narcotics. The use of naloxone is limited by a short duration of action and the necessity of parenteral route of administration.



naloxone (Narcan®)



naltrexone (Trexan®)

Naltrexone became clinically available in 1985 as a new narcotic antagonist. Its actions resemble those of naloxone, but naltrexone is well is well absorbed orally and is long acting, necessitating only a dose of 50 to 100 mg. Therefore, it is useful in narcotic treatment programs where it is desired to maintain an individual on chronic therapy with a narcotic antagonist. In individuals taking naltrexone, subsequent injection of an opiate will produce little or no effect. naltrexone appears to be particularly effective for the treatment of narcotic dependence in addicts who have more to gain by being drug-free rather than drug dependant.

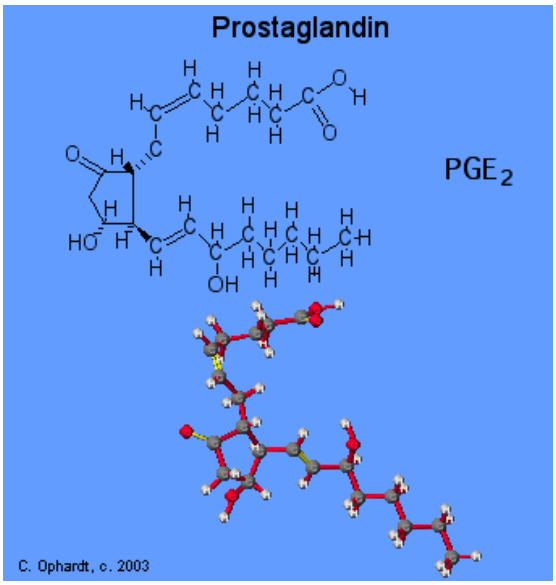
Contributors and Attributions

- Edward B. Walker (Weber State University)

Prostaglandins were first discovered and isolated from human semen in the 1930s by Ulf von Euler of Sweden. Thinking they had come from the prostate gland, he named them prostaglandins. It has since been determined that they exist and are synthesized in virtually every cell of the body. Prostaglandins, are like hormones in that they act as chemical messengers, but do not move to other sites, but work right within the cells where they are synthesized.

Introduction

Prostaglandins are unsaturated carboxylic acids, consisting of of a 20 carbon skeleton that also contains a five member ring. They are biochemically synthesized from the fatty acid, arachidonic acid. See the graphic on the left. The unique shape of the arachidonic acid caused by a series of cis double bonds helps to put it into position to make the five member ring. See the prostaglandin in the next panel



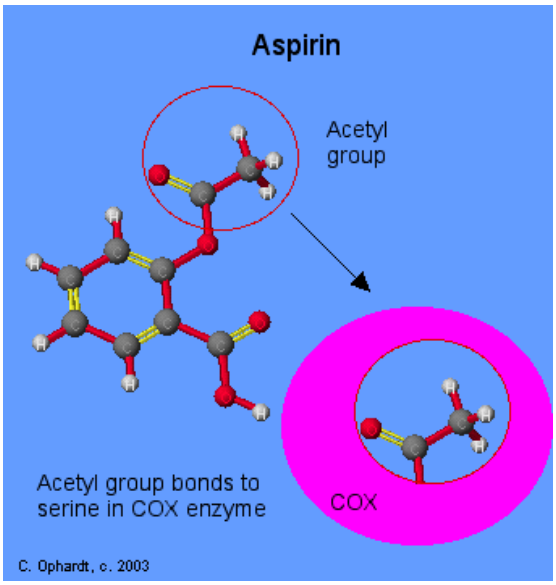
Prostaglandin Structure

Prostaglandins are unsaturated carboxylic acids, consisting of of a 20 carbon skeleton that also contains a five member ring and are based upon the fatty acid, arachidonic acid. There are a variety of structures one, two, or three double bonds. On the five member ring there may also be double bonds, a ketone, or alcohol groups. A typical structure is on the left graphic.

Functions of Prostaglandins

There are a variety of physiological effects including:

1. Activation of the inflammatory response, production of pain, and fever. When tissues are damaged, white blood cells flood to the site to try to minimize tissue destruction. Prostaglandins are produced as a result.
2. Blood clots form when a blood vessel is damaged. A type of prostaglandin called thromboxane stimulates constriction and clotting of platelets. Conversely, PGI2, is produced to have the opposite effect on the walls of blood vessels where clots should not be forming.
3. Certain prostaglandins are involved with the induction of labor and other reproductive processes. PGE2 causes uterine contractions and has been used to induce labor.
4. Prostaglandins are involved in several other organs such as the gastrointestinal tract (inhibit acid synthesis and increase secretion of protective mucus), increase blood flow in kidneys, and leukotriens promote constriction of bronchi associated with asthma.



Effects of Aspirin and other Pain Killers

When you see that prostaglandins induce inflammation, pain, and fever, what comes to mind but aspirin. Aspirin blocks an enzyme called cyclooxygenase, COX-1 and COX-2, which is involved with the ring closure and addition of oxygen to arachidonic acid converting to prostaglandins. The acetyl group on aspirin is hydrolzed and then bonded to the alcohol group of serine as an ester. This has the effect of blocking the channel in the enzyme and arachidonic can not enter the active site of the enzyme. By inhibiting or blocking this enzyme, the synthesis of prostaglandins is blocked, which in turn relives some of the effects of pain and fever. Aspirin is also thought to inhibit the prostaglandin synthesis involved with unwanted blood clotting in coronary heart disease. At the same time an injury while taking aspirin may cause more extensive bleeding.

Contributors

Charles Ophardt (Professor Emeritus, Elmhurst College); Virtual Chembook

This page titled 10.12: Analgesics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

- Analgesics and Anti-Inflammatory Agents is licensed CC BY-NC-SA 4.0.

10.13: Narcotic Analgesics

Narcotic agents are potent analgesics which are effective for the relief of severe pain. Analgesics are selective central nervous system depressants used to relieve pain. The term analgesic means "without pain". Even in therapeutic doses, narcotic analgesics can cause respiratory depression, nausea, and drowsiness. Long term administration produces tolerance, psychic, and physical dependence called addiction.

Introduction

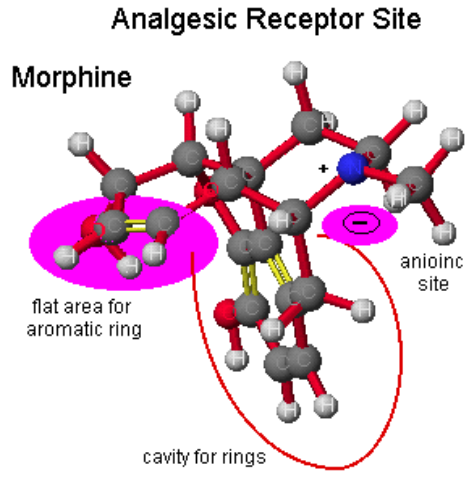
Narcotic agents may be classified into four categories:

1. Morphine and codeine - natural alkaloids of opium.
2. Synthetic derivatives of morphine such as heroin.
3. Synthetic agents which resemble the morphine structure.
4. Narcotic antagonists which are used as antidotes for overdoses of narcotic analgesics.

The main pharmacological action of analgesics is on the cerebrum and medulla of the central nervous system. Another effect is on the smooth muscle and glandular secretions of the respiratory and gastro-intestinal tract. The precise mechanism of action is unknown although the narcotics appear to interact with specific receptor sites to interfere with pain impulses.

Receptor Site

A schematic for an analgesic receptor site may look as shown in the graphic below with morphine. Three areas are needed: a flat areas to accommodate a flat nonpolar aromatic ring, a cavity to accept another series of rings perpendicular, and an anionic site for polar interaction of the amine group.



C. Ophardt, c. 2003

Enkephalins

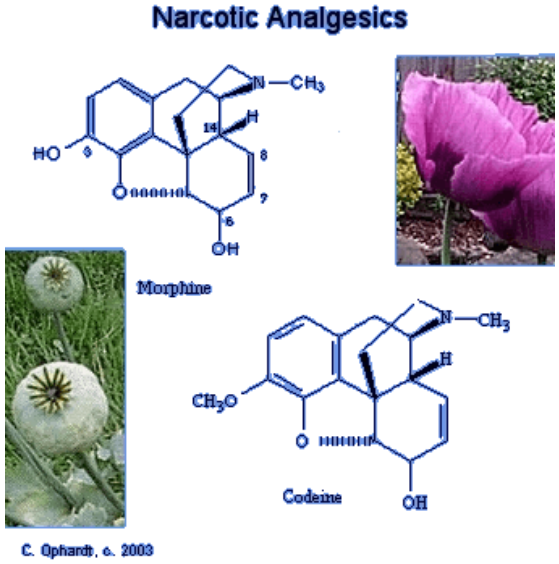
Recently investigators have discovered two compounds in the brain called enkephalins which resemble morphine in structure. Each one is a peptide composed of 5 amino acids and differ only in the last amino acid. The peptide sequences are: tyr-gly-gly-phe-leu and tyr-gly-gly-phe-met. Molecular models show that the structures of the enkephalins has some similarities with morphine. The main feature in common appears to be the aromatic ring with the -OH group attached (tyr). Methadone and other similar analgesics have 2 aromatic rings which would be similar to the enkephalins (tyr and phe).

Analgesics may relieve pain by preventing the release of acetylcholine. Enkephalin molecules are released from a nerve cell and bind to analgesic receptor sites on the nerve cell sending the impulse. The binding of enkephalin or morphine-like drugs changes the shape of the nerve sending the impulse in such a fashion as to prevent the cell from releasing acetylcholine. As a result, the pain impulse cannot be transmitted and the brain does not preceive pain.

Morphine and Codeine

Morphine exerts a narcotic action manifested by analgesia, drowsiness, changes in mood, and mental clouding. The major medical action of morphine sought in the CNS is analgesia. Opiates suppress the "cough center" which is also located in the brain stem, the medulla. Such an action is thought to underlie the use of opiate narcotics as cough suppressants. Codeine appears to be particularly effective in this action and is widely used for this purpose.

Narcotic analgesics cause an addictive physical dependence. If the drug is discontinued, withdrawal symptoms are experienced. Although the reasons for addiction and withdrawal symptoms are not completely known, recent experiments have provided some information. A nucleotide known as cyclicadenosine monophosphate (cAMP) is synthesized with the aid of the enzyme adenylate cyclase. Enkephalin and morphine-like drugs inhibit this enzyme and thus decrease the amount of cAMP in the cells. In order to compensate for the decreased cAMP, the cells synthesize more enzyme in an attempt to produce more cAMP. Since more enzyme has been produced, more morphine is required as an inhibitor to keep the cAMP at a low level. This cycle repeats itself causing an increase in the tolerance level and increasing the amounts of morphine required. If morphine is suddenly withheld, withdrawal symptoms are probably caused by a high concentration of cAMP since the synthesizing enzyme, adenylate cyclase, is no longer being inhibited.



Morphine and codeine are contained in opium from the poppy (*Papaver Somniterum*) plant found in Turkey, Mexico, Southeast Asia, China, and India. This plant is 3-4 feet tall with 5-8 egg shaped capsules on top. Ten days after the poppy blooms in June, incisions are made in the capsules permitting a milky fluid to ooze out. The following day the gummy mass (now brown) is carefully scraped off and pressed into cakes of raw opium to dry.

Opium contains over 20 compounds but only morphine (10%) and codeine (0.5%) are of any importance. Morphine is extracted from the opium and isolated in a relatively pure form. Since codeine is in such low concentration, it is synthesized from morphine by an ether-type methylation of an alcohol group. Codeine has only a fraction of the potency compared to morphine. It is used with aspirin and as a cough suppressant.

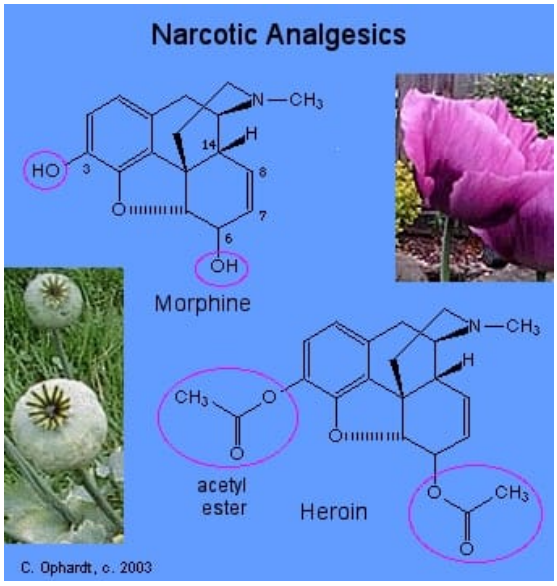
Heroin

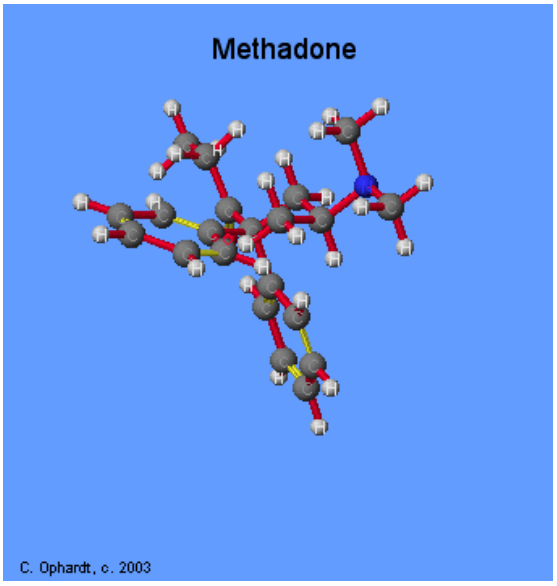
Heroin is synthesized from morphine by a relatively simple esterification reaction of two alcohol (phenol) groups with acetic anhydride (equivalent to acetic acid). Heroin is much more potent than morphine but without the respiratory depression effect. A possible reason may be that heroin passes the blood-brain barrier much more rapidly than morphine. Once in the brain, the

heroin is hydrolyzed to morphine which is responsible for its activity.

Synthetic narcotic analgesics include meperidine and methadone. Meperidine is the most common subsitute for morphine. It exerts several pharmacological effects: analgesic, local anesthetic, and mild antihistamine. This multiple activity may be explained by its structural resemblance to morphine, atropine, and histamine.

Methadone is more active and more toxic than morphine. It can be used for the relief of may types of pain. In addition it is used as a narcotic substitute in addiction treatment because it prevents morphine abstinence syndrome. Methadone was synthesized by German chemists during Wold War II when the United States and our allies cut off their opium supply. And it is difficult to fight a war without analgesics so the Germans went to work and synthesized a number of medications in use today, including demerol and darvon which is structurally simular to methadone. And before we go further lets clear up another myth. Methadone, or dolophine was not named after Adolf Hitler. The "dol" in dolophine comes from the latin root "dolor." The female name Dolores is derived from it and the term dol is used in pain research to measure pain e.g., one dol is 1 unit of pain.

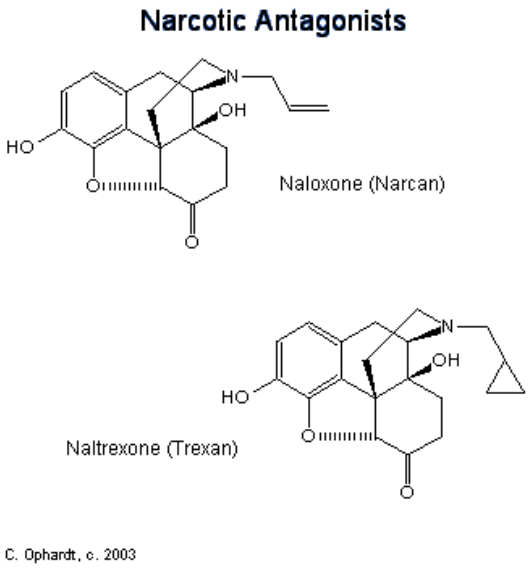




Even methadone, which looks strikingly different from other opioid agonists, has steric forces which produce a configuration that closely resembles that of other opiates. See the graphic on the left and the top graphic on this page. In other words, steric forces bend the molecule of methadone into the correct configuration to fit into the opiate receptor. When you take methadone it first must be metabolized in the liver to a product that your body can use. Excess methadone is also stored in the liver and blood stream and this is how methadone works its 'time release trick' and last for 24 hours or more. Once in the blood stream metabolized methadone is slowly passed to the brain when it is needed to fill opiate receptors. Methadone is the only effective treatment for heroin addiction. It works to smooth the ups and down of heroin craving and allows the person to function normally.

Narcotic Antagonists

Narcotic Antagonists prevent or abolish excessive respiratory depression caused by the administration of morphine or related compounds. They act by competing for the same analgesic receptor sites. They are structurally related to morphine with the exception of the group attached to nitrogen.



Nalorphine precipitates withdrawal symptoms and produces behavioral disturbances in addition to the antogonism action. Naloxane is a pure antagonist with no morphine like effects. It blocks the euphoric effect of heroin when given before heroin. Naltrexone became clinically available in 1985 as a new narcotic antagonist. Its actions resemble those of naloxone, but naltrexone is well is well absorbed orally and is long acting, necessitating only a dose of 50 to 100 mg. Therefore, it is useful in narcotic treatment programs where it is desired to maintain an individual on chronic therapy with a narcotic antagonist. In individuals taking naltrexone, subsequent injection of an opiate will produce little or no effect. Naltrexone appears to be particularly effective for the treatment of narcotic dependence in addicts who have more to gain by being drug-free rather than drug dependant

Contributors and Attributions

- Charles Ophardt, Professor Emeritus, Elmhurst College; Virtual Chembook
- Poppies image from: leda.lycaeum.org

This page titled 10.13: Narcotic Analgesics is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

- Narcotic Analgesic Drugs by Charles Ophardt is licensed CC BY-NC-SA 4.0.

10.14: Antidepressant medications

Antidepressant drugs are used to restore mentally depressed patients to an improved mental status. Depression results from a deficiency of norepinephrine at receptors in the brain. Mechanisms that increase their effective concentration at the receptor sites should alleviate depression.

Introduction

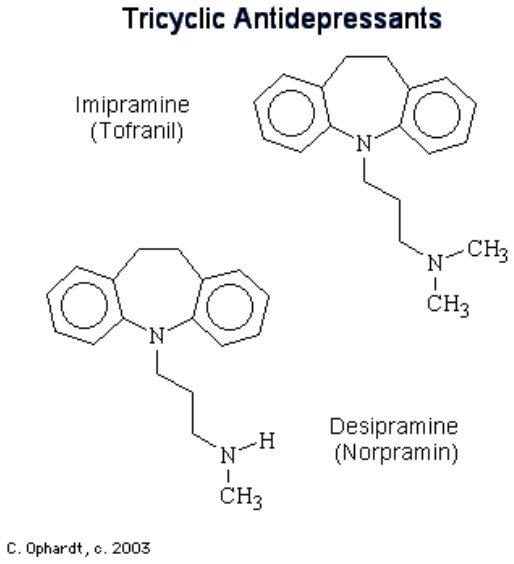
Antidepressant drugs act by one or more of the following stimulation type mechanisms:

1. Increase release of norepinephrine:Amphetamines and electroconvulsive therapy act by this mechanism. Amphetamines mimic norepinephrine.
2. Prevent inactivation of norepinephrine:Monoamine oxidase (MAO) inhibitors are thought to act as antidepressant agents in part by preventing the breakdown and inactivation of norepinephrine.
3. Prevent the re-uptake of norepinephrine:The action of norepinephrine at the receptor site is terminated by the re-uptake of norepinephrine by the neuron from which it was originally released.

Tricyclic Antidepressants

The tricyclic antidepressants are the most effective drugs presently available for the treatment of depression. These act by increasing the release of norepinephrine. Amphetamine and cocaine can also act in this manner. Imipramine, amitriptylin, and other closely related drugs are among the drugs currently most widely used for the treatment of major depression.

- imipramine (Tofranil)
- desipramine (Norpramin)



The activity of the tricyclic drugs depends on the central ring of seven or eight atoms which confers an angled or twisted conformation. The side chain must have at least 2 carbons although 3 appear to be better. The amine group may be either tertiary or secondary. All tricyclic antidepressants block the re-uptake of norepinephrine at nerve terminals. However, the potency and selectivity for the inhibition of the uptake of norepinephrine, serotonin, and dopamine vary greatly among the agents. The tertiary amine tricyclics seem to inhibit the serotonin uptake pump, whereas the secondary amine ones seem better in switching off the NE pump. For instance, imipramine is a potent and selective blocker of serotonin transport, while desipramine inhibits the uptake of norepinephrine.

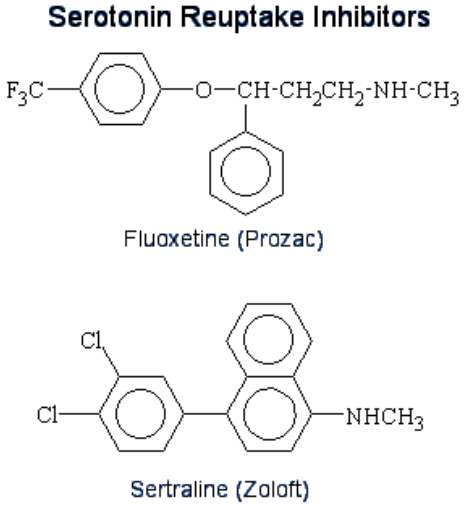
Serotonin

Serotonin (5-hydroxytryptamine or 5-HT) is a monoamine neurotransmitter found in cardiovascular tissue, in endothelial cells, in blood cells, and in the central nervous system. The role of serotonin in neurological function is diverse, and there is little doubt that serotonin is an important CNS neurotransmitter. Although some of the serotonin is metabolized by monoamine oxidase, most of the serotonin released into the post-synaptic space is removed by the neuron through a re-uptake mechanism inhibited by the tricyclic antidepressants and the newer, more selective antidepressant re-uptake inhibitors such as fluoxetine and sertraline.



Selective Serotonin Reuptake Inhibitors

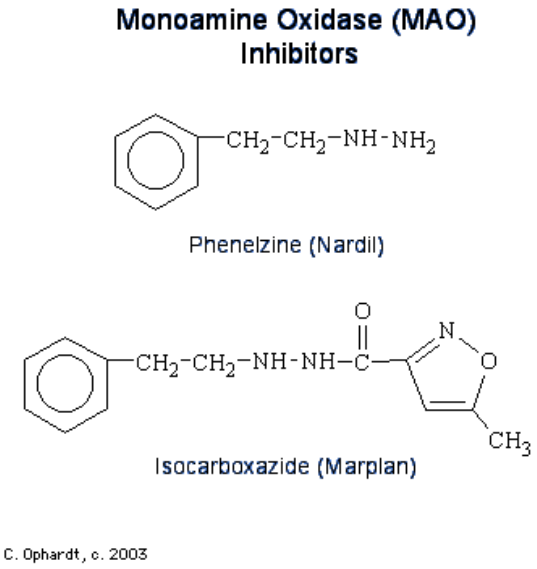
In recent years, selective serotonin reuptake inhibitors have been introduced for the treatment of depression. Prozac is the most famous drug in this class. Clomiprimine, fluoxetine (Prozac), sertraline and paroxetine selectively block the re-uptake of serotonin, thereby increasing the levels of serotonin in the central nervous system. Note the similarities and differences between the tricyclic antidepressants and the selective serotonin re-uptake inhibitors. Clomipramine has been useful in the treatment of obsessive-compulsive disorders.



Monoamine Oxidase Inhibitors

Monoamine oxidase (MAO) causes the oxidative deamination of norephinephrine, serotonin, and other amines. This oxidation is the method of reducing the concentration of the neurotransmitter after it has sent the signal at the receptor site. A drug which inhibits this enzyme has the effect of **increasing the concentration of the norepinephrine which in turn causes a stimulation effect.** Most MAO inhibitors are hydrazine derivatives. Hydrazine is highly reactive and may form a strong covalent bond with MAO with consequent inhibition for up to 5 days.

These drugs are less effective and produce more side effects than the tricyclic antidepressants. For example, they lower blood pressure and were at one time used to treat hypertension. Their use in psychiatry has also become very limited as the tricyclic antidepressants have come to dominate the treatment of depression and allied conditions. Thus, MAOIs are used most often when tricyclic antidepressants give unsatisfactory results.



Phenelzine is the hydrazine analog of phenylethylamine, a substrate of MAO. This and several other MAOIs, such as isocarboxazide, are structurally related to amphetamine and were synthesized in an attempt to enhance central stimulant properties.

- phenelzine (Nardil)
- isocarboxazid (Marplan)

Contributors

Charles Ophardt (Professor Emeritus, Elmhurst College); Virtual Chembook

This page titled 10.14: Antidepressant medications is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

11: INTRODUCTION TO CHEMICAL REACTIONS

Chemical change is a central concept in chemistry. The goal of chemists is to know how and why a substance changes in the presence of another substance or even by itself. Because there are tens of millions of known substances, there are a huge number of possible chemical reactions. In this chapter, we will find that many of these reactions can be classified into a small number of categories according to certain shared characteristics.

11.1: Prelude to Introduction to Chemical Reactions

11.2: The Law of Conservation of Matter

11.3: Chemical Equations

11.4: Quantitative Relationships Based on Chemical Equations

11.5: Some Types of Chemical Reactions

11.E: Introduction to Chemical Reactions (Exercises)

11.S: Introduction to Chemical Reactions (Summary)

This page titled 11: Introduction to Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.1: Prelude to Introduction to Chemical Reactions

Although yeast has been used for thousands of years, its true nature has been known only for the last two centuries. Yeasts are single-celled fungi. About 1,000 species are recognized, but the most common species is *Saccharomyces cerevisiae*, which is used in bread making. Other species are used for the fermentation of alcoholic beverages. Some species can cause infections in humans.

Yeasts live primarily on sugars, such as glucose ($\mathrm{C_6H_{12}O_6}$). They convert glucose into carbon dioxide ($\mathrm{CO_2}$) and ethanol ($\mathrm{C_2H_5OH}$) in a chemical transformation that is represented as follows:

$$\mathrm{C_6H_{12}O_6} \longrightarrow 2\,\mathrm{CO_2\,(g)} + 2\,\mathrm{C_2H_5OH}(\ell) \tag{11.1.1}$$

Bread making depends on the production of carbon dioxide. The gas, which is produced in tiny pockets in bread dough, acts as a leavening agent: it expands during baking and makes the bread rise. Leavened bread is softer, lighter, and easier to eat and chew than unleavened bread. The other major use of yeast, fermentation, depends on the production of ethanol, which results from the same chemical transformation. Some alcoholic beverages, such as champagne, can also be carbonated using the carbon dioxide produced by the yeast.

Yeast is among the simplest life forms on Earth, yet it is absolutely necessary for at least two major food industries. Without yeast to turn dough into bread and juice into wine, these foods and food industries would not exist today.

This page titled 11.1: Prelude to Introduction to Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.2: The Law of Conservation of Matter

Learning Objectives

- Correctly define a law as it pertains to science.
- State the law of conservation of matter.

In science, a law is a general statement that explains a large number of observations. Before being accepted, a law must be verified many times under many conditions. Laws are therefore considered the highest form of scientific knowledge and are generally thought to be inviolable. Scientific laws form the core of scientific knowledge. One scientific law that provides the foundation for understanding in chemistry is the law of conservation of matter. It states that in any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant. A concise way of expressing this law is to say that the amount of matter in a system is *conserved*.

With the development of more precise ideas on elements, compounds and mixtures, scientists began to investigate how and why substances react. French chemist A. Lavoisier laid the foundation to the scientific investigation of matter by describing that substances react by following certain laws. These laws are called the laws of chemical combination. These eventually formed the basis of Dalton's Atomic Theory of Matter.

Law of Conservation of Mass

According to this law, during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants.

$$\overbrace{\underbrace{\mathrm{HgO\,(s)}}_{100\ \mathrm{g}}}^{\text{Mercuric oxide}} \rightarrow \overbrace{\underbrace{\mathrm{Hg\,(l)}}_{92.6\ \mathrm{g}}}^{\text{Mercury}} + \overbrace{\underbrace{\mathrm{O_2\,(g)}}_{7.4\ \mathrm{g}}}^{\text{Oxygen}} \tag{11.2.1}$$

The law of conservation of mass is also known as the "law of indestructibility of matter."

Example $11.2.1$

If heating 10 grams of $\mathrm{CaCO_3}$ produces 4.4 g of $\mathrm{CO_2}$ and 5.6 g of CaO, show that these observations are in agreement with the law of conservation of mass.



A sample of calcium carbonate (CaCO3). Image used with permission (Public Domain; Walkerma).

SOLUTION

- Mass of the reactants: $10\ g$
- Mass of the products: $4.4\ g + 5.6\ g = 10\ g$.

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

What does this mean for chemistry? In any chemical change, one or more initial substances change into a different substance or substances. Both the initial and final substances are composed of atoms because all matter is composed of atoms. According to the law of conservation of matter, matter is neither created nor destroyed, so we must have the same number and type of atoms after the chemical change as were present before the chemical change.

Before looking at explicit examples of the law of conservation of matter, we need to examine the method chemists use to represent chemical changes.

Exercise $11.2.1$

a. What is the law of conservation of matter?
b. How does the law of conservation of matter apply to chemistry?

Answer a

The law of conservation of matter states that in any given system that is closed to the transfer of matter, the amount of matter in the system stays constant

Answer b

The law of conservation of matter says that in chemical reactions, the total mass of the products must equal the total mass of the reactants.

Summary

The amount of matter in a closed system is conserved.

Contributors

- Binod Shrestha (University of Lorraine)

Exercises

1. Express the law of conservation of matter in your own words.
2. Explain why the concept of conservation of matter is considered a scientific law.

Answer

1. Matter may not be created or destroyed.

This page titled 11.2: The Law of Conservation of Matter is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.3: Chemical Equations

Learning Objectives

- Define *chemical reaction*.
- Use a balanced chemical equation to represent a chemical reaction.

Water (H_2O) is composed of hydrogen and oxygen. Suppose we imagine a process in which we take some elemental hydrogen (H_2) and elemental oxygen (O_2) and let them react to make water. The statement

"*hydrogen and oxygen react to make water*"

is one way to represent that process, which is called a chemical reaction. Figure $11.3.1$ shows a rather dramatic example of this very reaction.

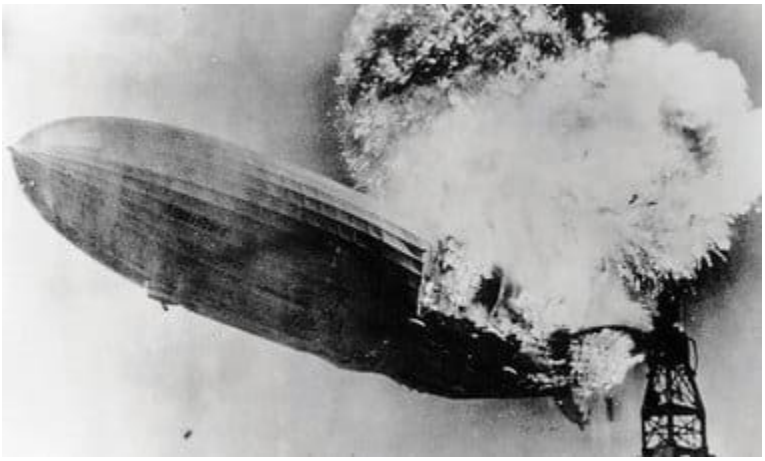


Figure $11.3.1$: The Formation of Water. Hydrogen and oxygen combine to form water. Here, the hydrogen gas in the zeppelin SS *Hindenburg* reacts with oxygen in the air to make water. Source: Photo courtesy of the US Navy. For a video of this see https://www.youtube.com/watch?v=CgWHbpMVQ1U.

To simplify the writing of reactions, we use formulas instead of names when we describe a reaction. We can also use symbols to represent other words in the reaction. A plus sign connects the initial substances (and final substances, if there is more than one), and an arrow ($\rightarrow$) represents the chemical change:

$$\mathrm{H_2 + O_2 \ \rightarrow H_2O} \tag{11.3.1}$$

This statement is one example of a chemical equation, an abbreviated way of using symbols to represent a chemical change. The substances on the left side of the arrow are called reactants, and the substances on the right side of the arrow are called products. It is not uncommon to include a phase label with each formula—(s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for a substance dissolved in water, also known as an *aqueous solution*. If we included phase labels for the reactants and products, under normal environmental conditions, the reaction would be as follows:

$$\mathrm{H_2\,(g) + O_2\,(g) \ \rightarrow H_2O\ (\ell)} \tag{11.3.2}$$

Chemical equations can also be used to describe physical changes. We will see examples of this soon.

This equation is still not complete because *it does not satisfy the law of conservation of matter*. Count the number of atoms of each element on each side of the arrow. On the reactant side, there are two H atoms and two O atoms; on the product side, there are two H atoms and only one oxygen atom. The equation is not balanced because the number of oxygen atoms on each side is not the same (Figure $11.3.2$).

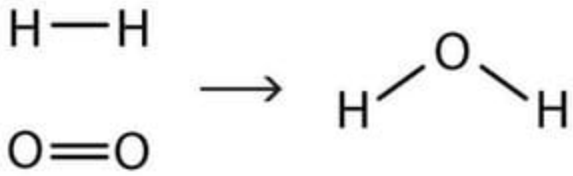


Figure 11.3.2: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is not balanced as written.

To make this chemical equation conform to the law of conservation of matter, we must revise the amounts of the reactants and the products as necessary to get the same number of atoms of a given element on each side. Because every substance has a characteristic chemical formula, we cannot change the chemical formulas of the individual substances. For example, we cannot change the formula for elemental oxygen to O. However, we can assume that different numbers of reactant molecules or product molecules may be involved. For instance, perhaps two water molecules are produced, not just one:

$$\mathrm{H_2\,(g) + O_2\,(g)\ \rightarrow 2\,H_2O\ (\ell)} \tag{11.3.3}$$

The 2 preceding the formula for water is called a coefficient. It implies that two water molecules are formed. There are now two oxygen atoms on each side of the equation.

This point is so important that we should repeat it. You **cannot** change the formula of a chemical substance to balance a chemical reaction! You **must** use the proper chemical formula of the substance.

Unfortunately, by inserting the coefficient 2 in front of the formula for water, we have also changed the number of hydrogen atoms on the product side as well. As a result, we no longer have the same number of hydrogen atoms on each side. This can be easily fixed, however, by putting a coefficient of 2 in front of the diatomic hydrogen reactant:

$$\mathrm{2\,H_2\,(g) + O_2\,(g)\ \rightarrow 2\,H_2O\ (\ell)} \tag{11.3.4}$$

Now we have four hydrogen atoms and two oxygen atoms on each side of the equation. The law of conservation of matter is satisfied because we now have the same number of atoms of each element in the reactants and in the products. We say that the reaction is balanced (Figure 11.3.3). The diatomic oxygen has a coefficient of 1, which typically is not written but assumed in balanced chemical equations.

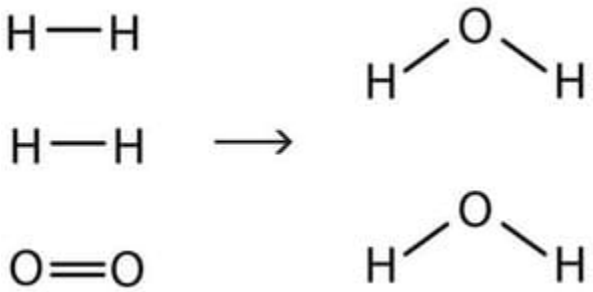


Figure 11.3.3: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is now balanced.

Proper chemical equations should be balanced. Writing balanced reactions is a chemist's way of acknowledging the law of conservation of matter.

Example 11.3.1

Is each chemical equation balanced?

a. $2Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
b. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$
c. $AgNO_3(aq) + 2KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

SOLUTION

a. By counting, we find two sodium atoms and two oxygen atoms in the reactants and four sodium atoms and two oxygen atoms in the products. This equation is not balanced.
b. The reactants have one carbon atom, four hydrogen atoms, and four oxygen atoms. The products have one carbon atom, four hydrogen atoms, and four oxygen atoms. This equation is balanced.
c. The reactants have one silver atom, one nitrogen atom, three oxygen atoms, two potassium atoms, and two chlorine atoms. The products have one silver atom, one chlorine atom, one potassium atom, one nitrogen atom, and three oxygen atoms. Because there are different numbers of chlorine and potassium atoms, this equation is not balanced.

Exercise 11.3.1

Is each chemical equation balanced?

a. $2Hg_{(\ell)} + O_{2(g)} \rightarrow Hg_2O_{2(s)}$
b. $C_2H_{4(g)} + 2O_{2(g)} \rightarrow 2CO_{2(g)} + 2H2O_{(\ell)}$
c. $Mg(NO_3)_{2(s)} + 2Li_{(s)} \rightarrow Mg_{(s)} + 2LiNO_{3(s)}$

How does one balance a chemical equation, starting with the correct formulas of the reactants and products? Basically, a back-and-forth approach is adopted, counting the number of atoms of one element on one side, checking the number of atoms of that element on the other side, and changing a coefficient if necessary. Then check another element, going back and forth from one side of the equation to another, until each element has the same number of atoms on both sides of the arrow. In many cases, it does not matter which element is balanced first and which is balanced last, as long as all elements have the same number of atoms on each side of the equation.

For example, to balance the equation

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + HCl} \tag{11.3.5}$$

we might choose to count the carbon atoms first, finding that both sides are balanced with one carbon atom. The reactant side has four hydrogen atoms, so the product side must also have four hydrogen atoms. We fix this by putting a 4 in front of the HCl:

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{11.3.6}$$

Now each side has four hydrogen atoms. The product side has a total of eight chlorine atoms (four from the CCl_4 and four from the four molecules of HCl), so we need eight chlorine atoms as reactants. Because elemental chlorine is a diatomic molecule, we need four chlorine molecules to get a total of eight chlorine atoms. We add another 4 in front of the Cl_2 reactant:

$$\mathrm{CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{11.3.7}$$

Now we check: each side has one carbon atom, four hydrogen atoms, and eight chlorine atoms. The chemical equation is balanced.

Exercise 11.3.2

a. What are the parts of a chemical equation?
b. Explain why chemical equations need to be balanced.

Answer a

reactants and products

Answer b

Chemical equations need to be balanced to satisfy the law of conservation of matter.

Summary

Chemical reactions are represented by chemical equations that list reactants and products. Proper chemical equations are balanced; the same number of each element's atoms appears on each side of the equation.

Exercises

1. Write a chemical equation to express the fact that hydrogen gas and solid iodine react to make gaseous hydrogen iodide.

Make sure the equation satisfies the law of conservation of matter.

2. Write a chemical equation to express the fact that sodium metal and chlorine gas react to make solid sodium chloride. Make sure the equation satisfies the law of conservation of matter.

3. Write an equation expressing the fact that hydrogen gas and fluorine gas react to make gaseous hydrogen fluoride. Make sure the equation satisfies the law of conservation of matter.

4. Write an equation expressing the fact that solid potassium and fluorine gas react to make solid potassium fluoride. Make sure the equation satisfies the law of conservation of matter.

5. Mercury reacts with oxygen to make mercury(II) oxide. Write a balanced chemical equation that summarizes this reaction.

6. Octane (C_8H_{18}) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.

7. Propyl alcohol (C_3H_7OH) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.

8. Sulfuric acid reacts with iron metal to make iron(III) sulfate and hydrogen gas. Write a balanced chemical equation that summarizes this reaction.

9. Balance each equation.

    a. $MgCl_2 + K \rightarrow KCl + Mg$
    b. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
    c. $NaN_3 \rightarrow Na + N_2$ (This is the reaction used to inflate airbags in cars.)

10. Balance each equation.

    a. $NH_4NO_3 \rightarrow N_2O + H_2O$
    b. $TiBr_4 + H_2O \rightarrow TiO_2 + HBr$
    c. $C_3H_5N_3O_9 \rightarrow CO_2 + N_2 + O_2 + H_2O$ (This reaction represents the decomposition of nitroglycerine.)

11. Balance each equation.

    a. $NH_3 + O_2 \rightarrow NO + H_2O$
    b. $Li + N_2 \rightarrow Li_3N$
    c. $AuCl \rightarrow Au + AuCl_3$

12. Balance each equation.

    a. $NaOH + H_3PO_4 \rightarrow Na_3PO_4 + H_2O$
    b. $N_2H_4 + Cl_2 \rightarrow N_2 + HCl$
    c. $Na_2S + H_2S \rightarrow NaSH$

13. Chromium(III) oxide reacts with carbon tetrachloride to make chromium(III) chloride and phosgene ($COCl_2$). Write the balanced chemical equation for this reaction.

14. The reaction that occurs when an Alka-Seltzer tablet is dropped into a glass of water has sodium bicarbonate reacting with citric acid ($H_3C_6H_5O_7$) to make carbon dioxide, water, and sodium citrate ($Na_3C_6H_5O_7$). Write the balanced chemical equation for this reaction.

15. When sodium hydrogen carbonate is used to extinguish a kitchen fire, it decomposes into sodium carbonate, water, and carbon dioxide. Write a balanced chemical equation for this reaction.

16. Elemental bromine gas can be generated by reacting sodium bromide with elemental chlorine. The other product is sodium chloride. Write a balanced chemical equation for this reaction.

Answers

1. $H_2(g) + I_2(s) \rightarrow 2HI(g)$

3. $H_2(g) + F_2(g) \rightarrow 2HF(g)$

5. $2Hg + O_2 \rightarrow 2HgO$

7. $2C_3H_7OH + 9O_2 \rightarrow 6CO_2 + 8H_2O$

9. a. $MgCl_2 + 2K \rightarrow 2KCl + Mg$
   b. $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$
   c. $2NaN_3 \rightarrow 2Na + 3N_2$

11. a. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
    b. $6Li + N_2 \rightarrow 2Li_3N$
    c. $3AuCl \rightarrow 2Au + AuCl_3$

13. $Cr_2O_3 + 3CCl_4 \rightarrow 2CrCl_3 + 3COCl_2$

15. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$)

This page titled 11.3: Chemical Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.4: Quantitative Relationships Based on Chemical Equations

Learning Objectives

- To calculate the amount of one substance that will react with or be produced from a given amount of another substance.

A balanced chemical equation not only describes some of the chemical properties of substances—by showing us what substances react with what other substances to make what products—but also shows numerical relationships between the reactants and the products. The study of these numerical relationships is called stoichiometry. The stoichiometry of chemical equations revolves around the coefficients in the balanced chemical equation because these coefficients determine the molecular ratio in which reactants react and products are made.

> The word *stoichiometry* is pronounced “stow-eh-key-OM-et-tree.” It is of mixed Greek and English origins, meaning roughly “measure of an element.”

Looking Closer: Stoichiometry in Cooking

Let us consider a stoichiometry analogy from the kitchen. A recipe that makes 1 dozen biscuits needs 2 cups of flour, 1 egg, 4 tablespoons of shortening, 1 teaspoon of salt, 1 teaspoon of baking soda, and 1 cup of milk. If we were to write this as a chemical equation, we would write

$$\text{2 c flour + 1 egg + 4 tbsp shortening + 1 tsp salt + 1 tsp baking soda + 1 c milk} \rightarrow \text{12 biscuits}$$

(Unlike true chemical reactions, this one has all 1 coefficients written explicitly—partly because of the many different units here.) This equation gives us ratios of how much of what reactants are needed to make how much of what product. Two cups of flour, when combined with the proper amounts of the other ingredients, will yield 12 biscuits. One teaspoon of baking soda (when also combined with the right amounts of the other ingredients) will make 12 biscuits. One egg must be combined with 1 cup of milk to yield the product food. Other relationships can also be expressed.

We can use the ratios we derive from the equation for predictive purposes. For instance, if we have 4 cups of flour, how many biscuits can we make if we have enough of the other ingredients? It should be apparent that we can make a double recipe of 24 biscuits.

But how would we find this answer formally, that is, mathematically? We would set up a conversion factor, much like we did in Chapter 1 "Chemistry, Matter, and Measurement". Because 2 cups of flour make 12 biscuits, we can set up an equivalency ratio:

$$\frac{\text{12 biscuits}}{\text{2 c flour}} \tag{11.4.1}$$

We then can use this ratio in a formal conversion of flour to biscuits:

$$\text{4 c flour} \times \frac{\text{12 biscuits}}{\text{2 c flour}} = \text{24 biscuits} \tag{11.4.2}$$

Similarly, by constructing similar ratios, we can determine how many biscuits we can make from any amount of ingredient. When you are doubling or halving a recipe, you are doing a type of stoichiometry. Applying these ideas to chemical reactions should not be difficult if you use recipes when you cook.



A recipe shows how much of each ingredient is needed for the proper reaction to take place. Image used with permission from Wikipedia.

Consider the following balanced chemical equation:

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O \tag{11.4.3}$$

The coefficients on the chemical formulas give the ratios in which the reactants combine and the products form. Thus, we can make the following statements and construct the following ratios:

Statement from the Balanced Chemical Reaction	Ratio	Inverse Ratio
two C_2H_2 molecules react with five O_2 molecules	$\frac{2C_2H_2}{5O_2}$	$\frac{5O_2}{2C_2H_2}$
two C_2H_2 molecules react to make four CO_2 molecules	$\frac{2C_2H_2}{4CO_2}$	$\frac{4CO_2}{2C_2H_2}$
five O_2 molecules react to make two H_2O molecules	$\frac{5O_2}{2H_2O}$	$\frac{2H_2O}{5O_2}$
four CO_2 molecules are made at the same time as two H_2O molecules	$\frac{2H_2O}{4CO_2}$	$\frac{4CO_2}{2H_2O}$

Other relationships are possible; in fact, 12 different conversion factors can be constructed from this balanced chemical equation. In each ratio, the unit is assumed to be molecules because that is how we are interpreting the chemical equation.

Any of these fractions can be used as a conversion factor to relate an amount of one substance to an amount of another substance. For example, suppose we want to know how many CO_2 molecules are formed when 26 molecules of C_2H_2 are reacted. As usual with a conversion problem, we start with the amount we are given—$26C_2H_2$—and multiply it by a conversion factor that cancels out our original unit and introduces the unit we are converting to—in this case, CO_2. That conversion factor is $\frac{4CO_2}{2C_2H_2}$, which is composed of terms that come directly from the balanced chemical equation. Thus, we have

$$26C_2H_2 \times \frac{4CO_2}{2C_2H_2} \tag{11.4.4}$$

The molecules of C_2H_2 cancel, and we are left with molecules of CO_2. Multiplying through, we get

$$26\mathrm{C_2H_2} \times \frac{4\mathrm{CO_2}}{2\mathrm{C_2H_2}} = 52\mathrm{CO_2} \tag{11.4.5}$$

Thus, 52 molecules of CO_2 are formed.

This application of stoichiometry is extremely powerful in its predictive ability, as long as we begin with a balanced chemical equation. Without a balanced chemical equation, the predictions made by simple stoichiometric calculations will be incorrect.

Example 11.4.1

Start with this balanced chemical equation.

$$KMnO_4 + 8HCl + 5FeCl_2 \rightarrow 5\ FeCl_3 + MnCl_2 + 4H_2O + KCl$$

1. Verify that the equation is indeed balanced.
2. Give 2 ratios that give the relationship between HCl and $FeCl_3$.

SOLUTION

1. Each side has 1 K atom and 1 Mn atom. The 8 molecules of HCl yield 8 H atoms, and the 4 molecules of H_2O also yield 8 H atoms, so the H atoms are balanced. The Fe atoms are balanced, as we count 5 Fe atoms from 5 $FeCl_2$ reactants and 5 $FeCl_3$ products. As for Cl, on the reactant side, there are 8 Cl atoms from HCl and 10 Cl atoms from the 5 $FeCl_2$ formula units, for a total of 18 Cl atoms. On the product side, there are 15 Cl atoms from the 5 $FeCl_3$ formula units, 2 from the $MnCl_2$ formula unit, and 1 from the KCl formula unit. This is a total of 18 Cl atoms in the products, so the Cl atoms are balanced. All the elements are balanced, so the entire chemical equation is balanced.
2. Because the balanced chemical equation tells us that 8 HCl molecules react to make 5 $FeCl_3$ formula units, we have the following 2 ratios: $\frac{8\text{HCl}}{5\text{FeCl}_3}$ and $\frac{5\text{FeCl}_3}{8\text{HCl}}$. There are a total of 42 possible ratios. Can you find the other 40 relationships?

Exercise 11.4.1

Start with this balanced chemical equation.

$$2\text{KMnO}_4 + 3\text{CH}_2{=}\text{CH}_2 + 4\text{H}_2\text{O} \rightarrow 2\text{MnO}_2 + 3\text{HOCH}_2\text{CH}_2\text{OH} + 2\text{KOH}$$

a. Verify that the equation is balanced.
b. Give 2 ratios that give the relationship between $KMnO_4$ and $CH_2{=}CH_2$. (A total of 30 relationships can be constructed from this chemical equation. Can you find the other 28?)

Summary

A balanced chemical equation gives the ratios in which molecules of substances react and are produced in a chemical reaction.

Concept Review Exercises

1. Explain how stoichiometric ratios are constructed from a chemical equation.
2. Why is it necessary for a chemical equation to be balanced before it can be used to construct conversion factors?

Answers

1. Stoichiometric ratios are made using the coefficients of the substances in the balanced chemical equation.
2. A balanced chemical equation is necessary so one can construct the proper stoichiometric ratios.

Exercises

1. Balance this equation and write every stoichiometric ratio you can from it.

$$NH_4NO_3 \rightarrow N_2O + H_2O$$

2. Balance this equation and write every stoichiometric ratio you can from it.

$$N_2 + H_2 \rightarrow NH_3$$

3. Balance this equation and write every stoichiometric ratio you can from it.

$$Fe_2O_3 + C \rightarrow Fe + CO_2$$

4. Balance this equation and write every stoichiometric ratio you can from it.

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

5. Balance this equation and determine how many molecules of CO_2 are formed if 15 molecules of C_6H_6 are reacted.

$$C_6H_6 + O_2 \rightarrow CO_2 + H_2O$$

6. Balance this equation and determine how many molecules of $Ag_2CO_3(s)$ are produced if 20 molecules of Na_2CO_3 are reacted.

$$Na_2CO_3(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag_2CO_3(s)$$

7. Copper metal reacts with nitric acid according to this equation:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Cu atoms will react if 488 molecules of aqueous HNO_3 are reacted?

8. Gold metal reacts with a combination of nitric acid and hydrochloric acid according to this equation:

$$Au(s) + 3HNO_3(aq) + 4HCl(aq) \rightarrow HAuCl_4(aq) + 3NO_2(g) + 3H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Au atoms react with 639 molecules of aqueous HNO_3?

9. Sulfur can be formed by reacting sulfur dioxide with hydrogen sulfide at high temperatures according to this equation:

$$SO_2(g) + 2H_2S(g) \rightarrow 3S(g) + 2H_2O(g)$$

   a. Verify that this equation is balanced.
   b. How many S atoms will be formed from by reacting 1,078 molecules of H_2S?

10. Nitric acid is made by reacting nitrogen dioxide with water:

$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

   a. Verify that this equation is balanced.
   b. How many molecules of NO will be formed by reacting 2,268 molecules of NO_2?

Answers

1. $NH_4NO_3 \rightarrow N_2O + 2H_2O$; the stoichiometric ratios are $\frac{1NH_4NO_3}{1N_2O}$, $\frac{1NH_4NO_3}{2H_2O}$, $\frac{1N_2O}{2H_2O}$, and their reciprocals.

3. $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$; the stoichiometric ratios are $\frac{2Fe_2O_3}{3C}$, $\frac{2Fe_2O_3}{4Fe}$, $\frac{2Fe_2O_3}{3CO_2}$, $\frac{3C}{4Fe}$, $\frac{3C}{3CO_2}$, $\frac{4Fe}{3CO_2}$, and their reciprocals.

5. $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$; 90 molecules

7. a. It is balanced.
   b. 183 atoms

9. a. It is balanced.
   b. 1,617 atoms

This page titled 11.4: Quantitative Relationships Based on Chemical Equations is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.5: Some Types of Chemical Reactions

Learning Objectives

- To classify a given chemical reaction into a variety of types.

Although there are untold millions of possible chemical reactions, most can be classified into a small number of general reaction types. Classifying reactions has two purposes: it helps us to recognize similarities among them, and it enables us to predict the products of certain reactions. A particular reaction may fall into more than one of the categories that we will define in this book.

Combination (composition) Reactions

A combination (composition) reaction is a chemical reaction that makes a single substance from two or more reactants. There may be more than one molecule of product in the balanced chemical equation, but there is only one substance produced. For example, the equation

$$4\,\mathrm{Fe} + 3\,\mathrm{O_2} \rightarrow 2\,\mathrm{Fe_2O_3} \tag{11.5.1}$$

is a combination reaction that produces Fe_2O_3 from its constituent elements—Fe and O_2. Combination reactions do not have to combine elements, however. The chemical equation

$$\mathrm{Fe_2O_3} + 3\,\mathrm{SO_3} \rightarrow \mathrm{Fe_2(SO_4)_3} \tag{11.5.2}$$

shows a combination reaction in which Fe_2O_3 combines with three molecules of SO_3 to make $Fe_2(SO_4)_3$.

Example 11.5.1

Which equations are combination reactions?

a. $Co(s) + Cl_2(g) \rightarrow CoCl_2(s)$
b. $CO(g) + Cl_2(g) \rightarrow COCl_2(g)$
c. $N_2H_4(\ell) + O_2(g) \rightarrow N_2(g) + 2H_2O(\ell)$

SOLUTION

a. This is a combination reaction.
b. This is a combination reaction. (The compound $COCl_2$ is called phosgene and, in the past, was used as a gassing agent in chemical warfare.)
c. This is not a combination reaction.

Exercise 11.5.1

Which equations are combination reactions?

a. $P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(g)$
b. $SO_3(\ell) + H_2O(\ell) \rightarrow H_2SO_4(\ell)$
c. $NaOH(s) + HCl(g) \rightarrow NaCl(s) + H_2O(\ell)$

Decomposition Reactions

A decomposition reaction is the reverse of a combination reaction. In a decomposition reaction, a single substance is converted into two or more products. There may be more than one molecule of the reactant, but there is only one substance initially. For example, the equation

$$2\,\mathrm{NaHCO_3\,(s)} \rightarrow \mathrm{Na_2CO_3\,(s) + CO_2\,(g) + H_2O\ (\ell)} \tag{11.5.3}$$

is a decomposition reaction that occurs when $NaHCO_3$ is exposed to heat. Another example is the decomposition of $KClO_3$:

$$2\,\mathrm{KClO_3\,(s)} \rightarrow 2\,\mathrm{KCl\,(s) + 3\,O_2\,(g)} \tag{11.5.4}$$

This reaction was once commonly used to generate small amounts of oxygen in the chemistry lab.

The decomposition reaction of $NaHCO_3$ is the reaction that occurs when baking soda is poured on a small kitchen fire. The intent is that the H_2O and CO_2 produced by the decomposition will smother the flames.

Combustion Reactions

A combustion reaction occurs when a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction. One example is the burning of acetylene (C_2H_2) in torches:

$$2\,\mathrm{C_2H_2 + 5\,O_2} \rightarrow 4\,\mathrm{CO_2 + 2\,H_2O} \tag{11.5.5}$$

Oxygen (in its elemental form) is a crucial reactant in combustion reactions, and it is also present in the products.

Energy in the form of heat is usually given off as a product in a combustion reaction as well.

Example 11.5.2

Identify each type of reaction.

a. $2K(s) + S(s) + 2O_2(g) \rightarrow K_2SO_4(s)$
b. $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + Cr_2O_3(s) + 4H_2O(\ell)$
c. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$

SOLUTION

a. Multiple reactants are combining to make a single product, so this reaction is a combination reaction.
b. A single substance reacts to make several products, so we have a decomposition reaction.
c. Oxygen reacts with a compound to make carbon dioxide (an oxide of carbon) and water (an oxide of hydrogen). This is a combustion reaction.

Exercise 11.5.2

Identify each type of reaction.

a. $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
b. $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$
c. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Summary

There are several recognizable types of chemical reactions: combination, decomposition, and combustion reactions are examples.

Concept Review Exercises

1. What is the difference between a combination reaction and a combustion reaction?
2. Give the distinguishing characteristic(s) of a decomposition reaction.
3. How do we recognize a combustion reaction?

Answers

1. A combination reaction produces a certain substance; a combustion reaction is a vigorous reaction, usually a combination with oxygen, that is accompanied by the production of light and/or heat.
2. In a decomposition reaction, a single substance reacts to make multiple substances as products.
3. A combustion reaction is typically a vigorous reaction accompanied by light and/or heat, usually because of reaction with oxygen.

Exercises

1. Identify each type of reaction.

    a. $C_6H_5CH_3 + 9O_2 \rightarrow 7CO_2 + 4H_2O$
    b. $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
    c. $C + 2H_2 \rightarrow CH_4$

2. Identify each type of reaction.

    a. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$
    b. $FeO + SO_3 \rightarrow FeSO_4$
    c. $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

3. Identify each type of reaction.

    a. $2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$
    b. $Hg(\ell) + ½O_2\,(g) \rightarrow HgO(s)$
    c. $CH_2CH_2(g) + Br_2(\ell) \rightarrow CH_2BrCH_2Br$

4. Identify each type of reaction.

    a. $Ti(s) + O_2(g) \rightarrow TiO_2(s)$
    b. $H_2SO_3(aq) \rightarrow H_2O(\ell) + SO_2(g)$
    c. $3O_2(g) \rightarrow 2O_3(g)$

Answers

1. a. combustion
    b. decomposition
    c. combination

3. a. decomposition
    b. combustion or combination
    c. combination

This page titled 11.5: Some Types of Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

11.E: Introduction to Chemical Reactions (Exercises)

5.1: The Law of Conservation of Matter

1. What is the law of conservation of matter?
2. How does the law of conservation of matter apply to chemistry?

ANSWERS

1. The law of conservation of matter states that in any given system that is closed to the transfer of matter, the amount of matter in the system stays constant
2. The law of conservation of matter says that in chemical reactions, the total mass of the products must equal the total mass of the reactants.

Exercises

1. Express the law of conservation of matter in your own words.

2. Explain why the concept of conservation of matter is considered a scientific law.

3. Potassium hydroxide (KOH) readily reacts with carbon dioxide ($\mathrm{CO_2}$) to produce potassium carbonate ($\mathrm{K_2CO_3}$) and water ($\mathrm{H_2O}$). How many grams of potassium carbonate is produced if 224.4 g of KOH reacted with 88.0 g of $\mathrm{CO_2}$. The reaction also produced 36.0 g of water.

Answers

1. Matter may not be created or destroyed.

2. The concept is a scientific law because it is based on experimentation.

3. 276.4 g

5.2: Chemical Equations

Concept Review Exercises

a. What are the parts of a chemical equation?
b. Explain why chemical equations need to be balanced.

Answers

a. reactants and products

b. Chemical equations need to be balanced to satisfy the law of conservation of matter.

Exercises

1. Write a chemical equation to express the fact that hydrogen gas and solid iodine react to make gaseous hydrogen iodide. Make sure the equation satisfies the law of conservation of matter.

2. Write a chemical equation to express the fact that sodium metal and chlorine gas react to make solid sodium chloride. Make sure the equation satisfies the law of conservation of matter.

3. Write an equation expressing the fact that hydrogen gas and fluorine gas react to make gaseous hydrogen fluoride. Make sure the equation satisfies the law of conservation of matter.

4. Write an equation expressing the fact that solid potassium and fluorine gas react to make solid potassium fluoride. Make sure the equation satisfies the law of conservation of matter.

5. Mercury reacts with oxygen to make mercury(II) oxide. Write a balanced chemical equation that summarizes this reaction.

6. Octane (C_8H_{18}) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.

7. Propyl alcohol (C_3H_7OH) reacts with oxygen to make carbon dioxide and water. Write a balanced chemical equation that summarizes this reaction.

8. Sulfuric acid (H_2SO_4) reacts with iron metal to make iron(III) sulfate and hydrogen gas. Write a balanced chemical equation that summarizes this reaction.

9. Balance each equation.

    a. $MgCl_2 + K \rightarrow KCl + Mg$
    b. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
    c. $NaN_3 \rightarrow Na + N_2$ (This is the reaction used to inflate airbags in cars.)

10. Balance each equation.

    a. $NH_4NO_3 \rightarrow N_2O + H_2O$
    b. $TiBr_4 + H_2O \rightarrow TiO_2 + HBr$
    c. $C_3H_5N_3O_9 \rightarrow CO_2 + N_2 + O_2 + H_2O$ (This reaction represents the decomposition of nitroglycerine.)

11. Balance each equation.

    a. $NH_3 + O_2 \rightarrow NO + H_2O$
    b. $Li + N_2 \rightarrow Li_3N$
    c. $AuCl \rightarrow Au + AuCl_3$

12. Balance each equation.

    a. $NaOH + H_3PO_4 \rightarrow Na_3PO_4 + H_2O$
    b. $N_2H_4 + Cl_2 \rightarrow N_2 + HCl$
    c. $Na_2S + H_2S \rightarrow NaSH$

13. Chromium(III) oxide reacts with carbon tetrachloride to make chromium(III) chloride and phosgene ($COCl_2$). Write the balanced chemical equation for this reaction.

14. The reaction that occurs when an Alka-Seltzer tablet is dropped into a glass of water has sodium bicarbonate reacting with citric acid ($H_3C_6H_5O_7$) to make carbon dioxide, water, and sodium citrate ($Na_3C_6H_5O_7$). Write the balanced chemical equation for this reaction.

15. When sodium hydrogen carbonate is used to extinguish a kitchen fire, it decomposes into sodium carbonate, carbon dioxide and water. Write a balanced chemical equation for this reaction.

16. Elemental bromine gas can be generated by reacting sodium bromide with elemental chlorine. The other product is sodium chloride. Write a balanced chemical equation for this reaction.

Answers

1. $H_2(g) + I_2(s) \rightarrow 2HI(g)$

2. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

3. $H_2(g) + F_2(g) \rightarrow 2HF(g)$

4. $2K(s) + F_2(g) \rightarrow 2KF(s)$

5. $2Hg + O_2 \rightarrow 2HgO$

6. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

7. $2C_3H_7OH + 9O_2 \rightarrow 6CO_2 + 8H_2O$

8. $3H_2SO_4 + 2Fe \rightarrow Fe_2(SO_4)_3 + 3H_2$

9. a. $MgCl_2 + 2K \rightarrow 2KCl + Mg$
   b. $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$
   c. $2NaN_3 \rightarrow 2Na + 3N_2$

10.

   a. $NH_4NO_3 \rightarrow N_2O + 2H_2O$
   b. $TiBr_4 + 2H_2O \rightarrow TiO_2 + 4HBr$
   c. $4C_3H_5N_3O_9 \rightarrow 12CO_2 + 6N_2 + O_2 + 10H_2O$

11.

   a. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
   b. $6Li + N_2 \rightarrow 2Li_3N$
   c. $3AuCl \rightarrow 2Au + AuCl_3$

12.

   a. $3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$
   b. $N_2H_4 + 2Cl_2 \rightarrow N_2 + 4HCl$
   c. $Na_2S + H_2S \rightarrow 2NaSH$

13. $Cr_2O_3 + 3CCl_4 \rightarrow 2CrCl_3 + 3COCl_2$

14. $3NaHCO_3 + H_3C_6H_5O_7 \rightarrow 3CO_2 + 3H_2O + Na_3C_6H_5O_7$

15. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

16. $2NaBr + Cl_2 \rightarrow Br_2 + 2NaCl$

5.3: Quantitative Relationships Based on Chemical Equations

Concept Review Exercises

1. Explain how stoichiometric ratios are constructed from a chemical equation.
2. Why is it necessary for a chemical equation to be balanced before it can be used to construct conversion factors?

Answers

1. Stoichiometric ratios are made using the coefficients of the substances in the balanced chemical equation.
2. A balanced chemical equation is necessary so one can construct the proper stoichiometric ratios.

Exercises

1. Balance this equation and write every stoichiometric ratio you can from it.

$$NH_4NO_3 \rightarrow N_2O + H_2O$$

2. Balance this equation.

$$N_2 + H_2 \rightarrow NH_3$$

3. Balance this equation and write every stoichiometric ratio you can from it.

$$Fe_2O_3 + C \rightarrow Fe + CO_2$$

4. Balance this equation.

$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

5. Balance this equation and determine how many molecules of CO_2 are formed if 15 molecules of C_6H_6 are reacted.

$$C_6H_6 + O_2 \rightarrow CO_2 + H_2O$$

6. Balance this equation and determine how many formula units of $Ag_2CO_3(s)$ are produced if 20 formula units of Na_2CO_3 are reacted.

$$Na_2CO_3(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + Ag_2CO_3(s)$$

7. Copper metal reacts with nitric acid according to this equation:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Cu atoms will react if 488 molecules of aqueous HNO_3 are reacted?

8. Gold metal reacts with a combination of nitric acid and hydrochloric acid according to this equation:

$$Au(s) + 3HNO_3(aq) + 4HCl(aq) \rightarrow HAuCl_4(aq) + 3NO_2(g) + 3H_2O(\ell)$$

   a. Verify that this equation is balanced.
   b. How many Au atoms react with 639 molecules of aqueous HNO_3?

9. Sulfur can be formed by reacting sulfur dioxide with hydrogen sulfide at high temperatures according to this equation:

$$SO_2(g) + 2H_2S(g) \rightarrow 3S(g) + 2H_2O(g)$$

a. Verify that this equation is balanced.
b. How many S atoms will be formed from by reacting 1,078 molecules of H_2S?

10. Nitric acid is made by reacting nitrogen dioxide with water:

$$3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$$

a. Verify that this equation is balanced.
b. How many molecules of NO will be formed by reacting 2,268 molecules of NO_2?

Answers

1. $NH_4NO_3 \rightarrow N_2O + 2H_2O$; the stoichiometric ratios are $\frac{1NH_4NO_3}{1N_2O}$, $\frac{1NH_4NO_3}{2H_2O}$, $\frac{1N_2O}{2H_2O}$, and their reciprocals.

2. $N_2 + 3H_2 \rightarrow 2NH_3$

3. $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$; the stoichiometric ratios are $\frac{2Fe_2O_3}{3C}$, $\frac{2Fe_2O_3}{4Fe}$, $\frac{2Fe_2O_3}{3CO_2}$, $\frac{3C}{4Fe}$, $\frac{3C}{3CO_2}$, $\frac{4Fe}{3CO_2}$, and their reciprocals.

4. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

5. $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$; 90 molecules

6. $Na_2CO_3(aq) + 2AgNO_3(aq) \rightarrow 2NaNO_3(aq) + Ag_2CO_3(s)$; 20 formula units

7.

a. It is balanced.
b. 183 atoms

8.

a. It is balanced.
b. 213 atoms

9.

a. It is balanced.
b. 1,617 atoms

10.

a. It is balanced.
b. 756 molecules

5.4: Some Types of Chemical Reactions

Concept Review Exercises

1. What is the difference between a combination reaction and a combustion reaction?
2. Give the distinguishing characteristic(s) of a decomposition reaction.
3. How do we recognize a combustion reaction?

Answers

1. A combination reaction produces a certain substance; a combustion reaction is a vigorous reaction, usually a combination with oxygen, that is accompanied by the production of light and/or heat.
2. In a decomposition reaction, a single substance reacts to make multiple substances as products.
3. A combustion reaction is typically a vigorous reaction accompanied by light and/or heat, usually because of reaction with oxygen.

Exercises

1. Identify each type of reaction.

    a. $C_6H_5CH_3 + 9O_2 \rightarrow 7CO_2 + 4H_2O$
    b. $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$
    c. $C + 2H_2 \rightarrow CH_4$

2. Identify each type of reaction.

    a. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$
    b. $FeO + SO_3 \rightarrow FeSO_4$
    c. $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

3. Identify each type of reaction.

    a. $2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$
    b. $Hg(\ell) + ½O_2\,(g) \rightarrow HgO(s)$
    c. $CH_2CH_2(g) + Br_2(\ell) \rightarrow CH_2BrCH_2Br$

4. Identify each type of reaction.

    a. $Ti(s) + O_2(g) \rightarrow TiO_2(s)$
    b. $H_2SO_3(aq) \rightarrow H_2O(\ell) + SO_2(g)$
    c. $3O_2(g) \rightarrow 2O_3(g)$

Answers

1.

    a. combustion

    b. decomposition

    c. combination

2.

    a. combination

    b. combination

    c. decomposition

3.

    a. decomposition

    b. combustion (also combination)

c. combination

4.

a. combination

b. decomposition

c. combination

5.5: Oxidation-Reduction (Redox) Reactions

Concept Review Exercises

1. Give two different definitions for oxidation and reduction.
2. Give an example of each definition of oxidation and reduction.

Answers

1. Oxidation is the loss of electrons or the addition of oxygen; reduction is the gain of electrons or the addition of hydrogen.
2. $Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation); $C_2H_4 + H_2 \rightarrow C_2H_6$ (reduction) (answers will vary)

Exercises

1. Which reactions are redox reactions? For those that are redox reactions, identify the oxidizing and reducing agents.
   a. $NaOH + HCl \rightarrow H_2O + NaCl$
   b. $3Mg + 2AlCl_3 \rightarrow 2Al + 3MgCl_2$
   c. $H_2O_2 + H_2 \rightarrow 2H_2O$
   d. $KCl + AgNO_3 \rightarrow AgCl + KNO_3$
2. Which reactions are redox reactions? For those that are redox reactions, identify the oxidizing and reducing agents.
   a. $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$
   b. $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
   c. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
   d. $2K + 2H_2O \rightarrow 2KOH + H_2$
3. Balance each redox reaction by writing appropriate half reactions and combining them to cancel the electrons.
   a. $Ca(s) + H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   b. $I^-(aq) + Br_2(\ell) \rightarrow Br^-(aq) + I_2(s)$
4. Balance each redox reaction by writing appropriate half reactions and combining them to cancel the electrons.
   a. $Fe(s) + Sn^{4+}(aq) \rightarrow Fe^{3+}(aq) + Sn^{2+}(aq)$
   b. $Pb(s) + Pb^{4+}(aq) \rightarrow Pb^{2+}(aq)$ (Hint: both half reactions will yield the same product.)

Answers

1.

a. no

b. yes; oxidizing agent: $AlCl_3$; reducing agent: Mg

c. yes; oxidizing agent: H_2O_2; reducing agent: H_2

d. no

2.

a. yes; oxidizing agent: HNO_3; reducing agent: Cu
b. yes; oxidizing agent: O_2; reducing agent: C_2H_6
c. no
d. yes; oxidizing agent: H_2O; reducing agent: K

3.

a. Combined: $Ca + 2H^+ \rightarrow Ca^{2+} + H_2$
b. Combined: $Br_2 + 2I^- \rightarrow 2Br^- + I_2$

4.

a. $(Fe \rightarrow Fe^{3+} + 3e^-) \times 2$

$(Sn^{4+} + 2e^- \rightarrow Sn^{2+}) \times 3$

Combined: $2Fe + 3Sn^{4+} \rightarrow 2Fe^{3+} + 3Sn^{2+}$

b. $Pb \rightarrow Pb^{2+} + 2e^-$

$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$

Combined: $Pb + Pb^{4+} \rightarrow 2Pb^{2+}$

5.6: Redox Reactions in Organic Chemistry and Biochemistry

Concept Review Exercise

1. Give some biochemical examples of oxidation and reduction reactions.

Answer

1. photosynthesis and antioxidants in foods (answers will vary)

Exercises

1. A typical respiratory reaction discussed in the text is the oxidation of glucose ($C_6H_{12}O_6$):

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

Is this a redox reaction? If so, what are the oxidizing and reducing agents?

2. The major net reaction in photosynthesis is as follows:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Is this a redox reaction? If so, what are the oxidizing and reducing agents?

3. What would be the ultimate organic product if $CH_3CH_2CH_2OH$ were to react with a solution of $K_2Cr_2O_7$?

4. What would be the ultimate organic product if $CH_3CH_2CH_2CH_2OH$ were to react with a solution of $K_2Cr_2O_7$?

5. What would be the final organic product if $CH_3CH_2CHOHCH_3$ were to react with a solution of $K_2Cr_2O_7$?

6. What would be the major organic product if $CH_3CH_2CHOHCH_2CH_3$ were to react with a solution of $K_2Cr_2O_7$?

7. What alcohol is produced in the reduction of acetone [$(CH_3)_2CO$]?

8. What alcohol is produced in the reduction of propanal (CH_3CH_2CHO)?

Answers

1. yes; oxidizing agent: O_2; reducing agent: $C_6H_{12}O_6$

2. yes; oxidizing agent: CO_2; reducing agent: H_2O

3. CH_3CH_2COOH

4. $CH_3CH_2CH_2COOH$

5. $CH_3CH_2C(O)CH_3$, where the carbon is double bonded to the oxygen

6. $CH_3CH_2C(O)CH_2CH_3$, carbon #3 is double bonded to the oxygen

7. $CH_3CHOHCH_3$, or isopropyl alcohol

8. $CH_3CH_2CH_2OH$

5.

Additional Exercises

1. Isooctane (C_8H_{18}) is used as a standard for comparing gasoline performance. Write a balanced chemical equation for the combustion of isooctane.

2. Heptane (C_7H_{16}), like isooctane (see Exercise 1), is also used as a standard for determining gasoline performance. Write a balanced chemical equation for the combustion of heptane.

3. What is the difference between a combination reaction and a redox reaction? Are all combination reactions also redox reactions? Are all redox reactions also combination reactions?

4. Are combustion reactions always redox reactions as well? Explain.

5. A friend argues that the equation

$$Fe^{2+} + Na \rightarrow Fe + Na^{+}$$

is balanced because each side has one iron atom and one sodium atom. Explain why your friend is incorrect.

6. Some antacids contain aluminum hydroxide [$Al(OH)_3$]. This compound reacts with excess hydrochloric acid (HCl) in the stomach to neutralize it. If the products of this reaction are water and aluminum chloride, what is the balanced chemical equation for this reaction? What is the stoichiometric ratio between the number of HCl molecules made to the number of H_2O molecules made?

7. Sulfuric acid is made in a three-step process: (1) the combustion of elemental sulfur to produce sulfur dioxide, (2) the continued reaction of sulfur dioxide with oxygen to produce sulfur trioxide, and (3) the reaction of sulfur trioxide with water to make sulfuric acid (H_2SO_4). Write balanced chemical equations for all three reactions.

8. If the products of glucose metabolism are carbon dioxide and water, what is the balanced chemical equation for the overall process? What is the stoichiometric ratio between the number of CO_2 molecules made to the number of H_2O molecules made?

9. Historically, the first true battery was the Leclanché cell, named after its discoverer, Georges Leclanché. It was based on the following reaction:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Identify what is being oxidized, what is being reduced, and the respective reducing and oxidizing agents.

Answers

1. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

2. $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

3. A combination reaction makes a new substance from more than one reactant; a redox reaction rearranges electrons. Most (not all) combination reactions are redox reactions. Not all redox reactions are combination reactions.

4. All combustion reactions are redox reactions. In combustion a chemical is combining with oxygen and that chemical is being oxidized. Oxygen, on the other hand, is being reduced.

5. Your friend is incorrect because the number of electrons transferring is not balanced. A balanced equation must not only have the same number of atoms of each element on each side of the equation but must also have the same charge on both sides.

6. $Al(OH)_3 + 3HCl \rightarrow 3H_2O + AlCl_3$; 1:1

7. (1) $S + O_2 \rightarrow SO_2$; (2) $2SO_2 + O_2 \rightarrow 2SO_3$; (3) $SO_3 + H_2O \rightarrow H_2SO_4$

8. $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$; 1:1

9. oxidized and reducing agent: Zn; reduced and oxidizing agent: Cu^{2+}

This page titled 11.E: Introduction to Chemical Reactions (Exercises) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- 5.E: Introduction to Chemical Reactions (Exercises) by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.1: The Law of Conservation of Matter by Binod Shrestha is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.2: Chemical Equations by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.3: Quantitative Relationships Based on Chemical Equations by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.4: Some Types of Chemical Reactions by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.5: Oxidation-Reduction (Redox) Reactions by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 5.6: Redox Reactions in Organic Chemistry and Biochemistry by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

11.S: Introduction to Chemical Reactions (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

Scientific **laws** are general statements that apply to a wide variety of circumstances. One important law in chemistry is the **law of conservation of matter**, which states that in any closed system, the amount of matter stays constant.

Chemical equations are used to represent **chemical reactions**. **Reactants** change chemically into **products**. The law of conservation of matter requires that a proper chemical equation be **balanced**. **Coefficients** are used to show the relative numbers of reactant and product molecules.

In **stoichiometry**, quantities of reactants and/or products can be related to each other using the balanced chemical equation. The coefficients in a balanced chemical reaction are used to devise the proper ratios that relate the number of molecules of one substance to the number of molecules of another substance.

Chemical reactions can be classified by type. **Combination reactions** (also called **composition reactions**) make a substance from other substances. **Decomposition reactions** break one substance down into multiple substances. **Combustion reactions** combine molecular oxygen with the atoms of another reactant.

Oxidation reactions are reactions in which an atom loses an electron. **Reduction reactions** are reactions in which an atom gains an electron. These two processes always occur together, so they are collectively referred to as **oxidation-reduction** (or **redox**) **reactions**. The species being oxidized it called the **reducing agent**, while the species being reduced is the **oxidizing agent**. Alternate definitions of oxidation and reduction focus on the gain or loss of oxygen atoms, or the loss or gain of hydrogen atoms. Redox reactions are easily balanced if the overall reaction is first separated into **half reactions**, which are individually balanced.

Oxidation-reduction reactions are common in organic and biological chemistry. **Respiration**, the process by which we inhale and metabolize oxygen, is a series of redox reactions. In the absence of oxygen, redox reactions still occur in a process called **anaerobic metabolism**. **Antioxidants** such as ascorbic acid also play a part in the human diet, acting as reducing agents in various biochemical reactions. **Photosynthesis**, the process by which plants convert water and carbon dioxide to glucose, is also based on redox reactions.

This page titled 11.S: Introduction to Chemical Reactions (Summary) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12: QUANTITIES IN CHEMICAL REACTIONS

So far, we have talked about chemical reactions in terms of individual atoms and molecules. Although this works, most of the reactions occurring around us involve much larger amounts of chemicals. Even a tiny sample of a substance will contain millions, billions, or a hundred billion billions of atoms and molecules. How do we compare amounts of substances to each other in chemical terms when it is so difficult to count to a hundred billion billion? Actually, there are ways to do this, which we will explore in this chapter. In doing so, we will increase our understanding of stoichiometry, which is the study of the numerical relationships between the reactants and the products in a balanced chemical reaction.

12.1: Prelude to Quantities in Chemical Reactions

12.2: The Mole

12.3: Atomic and Molar Masses

12.4: Mole-Mass Conversions

12.5: Mole-Mole Relationships in Chemical Reactions

12.6: Mole-Mass and Mass-Mass Problems

12.E: Quantities in Chemical Reactions (Exercise)

12.S: Quantities in Chemical Reactions (Summary)

This page titled 12: Quantities in Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.1: Prelude to Quantities in Chemical Reactions

When the disengaged gasses are carefully examined, they are found to weigh **113.7 grs.**; these are of two kinds, viz. **144 cubical inches** of carbonic acid gas, weighing **100 grs.** and **380 cubical inches** of a very light gas, weighing only **13.7 grs.**...and, when the water which has passed over into the bottle [labeled] H is carefully examined, it is found to have lost **85.7 grs.** of its weight. Thus, in this experiment, **85.7 grs.** of water, joined to **28 grs.** of charcoal, have combined in such a way as to form **100 grs.** of carbonic acid, and **13.7 grs.** of a particular gas capable of being burnt. (Bold emphasis added.)

In this paragraph from the *Elements of Chemistry*, Antoine Lavoisier (1743–94) is explaining an experiment in which he was trying to demonstrate that water is not an element but instead is composed of hydrogen (the gas "capable of being burnt") and oxygen. This is a historical account of a groundbreaking experiment and illustrates the importance of amounts in chemistry. Lavoisier was pointing out that the initial total mass of water and charcoal, 85.7 g plus 28 g, equals the final total mass of carbonic acid and the particular gas, 100 g plus 13.7 g. In this way, he was illustrating the law of conservation of matter. It is another way of saying that *amounts matter*.

Amounts do matter and in a variety of circumstances. The chapter-opening essay in Chapter 1 tells the story of a nurse who mistakenly read "2–3 mg" as "23 mg" and administered the higher and potentially fatal dose of morphine to a child. Food scientists who work in test kitchens must keep track of specific amounts of ingredients as they develop new products for us to eat. Quality control technicians measure amounts of substances in manufactured products to ensure that the products meet company or government standards. Supermarkets routinely weigh meat and produce and charge consumers by the ounce or the pound.

This page titled 12.1: Prelude to Quantities in Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.2: The Mole

Learning Objectives

- To define the mole unit.

Figure 12.2.1 shows that we need 2 hydrogen atoms and 1 oxygen atom to make 1 water molecule. If we want to make 2 water molecules, we will need 4 hydrogen atoms and 2 oxygen atoms. If we want to make 5 molecules of water, we need 10 hydrogen atoms and 5 oxygen atoms. The ratio of atoms we will need to make any number of water molecules is the same: 2 hydrogen atoms to 1 oxygen atom.

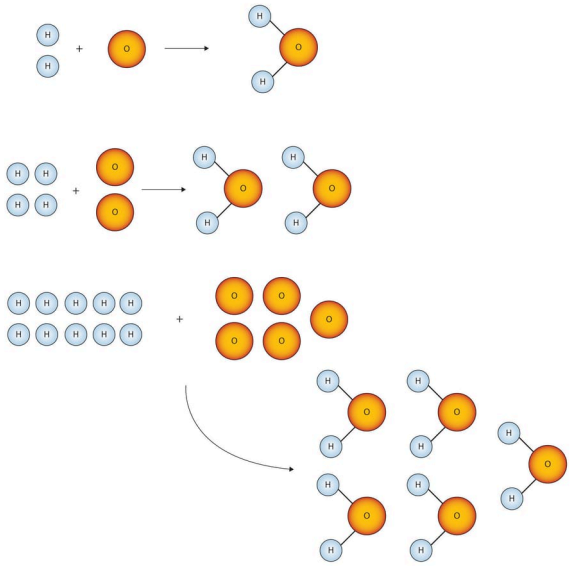


Figure 12.2.1 Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

One problem we have, however, is that it is extremely difficult, if not impossible, to organize atoms one at a time. As stated in the introduction, we deal with billions of atoms at a time. How can we keep track of so many atoms (and molecules) at a time? We do it by using mass rather than by counting individual atoms.

A hydrogen atom has a mass of approximately 1 u. An oxygen atom has a mass of approximately 16 u. The ratio of the mass of an oxygen atom to the mass of a hydrogen atom is therefore approximately 16:1.

If we have 2 atoms of each element, the ratio of their masses is approximately 32:2, which reduces to 16:1—the same ratio. If we have 12 atoms of each element, the ratio of their total masses is approximately (12 × 16):(12 × 1), or 192:12, which also reduces to 16:1. If we have 100 atoms of each element, the ratio of the masses is approximately 1,600:100, which again reduces to 16:1. As long as we have equal numbers of hydrogen and oxygen atoms, the ratio of the masses will always be 16:1.

The same consistency is seen when ratios of the masses of other elements are compared. For example, the ratio of the masses of silicon atoms to equal numbers of hydrogen atoms is always approximately 28:1, while the ratio of the masses of calcium atoms to equal numbers of lithium atoms is approximately 40:7.

So we have established that the masses of atoms are constant with respect to each other, as long as we have the same number of each type of atom. Consider a more macroscopic example. If a sample contains 40 g of Ca, this sample has the same number of atoms as there are in a sample of 7 g of Li. What we need, then, is a number that represents a convenient quantity of atoms so we can relate macroscopic quantities of substances. Clearly even 12 atoms are too few because atoms themselves are so small. We need a number that represents billions and billions of atoms.

Chemists use the term mole to represent a large number of atoms or molecules. Just as a dozen implies 12 things, a mole (mol) represents 6.022×10^{23} things. The number 6.022×10^{23}, called Avogadro's number after the 19th-century chemist Amedeo Avogadro, is the number we use in chemistry to represent macroscopic amounts of atoms and molecules. Thus, if we have 6.022×10^{23} O atoms, we say we have 1 mol of O atoms. If we have 2 mol of Na atoms, we have $2 \times (6.022 \times 10^{23})$ Na atoms, or 1.2044×10^{24} Na atoms. Similarly, if we have 0.5 mol of benzene (C_6H_6) molecules, we have $0.5 \times (6.022 \times 10^{23})$ C_6H_6 molecules, or 3.011×10^{23} C_6H_6 molecules.

A mole represents a very large number! If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Notice that we are applying the mole unit to different types of chemical entities. In these examples, we cited moles of atoms *and* moles of molecules. The word *mole* represents a number of things—6.022×10^{23} of them—but does not by itself specify what "they" are. They can be atoms, formula units (of ionic compounds), or molecules. That information still needs to be specified.

Because 1 H_2 molecule contains 2 H atoms, 1 mol of H_2 molecules (6.022×10^{23} molecules) has 2 mol of H atoms. Using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of moles of molecules to the number of moles of atoms. For example, in 1 mol of ethanol (C_2H_6O), we can construct the following relationships (Table 12.2.1):

Table 12.2.1*: Molecular Relationships*

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular Relationships
2 C atoms	2 mol of C atoms	$\frac{2\text{ mol C atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{2\text{ mol C atoms}}$
6 H atoms	6 mol of H atoms	$\frac{6\text{ mol H atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{6\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}$ or $\frac{1\text{ mol }C_2H_6O\text{ molecules}}{1\text{ mol O atoms}}$

The following example illustrates how we can use these relationships as conversion factors.

Example 12.2.1

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms, hydrogen atoms, and oxygen atoms does it have?

SOLUTION

Using the relationships in Table 12.2.1, we apply the appropriate conversion factor for each element:

$$2.5\text{ mol }C_2H_6O\text{ molecules} \times \underbrace{\frac{2\text{ mol C atoms}}{1\text{ mol }C_2H_6O\text{ molecules}}}_{\text{conversion factor}} = 5.0\text{ mol C atoms}$$

Note how the unit *mol C_2H_6O molecules* cancels algebraically. Similar equations can be constructed for determining the number of H and O atoms:

$$2.5\text{ mol }C_2H_6O\text{ molecules} \times \frac{6\text{ mol H atoms}}{1\text{ mol }C_2H_6O\text{ molecules}} = 15\text{ mol H atoms}$$

$$2.5\text{ mol }C_2H_6O\text{ molecules} \times \frac{1\text{ mol O atoms}}{1\text{ mol }C_2H_6O\text{ molecules}} = 2.5\text{ mol O atoms}$$

Exercise 12.2.1

If a sample contains 6.75 mol of Na_2SO_4, how many moles of sodium atoms, sulfur atoms, and oxygen atoms does it have?

The fact that 1 mol equals 6.022×10^{23} items can also be used as a conversion factor.

Example 12.2.2

How many formula units are present in 2.34 mol of NaCl? How many ions are in 2.34 mol?

SOLUTION

Typically in a problem like this, we start with what we are given and apply the appropriate conversion factor. Here, we are given a quantity of 2.34 mol of NaCl, to which we can apply the definition of a mole as a conversion factor:

$$2.34 \text{ mol NaCl} \times \frac{6.022 \times 10^{23} \text{ NaCl units}}{1 \text{ mol NaCl}} = 1.41 \times 10^{24} \text{ NaCl units}$$

Because there are two ions per formula unit, there are

$$1.41 \times 10^{24} \text{ NaCl units} \times \frac{2 \text{ ions}}{\text{NaCl units}} = 2.82 \times 10^{24} \text{ ions}$$

in the sample.

Exercise 12.2.2

How many molecules are present in 16.02 mol of C_4H_{10}? How many atoms are in 16.02 mol?

Concept Review Exercise

1. What is a mole?

Answer

1. A mole is 6.022×10^{23} things.

Key Takeaway

- A mole is 6.022×10^{23} things.

Exercises

1. How many dozens are in 1 mol? Express your answer in proper scientific notation.
2. A gross is a dozen dozen, or 144 things. How many gross are in 1 mol? Express your answer in proper scientific notation.
3. How many moles of each type of atom are in 1.0 mol of $C_6H_{12}O_6$?
4. How many moles of each type of atom are in 1.0 mol of $K_2Cr_2O_7$?

5. How many moles of each type of atom are in 2.58 mol of Na_2SO_4?
6. How many moles of each type of atom are in 0.683 mol of $C_{34}H_{32}FeN_4O_4$? (This is the formula of heme, a component of hemoglobin.)
7. How many molecules are in 16.8 mol of H_2O?
8. How many formula units are in 0.778 mol of iron(III) nitrate?
9. A sample of gold contains 7.02×10^{24} atoms. How many moles of gold is this?
10. A flask of mercury contains 3.77×10^{22} atoms. How many moles of mercury are in the flask?
11. An intravenous solution of normal saline may contain 1.72 mol of sodium chloride (NaCl). How many sodium and chlorine atoms are present in the solution?
12. A lethal dose of arsenic is 1.00×10^{21} atoms. How many moles of arsenic is this?

Answers

1. 5.018×10^{22} dozens
3. 6.0 mol of C atoms, 12.0 mol of H atoms, and 6.0 mol of O atoms
5. 5.16 mol of Na atoms, 2.58 mol of S atoms, and 10.32 mol of O atoms
7. 1.012×10^{25} molecules
9. 11.7 mol
11. 1.04×10^{24} Na atoms and 1.04×10^{24} Cl atoms

This page titled 12.2: The Mole is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.3: Atomic and Molar Masses

Learning Objectives

- To learn how the masses of moles of atoms and molecules are expressed.

Now that we have introduced the mole and practiced using it as a conversion factor, we ask the obvious question: why is the mole *that particular* number of things? Why is it 6.022×10^{23} and not 1×10^{23} or even 1×10^{20}?

The number in a mole, Avogadro's number, is related to the relative sizes of the atomic mass unit and gram mass units. Whereas one hydrogen atom has a mass of approximately 1 u, 1 mol of H atoms has a mass of approximately 1 *gram*. And whereas one sodium atom has an approximate mass of 23 u, 1 mol of Na atoms has an approximate mass of 23 *grams*.

One mole of a substance has the same mass in grams that one atom or molecule has in atomic mass units. The numbers in the periodic table that we identified as the atomic masses of the atoms not only tell us the mass of one atom in u but also tell us the mass of 1 mol of atoms in grams.

> One mole of a substance has the same mass in grams that one atom or molecule has in atomic mass units.

Example $12.3.1$: Moles to Mass Conversion with Elements

What is the mass of each quantity?

a. 1 mol of Al atoms
b. 2 mol of U atoms

SOLUTION

a. One mole of Al atoms has a mass in grams that is numerically equivalent to the atomic mass of aluminum. The periodic table shows that the atomic mass (rounded to two decimal points) of Al is 26.98, so 1 mol of Al atoms has a mass of 26.98 g.
b. According to the periodic table, 1 mol of U has a mass of 238.03 g, so the mass of 2 mol is twice that, or 476.06 g.

Exercise $12.3.1$: Moles to Mass Conversion with Elements

What is the mass of each quantity?

a. 1 mol of Au atoms
b. 5 mol of Br atoms

The mole concept can be extended to masses of formula units and molecules as well. The mass of 1 mol of molecules (or formula units) in grams is numerically equivalent to the mass of one molecule (or formula unit) in atomic mass units. For example, a single molecule of O_2 has a mass of 32.00 u, and 1 mol of O_2 molecules has a mass of 32.00 g. As with atomic mass unit–based masses, to obtain the mass of 1 mol of a substance, we simply sum the masses of the individual atoms in the formula of that substance. The mass of 1 mol of a substance is referred to as its molar mass, whether the substance is an element, an ionic compound, or a covalent compound.

Example $12.3.2$: Moles to Mass Conversion with Compounds

What is the mass of 1 mol of each substance?

1. NaCl
2. bilirubin ($C_{33}H_{36}N_4O_6$), the principal pigment present in bile (a liver secretion)

SOLUTION

1. Summing the molar masses of the atoms in the NaCl formula unit gives

1 Na molar mass:	23.00 g
1 Cl molar mass:	35.45 g
Total:	58.45 g

The mass of 1 mol of NaCl is 58.45 g.

2. Multiplying the molar mass of each atom by the number of atoms of that type in bilirubin's formula and adding the results, we get

33 C molar mass:	33 × 12.01 g	396.33 g
36 H molar mass:	36 × 1.01 =	36.36 g
4 N molar mass:	4 × 14.00 =	56.00 g
6 O molar mass:	6 × 16.00 =	96.00 g
Total:		584.69 g

The mass of 1 mol of bilirubin is 584.69 g.

Exercise 12.3.2: Moles to Mass Conversion with Compounds

What is the mass of 1 mol of each substance?

a. barium sulfate ($BaSO_4$), used to take X rays of the gastrointestional tract
b. adenosine ($C_{10}H_{13}N_5O_4$), a component of cell nuclei crucial for cell division

Be careful when counting atoms. In formulas with polyatomic ions in parentheses, the subscript outside the parentheses is applied to every atom inside the parentheses. For example, the molar mass of $Ba(OH)_2$ requires the sum of 1 mass of Ba, 2 masses of O, and 2 masses of H:

1 Ba molar mass:	1 × 137.33 g =	137.33 g
2 O molar mass:	2 × 16.00 g =	32.00 g
2 H molar mass:	2 × 1.01 g =	2.02 g
Total:		171.35 g

Because molar mass is defined as the mass for 1 mol of a substance, we can refer to molar mass as grams per mole (g/mol). The division sign (/) implies "per," and "1" is implied in the denominator. Thus, the molar mass of bilirubin can be expressed as 584.05 g/mol, which is read as "five hundred eighty four point zero five grams per mole."

Concept Review Exercises

1. How are molar masses of the elements determined?
2. How are molar masses of compounds determined?

Answers

1. Molar masses of the elements are the same numeric value as the masses of a single atom in atomic mass units but in units of grams instead.
2. Molar masses of compounds are calculated by adding the molar masses of their atoms.

Key Takeaway

- The mass of moles of atoms and molecules is expressed in units of grams.

Exercises

1. What is the molar mass of Si? What is the molar mass of U?
2. What is the molar mass of Mn? What is the molar mass of Mg?
3. What is the molar mass of $FeCl_2$? What is the molar mass of $FeCl_3$?
4. What is the molar mass of C_6H_6? What is the molar mass of $C_6H_5CH_3$?
5. What is the molar mass of $(NH_4)_2S$? What is the molar mass of $Ca(OH)_2$?
6. What is the molar mass of $(NH_4)_3PO_4$? What is the molar mass of $Sr(HCO_3)_2$?
7. Aspirin ($C_9H_8O_4$) is an analgesic (painkiller) and antipyretic (fever reducer). What is the molar mass of aspirin?
8. Ibuprofen ($C_{13}H_{18}O_2$) is an analgesic (painkiller). What is the molar mass of ibuprofen?
9. Morphine ($C_{17}H_{19}NO_3$) is a narcotic painkiller. What is the mass of 1 mol of morphine?
10. Heroin ($C_{21}H_{23}NO_5$) is a narcotic drug that is a derivative of morphine. What is the mass of 1 mol of heroin?

Answers

1. 28.09 g/mol; 238.00 g/mol
3. 126.75 g/mol; 162.20 g/mol
5. 68.15 g/mol; 74.10 g/mol
7. 180.17 g/mol

9 285.36 g

This page titled 12.3: Atomic and Molar Masses is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.4: Mole-Mass Conversions

Learning Objectives

- s to convert quantities between mass units and mole units.

A previous Example stated that the mass of 2 mol of U is twice the molar mass of uranium. Such a straightforward exercise does not require any formal mathematical treatment. Many questions concerning mass are not so straightforward, however, and require some mathematical manipulations.

The simplest type of manipulation using molar mass as a conversion factor is a mole-mass conversion (or its reverse, a mass-mole conversion). In such a conversion, we use the molar mass of a substance as a conversion factor to convert mole units into mass units (or, conversely, mass units into mole units).

We also established that 1 mol of Al has a mass of 26.98 g (Example). Stated mathematically,

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

We can divide both sides of this expression by either side to get one of two possible conversion factors:

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \text{ and } \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \tag{12.4.1}$$

The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do “by eye.”

Example 12.4.1

What is the mass of 3.987 mol of Al?

SOLUTION

The first step in a conversion problem is to decide what conversion factor to use. Because we are starting with mole units, we want a conversion factor that will cancel the mole unit and introduce the unit for mass in the numerator. Therefore, we should use the $\frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$ conversion factor. We start with the given quantity and multiply by the conversion factor:

$$3.987 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$$

Note that the mol units cancel algebraically. (The quantity 3.987 mol is understood to be in the numerator of a fraction that has 1 in the unwritten denominator.) Canceling and solving gives

$$3.987 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 107.6 \text{ g Al}$$

Our final answer is expressed to four significant figures.

Exercise 12.4.1

How many moles are present in 100.0 g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 12.4.1 is a chart for determining what conversion factor is needed,

and Figure 12.4.2 is a flow diagram for the steps needed to perform a conversion.

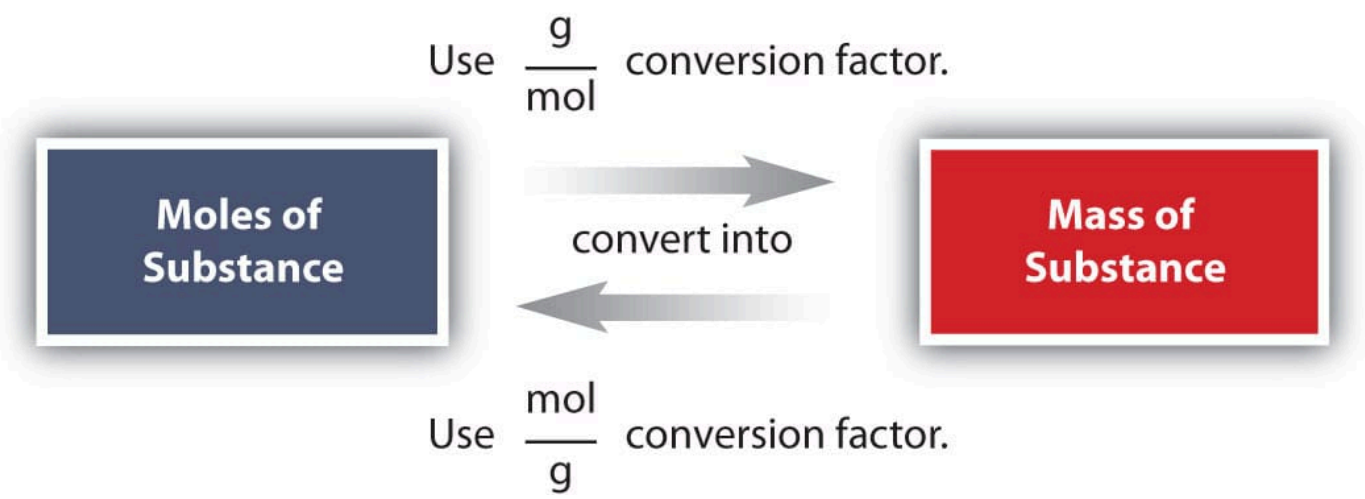


Figure 12.4.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

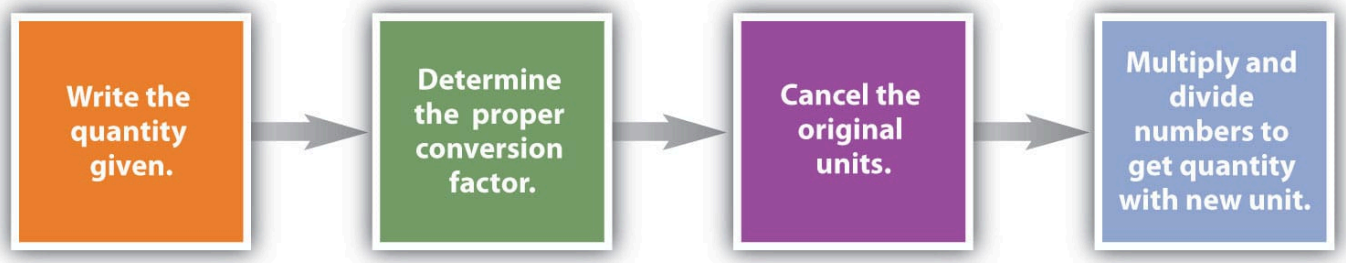


Figure 12.4.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Example 12.4.2

A biochemist needs 0.00655 mol of bilirubin ($C_{33}H_{36}N_4O_6$) for an experiment. How many grams of bilirubin will that be?

SOLUTION

To convert from moles to mass, we need the molar mass of bilirubin, which we can determine from its chemical formula:

33 C molar mass:	33 × 12.01 g =	396.33 g
36 H molar mass:	36 × 1.01 g =	36.36 g
4 N molar mass:	4 × 14.00 g =	56.00 g
6 O molar mass:	6 × 16.00 g =	96.00 g
Total:		584.69 g

The molar mass of bilirubin is 584.69 g. Using the relationship

1 mol bilirubin = 584.69 g bilirubin

we can construct the appropriate conversion factor for determining how many grams there are in 0.00655 mol. Following the steps from Figure 12.4.2:

$$0.00655 \text{ mol bilirubin} \times \frac{584.69 \text{ g bilirubin}}{\text{mol bilirubin}} = 3.83 \text{ g bilirubin}$$

The mol bilirubin unit cancels. The biochemist needs 3.83 g of bilirubin.

Exercise 12.4.2

A chemist needs 457.8 g of $KMnO_4$ to make a solution. How many moles of $KMnO_4$ is that?

To Your Health: Minerals

For our bodies to function properly, we need to ingest certain substances from our diets. Among our dietary needs are minerals, the noncarbon elements our body uses for a variety of functions, such developing bone or ensuring proper nerve transmission. The US Department of Agriculture has established some recommendations for the RDIs of various minerals. The accompanying table lists the RDIs for minerals, both in mass and moles, assuming a 2,000-calorie daily diet.

Table 12.4.1: Essential Minerals and their Composition in Humans

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
Ca	1,000 mg	0.025 mol	1,000 mg	0.025 mol
Cr	35 µg	6.7×10^{-7} mol	25 µg	4.8×10^{-7} mol
Cu	900 µg	1.4×10^{-5} mol	900 µg	1.4×10^{-5} mol
F	4 mg	2.1×10^{-4} mol	3 mg	1.5×10^{-4} mol
I	150 µg	1.2×10^{-6} mol	150 µg	1.2×10^{-6} mol
Fe	8 mg	1.4×10^{-4} mol	18 mg	3.2×10^{-4} mol
K	3,500 mg	9.0×10^{-2} mol	3,500 mg	9.0×10^{-2} mol
Mg	400 mg	1.6×10^{-2} mol	310 mg	1.3×10^{-2} mol
Mn	2.3 mg	4.2×10^{-5} mol	1.8 mg	3.3×10^{-5} mol
Mo	45 mg	4.7×10^{-7} mol	45 mg	4.7×10^{-7} mol
Na	2,400 mg	1.0×10^{-1} mol	2,400 mg	1.0×10^{-1} mol
P	700 mg	2.3×10^{-2} mol	700 mg	2.3×10^{-2} mol
Se	55 µg	7.0×10^{-7} mol	55 µg	7.0×10^{-7} mol
Zn	11 mg	1.7×10^{-4} mol	8 mg	1.2×10^{-4} mol

Table 12.4.1 illustrates several things. First, the needs of men and women for some minerals are different. The extreme case is for iron; women need over twice as much as men do. In all other cases where there is a different RDI, men need more than women.

Second, the amounts of the various minerals needed on a daily basis vary widely—both on a mass scale and a molar scale. The average person needs 0.1 mol of Na a day, which is about 2.5 g. On the other hand, a person needs only about 25–35 µg of Cr per day, which is under one millionth of a mole. As small as this amount is, a deficiency of chromium in the diet can lead to diabetes-like symptoms or neurological problems, especially in the extremities (hands and feet). For some minerals, the body does not require much to keep itself operating properly.

Although a properly balanced diet will provide all the necessary minerals, some people take dietary supplements. However, too much of a good thing, even minerals, is not good. Exposure to too much chromium, for example, causes a skin irritation, and certain forms of chromium are known to cause cancer (as presented in the movie *Erin Brockovich*).

Concept Review Exercises

1. What relationship is needed to perform mole-mass conversions?
2. What information determines which conversion factor is used in a mole-mass conversion?

Answers

1. The atomic or molar mass is needed for a mole-mass conversion.
2. The unit of the initial quantity determines which conversion factor is used.

Key Takeaway

- It is possible to convert between moles of material and mass of material.

Exercises

1. What is the mass of 8.603 mol of Fe metal?
2. What is the mass of 0.552 mol of Ag metal?
3. What is the mass of 6.24×10^{4} mol of Cl_2 gas?
4. What is the mass of 0.661 mol of O_2 gas?
5. What is the mass of 20.77 mol of $CaCO_3$?
6. What is the mass of 9.02×10^{-3} mol of the hormone epinephrine ($C_9H_{13}NO_3$)?
7. How many moles are present in 977.4 g of $NaHCO_3$?
8. How many moles of erythromycin ($C_{37}H_{67}NO_{13}$), a widely used antibiotic, are in 1.00×10^{3} g of the substance?
9. Cortisone ($C_{21}H_{28}O_5$) is a synthetic steroid that is used as an anti-inflammatory drug. How many moles of cortisone are present in one 10.0 mg tablet?
10. Recent research suggests that the daily ingestion of 85 mg of aspirin (also known as acetylsalicylic acid, $C_9H_8O_4$) will reduce a person's risk of heart disease. How many moles of aspirin is that?

Answers

1. 480.5 g
3. 4.42×10^{6} g
5. 2,079 g
7. 11.63 mol

9. 2.77×10^{-5} mol

This page titled 12.4: Mole-Mass Conversions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.5: Mole-Mole Relationships in Chemical Reactions

Learning Objectives

- To use a balanced chemical reaction to determine molar relationships between the substances.

In Chapter 5, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). Here, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O \tag{12.5.1}$$

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4\,H_2 + 2\,O_2 \longrightarrow 4\,H_2O \tag{12.5.2}$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22\,H_2 + 11\,O_2 \longrightarrow 22\,H_2O \tag{12.5.3}$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O \tag{12.5.4}$$

These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ \text{mol}\ \ H_2 + 1\ \ \text{mol}\ \ O_2 \longrightarrow 2\ \text{mol}\ \ H_2O \tag{12.5.5}$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O \tag{12.5.6}$$

Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level but also in terms of molar amounts of reactants and products*. Thus, we can read this reaction as “two moles of hydrogen react with one mole of oxygen to produce two moles of water.”

By the same token, the ratios we constructed to describe molecules reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2} \tag{12.5.7}$$

$$\frac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} \tag{12.5.8}$$

$$\frac{2\ \text{mol}\ H_2}{2\ \text{mol}\ H_2O}\ \text{or}\ \frac{2\ \text{mol}\ H_2O}{2\ \text{mol}\ H_2} \tag{12.5.9}$$

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical

reactions is called *stoichiometry*.

Example 12.5.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O? The balanced equation is as follows:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

SOLUTION

Because we are dealing with quantities of H_2O and O_2, we will use a ratio that relates those two substances. Because we are given an amount of H_2O and want to determine an amount of O_2, we will use the ratio that has H_2O in the denominator (so it cancels) and O_2 in the numerator (so it is introduced in the answer). Thus,

$$27.6\ \text{mol}\ H_2O \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} = 13.8\ \text{mol}\ O_2$$

To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.

Exercise 12.5.1

Using $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of hydrogen react with 3.07 mol of oxygen to produce H_2O?

Concept Review Exercise

1. How do we relate molar amounts of substances in chemical reactions?

Answer

1. Amounts of substances in chemical reactions are related by their coefficients in the balanced chemical equation.

Key Takeaway

- The balanced chemical reaction can be used to determine molar relationships between substances.

Exercises

1. List the molar ratios you can derive from this balanced chemical equation:

   $$NH_3 + 2O_2 \rightarrow HNO_3 + H_2O$$

2. List the molar ratios you can derive from this balanced chemical equation

   $$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

3. Given the following balanced chemical equation,

   $$6NaOH + 3Cl_2 \rightarrow NaClO_3 + 5NaCl + 3H_2O$$

   how many moles of NaCl can be formed if 3.77 mol of NaOH were to react?

4. Given the following balanced chemical equation,

$$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$$

how many moles of H_2O can be formed if 0.0652 mol of C_5H_{12} were to react?

5. Balance the following unbalanced equation and determine how many moles of H_2O are produced when 1.65 mol of NH_3 react.

$$NH_3 + O_2 \rightarrow N_2 + H_2O$$

6. Trinitrotoluene [$C_6H_2(NO_2)_2CH_3$], also known as TNT, is formed by reacting nitric acid (HNO_3) with toluene ($C_6H_5CH_3$):

$$HNO_3 + C_6H_5CH_3 \rightarrow C_6H_2(NO_2)_2CH_3 + H_2O$$

Balance the equation and determine how many moles of TNT are produced when 4.903 mol of HNO_3 react.

7. Chemical reactions are balanced in terms of molecules and in terms of moles. Are they balanced in terms of dozens? Defend your answer.

8. Explain how a chemical reaction balanced in terms of moles satisfies the law of conservation of matter.

Answers

1. 1 mol NH_3:2 mol O_2:1 mol HNO_3:1 mol H_2O

3. 3.14 mol

5. $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol

7. Yes, they are still balanced.

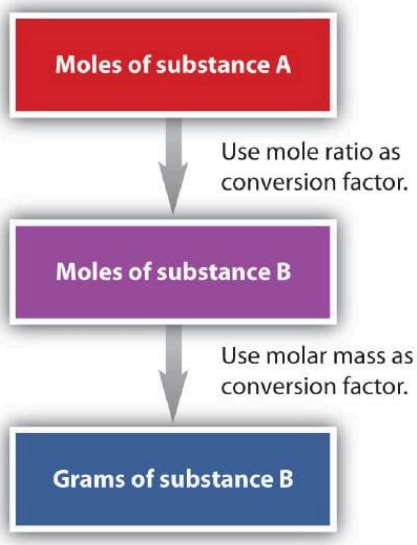
This page titled 12.5: Mole-Mole Relationships in Chemical Reactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.6: Mole-Mass and Mass-Mass Problems

Learning Objectives

- To convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

We have established that a balanced chemical equation is balanced in terms of moles as well as atoms or molecules. We have used balanced equations to set up ratios, now in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions, such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that ability to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations.

As an example, consider the balanced chemical equation

$$Fe_2O_3 + 3SO_3 \rightarrow Fe_2(SO_4)_3 \tag{12.6.1}$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

The first step resembles the exercises we did in Section 6.4 "Mole-Mole Relationships in Chemical Reactions". As usual, we start with the quantity we were given:

$$3.59 \text{ mol } Fe_2O_3 \times \frac{3 \text{ mol } SO_3}{1 \text{ mol } Fe_2O_3} = 10.77 \text{ mol } SO_3 \tag{12.6.2}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77 \text{ mol } SO_3 \times \frac{80.06 \text{ g } SO_3}{1 \text{ mol } SO_3} = 862 \text{ g } SO_3 \tag{12.6.3}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:

$$3.59 \text{ mol } Fe_2O_3 \times \underbrace{\frac{3 \text{ mol } SO_3}{1 \text{ mol } Fe_2O_3}}_{\text{converts to moles of } SO_3} \times \underbrace{\frac{80.06 \text{ g } SO_3}{1 \text{ mol } SO_3}}_{\text{converts to grams of } SO_3} = 862 \text{ g } SO_3$$

We get exactly the same answer when combining all the math steps together as we do when we calculate one step at a time.

Example 12.6.1

How many grams of CO_2 are produced if 2.09 mol of HCl are reacted according to this balanced chemical equation?

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O \tag{12.6.4}$$

SOLUTION

Our strategy will be to convert from moles of HCl to moles of CO_2 and then from moles of CO_2 to grams of CO_2. We will need the molar mass of CO_2, which is 44.01 g/mol. Performing these two conversions in a single-line gives 46.0 g of CO_2:

$$2.09 \text{ mol HCl} \times \underbrace{\frac{1 \text{ mol } CO_2}{2 \text{ mol HCl}}}_{\text{from the coefficients of the balanced equation}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 46.0 \text{ g } CO_2$$

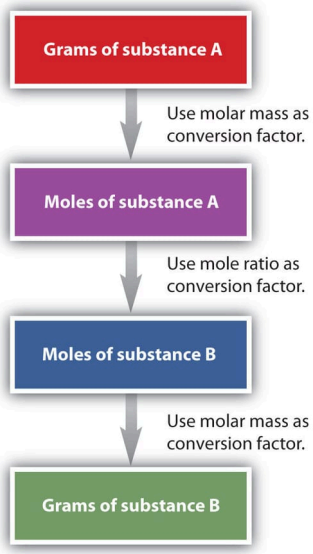
The molar ratio between CO_2 and HCl comes from the balanced chemical equation.

Exercise

How many grams of glucose ($C_6H_{12}O_6$) are produced if 17.3 mol of H_2O are reacted according to this balanced chemical equation?

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2 \tag{12.6.5}$$

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

Example 12.6.2 :Chlorination of Carbon

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows:

$$CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl \tag{12.6.6}$$

How many grams of HCl are produced by the reaction of 100.0 g of CH_4?

SOLUTION

First, let us work the problem in stepwise fashion. We begin by converting the mass of CH_4 to moles of CH_4, using the molar mass of CH_4 (16.05 g/mol) as the conversion factor:

$$100.0 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.05 \text{ g } CH_4} = 6.231 \text{ mol } CH_4$$

Note that we inverted the molar mass so that the gram units cancel, giving us an answer in moles. Next, we use the balanced chemical equation to determine the ratio of moles CH_4 and moles HCl and convert our first result into moles of HCl:

$$6.231 \text{ mol } CH_4 \times \frac{4 \text{ mol HCl}}{1 \text{ mol } CH_4} = 24.92 \text{ mol HCl}$$

Finally, we use the molar mass of HCl (36.46 g/mol) as a conversion factor to calculate the mass of 24.92 mol of HCl:

$$24.92 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 908.5 \text{ g HCl}$$

In each step, we have limited the answer to the proper number of significant figures. If desired, we can do all three conversions on a single line:

$$100.0 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.05 \text{ g } CH_4} \times \frac{4 \text{ mol HCl}}{1 \text{ mol } CH_4} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 908.7 \text{ g HCl}$$

This final answer is slightly different from our first answer because only the final answer is restricted to the proper number of significant figures. In the first answer, we limited each intermediate quantity to the proper number of significant figures. As you can see, both answers are essentially the same.

Exercise 12.6.2: Oxidation of Propanal

The oxidation of propanal (CH_3CH_2CHO) to propionic acid (CH_3CH_2COOH) has the following chemical equation:

$$CH_3CH_2CHO + 2K_2Cr_2O_7 \rightarrow CH_3CH_2COOH + \text{other products}$$

How many grams of propionic acid are produced by the reaction of 135.8 g of $K_2Cr_2O_7$?

To Your Health: The Synthesis of Taxol

Taxol is a powerful anticancer drug that was originally extracted from the Pacific yew tree (*Taxus brevifolia*). As you can see from the accompanying figure, taxol is a very complicated molecule, with a molecular formula of $C_{47}H_{51}NO_{14}$. Isolating taxol from its natural source presents certain challenges, mainly that the Pacific yew is a slow-growing tree, and the equivalent of six trees must be harvested to provide enough taxol to treat a single patient. Although related species of yew trees also produce taxol in small amounts, there is significant interest in synthesizing this complex molecule in the laboratory.

After a 20-year effort, two research groups announced the complete laboratory synthesis of taxol in 1994. However, each synthesis required over 30 separate chemical reactions, with an overall efficiency of less than 0.05%. To put this in perspective, to obtain a single 300 mg dose of taxol, you would have to begin with 600 g of starting material. To treat the 26,000 women who are diagnosed with ovarian cancer each year with one dose, almost 16,000 kg (over 17 tons) of starting material must be converted to taxol. Taxol is also used to treat breast cancer, with which 200,000 women in the United States are diagnosed every year. This only increases the amount of starting material needed.

Clearly, there is intense interest in increasing the overall efficiency of the taxol synthesis. An improved synthesis not only will be easier but also will produce less waste materials, which will allow more people to take advantage of this potentially life-saving drug.

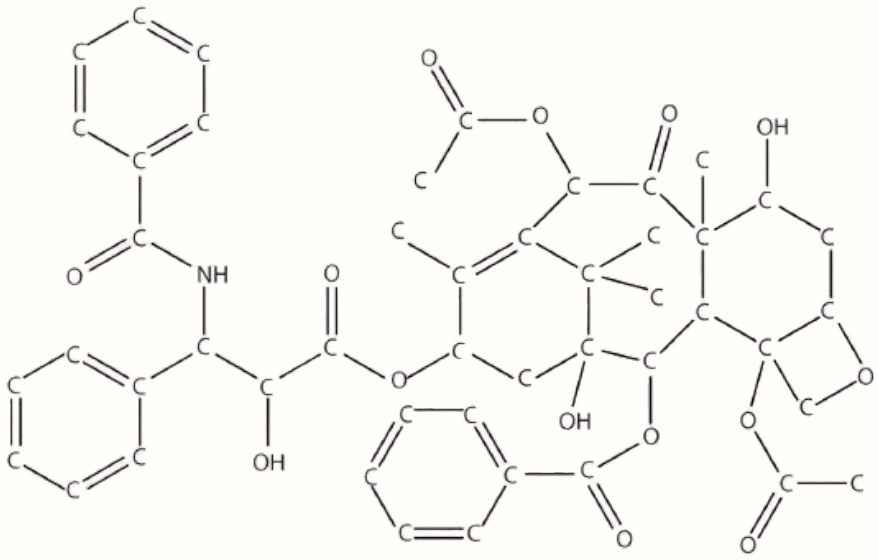


Figure 12.6.1 The Structure of the Cancer Drug Taxol. Because of the complexity of the molecule, hydrogen atoms are not shown, but they are present on every atom to give the atom the correct number of covalent bonds (four bonds for each carbon atom).

Concept Review Exercises

1. What is the general sequence of conversions for a mole-mass calculation?
2. What is the general sequence of conversions for a mass-mass calculation?

Answers

1. mol first substance → mol second substance → mass second substance
2. mass first substance → mol first substance → mol second substance → mass second substance

Key Takeaway

- A balanced chemical equation can be used to relate masses or moles of different substances in a reaction.

Exercises

1. Given the following unbalanced chemical equation,

   $$H_3PO_4 + NaOH \rightarrow H_2O + Na_3PO_4$$

   what mass of H_2O is produced by the reaction of 2.35 mol of H_3PO_4?

2. Given the following unbalanced chemical equation,

   $$C_2H_6 + Br_2 \rightarrow C_2H_4Br_2 + HBr$$

   what mass of HBr is produced if 0.884 mol of C_2H_6 is reacted?

3. Certain fats are used to make soap, the first step being to react the fat with water to make glycerol (also known as glycerin) and compounds called fatty acids. One example is as follows:

   $$\underset{\text{a fat}}{C_3H_5(OOC(CH_2)_{14}CH_3)_3} + 3H_2O \rightarrow \underset{\text{glycerol}}{C_3H_5(OH)_3} + 3\underset{\text{fatty acid}}{CH_3(CH_2)_{14}COOH}$$

   How many moles of glycerol can be made from the reaction of 1,000.0 g of $C_3H_5(OOC(CH_2)_{14}CH_3)_3$?

4. Photosynthesis in plants leads to the general overall reaction for producing glucose ($C_6H_{12}O_6$):

   $$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

   How many moles of glucose can be made from the reaction of 544 g of CO_2?

5. Precipitation reactions, in which a solid (called a precipitate) is a product, are commonly used to remove certain ions from solution. One such reaction is as follows:

   $$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$$

   How many grams of Na_2SO_4 are needed to precipitate all the barium ions produced by 43.9 g of $Ba(NO_3)_2$?

6. Nitroglycerin [$C_3H_5(ONO_2)_3$] is made by reacting nitric acid (HNO_3) with glycerol [$C_3H_5(OH)_3$] according to this reaction:

   $$C_3H_5(OH)_3 + 3HNO_3 \rightarrow C_3H_5(ONO_2)_3 + 3H_2O$$

   If 87.4 g of HNO_3 are reacted with excess glycerol, what mass of nitroglycerin can be made?

7. Antacids are bases that neutralize acids in the digestive tract. Magnesium hydroxide [$Mg(OH)_2$] is one such antacid. It reacts with hydrochloric acid in the stomach according to the following reaction:

   $$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$$

   How many grams of HCl can a 200 mg dose of $Mg(OH)_2$ neutralize?

8. Acid rain is caused by the reaction of nonmetal oxides with water in the atmosphere. One such reaction involves nitrogen dioxide (NO_2) and produces nitric acid (HNO_3):

$$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

If 1.82×10^{13} g of NO_2 enter the atmosphere every year due to human activities, potentially how many grams of HNO_3 can be produced annually?

9. A simplified version of the processing of iron ore into iron metal is as follows:

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

How many grams of C are needed to produce 1.00×10^9 g of Fe?

10. The SS *Hindenburg* contained about 5.33×10^5 g of H_2 gas when it burned at Lakehurst, New Jersey, in 1937. The chemical reaction is as follows:

$$2H_2 + O_2 \rightarrow 2H_2O$$

How many grams of H_2O were produced?

Answers

1. 127 g
3. 1.236 mol
5. 23.9 g
7. 0.251 g

9. 1.61×10^8 g

This page titled 12.6: Mole-Mass and Mass-Mass Problems is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.E: Quantities in Chemical Reactions (Exercise)

Additional Exercises

1. If the average male has a body mass of 70 kg, of which 60% is water, how many moles of water are in an average male?
2. If the average female is 60.0 kg and contains 0.00174% iron, how many moles of iron are in an average female?
3. How many moles of each element are present in 2.67 mol of each compound?
    a. HCl
    b. H_2SO_4
    c. $Al(NO_3)_3$
    d. $Ga_2(SO_4)_3$
4. How many moles of each element are present in 0.00445 mol of each compound?
    a. HCl
    b. H_2SO_4
    c. $Al_2(CO_3)_3$
    d. $Ga_2(SO_4)_3$
5. What is the mass of one hydrogen atom in grams? What is the mass of one oxygen atom in grams? Do these masses have a 1:16 ratio, as expected?
6. What is the mass of one sodium atom in grams?
7. If 6.63×10^{-6} mol of a compound has a mass of 2.151 mg, what is the molar mass of the compound?
8. Hemoglobin (molar mass is approximately 64,000 g/mol) is the major component of red blood cells that transports oxygen and carbon dioxide in the body. How many moles are in 0.034 g of hemoglobin?

Answers

1. 2,330 mol
3. a. 2.67 mol of H and 2.67 mol of Cl
    b. 5.34 mol of H, 2.67 mol of S, and 10.68 mol of O
    c. 2.67 mol of Al, 8.01 mol of N, and 24.03 mol of O
    d. 5.34 mol of Ga, 8.01 mol of S, and 32.04 mol of O
5. H = 1.66×10^{-24} g and O = 2.66×10^{-23} g; yes, they are in a 1:16 ratio.
7. 324 g/mol

This page titled 12.E: Quantities in Chemical Reactions (Exercise) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

12.S: Quantities in Chemical Reactions (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

Chemical reactions relate quantities of reactants and products. Chemists use the **mole** unit to represent 6.022×10^{23} things, whether the things are atoms of elements or molecules of compounds. This number, called **Avogadro's number**, is important because this number of atoms or molecules has the same mass in grams as one atom or molecule has in atomic mass units. **Molar masses** of substances can be determined by summing the appropriate masses from the periodic table; the final molar mass will have units of grams.

Because one mole of a substance will have a certain mass, we can use that relationship to construct conversion factors that will convert a mole amount into a mass amount, or vice versa. Such **mole-mass conversions** typically take one algebraic step.

Chemical reactions list reactants and products in molar amounts, not just molecular amounts. We can use the coefficients of a balanced chemical equation to relate moles of one substance in the reaction to moles of other substances (stoichiometry). Chemical reactions obey the Law of Conservation of Mass. To balance a chemical reaction, the coefficients in front of each compound can be adjusted until the total number of atoms of each elements is equal on both sides of the reaction arrow.

Collision Theory can be used to described the energetic aspects of chemical reactions. The reactants and products of a chemical reaction store potential energy in the form of chemical bonds and intermolecular forces. The energy difference in the bond energies of the reactants and products is called "Enthalpy". When the products are lower in potential energy than the reactants, then this excess energy is released as heat and the reaction is described as exothermic. When the products are higher in potential energy than the reactants, then energy must be added to the reaction as heat for it to occur and the reaction is described as endothermic.

The rate of a chemical reaction is influenced by its activation energy. The larger the activation energy, the slower the reaction rate. Catalysts can be added to reactions to lower the activation energy and increase the reaction rate.

Not all reactions go to completion. For some reactions, both the forward and reverse reaction can occur simultaneously. When the forward and reverse reaction rates are equal, the reaction is described as being "at equilibrium". Another way to recognize when a system is at equilibrium is when the concentration of reactants and products remain constant. Le Chatlier's principle states that the direction of an equilibrium reaction will shift to reduce the stress.

This page titled 12.S: Quantities in Chemical Reactions (Summary) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13: SOLIDS, LIQUIDS, AND GASES

Most of us are familiar with the three phases of matter: solid, liquid, and gas. Indeed, we addressed the energy changes involved in phase changes. The substance we are probably most familiar with has those three phases: water. In everyday life, we commonly come in contact with water as a solid (ice), as a liquid, and as a gas (steam). All we have to do is change the conditions of the substance—typically temperature—and we can change the phase from solid to liquid to gas and back again. Under the proper conditions of temperature and pressure, many substances—not only water—can experience the three different phases. An understanding of the phases of matter is important for our understanding of all matter. In this chapter, we will explore the three phases of matter.

13.1: Prelude to Solids, Liquids, and Gases

13.2: Intermolecular Interactions

13.3: Solids and Liquids

13.4: Gases and Pressure

13.5: Gas Laws

13.E: Solids, Liquids, and Gases (Exercises)

13.S: Solids, Liquids, and Gases (Summary)

This page titled 13: Solids, Liquids, and Gases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.1: Prelude to Solids, Liquids, and Gases

We normally experience carbon dioxide (CO_2) as a gas, but if it were cooled down to about –78°C, it would become a solid. The everyday term for solid carbon dioxide is dry ice.

Why “dry” ice? Solid carbon dioxide is called dry ice because it converts from a solid to a gas directly, without going through the liquid phase, in a process called sublimation. Thus, there is no messy liquid phase to worry about. Although it is a novelty, dry ice has some potential dangers. Because it is so cold, it can freeze living tissues very quickly, so people handling dry ice should wear special protective gloves. The cold carbon dioxide gas is also heavier than air (because it is cold and more dense), so people in the presence of dry ice should be in a well-ventilated area.



Figure 13.1.1: Dry ice. Image used with permission (CC BY SA 3.0 unported; ProjectManhattan).

Dry ice has several common uses. Because it is so cold, it is used as a refrigerant to keep other things cold or frozen (e.g., meats or ice cream). In the medical field, dry ice is used to preserve medical specimens, blood products, and drugs. It also has dermatological applications (e.g., freezing off warts). Organs for transplant are kept cool with dry ice until the recipient of the new organ is ready for surgery. In this respect, carbon dioxide is much like water—more than one phase of the same substance has significant uses in the real world.

This page titled 13.1: Prelude to Solids, Liquids, and Gases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.2: Intermolecular Interactions

Learning Objectives

- Define *phase*.
- Identify the types of interactions between molecules.

A phase is a certain form of matter that includes a specific set of physical properties. That is, the atoms, the molecules, or the ions that make up the phase do so in a consistent manner throughout the phase. Science recognizes three stable phases: the *solid phase*, in which individual particles can be thought of as in contact and held in place; the *liquid phase*, in which individual particles are in contact but moving with respect to each other; and the *gas phase*, in which individual particles are separated from each other by relatively large distances. Not all substances will readily exhibit all phases. For example, carbon dioxide does not exhibit a liquid phase unless the pressure is greater than about six times normal atmospheric pressure. Other substances, especially complex organic molecules, may decompose at higher temperatures, rather than becoming a liquid or a gas.

> For many substances, there are different arrangements the particles can take in the solid phase, depending on temperature and pressure.

Which phase a substance adopts depends on the pressure and the temperature it experiences. Of these two conditions, temperature variations are more obviously related to the phase of a substance. When it is very cold, H_2O exists in the solid form as ice. When it is warmer, the liquid phase of H_2O is present. At even higher temperatures, H_2O boils and becomes steam.

Pressure changes can also affect the presence of a particular phase (as we indicated for carbon dioxide), but its effects are less obvious most of the time. We will mostly focus on the temperature effects on phases, mentioning pressure effects only when they are important. Most chemical substances follow the same pattern of phases when going from a low temperature to a high temperature: the solid phase, then the liquid phase, and then the gas phase. However, the temperatures at which these phases are present differ for all substances and can be rather extreme. Table 13.2.1 shows the temperature ranges for solid, liquid, and gas phases for three substances. As you can see, there is extreme variability in the temperature ranges.

Table 13.2.1: Temperature Ranges for the Three Phases of Various Substances

Substance	Solid Phase Below	Liquid Phase Above	Gas Phase Above
hydrogen (H_2)	−259°C	−259°C	−253°C
water (H_2O)	0°C	0°C	100°C
sodium chloride (NaCl)	801°C	801°C	1413°C
The *melting point* of a substance is the temperature that separates a solid and a liquid. The *boiling point* of a substance is the temperature that separates a liquid and a gas.			

What accounts for this variability? Why do some substances become liquids at very low temperatures, while others require very high temperatures before they become liquids? It all depends on the strength of the intermolecular interactions between the particles of substances. (Although ionic compounds are not composed of discrete molecules, we will still use the term *intermolecular* to include interactions between the ions in such compounds.) Substances that experience strong intermolecular interactions require higher temperatures to become liquids and, finally, gases. Substances that experience weak intermolecular interactions do not need much energy (as measured by temperature) to become liquids and gases and will exhibit these phases at lower temperatures.

Substances with the highest melting and boiling points have covalent network bonding. This type of intermolecular interaction is actually a covalent bond. In these substances, all the atoms in a sample are covalently bonded to other atoms; in effect, the entire sample is essentially one large molecule. Many of these substances are solid over a large temperature range because it takes a lot of energy to disrupt all the covalent bonds at once. One example of a substance that shows covalent network bonding is diamond (Figure 13.2.1), which is a form of pure carbon. At temperatures over 3,500°C, diamond finally vaporizes into gas-phase

atoms.

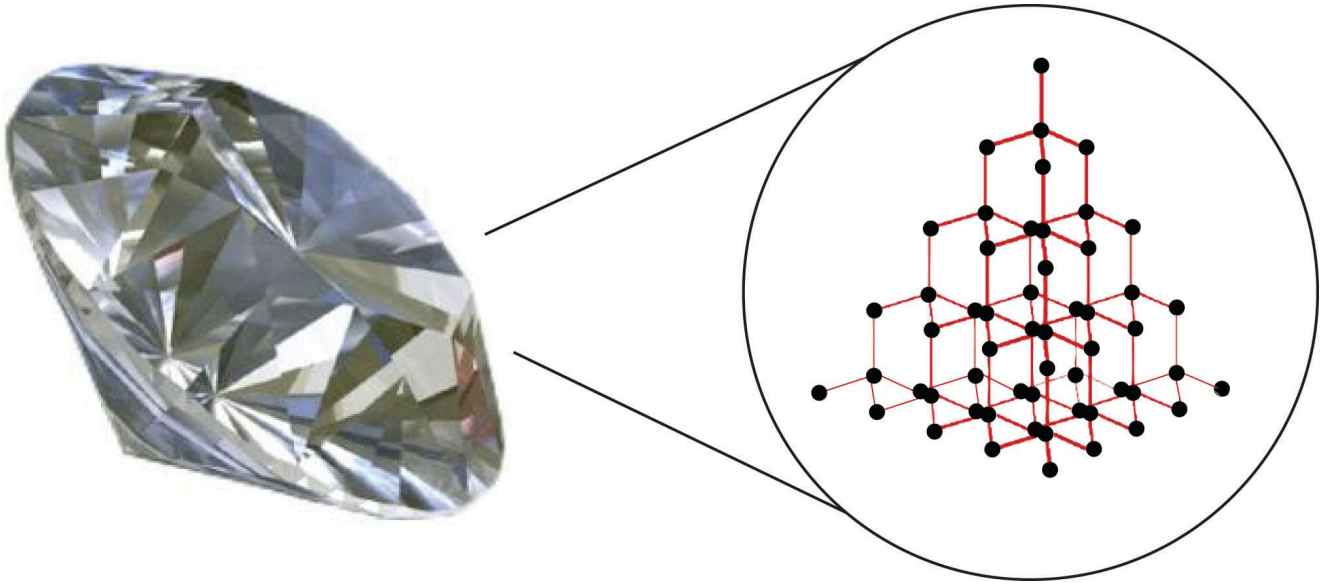


Figure 13.2.1: Diamond. Diamond, a form of pure carbon, has covalent network bonding. It takes a very high temperature—over 3,500°C—for diamond to leave the solid state. Source: Photo © Thinkstock

The strongest force between any two particles is the ionic bond, in which two ions of opposing charge are attracted to each other. Thus, ionic interactions between particles are another type of intermolecular interaction. Substances that contain ionic interactions are relatively strongly held together, so these substances typically have high melting and boiling points. Sodium chloride (Figure 13.2.2) is an example of a substance whose particles experience ionic interactions (Table 13.2.1).

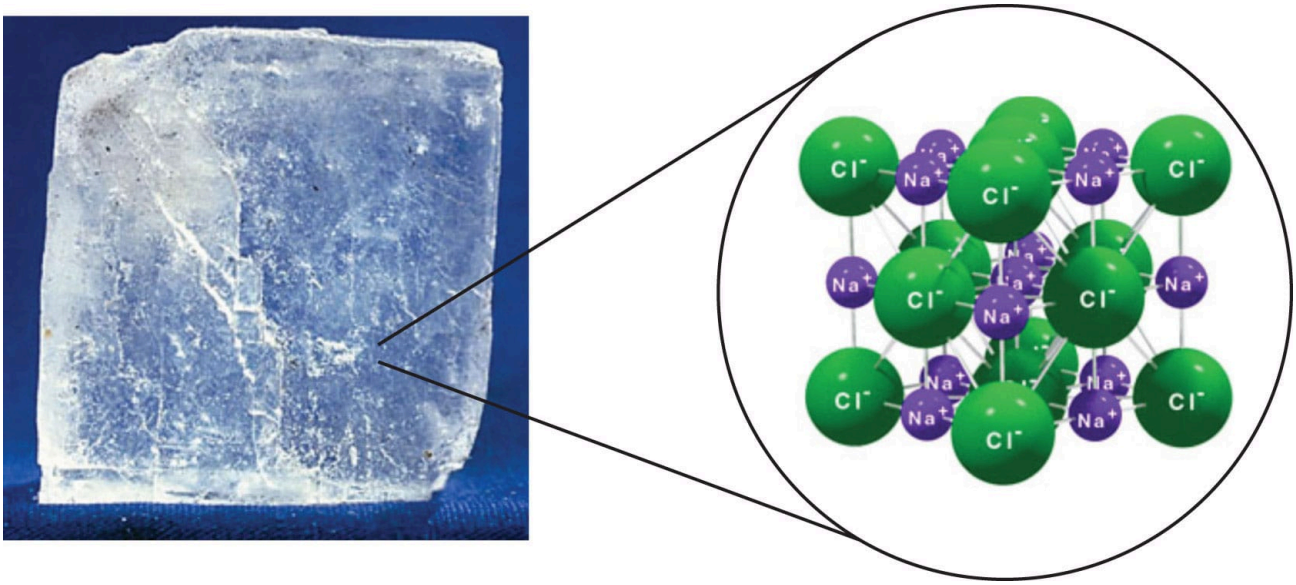


Figure 13.2.2: Sodium Chloride. Solid NaCl is held together by ionic interactions. Source: Photo © Thinkstock

Many substances that experience covalent bonding exist as discrete molecules. In many molecules, the electrons that are shared in a covalent bond are not shared equally between the two atoms in the bond. Typically, one of the atoms attracts the electrons more strongly than the other, leading to an unequal sharing of electrons in the bond. This idea is illustrated in Figure 13.2.3, which shows a diagram of the covalent bond in hydrogen fluoride (HF). The fluorine atom attracts the electrons in the bond more than the hydrogen atom does. The result is an unequal distribution of electrons in the bond, favoring the fluorine side of the covalent bond. Because of this unequal distribution, the fluorine side of the covalent bond actually takes on a partial negative charge (indicated by the δ– in Figure 13.2.3), while the hydrogen side of the bond, being electron deficient, takes on a partial positive charge (indicated by the δ+ in Figure 13.2.3). A covalent bond that has an unequal sharing of electrons is called a polar covalent bond. (A covalent bond that has an equal sharing of electrons, as in a covalent bond with the same atom on each side, is called a

nonpolar covalent bond.) A molecule with a net unequal distribution of electrons in its covalent bonds is a polar molecule. HF is an example of a polar molecule.

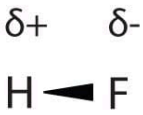


Figure 13.2.3: Polar Covalent Bonds. The electrons in the HF molecule are not equally shared by the two atoms in the bond. Because the fluorine atom has nine protons in its nucleus, it attracts the negatively charged electrons in the bond more than the hydrogen atom does with its one proton in its nucleus. Thus, electrons are more strongly attracted to the fluorine atom, leading to an imbalance in the electron distribution between the atoms. The fluorine side of the bond picks up a partial overall negative charge (represented by the δ– in the diagram), while the hydrogen side of the bond has an overall partial positive charge (represented by the δ+ in the diagram). Such a bond is called a *polar covalent bond*.

The charge separation in a polar covalent bond is not as extreme as is found in ionic compounds, but there is a related result: oppositely charged ends of different molecules will attract each other. This type of intermolecular interaction is called a dipole-dipole interaction. Many molecules with polar covalent bonds experience dipole-dipole interactions. The covalent bonds in some molecules are oriented in space in such a way that the bonds in the molecules cancel each other out. The individual bonds are polar, but the overall molecule is not polar; rather, the molecule is *nonpolar*. Such molecules experience little or no dipole-dipole interactions. Carbon dioxide (CO_2) and carbon tetrachloride (CCl_4) are examples of such molecules (Figure 13.2.4).

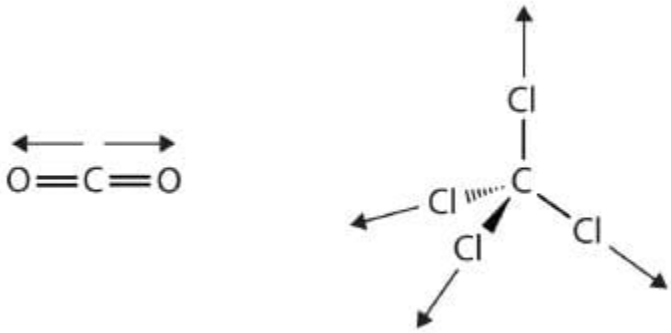


Figure 13.2.4: Nonpolar Molecules. Although the individual bonds in both CO_2 and CCl_4 are polar, their effects cancel out because of the spatial orientation of the bonds in each molecule. As a result, such molecules experience little or no dipole-dipole interaction.

The H–F, O–H, and N–H bonds are strongly polar; in molecules that have these bonds, particularly strong dipole-dipole interactions (as strong as 10% of a true covalent bond) can occur. Because of this strong interaction, hydrogen bonding is used to describe this dipole-dipole interaction. The physical properties of water, which has two O–H bonds, are strongly affected by the presence of hydrogen bonding between water molecules. Figure 13.2.5 shows how molecules experiencing hydrogen bonding can interact.

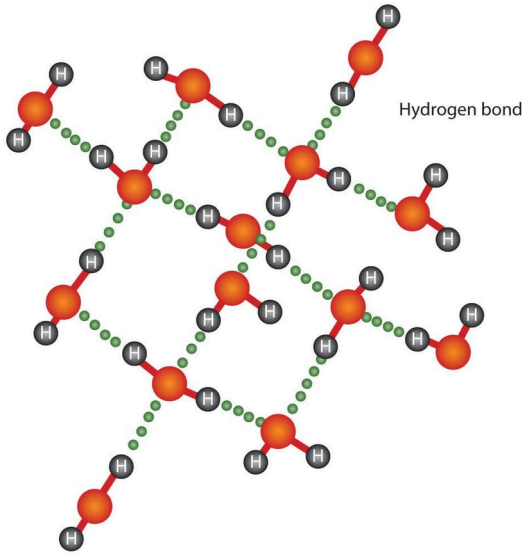


Figure 13.2.5: Hydrogen Bonding between Water Molecules. The presence of hydrogen bonding in molecules like water can have a large impact on the physical properties of a substance.

Finally, there are forces between all molecules that are caused by electrons being in different places in a molecule at any one time, which sets up a temporary separation of charge that disappears almost as soon as it appears. These are very weak intermolecular interactions and are called dispersion forces (or London forces). (An alternate name is London dispersion forces.) Molecules that experience no other type of intermolecular interaction will at least experience dispersion forces. Substances that experience only dispersion forces are typically soft in the solid phase and have relatively low melting points. Because dispersion forces are caused by the instantaneous distribution of electrons in a molecule, larger molecules with a large number of electrons can experience substantial dispersion forces. Examples include *waxes*, which are long hydrocarbon chains that are solids at room temperature because the molecules have so many electrons. The resulting dispersion forces between these molecules make them assume the solid phase at normal temperatures.

The phase that a substance adopts depends on the type and magnitude of the intermolecular interactions the particles of a substance experience. If the intermolecular interactions are relatively strong, then a large amount of energy—in terms of temperature—is necessary for a substance to change phases. If the intermolecular interactions are weak, a low temperature is all that is necessary to move a substance out of the solid phase.

Example 13.2.1: Intermolecular Forces

What intermolecular forces besides dispersion forces, if any, exist in each substance? Are any of these substances solids at room temperature?

a. potassium chloride (KCl)
b. ethanol (C_2H_5OH)
c. bromine (Br_2)

SOLUTION

a. Potassium chloride is composed of ions, so the intermolecular interaction in potassium chloride is ionic forces. Because ionic interactions are strong, it might be expected that potassium chloride is a solid at room temperature.
b. Ethanol has a hydrogen atom attached to an oxygen atom, so it would experience hydrogen bonding. If the hydrogen bonding is strong enough, ethanol might be a solid at room temperature, but it is difficult to know for certain. (Ethanol is actually a liquid at room temperature.)
c. Elemental bromine has two bromine atoms covalently bonded to each other. Because the atoms on either side of the covalent bond are the same, the electrons in the covalent bond are shared equally, and the bond is a nonpolar covalent bond. Thus, diatomic bromine does not have any intermolecular forces other than dispersion forces. It is unlikely to be a solid at room temperature unless the dispersion forces are strong enough. Bromine is a liquid at room temperature.

Exercise 13.2.1

What intermolecular forces besides dispersion forces, if any, exist in each substance? Are any of these substances solids at room temperature?

a. methylamine (CH_3NH_2)
b. calcium sulfate ($CaSO_4$)
c. carbon monoxide (CO)

Concept Review Exercise

1. What types of intermolecular interactions can exist in compounds?

Answer

1. polar and nonpolar covalent bonding, ionic bonding, dispersion forces, dipole-dipole interactions, and hydrogen bonding

Key Takeaways

- A phase is a form of matter that has the same physical properties throughout.
- Molecules interact with each other through various forces: ionic and covalent bonds, dipole-dipole interactions, hydrogen bonding, and dispersion forces.

Exercises

1. List the three common phases in the order you are likely to find them—from lowest temperature to highest temperature.
2. List the three common phases in the order they exist from lowest energy to highest energy.
3. List these intermolecular interactions from weakest to strongest: London forces, hydrogen bonding, and ionic interactions.
4. List these intermolecular interactions from weakest to strongest: covalent network bonding, dipole-dipole interactions, and dispersion forces.
5. What type of intermolecular interaction is predominate in each substance?
   a. water (H_2O)
   b. sodium sulfate (Na_2SO_4)
   c. decane ($C_{10}H_{22}$)
6. What type of intermolecular interaction is predominate in each substance?
   a. diamond (C, crystal)
   b. helium (He)
   c. ammonia (NH_3)
7. Explain how a molecule like carbon dioxide (CO_2) can have polar covalent bonds but be nonpolar overall.
8. Sulfur dioxide (SO_2) has a formula similar to that of carbon dioxide (see Exercise 7) but is a polar molecule overall. What can you conclude about the shape of the SO_2 molecule?
9. What are some of the physical properties of substances that experience covalent network bonding?
10. What are some of the physical properties of substances that experience only dispersion forces?

Answers

1. solid, liquid, and gas

3. London forces, hydrogen bonding, and ionic interactions

5. a. hydrogen bonding
   b. ionic interactions
   c. dispersion forces

7. The two covalent bonds are oriented in such a way that their dipoles cancel out.

9. very hard, high melting point

This page titled 13.2: Intermolecular Interactions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.3: Solids and Liquids

Learning Objectives

- To describe the solid and liquid phases.

Solids and liquids are collectively called *condensed phases* because their particles are in virtual contact. The two states share little else, however.

Solids

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



Figure 13.3.1: Crystalline Arrangement of Quartz crystal cluster. Some large crystals look the way they do because of the regular arrangement of atoms (ions) in their crystal structure. Image used with permission from Wikipedia.

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a crystal. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 13.3.1. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

Liquids

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



Figure 13.3.2: The formation of a spherical droplet of liquid water minimizes the surface area, which is the natural result of surface tension in liquids. Image used with permission from Wikipedia.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either.

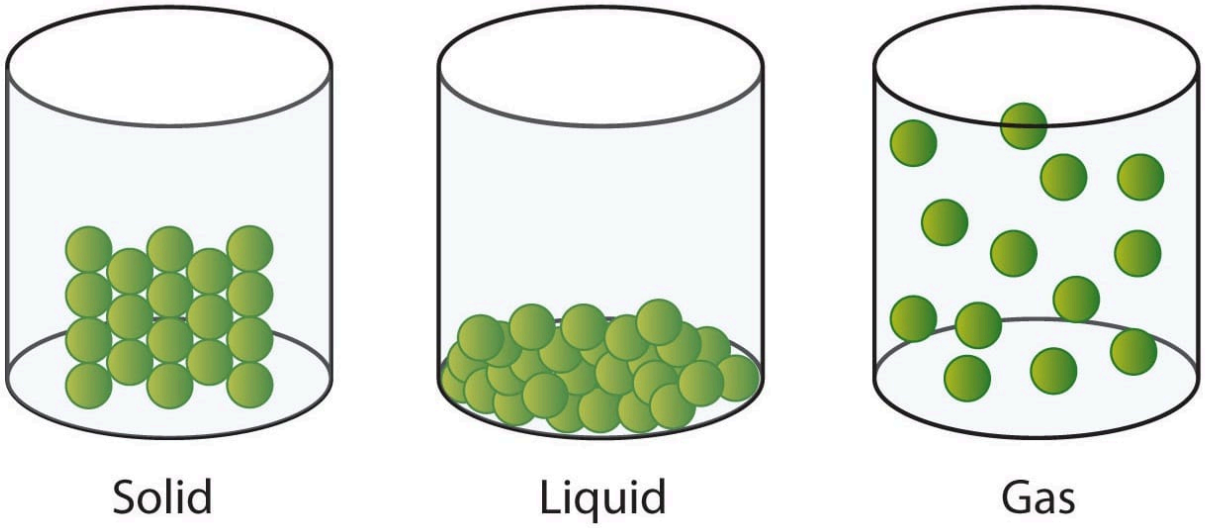


Figure 13.3.3: A Representation of the Solid, Liquid, and Gas States. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Figure 13.3.3 shows the differences among solids, liquids, and gases at the molecular level, while Table 13.3.1 lists the different characteristics of these states.

Table 13.3.1: Characteristics of the Three States of Matter

Characteristic	Solid	Liquid	Gas
shape	definite	indefinite	indefinite

Characteristic	Solid	Liquid	Gas
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

Example 13.3.1

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

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

SOLUTION

a. This statement describes either the liquid state or the solid state.
b. This statement describes either the liquid state or the gas state.
c. This statement describes the liquid state.

Exercise 13.3.1

What state or states of matter does each statement describe?

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

Looking Closer: Water, the Most Important Liquid

Earth is the only known body in our solar system that has liquid water existing freely on its surface. That is a good thing because life on Earth would not be possible without the presence of liquid water.

Water has several properties that make it a unique substance among substances. It is an excellent solvent; it dissolves many other substances and allows those substances to react when in solution. In fact, water is sometimes called the *universal solvent* because of this ability. Water has unusually high melting and boiling points (0°C and 100°C, respectively) for such a small molecule. The boiling points for similar-sized molecules, such as methane (BP = –162°C) and ammonia (BP = –33°C), are more than 100° lower. Though a liquid at normal temperatures, water molecules experience a relatively strong intermolecular interaction that allows them to maintain the liquid phase at higher temperatures than expected.

Unlike most substances, the solid form of water is less dense than its liquid form, which allows ice to float on water. In colder weather, lakes and rivers freeze from the top, allowing animals and plants to continue to live underneath. Water also requires an unusually large amount of energy to change temperature. While 100 J of energy will change the temperature of 1 g of Fe by 230°C, this same amount of energy will change the temperature of 1 g of H_2O by only 100°C. Thus, water changes its temperature slowly as heat is added or removed. This has a major impact on weather, as storm systems like hurricanes can be impacted by the amount of heat that ocean water can store.

Water's influence on the world around us is affected by these properties. Isn't it fascinating that such a small molecule can have such a big impact?

Key Takeaway

- Solids and liquids are phases that have their own unique properties.

Concept Review Exercise

1. How do the strengths of intermolecular interactions in solids and liquids differ?

Answer

1. Solids have stronger intermolecular interactions than liquids do.

Exercises

1. What are the general properties of solids?
2. What are the general properties of liquids
3. What are the general properties of gases?
4. What phase or phases have a definite volume? What phase or phases do not have a definite volume?
5. Name a common substance that forms a crystal in its solid state.
6. Name a common substance that forms an amorphous solid in its solid state.
7. Are substances with strong intermolecular interactions likely to be solids at higher or lower temperatures? Explain.
8. Are substances with weak intermolecular interactions likely to be liquids at higher or lower temperatures? Explain.
9. State two similarities between the solid and liquid states.
10. State two differences between the solid and liquid states.
11. If individual particles are moving around with respect to each other, a substance may be in either the _____ or _____ state but probably not in the _____ state.
12. If individual particles are in contact with each other, a substance may be in either the _____ or _____ state but probably not in the _____ state.

Answers

1. hard, specific volume and shape, high density, cannot be compressed
3. variable volume and shape, low density, compressible
5. sodium chloride (answers will vary)
7. At higher temperatures, their intermolecular interactions are strong enough to hold the particles in place.
9. high density; definite volume
11. liquid; gas; solid

This page titled 13.3: Solids and Liquids is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.4: Gases and Pressure

Learning Objectives

- To describe the gas phase.

The gas phase is unique among the three states of matter in that there are some simple models we can use to predict the physical behavior of all gases—independent of their identities. We cannot do this for the solid and liquid states. In fact, the development of this understanding of the behavior of gases represents the historical dividing point between alchemy and modern chemistry. Initial advances in the understanding of gas behavior were made in the mid 1600s by Robert Boyle, an English scientist who founded the Royal Society (one of the world's oldest scientific organizations).

How is it that we can model all gases independent of their chemical identity? The answer is in a group of statements called the kinetic theory of gases:

- Gases are composed of tiny particles that are separated by large distances.
- Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container.
- The velocity of gas particles is related to the temperature of a gas.
- Gas particles do not experience any force of attraction or repulsion with each other.

Did you notice that none of these statements relates to the identity of the gas? This means that all gases should behave similarly. A gas that follows these statements perfectly is called an *ideal gas*. Most gases show slight deviations from these statements and are called *real gases*. However, the existence of real gases does not diminish the importance of the kinetic theory of gases.

One of the statements of the kinetic theory mentions collisions. As gas particles are constantly moving, they are also constantly colliding with each other and with the walls of their container. There are forces involved as gas particles bounce off the container walls (Figure 13.4.1). The force generated by gas particles divided by the area of the container walls yields pressure. Pressure is a property we can measure for a gas, but we typically do not consider pressure for solids or liquids.

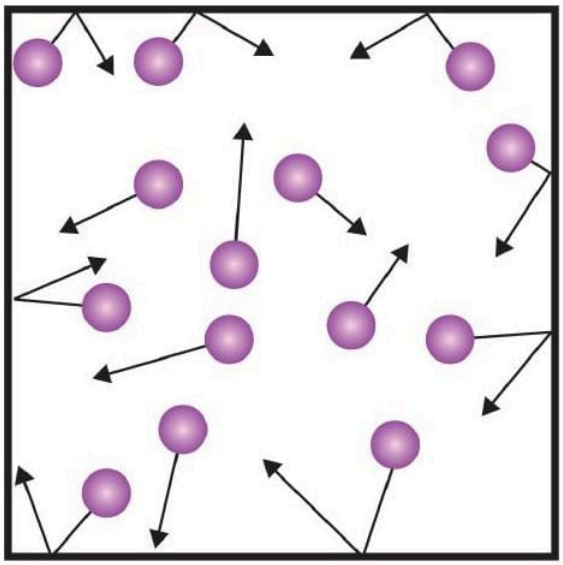


Figure 13.4.1: Gas Pressure. Pressure is what results when gas particles rebound off the walls of their container.

The basic unit of pressure is the newton per square meter (N/m^2). This combined unit is redefined as a pascal (Pa). One pascal is not a very large amount of pressure. A more useful unit of pressure is the bar, which is 100,000 Pa (1 bar = 100,000 Pa). Other common units of pressure are the atmosphere (atm), which was originally defined as the average pressure of Earth's atmosphere at sea level; and mmHg (millimeters of mercury), which is the pressure generated by a column of mercury 1 mm high. The unit millimeters of mercury is also called a torr, named after the Italian scientist Evangelista Torricelli, who invented the barometer in

the mid-1600s. A more precise definition of atmosphere, in terms of torr, is that there are exactly 760 torr in 1 atm. A bar equals 1.01325 atm. Given all the relationships between these pressure units, the ability to convert from one pressure unit to another is a useful skill.

Example 13.4.1: Converting Pressures

Write a conversion factor to determine how many atmospheres are in 1,547 mmHg.

SOLUTION

Because 1 mmHg equals 1 torr, the given pressure is also equal to 1,547 torr. Because there are 760 torr in 1 atm, we can use this conversion factor to do the mathematical conversion:

$$1,547 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 2.04 \text{ atm}$$

Note how the torr units cancel algebraically.

Exercise 13.4.1: Converting Pressures

Write a conversion factor to determine how many millimeters of mercury are in 9.65 atm.

The kinetic theory also states that there is no interaction between individual gas particles. Although we know that there are, in fact, intermolecular interactions in real gases, the kinetic theory assumes that gas particles are so far apart that the individual particles don't "feel" each other. Thus, we can treat gas particles as tiny bits of matter whose identity isn't important to certain physical properties.

Concept Review Exercise

1. What is pressure, and what units do we use to express it?

Answer

1. Pressure is the force per unit area; its units can be pascals, torr, millimeters of mercury, or atmospheres.

Key Takeaway

- The gas phase has certain general properties characteristic of that phase.

Exercises

1. What is the kinetic theory of gases?
2. According to the kinetic theory of gases, the individual gas particles are (always, frequently, never) moving.
3. Why does a gas exert pressure?
4. Why does the kinetic theory of gases allow us to presume that all gases will show similar behavior?
5. Arrange the following pressure quantities in order from smallest to largest: 1 mmHg, 1 Pa, and 1 atm.
6. Which unit of pressure is larger—the torr or the atmosphere?

7. How many torr are there in 1.56 atm?
8. Convert 760 torr into pascals.
9. Blood pressures are expressed in millimeters of mercury. What would be the blood pressure in atmospheres if a patient's systolic blood pressure is 120 mmHg and the diastolic blood pressure is 82 mmHg? (In medicine, such a blood pressure would be reported as "120/82," spoken as "one hundred twenty over eighty-two.")
10. In weather forecasting, barometric pressure is expressed in inches of mercury (in. Hg), where there are exactly 25.4 mmHg in every 1 in. Hg. What is the barometric pressure in millimeters of mercury if the barometric pressure is reported as 30.21 in. Hg?

Answers

1. Gases are composed of tiny particles that are separated by large distances. Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container. The velocity of gas particles is related to the temperature of a gas. Gas particles do not experience any force of attraction or repulsion with each other.
3. A gas exerts pressure as its particles rebound off the walls of its container.
5. 1 Pa, 1 mmHg, and 1 atm
7. 1,190 torr
9. 0.158 atm; 0.108 atm

This page titled 13.4: Gases and Pressure is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.5: Gas Laws

Learning Objectives

- To predict the properties of gases using the gas laws.

Experience has shown that several properties of a gas can be related to each other under certain conditions. The properties are pressure (*P*), volume (*V*), temperature (*T*, in kelvins), and amount of material expressed in moles (*n*). What we find is that a sample of gas cannot have any random values for these properties. Instead, only certain values, dictated by some simple mathematical relationships, will occur.

Boyle's Law

The first simple relationship, referred to as a gas law, is between the pressure of a gas and its volume. If the amount of gas in a sample and its temperature are kept constant, then as the pressure of a gas is increased, the volume of the gas decreases proportionately. Mathematically, this is written as

$$\mathrm{P} \propto \frac{1}{\mathrm{V}} \tag{13.5.1}$$

where the "$\propto$" symbol means "is proportional to." This is one form of Boyle's law, which relates the pressure of a gas to its volume.

A more useful form of Boyle's law involves a change in conditions of a gas. For a given amount of gas at a constant temperature, if we know the initial pressure and volume of a gas sample and the pressure or volume changes, we can calculate what the new volume or pressure will be. That form of Boyle's law is written

$$P_i V_i = P_f V_f \tag{13.5.2}$$

where the subscript i refers to initial conditions and the subscript f refers to final conditions.

To use 13.5.2, you need to know any three of the variables so that you can algebraically calculate the fourth variable. Also, the pressure quantities must have the same units, as must the two volume quantities. If the two similar variables don't have the same variables, one value must be converted to the other value's unit.

Example 13.5.1: Increasing Pressure in a Gas

What happens to the volume of a gas if its pressure is increased? Assume all other conditions remain the same.

SOLUTION

If the pressure of a gas is increased, the volume decreases in response.

Exercise 13.5.1: Increasing Volume in a Gas

What happens to the pressure of a gas if its volume is increased? Assume all other conditions remain the same.

Example 13.5.2: Gas Compression

If a sample of gas has an initial pressure of 1.56 atm and an initial volume of 7.02 L, what is the final volume if the pressure is reduced to 0.987 atm? Assume that the amount and the temperature of the gas remain constant.

SOLUTION

The key in problems like this is to be able to identify which quantities represent which variables from the relevant equation. The way the question is worded, you should be able to tell that 1.56 atm is P_i, 7.02 L is V_i, and 0.987 atm is P_f. What we are looking for is the final volume—V_f. Therefore, substituting these values into $P_iV_i = P_fV_f$:

$$(1.56 \text{ atm})(7.02 \text{ L}) = (0.987 \text{ atm}) \times V_f$$

The expression has atmospheres on both sides of the equation, so they cancel algebraically:

$$(1.56)(7.02 \text{ L}) = (0.987) \times V_f$$

Now we divide both sides of the expression by 0.987 to isolate V_f, the quantity we are seeking:

$$\frac{(1.56)(7.02 \text{ L})}{0.987} = \text{V}_\text{f}$$

Performing the multiplication and division, we get the value of V_f, which is 11.1 L. The volume increases. This should make sense because the pressure decreases, so pressure and volume are inversely related.

Exercise 13.5.2

If a sample of gas has an initial pressure of 3.66 atm and an initial volume of 11.8 L, what is the final pressure if the volume is reduced to 5.09 L? Assume that the amount and the temperature of the gas remain constant.

If the units of similar quantities are not the same, one of them must be converted to the other quantity's units for the calculation to work out properly. It does not matter which quantity is converted to a different unit; the only thing that matters is that the conversion and subsequent algebra are performed properly. The following example illustrates this process.

Example 13.5.3

If a sample of gas has an initial pressure of 1.56 atm and an initial volume of 7.02 L, what is the final volume if the pressure is changed to 1,775 torr? Does the answer make sense? Assume that the amount and the temperature of the gas remain constant.

SOLUTION

This example is similar to Example 13.5.2, except now the final pressure is expressed in torr. For the math to work out properly, one of the pressure values must be converted to the other unit. Let us change the initial pressure to torr:

$$1.56 \text{ atm} \times \frac{760 \text{ torr}}{1 \text{ atm}} = 1,190 \text{ torr}$$

Now we can use Boyle's law:

$$(1,190 \text{ torr})(7.02 \text{ L}) = (1,775 \text{ torr}) \times V_f$$

Torr cancels algebraically from both sides of the equation, leaving

$$(1,190)(7.02 \text{ L}) = (1,775) \times V_f$$

Now we divide both sides of the equation by 1,775 to isolate V_f on one side. Solving for the final volume,

$$\text{V}_\text{f} = \frac{(1,190)(7.02 \text{ L})}{1,775} = 4.71 \text{ L}$$

Because the pressure increases, it makes sense that the volume decreases.

The answer for the final volume is essentially the same if we converted the 1,775 torr to atmospheres: $1,775\ \text{torr} \times \dfrac{1\ \text{atm}}{760\ \text{torr}} = 2.336\ \text{atm}$. Using Boyle's law: (1.56 atm)(7.02 L) = (2.335 atm) × V_f; $\mathrm{V_f} = \dfrac{(1.56\ \text{atm})(7.02\ \text{L})}{2.336\ \text{atm}} = 4.69\ \text{L}$.

Exercise $13.5.3$

If a sample of gas has an initial pressure of 375 torr and an initial volume of 7.02 L, what is the final pressure if the volume is changed to 4,577 mL? Does the answer make sense? Assume that amount and the temperature of the gas remain constant.

To Your Health: Breathing

Breathing certainly is a major contribution to your health! Without breathing, we could not survive. Curiously, the act of breathing itself is little more than an application of Boyle's law.

The lungs are a series of ever-narrowing tubes that end in a myriad of tiny sacs called alveoli. It is in the alveoli that oxygen from the air transfers to the bloodstream and carbon dioxide from the bloodstream transfers to the lungs for exhalation. For air to move in and out of the lungs, the pressure inside the lungs must change, forcing the lungs to change volume—just as predicted by Boyle's law.

The pressure change is caused by the diaphragm, a muscle that covers the bottom of the lungs. When the diaphragm moves down, it expands the size of our lungs. When this happens, the air pressure inside our lungs decreases slightly. This causes new air to rush in, and we inhale. The pressure decrease is slight—only 3 torr, or about 0.4% of an atmosphere. We inhale only 0.5–1.0 L of air per normal breath.

Exhaling air requires that we relax the diaphragm, which pushes against the lungs and slightly decreases the volume of the lungs. This slightly increases the pressure of the air in the lungs, and air is forced out; we exhale. Only 1–2 torr of extra pressure is needed to exhale. So with every breath, our own bodies are performing an experimental test of Boyle's law.

Charles's Law

Another simple gas law relates the volume of a gas to its temperature. Experiments indicate that as the temperature of a gas sample is increased, its volume increases as long as the pressure and the amount of gas remain constant. The way to write this mathematically is

$$\mathrm{V} \propto T \tag{13.5.3}$$

At this point, the concept of temperature must be clarified. Although the Kelvin scale is the preferred temperature scale, the Celsius scale is also a common temperature scale used in science. The Celsius scale is based on the melting and boiling points of water and is actually the common temperature scale used by most countries around the world (except for the United States, which still uses the Fahrenheit scale). The value of a Celsius temperature is directly related to its Kelvin value by a simple expression:

Kelvin temperature = Celsius temperature + 273

Thus, it is easy to convert from one temperature scale to another.

The Kelvin scale is sometimes referred to as the absolute scale because the zero point on the Kelvin scale is at absolute zero, the coldest possible temperature. On the other temperature scales, absolute zero is –260°C or –459°F.

The expression relating a gas volume to its temperature begs the following question: to which temperature scale is the volume of a gas related? The answer is that gas volumes are directly related to the *Kelvin temperature*. Therefore, the temperature of a gas sample should always be expressed in (or converted to) a Kelvin temperature.

Example 13.5.4: Increasing Temperature

What happens to the volume of a gas if its temperature is decreased? Assume that all other conditions remain constant.

SOLUTION

If the temperature of a gas sample is decreased, the volume decreases as well.

Exercise 13.5.4

What happens to the temperature of a gas if its volume is increased? Assume that all other conditions remain constant.

As with Boyle's law, the relationship between volume and temperature can be expressed in terms of initial and final values of volume and temperature, as follows:

$$\frac{\mathrm{V_i}}{\mathrm{T_i}} = \frac{\mathrm{V_f}}{\mathrm{T_f}} \tag{13.5.4}$$

where V_i and T_i are the initial volume and temperature, and V_f and T_f are the final volume and temperature. This is Charles's law. The restriction on its use is that the pressure of the gas and the amount of gas must remain constant. (Charles's law is sometimes referred to as Gay-Lussac's law, after the scientist who promoted Charles's work.)

Example 13.5.5

A gas sample at 20°C has an initial volume of 20.0 L. What is its volume if the temperature is changed to 60°C? Does the answer make sense? Assume that the pressure and the amount of the gas remain constant.

SOLUTION

Although the temperatures are given in degrees Celsius, we must convert them to the kelvins before we can use Charles's law. Thus,

20°C + 273 = 293 K = T_i 60°C + 273 = 333 K = T_f

Now we can substitute these values into Charles's law, along with the initial volume of 20.0 L:

$$\frac{20.0\ \mathrm{L}}{293\ \mathrm{K}} = \frac{\mathrm{V_f}}{333\ \mathrm{K}}$$

Multiplying the 333 K to the other side of the equation, we see that our temperature units will cancel:

$$\frac{(333\ \mathrm{K})(20.0\ \mathrm{L})}{293\ \mathrm{K}} = \mathrm{V_f}$$

Solving for the final volume, V_f = 22.7 L. So, as the temperature is increased, the volume increases. This makes sense because volume is directly proportional to the absolute temperature (as long as the pressure and the amount of the remain constant).

Exercise 13.5.5

A gas sample at 35°C has an initial volume of 5.06 L. What is its volume if the temperature is changed to −35°C? Does the answer make sense? Assume that the pressure and the amount of the gas remain constant.

Combined Gas Law

Other gas laws can be constructed, but we will focus on only two more. The combined gas law brings Boyle's and Charles's laws together to relate pressure, volume, and temperature changes of a gas sample:

Double subscripts: use braces to clarify

To apply this gas law, the amount of gas should remain constant. As with the other gas laws, the temperature must be expressed in kelvins, and the units on the similar quantities should be the same. Because of the dependence on three quantities at the same time, it is difficult to tell in advance what will happen to one property of a gas sample as two other properties change. The best way to know is to work it out mathematically.

Example 13.5.6

A sample of gas has P_i = 1.50 atm, V_i = 10.5 L, and T_i = 300 K. What is the final volume if P_f = 0.750 atm and T_f = 350 K?

SOLUTION

Using the combined gas law, substitute for five of the quantities:

$$\frac{(1.50\ \text{atm})(10.5\ \text{L})}{300\ \text{K}} = \frac{(0.750\ \text{atm})(\text{V}_\text{f})}{350\ \text{K}}$$

We algebraically rearrange this expression to isolate V_f on one side of the equation:

$$\text{V}_\text{f} = \frac{(1.50\ \text{atm})(10.5\ \text{L})(350\ \text{K})}{(300\ \text{K})(0.750\ \text{atm})} = 24.5\ \text{L}$$

Note how all the units cancel except the unit for volume.

Exercise 13.5.6

A sample of gas has P_i = 0.768 atm, V_i = 10.5 L, and T_i = 300 K. What is the final pressure if V_f = 7.85 L and T_f = 250 K?

Example 13.5.7

A balloon containing a sample of gas has a temperature of 22°C and a pressure of 1.09 atm in an airport in Cleveland. The balloon has a volume of 1,070 mL. The balloon is transported by plane to Denver, where the temperature is 11°C and the pressure is 655 torr. What is the new volume of the balloon?

SOLUTION

The first task is to convert all quantities to the proper and consistent units. The temperatures must be expressed in kelvins, and the pressure units are different so one of the quantities must be converted. Let us convert the atmospheres to torr:

22°C + 273 = 295 K = T_i

11°C + 273 = 284 K = T_f

$$1.09\ \text{atm} \times \frac{760\ \text{torr}}{1\ \text{atm}} = 828\ \text{torr} = \text{P}_\text{i}$$

Now we can substitute the quantities into the combined has law:

$$\frac{(828\ \text{torr})(1,070\ \text{mL})}{295\ \text{K}} = \frac{(655\ \text{torr}) \times \text{V}_\text{f}}{284\ \text{K}}$$

To solve for V_f, we multiply the 284 K in the denominator of the right side into the numerator on the left, and we divide 655 torr in the numerator of the right side into the denominator on the left:

$$\frac{(828\ \text{torr})(1,070\ \text{mL})(284\ \text{K})}{(295\ \text{K})(655\ \text{torr})} = \text{V}_\text{f}$$

Notice that torr and kelvins cancel, as they are found in both the numerator and denominator. The only unit that remains is milliliters, which is a unit of volume. So V_f = 1,300 mL. The overall change is that the volume of the balloon has increased by 230 mL.

Exercise 13.5.7

A balloon used to lift weather instruments into the atmosphere contains gas having a volume of 1,150 L on the ground, where the pressure is 0.977 atm and the temperature is 18°C. Aloft, this gas has a pressure of 6.88 torr and a temperature of –15°C. What is the new volume of the gas?

The Ideal Gas Law

So far, the gas laws we have used have focused on changing one or more properties of the gas, such as its volume, pressure, or temperature. There is one gas law that relates all the independent properties of a gas under any particular condition, rather than a change in conditions. This gas law is called the ideal gas law. The formula of this law is as follows:

$$PV = nRT \tag{13.5.5}$$

In this equation, *P* is pressure, *V* is volume, *n* is amount of moles, and *T* is temperature. *R* is called the ideal gas law constant and is a proportionality constant that relates the values of pressure, volume, amount, and temperature of a gas sample. The variables in this equation do not have the subscripts *i* and *f* to indicate an initial condition and a final condition. The ideal gas law relates the four independent properties of a gas under *any* conditions.

The value of *R* depends on what units are used to express the other quantities. If volume is expressed in liters and pressure in atmospheres, then the proper value of *R* is as follows:

$$\text{R} = 0.08205\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \tag{13.5.6}$$

This may seem like a strange unit, but that is what is required for the units to work out algebraically.

Example 13.5.8

What is the volume in liters of 1.45 mol of N_2 gas at 298 K and 3.995 atm?

SOLUTION

Using the ideal gas law where *P* = 3.995 atm, *n* = 1.45, and *T* = 298,

$$(3.995\ \text{atm}) \times \text{V} = (1.45\ \text{mol}) \left(0.08205\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298\ \text{K})$$

On the right side, the moles and kelvins cancel. Also, because atmospheres appear in the numerator on both sides of the equation, they also cancel. The only remaining unit is liters, a unit of volume. So

$$3.995 \times V = (1.45)(0.08205)(298)\ \text{L}$$

Dividing both sides of the equation by 3.995 and evaluating, we get *V* = 8.87 L. Note that the conditions of the gas are not changing. Rather, the ideal gas law allows us to determine what the fourth property of a gas (here, volume) *must* be if three

other properties (here, amount, pressure, and temperature) are known.

Exercise 13.5.8

What is the pressure of a sample of CO_2 gas if 0.557 mol is held in a 20.0 L container at 451 K?

For convenience, scientists have selected 273 K (0°C) and 1.00 atm pressure as a set of standard conditions for gases. This combination of conditions is called standard temperature and pressure (STP). Under these conditions, 1 mol of any gas has about the same volume. We can use the ideal gas law to determine the volume of 1 mol of gas at STP:

$$(1.00\,\text{atm}) \times \text{V} = (1.00\,\text{mol}) \left(0.08205\,\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (273\,\text{K}) \tag{13.5.7}$$

This volume is 22.4 L. Because this volume is independent of the identity of a gas, the idea that 1 mol of gas has a volume of 22.4 L at STP makes a convenient conversion factor:

1 mol gas = 22.4 L (at STP)

Example 13.5.9

Cyclopropane (C_3H_6) is a gas that formerly was used as an anesthetic. How many moles of gas are there in a 100.0 L sample if the gas is at STP?

SOLUTION

We can set up a simple, one-step conversion that relates moles and liters:

$$100.0\,\text{L}\,\text{C}_3\text{H}_6 \times \frac{1\,\text{mol}}{22.4\,\text{L}} = 4.46\,\text{mol}\,\text{C}_3\text{H}_6$$

There are almost 4.5 mol of gas in 100.0 L.

Note: Because of its flammability, cyclopropane is no longer used as an anesthetic gas.

Exercise 13.5.9

Freon is a trade name for a series of fluorine- and chlorine-containing gases that formerly were used in refrigeration systems. What volume does 8.75 mol of Freon have at STP?

Note: Many gases known as Freon are no longer used because their presence in the atmosphere destroys the ozone layer, which protects us from ultraviolet light from the sun.

Dalton's Law of Partial Pressures

The ideal gas equation of state applies to mixtures just as to pure gases. It was in fact with a gas mixture, ordinary air, that Boyle, Gay-Lussac and Charles did their early experiments. The only new concept we need in order to deal with gas mixtures is the *partial pressure*, a concept invented by the famous English chemist John Dalton (1766-1844). Dalton reasoned that the low density and high compressibility of gases indicates that they consist mostly of empty space; from this it follows that when two or more different gases occupy the same volume, they behave entirely independently. The contribution that each component of a gaseous mixture makes to the total pressure of the gas is known as the *partial pressure* of that gas.

The definition of Dalton's Law of Partial Pressures that address this is:

The total pressure of a gas is the sum of the partial pressures of its components

which is expressed algebraically as

$$P_{total} = P_1 + P_2 + P_3 ... = \sum_i P_i \tag{13.5.8}$$

or, equivalently

$$P_{total} = \frac{RT}{V} \sum_i n_i \tag{13.5.9}$$

There is also a similar relationship based on *volume fractions*, known as *Amagat's law of partial volumes*. It is exactly analogous to Dalton's law, in that it states that the total volume of a mixture is just the sum of the partial volumes of its components. But there are two important differences: Amagat's law holds only for ideal gases which must all be at the same temperature and pressure. Dalton's law has neither of these restrictions. Although Amagat's law seems intuitively obvious, it sometimes proves useful in chemical engineering applications. We will make no use of it in this course.

Example 13.5.10

Three flasks having different volumes and containing different gases at various pressures are connected by stopcocks as shown. When the stopcocks are opened,

a. What will be the pressure in the system?
b. Which gas will be most abundant in the mixture?

Assume that the temperature is uniform and that the volume of the connecting tubes is negligible.

SOLUTION

The trick here is to note that the total number of moles n_T and the temperature remain unchanged, so we can make use of Boyle's law PV = constant. We will work out the details for CO_2 only, denoted by subscripts *a*.

For CO_2,

$$P_a V_a = (2.13\ atm)(1.50\ L) = 3.19\ L \cdot atm \tag{13.5.10}$$

Adding the PV products for each separate container, we obtain

$$\sum_i P_i V_i = 6.36\ L \cdot atm = n_T RT \tag{13.5.11}$$

We will call this sum P_1V_1. After the stopcocks have been opened and the gases mix, the new conditions are denoted by P_2V_2.

From Boyle's law (13.5.2),

$$P_1 V_1 = P_2 V_2 = 6.36\ L \cdot atm \tag{13.5.12}$$

$$V_2 = \sum_i V_i = 4.50\ L \tag{13.5.13}$$

Solving for the final pressure P_2 we obtain (6.36 L-atm)/(4.50 L) = **1.41 atm**.

For part *(b)*, note that the number of moles of each gas is $n = PV/RT$. The mole fraction if any one gas is $X_i = n_i / n_T$. For CO_2, this works out to (3.19/*RT*) / (6.36/*RT*) = 0.501. Because this exceeds 0.5, we know that this is the most abundant gas in the final mixture.

Dalton's law states that in a gas mixture (P_{total}) each gas will exert a pressure independent of the other gases (P_n) and each gas will behave as if it alone occupies the total volume. By extension, the partial pressure of each gas can be calculated by multiplying the total pressure (P_{total}) by the gas percentage (%).

$$P_{Total} = P_1 + P_2 + P_3 + P_4 + ... + P_n \tag{13.5.14}$$

or

$$P_n = \frac{^n}{P_{Total}} \tag{13.5.15}$$

Table 13.5.1: Partial Pressures for the gases in air on a typical day

Gas	Partial Pressure (mm Hg)	Percentage (%)
Nitrogen, (N_2\)	P_{N_2} = 594	78
Oxygen, O_2	P_{O_2}= 160	21
Carbon Dioxide, CO_2	P_{CO_2} = 0.25	0.033
Water Vapor, H_2O	P_{H_2O} = 5.7	0.75
Other trace gases	P_{Other} = 0.05	0.22
Total air	P_{Total} = 760	1

Henry's Law

Henry's law is one of the gas laws formulated by William Henry in 1803. It states: "At a constant temperature, the amount of a given gas that dissolves in a given type and volume of liquid is directly proportional to the partial pressure of that gas in equilibrium with that liquid." An equivalent way of stating the law is that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.

To explain this law, Henry derived the equation:

$$C = kP_{gas} \tag{13.5.16}$$

where

- C is the solubility of a gas at a fixed temperature in a particular solvent (in units of M or mL gas/L)
- k is Henry's law constant (often in units of M/atm)
- P_{gas} is the partial pressure of the gas (often in units of Atm)

Henry's Law tells us that the greater the pressure of gas above the surface of a liquid, the higher the concentration of the gas in the liquid. Also, Henry's law tells us that gases diffuse from areas of high gas concentration to areas of low gas concentration.

Applicability of Henry's Law

- Henry's law only works if the molecules are at equilibrium.
- Henry's law does not work for gases at high pressures (e.g., $N_{2\ (g)}$ at high pressure becomes very soluble and harmful when in the blood supply).
- Henry's law does not work if there is a chemical reaction between the solute and solvent (e.g., $HCl_{(g)}$ reacts with water by a dissociation reaction to generate H_3O^+ and Cl^- ions).

Collecting Gases over Water

A common laboratory method of collecting the gaseous product of a chemical reaction is to conduct it into an inverted tube or bottle filled with water, the opening of which is immersed in a larger container of water. This arrangement is called a *pneumatic trough*, and was widely used in the early days of chemistry. As the gas enters the bottle it displaces the water and becomes trapped in the upper part.

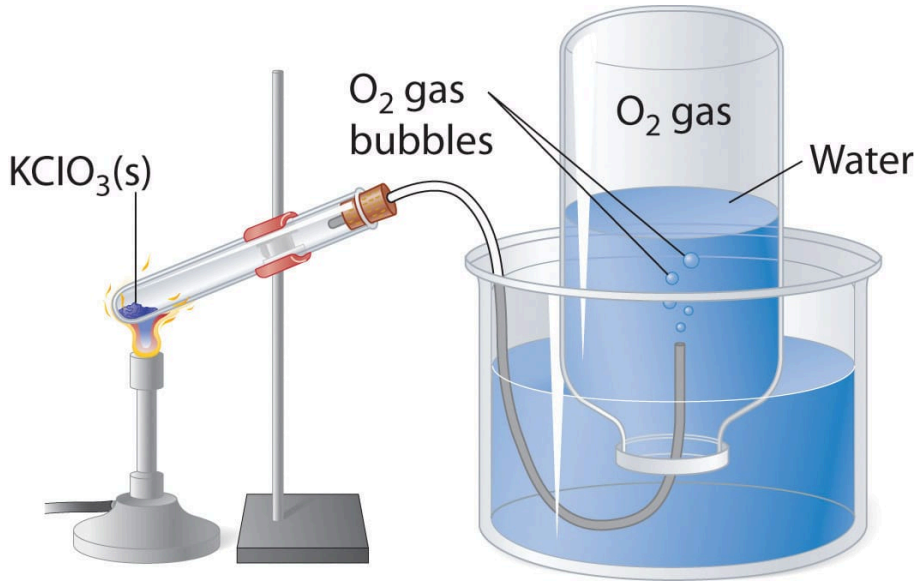


Figure 13.5.1:An Apparatus for Collecting Gases by the Displacement of Water

The volume of the gas can be observed by means of a calibrated scale on the bottle, but what about its pressure? The total pressure confining the gas is just that of the atmosphere transmitting its force through the water. (An exact calculation would also have to take into account the height of the water column in the inverted tube.) But liquid water itself is always in equilibrium with its vapor, so the space in the top of the tube is a mixture of two gases: the gas being collected, and gaseous H_2O. The partial pressure of H_2O is known as the vapor pressure of water and it depends on the temperature. In order to determine the quantity of gas we have collected, we must use Dalton's Law to find the partial pressure of that gas.

Example 13.5.11

Oxygen gas was collected over water as shown above. The atmospheric pressure was 754 torr, the temperature was 22°C, and the volume of the gas was 155 mL. The vapor pressure of water at 22°C is 19.8 torr. Use this information to estimate the number of moles of O_2 produced.

SOLUTION

From Dalton's law,

$$P_{O_2} = P_{total} – P_{H_2O} = 754 – 19.8 = 734\ torr = 0.966\ atm$$

Now use the Ideal Gas Law to convert to moles

$$n = \frac{PV}{RT} = \frac{(0.966\ atm)(0.155\ L)}{(0.082\ Latmmol^{-1}K^{-1})(295\ K)} = 0.00619\ mol$$

Scuba diving

Our respiratory systems are designed to maintain the proper oxygen concentration in the blood when the partial pressure of O_2 is 0.21 atm, its normal sea-level value. Below the water surface, the pressure increases by 1 atm for each 10.3 m increase in depth; thus a scuba diver at 10.3 m experiences a total of 2 atm pressure pressing on the body. In order to prevent the lungs from collapsing, the air the diver breathes should also be at about the same pressure.

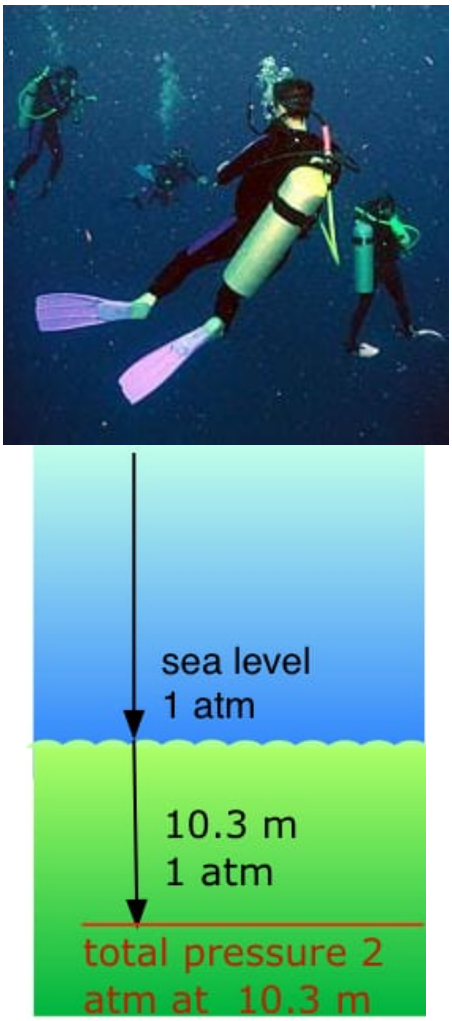


Figure 13.5.2: Scuba Dviging actively takes into account both Henry's and Dalton's Laws

But at a total pressure of 2 atm, the partial pressure of O_2 in ordinary air would be 0.42 atm; at a depth of 100 ft (about 30 m), the O_2 pressure of 0.8 atm would be far too high for health. For this reason, the air mixture in the pressurized tanks that scuba divers wear must contain a smaller fraction of O_2. This can be achieved most simply by raising the nitrogen content, but high partial pressures of N_2 can also be dangerous, resulting in a condition known as nitrogen narcosis. The preferred diluting agent for sustained deep diving is helium, which has very little tendency to dissolve in the blood even at high pressures.

Career Focus: Respiratory Therapist

Certain diseases—such as emphysema, lung cancer, and severe asthma—primarily affect the lungs. Respiratory therapists help patients with breathing-related problems. They can evaluate, help diagnose, and treat breathing disorders and even help provide emergency assistance in acute illness where breathing is compromised.

Most respiratory therapists must complete at least two years of college and earn an associate's degree, although therapists can assume more responsibility if they have a college degree. Therapists must also pass state or national certification exams. Once certified, respiratory therapists can work in hospitals, doctor's offices, nursing homes, or patient's homes. Therapists work with equipment such as oxygen tanks and respirators, may sometimes dispense medication to aid in breathing, perform tests, and educate patients in breathing exercises and other therapy.

Because respiratory therapists work directly with patients, the ability to work well with others is a must for this career. It is an important job because it deals with one of the most crucial functions of the body.

Concept Review Exercises

1. What properties do the gas laws help us predict?
2. What makes the ideal gas law different from the other gas laws?

Answers

1. Gas laws relate four properties: pressure, volume, temperature, and number of moles.
2. The ideal gas law does not require that the properties of a gas change.

Key Takeaway

- The physical properties of gases are predictable using mathematical formulas known as gas laws.

Exercises

1. What conditions of a gas sample should remain constant for Boyle's law to be used?
2. What conditions of a gas sample should remain constant for Charles's law to be used?
3. Does the identity of a gas matter when using Boyle's law? Why or why not?
4. Does the identity of a gas matter when using Charles's law? Why or why not?
5. A sample of nitrogen gas is confined to a balloon that has a volume of 1.88 L and a pressure of 1.334 atm. What will be the volume of the balloon if the pressure is changed to 0.662 atm? Assume that the temperature and the amount of the gas remain constant.
6. A sample of helium gas in a piston has a volume of 86.4 mL under a pressure of 447 torr. What will be the volume of the helium if the pressure on the piston is increased to 1,240 torr? Assume that the temperature and the amount of the gas remain constant.
7. If a gas has an initial pressure of 24,650 Pa and an initial volume of 376 mL, what is the final volume if the pressure of the gas is changed to 775 torr? Assume that the amount and the temperature of the gas remain constant.
8. A gas sample has an initial volume of 0.9550 L and an initial pressure of 564.5 torr. What would the final pressure of the gas be if the volume is changed to 587.0 mL? Assume that the amount and the temperature of the gas remain constant.
9. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is 18°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and amount of the gas remain constant.
10. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is –10°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and the amount of the gas remain constant.
11. An air/gas vapor mix in an automobile cylinder has an initial temperature of 450 K and a volume of 12.7 cm^3. The gas mix is heated to 565°C. If pressure and amount are held constant, what is the final volume of the gas in cubic centimeters?
12. Given the following conditions for a gas: V_i = 0.665 L, T_i = 23.6°C, V_f = 1.034 L. What is T_f in degrees Celsius and kelvins?
13. Assuming the amount remains the same, what must be the final volume of a gas that has an initial volume of 387 mL, an initial pressure of 456 torr, an initial temperature of 65.0°C, a final pressure of 1.00 atm, and a final temperature of 300 K?
14. When the nozzle of a spray can is depressed, 0.15 mL of gas expands to 0.44 mL, and its pressure drops from 788 torr to 1.00 atm. If the initial temperature of the gas is 22.0°C, what is the final temperature of the gas?
15. Use the ideal gas law to show that 1 mol of a gas at STP has a volume of about 22.4 L.

16. Use a standard conversion factor to determine a value of the ideal gas law constant *R* that has units of L•torr/mol•K.

17. How many moles of gas are there in a 27.6 L sample at 298 K and a pressure of 1.44 atm?

18. How many moles of gas are there in a 0.066 L sample at 298 K and a pressure of 0.154 atm?

19. A 0.334 mol sample of carbon dioxide gas is confined to a volume of 20.0 L and has a pressure of 0.555 atm. What is the temperature of the carbon dioxide in kelvins and degrees Celsius?

20. What must *V* be for a gas sample if *n* = 4.55 mol, *P* = 7.32 atm, and *T* = 285 K?

21. What is the pressure of 0.0456 mol of Ne gas contained in a 7.50 L volume at 29°C?

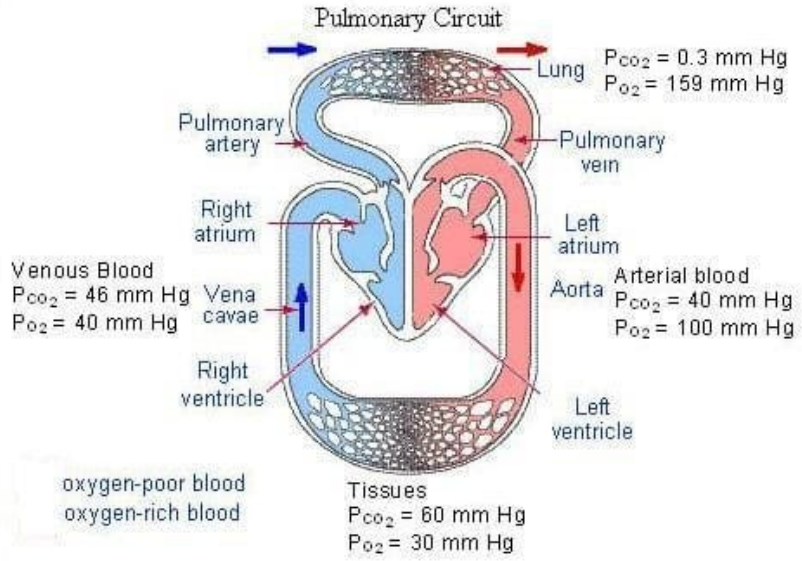
22. What is the pressure of 1.00 mol of Ar gas that has a volume of 843.0 mL and a temperature of −86.0°C?

23. A mixture of the gases N_2, O_2, and Ar has a total pressure of 760 mm Hg. If the partial pressure of N_2 is 220 mm Hg and of O_2 is 470 mm Hg, What is the partial pressure of Ar?

24. What percent of the gas above is Ar?

25. Apply Henry's Law to the diagram below to explain:

why oxygen diffuses from the alveoli of the lungs into the blood and from the blood into the tissues of the body. why carbon dioxide diffuses from the tissues into the blood and from the blood into the alveoli and then finally out into the atmosphere.



Answers

1. temperature and amount of the gas

3. The identity does not matter because the variables of Boyle's law do not identify the gas.

5. 3.89 L

7. 92.1 mL

9. 1.07 L

11. 23.7 cm^3

13. 206 mL

15. The ideal gas law confirms that 22.4 L equals 1 mol.

17. 1.63 mol

19. 405 K; 132°C

21. 0.151 atm

This page titled 13.5: Gas Laws is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

13.E: Solids, Liquids, and Gases (Exercises)

8.1: Intermolecular Interactions

Concept Review Exercise

1. What types of intermolecular interactions can exist in compounds?

2. What is the difference between covalent network and covalent molecular compounds?

Answer

1. ionic bonding, network covalent, dispersion forces, dipole-dipole interactions, and hydrogen bonding.

2. Covalent network compounds contain atoms that are covalently bonded to other individual atoms in a giant 3-dimensional network. Covalent molecular compounds contain individual molecules that are attracted to one another through dispersion, dipole-dipole or hydrogen bonding.

Exercises

1. List the three common phases in the order you are likely to find them—from lowest temperature to highest temperature.
2. List the three common phases in the order they exist from lowest energy to highest energy.
3. List these intermolecular interactions from weakest to strongest: London forces, hydrogen bonding, and ionic interactions.
4. List these intermolecular interactions from weakest to strongest: covalent network bonding, dipole-dipole interactions, and dispersion forces.
5. What type of intermolecular interaction is predominate in each substance?
    a. water (H_2O)
    b. sodium sulfate (Na_2SO_4)
    c. decane ($C_{10}H_{22}$)
6. What type of intermolecular interaction is predominate in each substance?
    a. diamond (C, crystal)
    b. helium (He)
    c. ammonia (NH_3)
7. Explain how a molecule like carbon dioxide (CO_2) can have polar covalent bonds but be nonpolar overall.
8. Sulfur dioxide (SO_2) has a formula similar to that of carbon dioxide (see Exercise 7) but is a polar molecule overall. What can you conclude about the shape of the SO_2 molecule?
9. What are some of the physical properties of substances that experience covalent network bonding?
10. What are some of the physical properties of substances that experience only dispersion forces?

Answers

1. solid, liquid, and gas

2. solid, liquid, and gas

3. London forces, hydrogen bonding, and ionic interactions

4. dispersion, dipole-dipole, network covalent

5. a. hydrogen bonding
   b. ionic interactions
   c. dispersion forces

6. a. network covalent
b. dispersion
c. hydrogen bonding

7. The two covalent bonds are oriented in such a way that their dipoles cancel out.

8. SO_2 is not a linear molecule. It has a bent or V-shape.

9. very hard, high melting point

10. very soft, very low melting point

8.2: Solids and Liquids

Concept Review Exercise

1. How do the strengths of intermolecular interactions in solids and liquids differ?

Answer

1. Solids have stronger intermolecular interactions than liquids do.

Exercises

1. What are the general properties of solids?
2. What are the general properties of liquids
3. What are the general properties of gases?
4. What phase or phases have a definite volume? What phase or phases do not have a definite volume?
5. Name a common substance that forms a crystal in its solid state.
6. Name a common substance that forms an amorphous solid in its solid state.
7. Are substances with strong intermolecular interactions likely to be solids at higher or lower temperatures? Explain.
8. Are substances with weak intermolecular interactions likely to be liquids at higher or lower temperatures? Explain.

9. State two similarities between the solid and liquid states.

10. State two differences between the solid and liquid states.

11. If individual particles are moving around with respect to each other, a substance may be in either the _____ or _____ state but probably not in the _____ state.

12. If individual particles are in contact with each other, a substance may be in either the _____ or _____ state but probably not in the _____ state.

Answers

1. hard, specific volume and shape, high density, cannot be compressed

2. fixed volume, no definite shape, high density, individual molecules touch each other but in a random way

3. variable volume and shape, low density, compressible

4. solid and liquid have definite volume; gas has no definite volume

5. sodium chloride (answers will vary)

6. glass

7. At higher temperatures, their intermolecular interactions are strong enough to hold the particles in place.

8. Substances with weak intermolecular interactions are likely to be liquids at lower temperatures. Their attractive forces are more easily broken hence they melt more readily.

9. high density; definite volume

10. Solids have definite shape while liquids don't. In solids, molecules occupy fixed positions in a pattern, while in liquids, the molecules are moving in a random arrangement.

11. liquid; gas; solid

12. solid; liquid; gas

8.3: Gases and Pressure

Concept Review Exercise

1. What is pressure, and what units do we use to express it?

Answer

1. Pressure is the force per unit area; its units can be pascals, torr, millimeters of mercury, or atmospheres.

Exercises

1. What is the kinetic theory of gases?

2. According to the kinetic theory of gases, the individual gas particles are (always, frequently, never) moving.

3. Why does a gas exert pressure?
4. Why does the kinetic theory of gases allow us to presume that all gases will show similar behavior?
5. Arrange the following pressure quantities in order from smallest to largest: 1 mmHg, 1 Pa, and 1 atm.
6. Which unit of pressure is larger—the torr or the atmosphere?
7. How many torr are there in 1.56 atm?
8. Convert 760 torr into pascals.
9. Blood pressures are expressed in millimeters of mercury. What would be the blood pressure in atmospheres if a patient's systolic blood pressure is 120 mmHg and the diastolic blood pressure is 82 mmHg? (In medicine, such a blood pressure would be reported as "120/82," spoken as "one hundred twenty over eighty-two.")
10. In weather forecasting, barometric pressure is expressed in inches of mercury (in. Hg), where there are exactly 25.4 mmHg in every 1 in. Hg. What is the barometric pressure in millimeters of mercury if the barometric pressure is reported as 30.21 in. Hg?

Answers

1. Gases are composed of tiny particles that are separated by large distances. Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container. The velocity of gas particles is related to the temperature of a gas. Gas particles do not experience any force of attraction or repulsion with each other.

2. always

3. A gas exerts pressure as its particles rebound off the walls of its container.

4. Because the molecules are far apart and don't have attractive forces between them

5. 1 Pa, 1 mmHg, and 1 atm

6. atm

7. 1,190 torr

8. 98,700 Pa

9. 0.158 atm; 0.108 atm

10. 767.3 mm Hg

8.4: Gas Laws

Concept Review Exercises

1. What properties do the gas laws help us predict?
2. What makes the ideal gas law different from the other gas laws?

Answers

1. Gas laws relate four properties: pressure, volume, temperature, and number of moles.
2. The ideal gas law does not require that the properties of a gas change.

Exercises

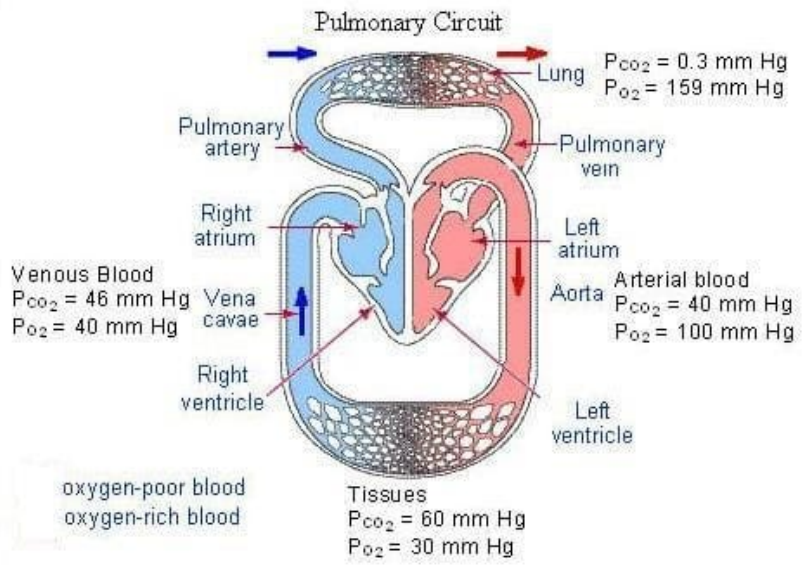
1. What conditions of a gas sample should remain constant for Boyle's law to be used?
2. What conditions of a gas sample should remain constant for Charles's law to be used?
3. Does the identity of a gas matter when using Boyle's law? Why or why not?
4. Does the identity of a gas matter when using Charles's law? Why or why not?
5. A sample of nitrogen gas is confined to a balloon that has a volume of 1.88 L and a pressure of 1.334 atm. What will be the volume of the balloon if the pressure is changed to 0.662 atm? Assume that the temperature and the amount of the gas remain constant.
6. A sample of helium gas in a piston has a volume of 86.4 mL under a pressure of 447 torr. What will be the volume of the helium if the pressure on the piston is increased to 1,240 torr? Assume that the temperature and the amount of the gas remain constant.
7. If a gas has an initial pressure of 24,650 Pa and an initial volume of 376 mL, what is the final volume if the pressure of the gas is changed to 775 torr? Assume that the amount and the temperature of the gas remain constant.
8. A gas sample has an initial volume of 0.9550 L and an initial pressure of 564.5 torr. What would the final pressure of the gas be if the volume is changed to 587.0 mL? Assume that the amount and the temperature of the gas remain constant.
9. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is 18°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and amount of the gas remain constant.
10. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is –10°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and the amount of the gas remain constant.
11. An air/gas vapor mix in an automobile cylinder has an initial temperature of 450 K and a volume of 12.7 cm^3. The gas mix is heated to 565°C. If pressure and amount are held constant, what is the final volume of the gas in cubic centimeters?
12. Given the following conditions for a gas: V_i = 0.665 L, T_i = 23.6°C, V_f = 1.034 L. What is T_f in degrees Celsius and kelvins?
13. Assuming the amount remains the same, what must be the final volume of a gas that has an initial volume of 387 mL, an initial pressure of 456 torr, an initial temperature of 65.0°C, a final pressure of 1.00 atm, and a final temperature of 300 K?
14. When the nozzle of a spray can is depressed, 0.15 mL of gas expands to 0.44 mL, and its pressure drops from 788 torr to 1.00 atm. If the initial temperature of the gas is 22.0°C, what is the final temperature of the gas?
15. Use the ideal gas law to show that 1 mol of a gas at STP has a volume of about 22.4 L.
16. Use a standard conversion factor to determine a value of the ideal gas law constant R that has units of L•torr/mol•K.
17. How many moles of gas are there in a 27.6 L sample at 298 K and a pressure of 1.44 atm?
18. How many moles of gas are there in a 0.066 L sample at 298 K and a pressure of 0.154 atm?
19. A 0.334 mol sample of carbon dioxide gas is confined to a volume of 20.0 L and has a pressure of 0.555 atm. What is the temperature of the carbon dioxide in kelvins and degrees Celsius?
20. What must V be for a gas sample if n = 4.55 mol, P = 7.32 atm, and T = 285 K?
21. What is the pressure of 0.0456 mol of Ne gas contained in a 7.50 L volume at 29°C?
22. What is the pressure of 1.00 mol of Ar gas that has a volume of 843.0 mL and a temperature of –86.0°C?
23. A mixture of the gases N_2, O_2, and Ar has a total pressure of 760 mm Hg. If the partial pressure of N_2 is 220 mm Hg and of

O_2 is 470 mm Hg, What is the partial pressure of Ar?

24. What percent of the gas above is Ar?

25. Apply Henry's Law to the diagram below to explain:

why oxygen diffuses from the alveoli of the lungs into the blood and from the blood into the tissues of the body. why carbon dioxide diffuses from the tissues into the blood and from the blood into the alveoli and then finally out into the atmosphere.



Answers

1. temperature and amount of the gas

2. pressure and amount of the gas
3. The identity does not matter because the variables of Boyle's law do not identify the gas.
4. The identity does not matter because the variables of Charles law do not identify the gas.
5. 3.79 L

6. 31.1 mL

7. 92.1 mL

8. 918.4 torr

9. 1.07 L

10. 1.18 L

11. 23.7 cm^3

12. 461 K; 188^{0}C

13. 206 mL

14. 835 K; 562^{0}C

15. The ideal gas law confirms that 22.4 L equals 1 mol.

16. $\frac{760\ torr}{1\ atm}$

17. 1.63 mol

18. 4.2 x 10^{-4} mol

19. 405 K; 132°C

20. 14.5 L

21. 0.151 atm

22. 18.2 atm

23. 70 mm Hg

24. 9.2%

25. Gases diffuse from high concentration to low concentration (Henry's Law). The partial pressure of **oxygen** is high in the **alveoli** and low in the **blood** of the pulmonary capillaries. As a result, **oxygen diffuses** across the respiratory membrane from the **alveoli** into the **blood**. It's also higher partial pressure in the blood than in the tissues, hence it transfers to the tissues. On the other hand, carbon dioxide **diffuses** from the tissues (highest CO_2 partial pressure) and across the respiratory membrane from the **blood** into the **alveoli** and out **to the atmosphere.**

Additional Exercises

1. How many grams of oxygen gas are needed to fill a 25.0 L container at 0.966 atm and 22°C?
2. A breath of air is about 1.00 L in volume. If the pressure is 1.00 atm and the temperature is 37°C, what mass of air is contained in each breath? Use an average molar mass of 28.8 g/mol for air.
3. The balanced chemical equation for the combustion of propane is as follows:

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(\ell)} \tag{13.E.1}$$

   a. If 100.0 g of propane are combusted, how many moles of oxygen gas are necessary for the reaction to occur?
   b. At STP, how many liters of oxygen gas would that be?

4. The equation for the formation of ammonia gas (NH_3) is as follows:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)} \tag{13.E.2}$$

   At 500°C and 1.00 atm, 10.0 L of N_2 gas are reacted to make ammonia.

   a. If the pressures and temperatures of H_2 and NH_3 were the same as those of N_2, what volume of H_2 would be needed to react with N_2, and what volume of NH_3 gas would be produced?
   b. Compare your answers to the balanced chemical equation. Can you devise a “shortcut” method to answer Exercise 4a?

5. At 20°C, 1 g of liquid H_2O has a volume of 1.002 mL. What volume will 1 g of water vapor occupy at 20°C if its pressure is 17.54 mmHg? By what factor has the water expanded in going from the liquid phase to the gas phase?
6. At 100°C, 1 g of liquid H_2O has a volume of 1.043 mL. What volume will 1 g of steam occupy at 100°C if its pressure is 760.0 mmHg? By what factor has the water expanded in going from the liquid phase to the gas phase?
7. Predict whether NaCl or NaI will have the higher melting point. Explain. (Hint: consider the relative strengths of the intermolecular interactions of the two compounds.)
8. Predict whether CH_4 or CH_3OH will have the lower boiling point. Explain. (Hint: consider the relative strengths of the

intermolecular interactions of the two compounds.)

9. A standard automobile tire has a volume of about 3.2 ft^3 (where 1 ft^3 equals 28.32 L). Tires are typically inflated to an absolute pressure of 45.0 pounds per square inch (psi), where 1 atm equals 14.7 psi. Using this information with the ideal gas law, determine the number of moles of air needed to fill a tire if the air temperature is 18.0°C.

10. Another gas law, Amontons's law, relates pressure and temperature under conditions of constant amount and volume:

$$\frac{P_i}{T_i} = \frac{P_f}{T_f}$$

If an automobile tire (see Exercise 9) is inflated to 45.0 psi at 18.0°C, what will be its pressure if the operating temperature (i.e., the temperature the tire reaches when the automobile is on the road) is 45.0°C? Assume that the volume and the amount of the gas remain constant.

Answers

1. 31.9 g

2. 1.13 g

3.

   a. 11.4 mol
   b. 255 L

4.

   a. 30.0 L H_2 and 20.0 L NH_3

   b. the mole ratio in the balanced equation is the same as the volume ratio.

$$10.0\ \text{L}\ \text{N}_2 \times \frac{3\ \text{L}\ \text{H}_2}{1\ \text{L}\ \text{N}_2} = 30.0\ \text{L}\ \text{H}_2$$

$$10.0\ \text{L}\ \text{N}_2 \times \frac{2\ \text{L}\ \text{NH}_3}{1\ \text{L}\ \text{N}_2} = 20.0\ \text{L}\ \text{NH}_3$$

5. 57.75 L; an expansion of 57,600 times

6. 1.698 L; an expansion of 1,628 times

7. NaCl; with smaller anions, NaCl likely experiences stronger ionic bonding.

8. CH_4 will have the lower boiling point because its intermolecular force (London dispersion force only) is weaker than those in CH_3OH. Aside from London dispersion, CH_3OH has dipole-dipole and hydrogen bonding.

9. 11.6 mol

10. 49.2 psi

This page titled 13.E: Solids, Liquids, and Gases (Exercises) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- 8.1: Intermolecular Interactions by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 8.2: Solids and Liquids by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 8.3: Gases and Pressure by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 8.4: Gas Laws by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

13.S: Solids, Liquids, and Gases (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

A **phase** is a certain form of matter that has the same physical properties throughout. Three phases are common: the solid, the liquid, and the gas phase. What determines the phase of a substance? Generally, the strength of the **intermolecular interactions** determines whether a substance is a solid, liquid, or gas under any particular conditions. **Covalent network bonding** is a very strong form of intermolecular interaction. Diamond is one example of a substance that has this intermolecular interaction. **Ionic interactions**, the forces of attraction due to oppositely charged ions, are also relatively strong. Covalent bonds are another type of interaction within molecules, but if the bonds are **polar covalent bonds**, then the unequal sharing of electrons can cause charge imbalances within molecules that cause interactions between molecules. These molecules are described as **polar**, and these interactions are called **dipole-dipole interactions**. A certain rather strong type of dipole-dipole interaction, involving a hydrogen atom, is called **hydrogen bonding**. On the other hand, equal sharing of electrons forms **nonpolar covalent bonds**, and the interactions between different molecules is less because the molecules are nonpolar. All substances have very weak **dispersion forces** (also called **London forces**) caused by the movement of electrons within the bonds themselves.

In the solid phase, intermolecular interactions are so strong that they hold the individual atoms or molecules in place. In many solids, the regular three-dimensional arrangement of particles makes a **crystal**. In other solids, the irregular arrangement of particles makes an **amorphous** solid. In liquids, the intermolecular interactions are strong enough to keep the particles of substance together but not in place. Thus, the particles are free to move over each other but still remain in contact.

In gases, the intermolecular interactions are weak enough that the individual particles are separated from each other in space. The **kinetic theory of gases** is a collection of statements that describe the fundamental behavior of all gases. Among other properties, gases exert a **pressure** on their container. Pressure is measured using units like **pascal**, **bar**, **atmosphere**, or **mmHg** (also called a **torr**).

There are several simple relationships between the variables used to describe a quantity of gas. These relationships are called **gas laws**. **Boyle's law** relates the pressure and volume of a gas, while **Charles's law** relates the volume and absolute temperature of a gas. The **combined gas law** relates the volume, pressure, and absolute temperature of a gas sample. All of these gas laws allow us to understand the changing conditions of a gas. The **ideal gas law** relates the pressure, volume, amount, and absolute temperature of a gas under any conditions. These four variables are related to the **ideal gas law constant**, which is the proportionality constant used to calculate the conditions of a gas. Because the conditions of a gas can change, a set of benchmark conditions called **standard temperature and pressure (STP)** is defined. Standard temperature is 0°C, and standard pressure is 1.00 atm.

This page titled 13.S: Solids, Liquids, and Gases (Summary) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14: SOLUTIONS

Solutions are all around us. Air, for example, is a solution. If you live near a lake, a river, or an ocean, that body of water is not pure H_2O but most probably a solution. Much of what we drink—for example, soda, coffee, tea, and milk—is at least in part a solution. Solutions are a large part of everyday life. A lot of the chemistry occurring around us happens in solution. In fact, much of the chemistry that occurs in our own bodies takes place in solution, and many solutions—such as the Ringer's lactate IV solution—are important for our health. In our understanding of chemistry, we need to understand a little bit about solutions. In this chapter, you will learn about the special characteristics of solutions, how solutions are characterized, and some of their properties.

14.1: Prelude to Solutions

14.2: Solutions

14.3: Concentration

14.4: The Dissolution Process

14.5: Properties of Solutions

14.E: Solutions (Exercises)

14.S: Solutions (Summary)

This page titled 14: Solutions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.1: Prelude to Solutions

If you watch any of the medical dramas on television, you may have heard a doctor (actually an actor) call for an intravenous solution of “Ringer’s lactate” (or “lactated Ringer’s”). So what is Ringer’s lactate?

Intravenous (IV) solutions are administered for two main reasons:

1. to introduce necessary substances into the bloodstream, such as ions for proper body function, sugar and other food substances for energy, or drugs to treat a medical condition, and
2. to increase the volume of the bloodstream.

Many people with acute or long-term medical conditions have received some type of an IV solution.

One basic IV solution, called a *normal saline solution*, is simply a dilute solution of NaCl dissolved in water. Normal saline is 9.0 g of NaCl dissolved in each liter of solution. *Ringer’s lactate* is a normal saline solution that also has small amounts of potassium and calcium ions mixed in. In addition, it contains about 2.5 g of lactate ions ($C_3H_5O_3^-$) per liter of solution. The liver metabolizes lactate ions into bicarbonate (HCO_3^-) ions, which help maintain the acid-base balance of blood. Many medical problems, such as heart attacks and shock, affect the acid-base balance of blood, and the presence of lactate in the IV solution eases problems caused by this imbalance.

Physicians can select from a range of premade IV solutions, in accordance with a patient’s particular needs. Ringer’s lactate is commonly used when a patient’s blood volume must be increased quickly. Another frequently used IV solution, called D5W, is a 5% solution of dextrose (a form of sugar) in water.

This page titled 14.1: Prelude to Solutions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.2: Solutions

Learning Objectives

- To understand what causes solutions to form.

A solution is another name for a homogeneous mixture. A *mixture* as a material composed of two or more substances. In a solution, the combination is so intimate that the different substances cannot be differentiated by sight, even with a microscope. Compare, for example, a mixture of salt and pepper and another mixture consisting of salt and water. In the first mixture, we can readily see individual grains of salt and the flecks of pepper. A mixture of salt and pepper is not a solution. However, in the second mixture, no matter how carefully we look, we cannot see two different substances. Salt dissolved in water is a solution.

The major component of a solution, called the solvent, is typically the same phase as the solution itself. Each minor component of a solution (and there may be more than one) is called the solute. In most of the solutions we will describe in this textbook, there will be no ambiguity about whether a component is the solvent or the solute.) For example, in a solution of salt in water, the solute is salt, and solvent is water.

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table 14.2.1 lists some common types of solutions, with examples of each.

Table 14.2.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	saltwater
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

What causes a solution to form? The simple answer is that the solvent and the solute must have similar intermolecular interactions. When this is the case, the individual particles of solvent and solute can easily mix so intimately that each particle of solute is surrounded by particles of solute, forming a solution. However, if two substances have very different intermolecular interactions, large amounts of energy are required to force their individual particles to mix intimately, so a solution does not form.

This process leads to a simple rule of thumb: *like dissolves like*. Solvents that are very polar will dissolve solutes that are very polar or even ionic. Solvents that are nonpolar will dissolve nonpolar solutes. Thus water, being polar, is a good solvent for ionic compounds and polar solutes like ethanol (C_2H_5OH). However, water does not dissolve nonpolar solutes, such as many oils and greases (Figure 14.2.1).



Figure 14.2.1: A beaker holds water with blue food dye (upper liquid layer) and a much more dense perfluoroheptane (a fluorocarbon) lower liquid layer. The two fluids cannot mix and the dye cannot dissolve in fluorocarbon. A goldfish and a crab have been introduced into the water. The goldfish cannot penetrate the dense fluorocarbon. The crab floats at the liquid boundary with only parts of his legs penetrating the fluorocarbon fluid, unable to sink to the bottom of the beaker. Quarter coins rest on the bottom of the beaker. Animals were rescued from their predicament after the photo was taken. Figure used with permission from Wikipedia (Sbharris (Steven B. Harris)).

We use the word soluble to describe a solute that dissolves in a particular solvent, and the word insoluble for a solute that does not dissolve in a solvent. Thus, we say that sodium chloride is soluble in water but insoluble in hexane (C_6H_{14}). If the solute and the solvent are both liquids and soluble in any proportion, we use the word miscible, and the word immiscible if they are not.

Example 14.2.1

Water is considered a polar solvent. Which substances should dissolve in water?

1. methanol (CH_3OH)
2. sodium sulfate (Na_2SO_4)
3. octane (C_8H_{18})

SOLUTION

Because water is polar, substances that are polar or ionic will dissolve in it.

1. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
2. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
3. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

Exercise 14.2.1

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Concept Review Exercises

1. What causes a solution to form?
2. How does the phrase *like dissolves like* relate to solutions?

Answers

1. Solutions form because a solute and a solvent have similar intermolecular interactions.
2. It means that substances with similar intermolecular interactions will dissolve in each other.

Key Takeaway

- Solutions form because a solute and a solvent experience similar intermolecular interactions.

Exercises

1. Define *solution*.
2. Give several examples of solutions.
3. What is the difference between a solvent and a solute?
4. Can a solution have more than one solute in it? Can you give an example?
5. Does a solution have to be a liquid? Give several examples to support your answer.
6. Give at least two examples of solutions found in the human body.
7. Which substances will probably be soluble in water, a very polar solvent?
    a. sodium nitrate ($NaNO_3$)
    b. hexane (C_6H_{14})
    c. isopropyl alcohol [$(CH_3)_2CHOH$]
    d. benzene (C_6H_6)
8. Which substances will probably be soluble in toluene ($C_6H_5CH_3$), a nonpolar solvent?
    a. sodium nitrate ($NaNO_3$)
    b. hexane (C_6H_{14})
    c. isopropyl alcohol [$(CH_3)_2CHOH$]
    d. benzene (C_6H_6)
9. The solubility of alcohols in water varies with the length of carbon chain. For example, ethanol (CH_3CH_2OH) is soluble in water in any ratio, while only 0.0008 mL of heptanol ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$) will dissolve in 100 mL of water. Propose an explanation for this behavior.
10. Dimethyl sulfoxide [$(CH_3)_2SO$] is a polar liquid. Based on the information in Exercise 9, which do you think will be more soluble in it—ethanol or heptanol?

Answers

1. a homogeneous mixture
3. A solvent is the majority component of a solution; a solute is the minority component of a solution.

5. A solution does not have to be liquid; air is a gaseous solution, while some alloys are solid solutions (answers will vary).

7. a. probably soluble
   b. probably not soluble
   c. probably soluble
   d. probably not soluble

9. Small alcohol molecules have strong polar intermolecular interactions, so they dissolve in water. In large alcohol molecules, the nonpolar end overwhelms the polar end, so they do not dissolve very well in water.

This page titled 14.2: Solutions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.3: Concentration

Learning Objectives

- Express the amount of solute in a solution in various concentration units.
- Use molarity to determine quantities in chemical reactions.
- Determine the resulting concentration of a diluted solution.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms whose meanings depend on various factors.

Solubility

There is usually a limit to how much solute will dissolve in a given amount of solvent. This limit is called the solubility of the solute. Some solutes have a very small solubility, while other solutes are soluble in all proportions. Table 14.3.1 lists the solubilities of various solutes in water. Solubilities vary with temperature, so Table 14.3.1 includes the temperature at which the solubility was determined.

Table 14.3.1: Solubilities of Various Solutes in Water at 25°C (Except as Noted)

Substance	Solubility (g in 100 mL of H2O)
AgCl(s)	0.019
$C_6H_6(\ell)$ (benzene)	0.178
$CH_4(g)$	0.0023
$CO_2(g)$	0.150
$CaCO_3(s)$	0.058
$CaF_2(s)$	0.0016
$Ca(NO_3)_2(s)$	143.9
$C_6H_{12}O_6$ (glucose)	120.3 (at 30°C)
KBr(s)	67.8
$MgCO_3(s)$	2.20
NaCl(s)	36.0
$NaHCO_3(s)$	8.41
$C_{12}H_{22}O_{11}$ (sucrose)	204.0 (at 20°C)

If a solution contains so much solute that its solubility limit is reached, the solution is said to be saturated, and its concentration is known from information contained in Table 14.3.1. If a solution contains less solute than the solubility limit, it is unsaturated. Under special circumstances, more solute can be dissolved even after the normal solubility limit is reached; such solutions are called *supersaturated* and are not stable. If the solute is solid, excess solute can easily recrystallize. If the solute is a gas, it can bubble out of solution uncontrollably, like what happens when you shake a soda can and then immediately open it.

Precipitation from Supersaturated Solutions

Recrystallization of excess solute from a supersaturated solution usually gives off energy as heat. Commercial heat packs containing supersaturated sodium acetate ($NaC_2H_3O_2$) take advantage of this phenomenon. You can probably find them at your local drugstore.

YouTube Video

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

View this YouTube Video online

Video 14.3.1**:** Watered-down sodium acetate trihydrate. Needle crystal is truly wonderful structures

Most solutions we encounter are unsaturated, so knowing the solubility of the solute does not accurately express the amount of solute in these solutions. There are several common ways of specifying the concentration of a solution.

Percent Composition

There are several ways of expressing the concentration of a solution by using a percentage. The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \mathrm{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{14.3.1}$$

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Example 14.3.1

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

SOLUTION

We can substitute the quantities given in the equation for mass/mass percent:

$$\%\ \mathrm{m/m} = \frac{36.5\ \mathrm{g}}{355\ \mathrm{g}} \times 100\% = 10.3\%$$

Exercise 14.3.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

For gases and liquids, volumes are relatively easy to measure, so the concentration of a liquid or a gas solution can be expressed as a volume/volume percent (% v/v): the volume of a solute divided by the volume of a solution times 100:

$$\% \text{ v/v} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{14.3.2}$$

Again, the units of the solute and the solution must be the same. A hybrid concentration unit, mass/volume percent (% m/v), is commonly used for intravenous (IV) fluids (Figure 14.3.1). It is defined as the mass in grams of a solute, divided by volume in milliliters of solution times 100:

$$\% \text{ m/v} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100\% \tag{14.3.3}$$



Figure 14.3.1: Mass/Volume Percent.The 0.9% NaCl concentration on this IV bag is mass/volume percent (left). Such solution is used for other purposes and available in bottles (right). Figures used with permission from Wikipedia

Each percent concentration can be used to produce a conversion factor between the amount of solute, the amount of solution, and the percent. Furthermore, given any two quantities in any percent composition, the third quantity can be calculated, as the following example illustrates.

Example 14.3.2

A sample of 45.0% v/v solution of ethanol (C_2H_5OH) in water has a volume of 115 mL. What volume of ethanol solute does the

sample contain?

SOLUTION

A percentage concentration is simply the number of parts of solute per 100 parts of solution. Thus, the percent concentration of 45.0% v/v implies the following:

$$45.0\%\ \text{v/v} \rightarrow \frac{45\ \text{mL}\ C_2H_5OH}{100\ \text{mL solution}}$$

That is, there are 45 mL of C_2H_5OH for every 100 mL of solution. We can use this fraction as a conversion factor to determine the amount of C_2H_5OH in 115 mL of solution:

$$115\ \text{mL solution} \times \frac{45\ \text{mL}\ C_2H_5OH}{100\ \text{mL solution}} = 51.8\ \text{mL}\ C_2H_5OH$$

The highest concentration of ethanol that can be obtained normally is 95% ethanol, which is actually 95% v/v.

Exercise 14.3.2

What volume of a 12.75% m/v solution of glucose ($C_6H_{12}O_6$) in water is needed to obtain 50.0 g of $C_6H_{12}O_6$?

Example 14.3.3

A normal saline IV solution contains 9.0 g of NaCl in every liter of solution. What is the mass/volume percent of normal saline?

SOLUTION

We can use the definition of mass/volume percent, but first we have to express the volume in milliliter units:

1 L = 1,000 mL

Because this is an exact relationship, it does not affect the significant figures of our result.

$$\%\ \text{m/v} = \frac{9.0\ \text{g NaCl}}{1,000\ \text{mL solution}} \times 100\% = 0.90\%\ \text{m/v}$$

Exercise 14.3.3

The chlorine bleach that you might find in your laundry room is typically composed of 27.0 g of sodium hypochlorite (NaOCl), dissolved to make 500.0 mL of solution. What is the mass/volume percent of the bleach?

In addition to percentage units, the units for expressing the concentration of extremely dilute solutions are parts per million (ppm) and parts per billion (ppb). Both of these units are mass based and are defined as follows:

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000 \tag{14.3.4}$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000,000 \tag{14.3.5}$$

Similar to parts per million and parts per billion, related units include parts per thousand (ppth) and parts per trillion (ppt).

Concentrations of *trace elements* in the body—elements that are present in extremely low concentrations but are nonetheless necessary for life—are commonly expressed in parts per million or parts per billion. Concentrations of poisons and pollutants are

also described in these units. For example, cobalt is present in the body at a concentration of 21 ppb, while the State of Oregon's Department of Agriculture limits the concentration of arsenic in fertilizers to 9 ppm.

In aqueous solutions, 1 ppm is essentially equal to 1 mg/L, and 1 ppb is equivalent to 1 µg/L.

Example 14.3.4

If the concentration of cobalt in a human body is 21 ppb, what mass in grams of Co is present in a body having a mass of 70.0 kg?

SOLUTION

A concentration of 21 ppb means "21 g of solute per 1,000,000,000 g of solution." Written as a conversion factor, this concentration of Co is as follows:

$$21 \text{ ppb Co} \rightarrow \frac{21 \text{ g Co}}{1,000,000,000 \text{ g solution}}$$

We can use this as a conversion factor, but first we must convert 70.0 kg to gram units:

$$70.0 \text{ kg} \times \frac{1,000 \text{ g}}{1 \text{ kg}} = 7.00 \times 10^4 \text{ g}$$

Now we determine the amount of Co:

$$7.00 \times 10^4 \text{ g solution} \times \frac{21 \text{ g Co}}{1,000,000,000 \text{ g solution}} = 0.0015 \text{ g Co}$$

This is only 1.5 mg.

Exercise 14.3.4

An 85 kg body contains 0.012 g of Ni. What is the concentration of Ni in parts per million?

Molarity

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

Molarity is defined as the number of moles of a solute dissolved per liter of solution:

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{14.3.6}$$

Molarity is abbreviated M (often referred to as "molar"), and the units are often abbreviated as mol/L. It is important to remember that "mol" in this expression refers to moles of solute and that "L" refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is therefore

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl} \tag{14.3.7}$$

which is read as "three point oh molar sodium chloride." Sometimes (aq) is added when the solvent is water, as in "3.0 M NaCl(aq)."

Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

Example 14.3.5

What is the molarity of an aqueous solution of 25.0 g of NaOH in 750 mL?

SOLUTION

Before we substitute these quantities into the definition of molarity, we must convert them to the proper units. The mass of NaOH must be converted to moles of NaOH. The molar mass of NaOH is 40.00 g/mol:

$$25.0\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.00\text{ g NaOH}} = 0.625\text{ mol NaOH}$$

Next, we convert the volume units from milliliters to liters:

$$750\text{ mL} \times \frac{1\text{ L}}{1,000\text{ mL}} = 0.750\text{ L}$$

Now that the quantities are expressed in the proper units, we can substitute them into the definition of molarity:

$$\text{M} = \frac{0.625\text{ mol NaOH}}{0.750\text{ L}} = 0.833\text{ M NaOH}$$

Exercise 14.3.5

If a 350 mL cup of coffee contains 0.150 g of caffeine ($C_8H_{10}N_4O_2$), what is the molarity of this caffeine solution?

The definition of molarity can also be used to calculate a needed volume of solution, given its concentration and the number of moles desired, or the number of moles of solute (and subsequently, the mass of the solute), given its concentration and volume. The following example illustrates this.

Example 14.3.6

1. What volume of a 0.0753 M solution of dimethylamine [$(CH_3)_2NH$] is needed to obtain 0.450 mol of the compound?
2. Ethylene glycol ($C_2H_6O_2$) is mixed with water to make auto engine coolants. How many grams of $C_2H_6O_2$ are in 5.00 L of a 6.00 M aqueous solution?

SOLUTION

In both parts, we will use the definition of molarity to solve for the desired quantity.

1. $0.0753\text{ M} = \dfrac{0.450\text{ mol }(CH_3)_2NH}{\text{volume of solution}}$

   To solve for the volume of solution, we multiply both sides by volume of solution and divide both sides by the molarity value to isolate the volume of solution on one side of the equation:

   $$\text{volume of solution} = \frac{0.450\text{ mol }(CH_3)_2NH}{0.0753\text{ M}} = 5.98\text{ L}$$

   Note that because the definition of molarity is mol/L, the division of mol by M yields L, a unit of volume.

2. The molar mass of $C_2H_6O_2$ is 62.08 g/mol., so

   $$6.00\text{ M} = \frac{\text{moles of solute}}{5.00\text{ L}}$$

   To solve for the number of moles of solute, we multiply both sides by the volume:

$$\text{moles of solute} = (6.00\ \text{M})(5.00\ \text{L}) = 30.0\ \text{mol}$$

Note that because the definition of molarity is mol/L, the product M × L gives mol, a unit of amount. Now, using the molar mass of $C_3H_8O_3$, we convert mol to g:

$$30.0\ \text{mol} \times \frac{62.08\ \text{g}}{\text{mol}} = 1,860\ \text{g}$$

Thus, there are 1,860 g of $C_2H_6O_2$ in the specified amount of engine coolant.

Note: Dimethylamine has a “fishy” odor. In fact, organic compounds called amines cause the odor of decaying fish.

Exercise 14.3.6

a. What volume of a 0.0902 M solution of formic acid (HCOOH) is needed to obtain 0.888 mol of HCOOH?
b. Acetic acid ($HC_2H_3O_2$) is the acid in vinegar. How many grams of $HC_2H_3O_2$ are in 0.565 L of a 0.955 M solution?

Using Molarity in Stoichiometry Problems

Of all the ways of expressing concentration, molarity is the one most commonly used in stoichiometry problems because it is directly related to the mole unit. Consider the following chemical equation:

$$HCl(aq) + NaOH(s) \rightarrow H_2O(\ell) + NaCl(aq)$$

Suppose we want to know how many liters of aqueous HCl solution will react with a given mass of NaOH. A typical approach to answering this question is as follows:

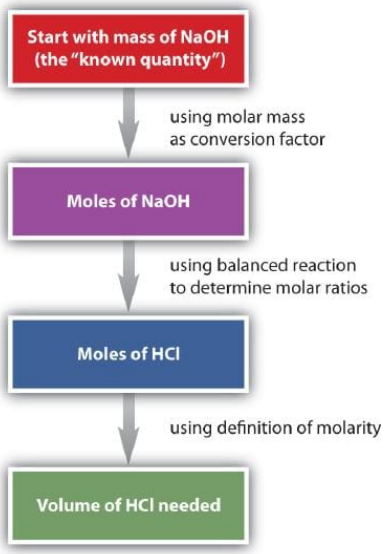


Figure 14.3.2*: Typical approach to solving Molarity problems*

In itself, each step is a straightforward conversion. It is the combination of the steps that is a powerful quantitative tool for problem solving.

Example 14.3.7

How many milliliters of a 2.75 M HCl solution are needed to react with 185 g of NaOH? The balanced chemical equation for

this reaction is as follows:

$$HCl(aq) + NaOH(s) \rightarrow H_2O(\ell) + NaCl(aq)$$

SOLUTION

We will follow the flowchart to answer this question. First, we convert the mass of NaOH to moles of NaOH using its molar mass, 40.00 g/mol:

$$185 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 4.63 \text{ mol NaOH}$$

Using the balanced chemical equation, we see that there is a one-to-one ratio of moles of HCl to moles of NaOH. We use this to determine the number of moles of HCl needed to react with the given amount of NaOH:

$$4.63 \text{ mol NaOH} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 4.63 \text{ mol HCl}$$

Finally, we use the definition of molarity to determine the volume of 2.75 M HCl needed:

$$2.75 \text{ M HCl} = \frac{4.63 \text{ mol HCl}}{\text{volume of HCl solution}}$$

$$\text{volume of HCl} = \frac{4.63 \text{ mol HCl}}{2.75 \text{ M HCl}} = 1.68 \text{ L} \times \frac{1,000 \text{ mL}}{1 \text{ L}} = 1,680 \text{ mL}$$

We need 1,680 mL of 2.75 M HCl to react with the NaOH.

Exercise 14.3.7

How many milliliters of a 1.04 M H_2SO_4 solution are needed to react with 98.5 g of $Ca(OH)_2$? The balanced chemical equation for the reaction is as follows:

$$H_2SO_{4(aq)} + Ca(OH)_{2(s)} \rightarrow 2H_2O_{(\ell)} + CaSO_{4(aq)}$$

The general steps for performing stoichiometry problems such as this are shown in Figure 14.3.3. You may want to consult this figure when working with solutions in chemical reactions. The double arrows in Figure 14.3.3 indicate that you can start at either end of the chart and, after a series of simple conversions, determine the quantity at the other end.

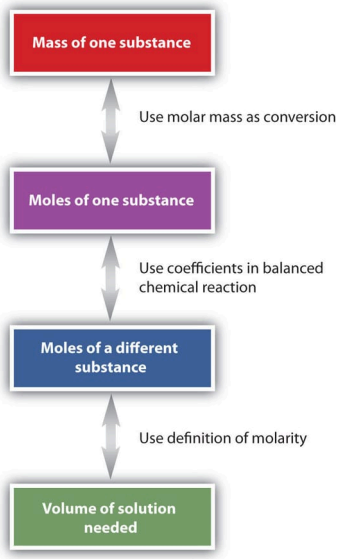


Figure 14.3.3: Diagram of Steps for Using Molarity in Stoichiometry Calculations. When using molarity in stoichiometry calculations, a specific sequence of steps usually leads you to the correct answer.

Many of the fluids found in our bodies are solutions. The solutes range from simple ionic compounds to complex proteins. Table 14.3.2 lists the typical concentrations of some of these solutes.

*Table 14.3.2: Approximate Concentrations of Various Solutes in Some Solutions in the Body**

Solution	Solute	Concentration (M)
blood plasma	Na^+	0.138
	K^+	0.005
	Ca^{2+}	0.004
	Mg^{2+}	0.003
	Cl^-	0.110
	HCO_3^-	0.030
stomach acid	HCl	0.10
urine	NaCl	0.15
	PO_4^{3-}	0.05
	NH_2CONH_2 (urea)	0.30
***Note: Concentrations are approximate and can vary widely.**		

Looking Closer: The Dose Makes the Poison

Why is it that we can drink 1 qt of water when we are thirsty and not be harmed, but if we ingest 0.5 g of arsenic, we might die? There is an old saying: *the dose makes the poison*. This means that what may be dangerous in some amounts may not be dangerous in other amounts.

Take arsenic, for example. Some studies show that arsenic deprivation limits the growth of animals such as chickens, goats, and

pigs, suggesting that arsenic is actually an essential trace element in the diet. Humans are constantly exposed to tiny amounts of arsenic from the environment, so studies of completely arsenic-free humans are not available; if arsenic is an essential trace mineral in human diets, it is probably required on the order of 50 ppb or less. A toxic dose of arsenic corresponds to about 7,000 ppb and higher, which is over 140 times the trace amount that may be required by the body. Thus, arsenic is not poisonous in and of itself. Rather, it is the amount that is dangerous: the dose makes the poison.

Similarly, as much as water is needed to keep us alive, too much of it is also risky to our health. Drinking too much water too fast can lead to a condition called water intoxication, which may be fatal. The danger in water intoxication is not that water itself becomes toxic. It is that the ingestion of too much water too fast dilutes sodium ions, potassium ions, and other salts in the bloodstream to concentrations that are not high enough to support brain, muscle, and heart functions. Military personnel, endurance athletes, and even desert hikers are susceptible to water intoxication if they drink water but do not replenish the salts lost in sweat. As this example shows, even the right substances in the wrong amounts can be dangerous!

Equivalents

Concentrations of ionic solutes are occasionally expressed in units called equivalents (Eq). One equivalent equals 1 mol of positive or negative charge. Thus, 1 mol/L of $Na^+(aq)$ is also 1 Eq/L because sodium has a 1+ charge. A 1 mol/L solution of $Ca^{2+}(aq)$ ions has a concentration of 2 Eq/L because calcium has a 2+ charge. Dilute solutions may be expressed in milliequivalents (mEq)—for example, human blood plasma has a total concentration of about 150 mEq/L. (For more information about the ions present in blood plasma, see Chapter 3 "Ionic Bonding and Simple Ionic Compounds", Section 3.3 "Formulas for Ionic Compounds".)

Dilution

When solvent is added to dilute a solution, the volume of the solution changes, but the amount of solute does not change. Before dilution, the amount of solute was equal to its original concentration times its original volume:

$$\text{amount in moles} = (\text{concentration} \times \text{volume})_{\text{initial}}$$

After dilution, the same amount of solute is equal to the final concentration times the final volume:

$$\text{amount in moles} = (\text{concentration} \times \text{volume})_{\text{final}}$$

To determine a concentration or amount after a dilution, we can use the following equation:

$$(\text{concentration} \times \text{volume})_{\text{initial}} = (\text{concentration} \times \text{volume})_{\text{final}}$$

Any units of concentration and volume can be used, as long as both concentrations and both volumes have the same unit.

Example 14.3.8

A 125 mL sample of 0.900 M NaCl is diluted to 1,125 mL. What is the final concentration of the diluted solution?

SOLUTION

Because the volume units are the same, and we are looking for the molarity of the final solution, we can use $(\text{concentration} \times \text{volume})_{\text{initial}} = (\text{concentration} \times \text{volume})_{\text{final}}$:

$$(0.900\ \text{M} \times 125\ \text{mL}) = (\text{concentration} \times 1{,}125\ \text{mL})$$

We solve by isolating the unknown concentration by itself on one side of the equation. Dividing by 1,125 mL gives

$$\text{concentration} = \frac{0.900\ \text{M} \times 125\ \text{mL}}{1{,}125\ \text{mL}} = 0.100\ \text{M}$$

as the final concentration.

Exercise 14.3.8

A nurse uses a syringe to inject 5.00 mL of 0.550 M heparin solution (heparin is an anticoagulant drug) into a 250 mL IV bag, for a final volume of 255 mL. What is the concentration of the resulting heparin solution?

Key Takeaways

- Various concentration units are used to express the amounts of solute in a solution.
- Concentration units can be used as conversion factors in stoichiometry problems.
- New concentrations can be easily calculated if a solution is diluted.

Concept Review Exercises

1. What are some of the units used to express concentration?
2. Distinguish between the terms *solubility* and *concentration*.

Answers

1. % m/m, % m/v, ppm, ppb, molarity, and Eq/L (answers will vary)
2. Solubility is typically a limit to how much solute can dissolve in a given amount of solvent. Concentration is the quantitative amount of solute dissolved at any concentration in a solvent.

Exercises

1. Define *solubility*. Do all solutes have the same solubility?
2. Explain why the terms *dilute* or *concentrated* are of limited usefulness in describing the concentration of solutions.
3. If the solubility of sodium chloride (NaCl) is 30.6 g/100 mL of H_2O at a given temperature, how many grams of NaCl can be dissolved in 250.0 mL of H_2O?
4. If the solubility of glucose ($C_6H_{12}O_6$) is 120.3 g/100 mL of H_2O at a given temperature, how many grams of $C_6H_{12}O_6$ can be dissolved in 75.0 mL of H_2O?
5. How many grams of sodium bicarbonate ($NaHCO_3$) can a 25.0°C saturated solution have if 150.0 mL of H_2O is used as the solvent?
6. If 75.0 g of potassium bromide (KBr) are dissolved in 125 mL of H_2O, is the solution saturated, unsaturated, or supersaturated?
7. Calculate the mass/mass percent of a saturated solution of NaCl. Use the data from Table 14.3.1 "Solubilities of Various Solutes in Water at 25°C (Except as Noted)", assume that masses of the solute and the solvent are additive, and use the density of H_2O (1.00 g/mL) as a conversion factor.
8. Calculate the mass/mass percent of a saturated solution of $MgCO_3$ Use the data from Table 14.3.1 "Solubilities of Various Solutes in Water at 25°C (Except as Noted)", assume that masses of the solute and the solvent are additive, and use the density of H_2O (1.00 g/mL) as a conversion factor.
9. Only 0.203 mL of C_6H_6 will dissolve in 100.000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of benzene in water.

10. Only 35 mL of aniline ($C_6H_5NH_2$) will dissolve in 1,000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of aniline in water.

11. A solution of ethyl alcohol (C_2H_5OH) in water has a concentration of 20.56% v/v. What volume of C_2H_5OH is present in 255 mL of solution?

12. What mass of KCl is present in 475 mL of a 1.09% m/v aqueous solution?

13. The average human body contains 5,830 g of blood. What mass of arsenic is present in the body if the amount in blood is 0.55 ppm?

14. The Occupational Safety and Health Administration has set a limit of 200 ppm as the maximum safe exposure level for carbon monoxide (CO). If an average breath has a mass of 1.286 g, what is the maximum mass of CO that can be inhaled at that maximum safe exposure level?

15. Which concentration is greater—15 ppm or 1,500 ppb?

16. Express the concentration 7,580 ppm in parts per billion.

17. What is the molarity of 0.500 L of a potassium chromate solution containing 0.0650 mol of K_2CrO_4?

18. What is the molarity of 4.50 L of a solution containing 0.206 mol of urea [$(NH_2)_2CO$]?

19. What is the molarity of a 2.66 L aqueous solution containing 56.9 g of NaBr?

20. If 3.08 g of $Ca(OH)_2$ is dissolved in enough water to make 0.875 L of solution, what is the molarity of the $Ca(OH)_2$?

21. What mass of HCl is present in 825 mL of a 1.25 M solution?

22. What mass of isopropyl alcohol (C_3H_8O) is dissolved in 2.050 L of a 4.45 M aqueous C_3H_8O solution?

23. What volume of 0.345 M NaCl solution is needed to obtain 10.0 g of NaCl?

24. How many milliliters of a 0.0015 M cocaine hydrochloride ($C_{17}H_{22}ClNO_4$) solution is needed to obtain 0.010 g of the solute?

25. Aqueous calcium chloride reacts with aqueous silver nitrate according to the following balanced chemical equation:

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

How many moles of AgCl(s) are made if 0.557 L of 0.235 M $CaCl_2$ react with excess $AgNO_3$? How many grams of AgCl are made?

26. Sodium bicarbonate ($NaHCO_3$) is used to react with acid spills. The reaction with sulfuric acid (H_2SO_4) is as follows:

$$2NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(\ell) + 2CO_2(g)$$

If 27.6 mL of a 6.25 M H_2SO_4 solution were spilled, how many moles of $NaHCO_3$ would be needed to react with the acid? How many grams of $NaHCO_3$ is this?

27. The fermentation of glucose to make ethanol and carbon dioxide has the following overall chemical equation:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

If 1.00 L of a 0.567 M solution of $C_6H_{12}O_6$ were completely fermented, what would be the resulting concentration of the C_2H_5OH solution? How many moles of CO_2 would be formed? How many grams is this? If each mole of CO_2 had a volume of 24.5 L, what volume of CO_2 is produced?

28. Aqueous sodium bisulfite gives off sulfur dioxide gas when heated:

$$2NaHSO_3(aq) \rightarrow Na_2SO_3(aq) + H_2O(\ell) + SO_2(g)$$

If 567 mL of a 1.005 M $NaHSO_3$ solution were heated until all the $NaHSO_3$ had reacted, what would be the resulting concentration of the Na_2SO_3 solution? How many moles of SO_2 would be formed? How many grams of SO_2 would be formed? If each mole of SO_2 had a volume of 25.78 L, what volume of SO_2 would be produced?

29. What is the concentration of a 1.0 M solution of $K^+(aq)$ ions in equivalents/liter?

30. What is the concentration of a 1.0 M solution of $SO_4^{2-}(aq)$ ions in equivalents/liter?

31. A solution having initial concentration of 0.445 M and initial volume of 45.0 mL is diluted to 100.0 mL. What is its final concentration?

32. A 50.0 mL sample of saltwater that is 3.0% m/v is diluted to 950 mL. What is its final mass/volume percent?

Answers

1. Solubility is the amount of a solute that can dissolve in a given amount of solute, typically 100 mL. The solubility of solutes varies widely.

3. 76.5 g

5. 12.6 g

7. 26.5%

9. 0.203%

11. 52.4 mL

13. 0.00321 g

15. 15 ppm

17. 0.130 M

19. 0.208 M

21. 37.6 g

23. 0.496 L

25. 0.262 mol; 37.5 g

27. 1.13 M C_2H_5OH; 1.13 mol of CO_2; 49.7 g of CO_2; 27.7 L of CO_2

29. 1.0 Eq/L

31. 0.200 M

This page titled 14.3: Concentration is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.4: The Dissolution Process

Learning Objectives

- To describe the dissolution process at the molecular level

The Dissolution Process

What occurs at the molecular level to cause a solute to dissolve in a solvent? The answer depends in part on the solute, but there are some similarities common to all solutes.

Recall the rule that *like dissolves like*. This means that substances must have similar intermolecular forces to form solutions. When a soluble solute is introduced into a solvent, the particles of solute can interact with the particles of solvent. In the case of a solid or liquid solute, the interactions between the solute particles and the solvent particles are so strong that the individual solute particles separate from each other and, surrounded by solvent molecules, enter the solution. (Gaseous solutes already have their constituent particles separated, but the concept of being surrounded by solvent particles still applies.) This process is called solvation and is illustrated in Figure 14.4.1. When the solvent is water, the word hydration, rather than solvation, is used.

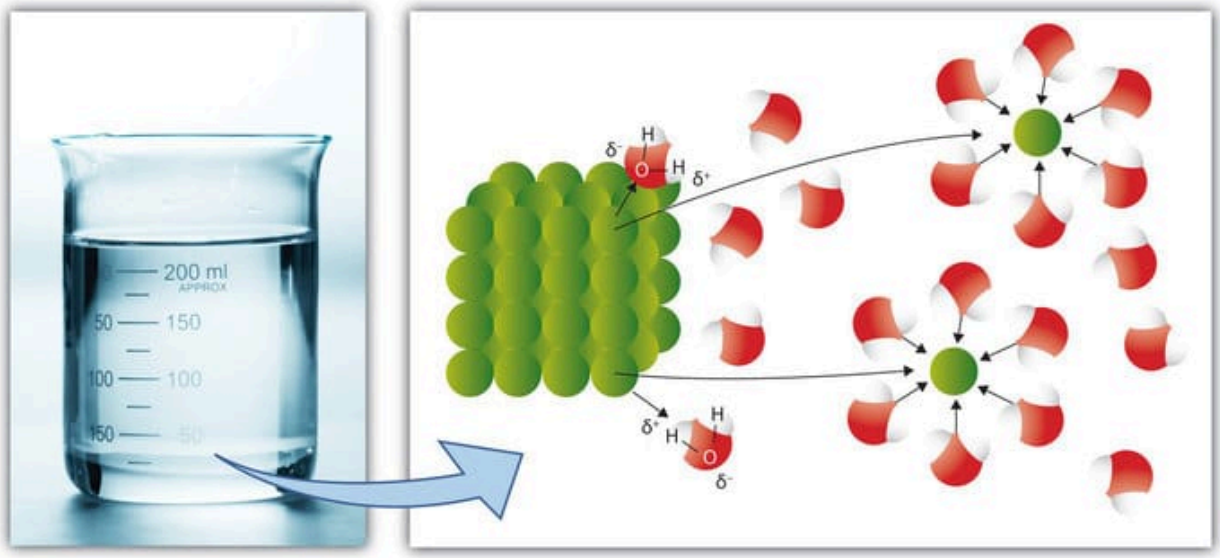


Figure 14.4.1: Solvation. When a solute dissolves, the individual particles of solute become surrounded by solvent particles. Eventually the particle detaches from the remaining solute, surrounded by solvent molecules in solution. Source: Photo © Thinkstock

Ionic Compounds and Covalent Compounds as Solutes

In the case of molecular solutes like glucose, the solute particles are individual molecules. However, if the solute is ionic, the individual ions separate from each other and become surrounded by solvent particles. That is, the cations and anions of an ionic solute separate when the solute dissolves. This process is referred to as dissociation (Figure 14.4.1).

The dissociation of soluble ionic compounds gives solutions of these compounds an interesting property: they conduct electricity. Because of this property, soluble ionic compounds are referred to as electrolytes. Many ionic compounds dissociate completely and are therefore called strong electrolytes. Sodium chloride is an example of a strong electrolyte. Some compounds dissolve but dissociate only partially, and solutions of such solutes may conduct electricity only weakly. These solutes are called weak electrolytes. Acetic acid (CH_3COOH), the compound in vinegar, is a weak electrolyte. Solutes that dissolve into individual neutral molecules without dissociation do not impart additional electrical conductivity to their solutions and are called nonelectrolytes. Table sugar ($C_{12}H_{22}O_{11}$) is an example of a nonelectrolyte.

The link below will connect you to an animation that illustrates the differences between ionic compounds (electroytes) and covalent compounds (non-electrolytes) dissolved in water.

> The term *electrolyte* is used in medicine to mean any of the important ions that are dissolved in aqueous solution in the body. Important physiological electrolytes include Na^+, K^+, Ca^{2+}, Mg^{2+}, and Cl^-.

Example 14.4.1

The following substances all dissolve to some extent in water. Classify each as an electrolyte or a nonelectrolyte.

1. potassium chloride (KCl)
2. fructose ($C_6H_{12}O_6$)
3. isopropyl alcohol [$CH_3CH(OH)CH_3$]
4. magnesium hydroxide [$Mg(OH)_2$]

SOLUTION

Each substance can be classified as an ionic solute or a nonionic solute. Ionic solutes are electrolytes, and nonionic solutes are nonelectrolytes.

1. Potassium chloride is an ionic compound; therefore, when it dissolves, its ions separate, making it an electrolyte.
2. Fructose is a sugar similar to glucose. (In fact, it has the same molecular formula as glucose.) Because it is a molecular compound, we expect it to be a nonelectrolyte.
3. Isopropyl alcohol is an organic molecule containing the alcohol functional group. The bonding in the compound is all covalent, so when isopropyl alcohol dissolves, it separates into individual molecules but not ions. Thus, it is a nonelectrolyte
4. Magnesium hydroxide is an ionic compound, so when it dissolves it dissociates. Thus, magnesium hydroxide is an electrolyte.

Exercise 14.4.1

The following substances all dissolve to some extent in water. Classify each as an electrolyte or a nonelectrolyte.

a. acetone (CH_3COCH_3)
b. iron(III) nitrate [$Fe(NO_3)_3$]
c. elemental bromine (Br_2)
d. sodium hydroxide (NaOH)

Concept Review Exercise

1. Explain how the solvation process describes the dissolution of a solute in a solvent.

Answer

1. Each particle of the solute is surrounded by particles of the solvent, carrying the solute from its original phase.

Key Takeaway

- When a solute dissolves, its individual particles are surrounded by solvent molecules and are separated from each other.

Exercises

1. Describe what happens when an ionic solute like Na_2SO_4 dissolves in a polar solvent.

2. Describe what happens when a molecular solute like sucrose ($C_{12}H_{22}O_{11}$) dissolves in a polar solvent.

3. Classify each substance as an electrolyte or a nonelectrolyte. Each substance dissolves in H_2O to some extent.

    a. NH_4NO_3
    b. CO_2
    c. NH_2CONH_2
    d. HCl

4. Classify each substance as an electrolyte or a nonelectrolyte. Each substance dissolves in H_2O to some extent.

    a. $CH_3CH_2CH_2OH$
    b. $Ca(CH_3CO_2)_2$
    c. I_2
    d. KOH

5. Will solutions of each solute conduct electricity when dissolved?

    a. $AgNO_3$
    b. $CHCl_3$
    c. $BaCl_2$
    d. Li_2O

6. Will solutions of each solute conduct electricity when dissolved?

    a. CH_3COCH_3
    b. $N(CH_3)_3$
    c. $CH_3CO_2C_2H_5$
    d. $FeCl_2$

Answers

1. Each ion of the ionic solute is surrounded by particles of solvent, carrying the ion from its associated crystal.

3. a. electrolyte
    b. nonelectrolyte
    c. nonelectrolyte
    d. electrolyte

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

This page titled 14.4: The Dissolution Process is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.5: Properties of Solutions

Learning Objectives

- To describe how the properties of solutions differ from those of pure solvents.

Solutions are likely to have properties similar to those of their major component—usually the solvent. However, some solution properties differ significantly from those of the solvent. Here, we will focus on liquid solutions that have a solid solute, but many of the effects we will discuss in this section are applicable to all solutions.

Colligative Properties

Solutes affect some properties of solutions that depend only on the concentration of the dissolved particles. These properties are called colligative properties. Four important colligative properties that we will examine here are vapor pressure depression, boiling point elevation, freezing point depression, and osmotic pressure.

Molecular compounds separate into individual molecules when they are dissolved, so for every 1 mol of molecules dissolved, we get 1 mol of particles. In contrast, ionic compounds separate into their constituent ions when they dissolve, so 1 mol of an ionic compound will produce more than 1 mol of dissolved particles. For example, every mole of NaCl that dissolves yields 1 mol of Na^+ ions and 1 mol of Cl^- ions, for a total of 2 mol of particles in solution. Thus, the effect on a solution's properties by dissolving NaCl may be twice as large as the effect of dissolving the same amount of moles of glucose ($C_6H_{12}O_6$).

Vapor Pressure Depression

All liquids evaporate. In fact, given enough volume, a liquid will turn completely into a vapor. If enough volume is not present, a liquid will evaporate only to the point where the rate of evaporation equals the rate of vapor condensing back into a liquid. The pressure of the vapor at this point is called the vapor pressure of the liquid.

The presence of a dissolved solid lowers the characteristic vapor pressure of a liquid so that it evaporates more slowly. (The exceptions to this statement are if the solute itself is a liquid or a gas, in which case the solute will also contribute something to the evaporation process. We will not discuss such solutions here.) This property is called vapor pressure depression and is depicted in Figure 14.5.1.

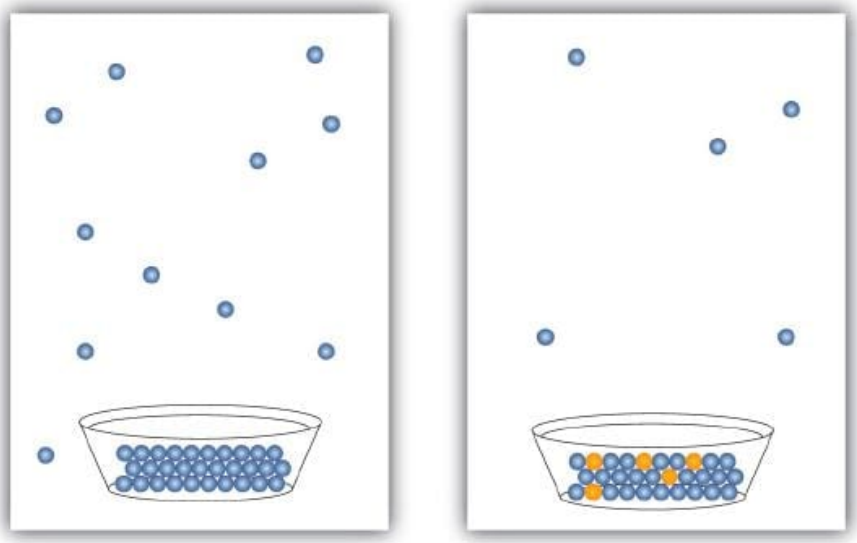


Figure 14.5.1: Vapor Pressure Depression. The presence of solute particles blocks some of the ability for liquid particles to evaporate. Thus, solutions of solid solutes typically have a lower vapor pressure than the pure solvent.

Boiling Point and Freezing Point Effects

A related property of solutions is that their boiling points are higher than the boiling point of the pure solvent. Because the presence of solute particles decreases the vapor pressure of the liquid solvent, a higher temperature is needed to reach the boiling point. This phenomenon is called boiling point elevation. For every mole of particles dissolved in a liter of water, the boiling point of water increases by about 0.5°C.

> Some people argue that putting a pinch or two of salt in water used to cook spaghetti or other pasta makes a solution that has a higher boiling point, so the pasta cooks faster. In actuality, the amount of solute is so small that the boiling point of the water is practically unchanged.

The presence of solute particles has the opposite effect on the freezing point of a solution. When a solution freezes, only the solvent particles come together to form a solid phase, and the presence of solute particles interferes with that process. Therefore, for the liquid solvent to freeze, more energy must be removed from the solution, which lowers the temperature. Thus, solutions have lower freezing points than pure solvents do. This phenomenon is called freezing point depression. For every mole of particles in a liter of water, the freezing point decreases by about 1.9°C.

Both boiling point elevation and freezing point depression have practical uses. For example, solutions of water and ethylene glycol ($C_2H_6O_2$) are used as coolants in automobile engines because the boiling point of such a solution is greater than 100°C, the normal boiling point of water. In winter, salts like NaCl and $CaCl_2$ are sprinkled on the ground to melt ice or keep ice from forming on roads and sidewalks (Figure 14.5.2). This is because the solution made by dissolving sodium chloride or calcium chloride in water has a lower freezing point than pure water, so the formation of ice is inhibited.



Figure 14.5.2: Effect of Freezing Point Depression. The salt sprinkled on this sidewalk makes the water on the sidewalk have a lower freezing point than pure water, so it does not freeze as easily. This makes walking on the sidewalk less hazardous in winter. © Thinkstock

Example 14.5.1

Which solution's freezing point deviates more from that of pure water—a 1 M solution of NaCl or a 1 M solution of $CaCl_2$?

SOLUTION

Colligative properties depend on the number of dissolved particles, so the solution with the greater number of particles in solution will show the greatest deviation. When NaCl dissolves, it separates into two ions, Na^+ and Cl^-. But when $CaCl_2$ dissolves, it separates into three ions—one Ca^{2+} ion and two Cl^- ions. Thus, mole for mole, $CaCl_2$ will have 50% more impact on freezing point depression than NaCl.

Exercise 14.5.1

Which solution's boiling point deviates more from that of pure water—a 1 M solution of $CaCl_2$ or a 1 M solution of $MgSO_4$?

Osmotic Pressure

The last colligative property of solutions we will consider is a very important one for biological systems. It involves osmosis, the process by which solvent molecules can pass through certain membranes but solute particles cannot. When two solutions of different concentration are present on either side of these membranes (called *semipermeable membranes*), there is a tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal. This tendency is called osmotic pressure. External pressure can be exerted on a solution to counter the flow of solvent; the pressure required to halt the osmosis of a solvent is equal to the osmotic pressure of the solution.

Osmolarity (osmol) is a way of reporting the total number of particles in a solution to determine osmotic pressure. It is defined as the molarity of a solute times the number of particles a formula unit of the solute makes when it dissolves (represented by *i*):

$$osmol = M \times i \tag{14.5.1}$$

If more than one solute is present in a solution, the individual osmolarities are additive to get the total osmolarity of the solution. Solutions that have the same osmolarity have the same osmotic pressure. If solutions of differing osmolarities are present on opposite sides of a semipermeable membrane, solvent will transfer from the lower-osmolarity solution to the higher-osmolarity solution. Counterpressure exerted on the high-osmolarity solution will reduce or halt the solvent transfer. An even higher pressure

can be exerted to force solvent from the high-osmolarity solution to the low-osmolarity solution, a process called *reverse osmosis*. Reverse osmosis is used to make potable water from saltwater where sources of fresh water are scarce.

Example 14.5.2

A 0.50 M NaCl aqueous solution and a 0.30 M $Ca(NO_3)_2$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

SOLUTION

The solvent will flow into the solution of higher osmolarity. The NaCl solute separates into two ions—Na^+ and Cl^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol (NaCl)} = 0.50\ \text{M} \times 2 = 1.0\ \text{osmol}$$

The $Ca(NO_3)_2$ solute separates into three ions—one Ca^{2+} and two NO_3^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol } [Ca(NO_3)_2] = 0.30\ \text{M} \times 3 = 0.90\ \text{osmol}$$

The osmolarity of the $Ca(NO_3)_2$ solution is lower than that of the NaCl solution, so water will transfer through the membrane from the $Ca(NO_3)_2$ solution to the NaCl solution.

Exercise 14.5.2

A 1.5 M $C_6H_{12}O_6$ aqueous solution and a 0.40 M $Al(NO_3)_3$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

To Your Health: Dialysis

The main function of the kidneys is to filter the blood to remove wastes and extra water, which are then expelled from the body as urine. Some diseases rob the kidneys of their ability to perform this function, causing a buildup of waste materials in the bloodstream. If a kidney transplant is not available or desirable, a procedure called dialysis can be used to remove waste materials and excess water from the blood.

In one form of dialysis, called *hemodialysis*, a patient's blood is passed though a length of tubing that travels through an *artificial kidney machine* (also called a *dialysis machine*). A section of tubing composed of a semipermeable membrane is immersed in a solution of sterile water, glucose, amino acids, and certain electrolytes. The osmotic pressure of the blood forces waste molecules and excess water through the membrane into the sterile solution. Red and white blood cells are too large to pass through the membrane, so they remain in the blood. After being cleansed in this way, the blood is returned to the body.



A patient undergoing hemodialysis depends on osmosis to cleanse the blood of waste products that the kidneys are incapable of removing due to disease. Image used with permission from Wikipedia.

Dialysis is a continuous process, as the osmosis of waste materials and excess water takes time. Typically, 5–10 lb of waste-containing fluid is removed in each dialysis session, which can last 2–8 hours and must be performed several times a week. Although some patients have been on dialysis for 30 or more years, dialysis is always a temporary solution because waste materials are constantly building up in the bloodstream. A more permanent solution is a kidney transplant.

Cell walls are semipermeable membranes, so the osmotic pressures of the body's fluids have important biological consequences. If solutions of different osmolarity exist on either side of the cells, solvent (water) may pass into or out of the cells, sometimes with disastrous results. Consider what happens if red blood cells are placed in a *hypotonic* solution, meaning a solution of lower osmolarity than the liquid inside the cells. The cells swell up as water enters them, disrupting cellular activity and eventually causing the cells to burst. This process is called *hemolysis*. If red blood cells are placed in a *hypertonic* solution, meaning one having a higher osmolarity than exists inside the cells, water leaves the cells to dilute the external solution, and the red blood cells shrivel and die. This process is called *crenation*. Only if red blood cells are placed in *isotonic* solutions that have the same osmolarity as exists inside the cells are they unaffected by negative effects of osmotic pressure. Glucose solutions of about 0.31 M, or sodium chloride solutions of about 0.16 M, are isotonic with blood plasma.

The concentration of an isotonic sodium chloride (NaCl) solution is only half that of an isotonic glucose ($C_6H_{12}O_6$) solution because NaCl produces two ions when a formula unit dissolves, while molecular $C_6H_{12}O_6$ produces only one particle when a formula unit dissolves. The osmolarities are therefore the same even though the concentrations of the two solutions are different.

Osmotic pressure explains why you should not drink seawater if you are abandoned in a life raft in the middle of the ocean. Its osmolarity is about three times higher than most bodily fluids. You would actually become thirstier as water from your cells was drawn out to dilute the salty ocean water you ingested. Our bodies do a better job coping with hypotonic solutions than with hypertonic ones. The excess water is collected by our kidneys and excreted.

Osmotic pressure effects are used in the food industry to make pickles from cucumbers and other vegetables and in brining meat to make corned beef. It is also a factor in the mechanism of getting water from the roots to the tops of trees!

Career Focus: Perfusionist

A perfusionist is a medical technician trained to assist during any medical procedure in which a patient's circulatory or breathing functions require support. The use of perfusionists has grown rapidly since the advent of open-heart surgery in 1953.

Most perfusionists work in operating rooms, where their main responsibility is to operate heart-lung machines. During many

heart surgeries, the heart itself must be stopped. In these situations, a heart-lung machine keeps the patient alive by aerating the blood with oxygen and removing carbon dioxide. The perfusionist monitors both the machine and the status of the blood, notifying the surgeon and the anesthetist of any concerns and taking corrective action if the status of the blood becomes abnormal.

Despite the narrow parameters of their specialty, perfusionists must be highly trained. Certified perfusion education programs require a student to learn anatomy, physiology, pathology, chemistry, pharmacology, math, and physics. A college degree is usually required. Some perfusionists work with other external artificial organs, such as hemodialysis machines and artificial livers.

Concept Review Exercises

1. What are the colligative properties of solutions?
2. Explain how the following properties of solutions differ from those of the pure solvent: vapor pressure, boiling point, freezing point, and osmotic pressure.

Answers

1. Colligative properties are characteristics that a solution has that depend on the number, not the identity, of solute particles.
2. In solutions, the vapor pressure is lower, the boiling point is higher, the freezing point is lower, and the osmotic pressure is higher.

Key Takeaway

- `Exercises

1. In each pair of aqueous systems, which will have the lower vapor pressure?
    a. pure water or 1.0 M NaCl
    b. 1.0 M NaCl or 1.0 M $C_6H_{12}O_6$
    c. 1.0 M $CaCl_2$ or 1.0 M $(NH_4)_3PO_4$
2. In each pair of aqueous systems, which will have the lower vapor pressure?
    a. 0.50 M $Ca(NO_3)_2$ or 1.0 M KBr
    b. 1.5 M $C_{12}H_{22}O_{11}$ or 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$ or pure water
3. In each pair of aqueous systems, which will have the higher boiling point?
    a. pure water or a 1.0 M NaCl
    b. 1.0 M NaCl or 1.0 M $C_6H_{12}O_6$
    c. 1.0 M $CaCl_2$ or 1.0 M $(NH_4)_3PO_4$
4. In each pair of aqueous systems, which will have the higher boiling point?
    a. 1.0 M KBr
    b. 1.5 M $C_{12}H_{22}O_{11}$ or 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$ or pure water
5. Estimate the boiling point of each aqueous solution. The boiling point of pure water is 100.0°C.
    a. 0.50 M NaCl

b. 1.5 M Na_2SO_4
c. 2.0 M $C_6H_{12}O_6$

6. Estimate the freezing point of each aqueous solution. The freezing point of pure water is 0.0°C.

   a. 0.50 M NaCl
   b. 1.5 M Na_2SO_4
   c. 2.0 M $C_6H_{12}O_6$

7. Explain why salt (NaCl) is spread on roads and sidewalks to inhibit ice formation in cold weather.

8. Salt (NaCl) and calcium chloride ($CaCl_2$) are used widely in some areas to minimize the formation of ice on sidewalks and roads. One of these ionic compounds is better, mole for mole, at inhibiting ice formation. Which is that likely to be? Why?

9. What is the osmolarity of each aqueous solution?

   a. 0.500 M NH_2CONH_2
   b. 0.500 M NaBr
   c. 0.500 M $Ca(NO_3)_2$

10. What is the osmolarity of each aqueous solution?

    a. 0.150 M KCl
    b. 0.450 M $(CH_3)_2CHOH$
    c. 0.500 M $Ca_3(PO_4)_2$

11. A 1.0 M solution of an unknown soluble salt has an osmolarity of 3.0 osmol. What can you conclude about the salt?

12. A 1.5 M NaCl solution and a 0.75 M $Al(NO_3)_3$ solution exist on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and the direction of solvent flow, if any, across the membrane.

Answers

1. a. 1.0 M NaCl
   b. 1.0 M NaCl
   c. 1.0 M $(NH_4)_3PO_4$

3. a. 1.0 M NaCl
   b. 1.0 M NaCl
   c. 1.0 M $(NH_4)_3PO_4$

5. a. 100.5°C
   b. 102.3°C
   c. 101°C

7. NaCl lowers the freezing point of water, so it needs to be colder for the water to freeze.

9. a. 0.500 osmol
   b. 1.000 osmol
   c. 1.500 osmol

11. It must separate into three ions when it dissolves.

This page titled 14.5: Properties of Solutions is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

14.E: Solutions (Exercises)

9.1: Solutions

Concept Review Exercises

1. What causes a solution to form?
2. How does the phrase *like dissolves like* relate to solutions?

Answers

1. Solutions form because a solute and a solvent have similar intermolecular interactions.
2. It means that substances with similar intermolecular interactions will dissolve in each other.

Exercises

1. Define *solution*.

2. Give several examples of solutions.

3. What is the difference between a solvent and a solute?

4. Can a solution have more than one solute in it? Can you give an example?

5. Does a solution have to be a liquid? Give several examples to support your answer.

6. Give at least two examples of solutions found in the human body.

7. Which substances will probably be soluble in water, a very polar solvent?

    a. sodium nitrate ($NaNO_3$)
    b. hexane (C_6H_{14})
    c. isopropyl alcohol [$(CH_3)_2CHOH$]
    d. benzene (C_6H_6)

8. Which substances will probably be soluble in toluene ($C_6H_5CH_3$), a nonpolar solvent?

    a. sodium nitrate ($NaNO_3$)
    b. hexane (C_6H_{14})
    c. isopropyl alcohol [$(CH_3)_2CHOH$]
    d. benzene (C_6H_6)

9. The solubility of alcohols in water varies with the length of carbon chain. For example, ethanol (CH_3CH_2OH) is soluble in water in any ratio, while only 0.0008 mL of heptanol ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$) will dissolve in 100 mL of water. Propose an explanation for this behavior.

10. Dimethyl sulfoxide [$(CH_3)_2SO$] is a polar liquid. Based on the information in Exercise 9, which do you think will be more soluble in it—ethanol or heptanol?

Answers

1. a homogeneous mixture

2. vinegar, dextrose IV, saline IV, coffee, tea, wine

3. A solvent is the majority component of a solution; a solute is the minority component of a solution.

4. yes. Coke or Pepsi has sugar, caffeine and carbon dioxide as solutes.

5. A solution does not have to be liquid; air is a gaseous solution, while some alloys are solid solutions (answers will vary).

6. Urine, plasma

7. a. probably soluble
   b. probably not soluble
   c. probably soluble
   d. probably not soluble

8.

   a. probably not soluble
   b. probably soluble
   c. probably not soluble
   d. probably soluble

9. Small alcohol molecules have strong polar intermolecular interactions, so they dissolve in water. In large alcohol molecules, the nonpolar end overwhelms the polar end, so they do not dissolve very well in water.

10. Ethanol is a smaller molecule. It will be more soluble in water than heptanol.

9.2: Concentration

Concept Review Exercises

1. What are some of the units used to express concentration?

2. Distinguish between the terms *solubility* and *concentration*.

Answers

1. % m/m, % m/v, ppm, ppb, molarity, and Eq/L (answers will vary)

2. Solubility is typically a limit to how much solute can dissolve in a given amount of solvent. Concentration is the quantitative amount of solute dissolved at any concentration in a solvent.

Exercises

1. Define *solubility*. Do all solutes have the same solubility?

2. Explain why the terms *dilute* or *concentrated* are of limited usefulness in describing the concentration of solutions.

3. If the solubility of sodium chloride (NaCl) is 30.6 g/100 mL of H_2O at a given temperature, how many grams of NaCl can be dissolved in 250.0 mL of H_2O?

4. If the solubility of glucose ($C_6H_{12}O_6$) is 120.3 g/100 mL of H_2O at a given temperature, how many grams of $C_6H_{12}O_6$ can be dissolved in 75.0 mL of H_2O?

5. How many grams of sodium bicarbonate ($NaHCO_3$) can a 25.0°C saturated solution have if 150.0 mL of H_2O is used as the solvent?

6. If 75.0 g of potassium bromide (KBr) are dissolved in 125 mL of H_2O, is the solution saturated, unsaturated, or supersaturated?

7. Calculate the mass/mass percent of a saturated solution of NaCl. Use the data from Table $14.E.1$ "Solubilities of Various Solutes in Water at 25°C (Except as Noted)", assume that masses of the solute and the solvent are additive, and use the density of H_2O (1.00 g/mL) as a conversion factor.

8. Calculate the mass/mass percent of a saturated solution of $MgCO_3$ Use the data from Table $14.E.1$ "Solubilities of Various Solutes in Water at 25°C (Except as Noted)", assume that masses of the solute and the solvent are additive, and use the density of H_2O (1.00 g/mL) as a conversion factor.

9. Only 0.203 mL of C_6H_6 will dissolve in 100.000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of benzene in water.

10. Only 35 mL of aniline ($C_6H_5NH_2$) will dissolve in 1,000 mL of H_2O. Assuming that the volumes are additive, find the volume/volume percent of a saturated solution of aniline in water.

11. A solution of ethyl alcohol (C_2H_5OH) in water has a concentration of 20.56% v/v. What volume of C_2H_5OH is present in 255 mL of solution?

12. What mass of KCl is present in 475 mL of a 1.09% m/v aqueous solution?

13. The average human body contains 5,830 g of blood. What mass of arsenic is present in the body if the amount in blood is 0.55 ppm?

14. The Occupational Safety and Health Administration has set a limit of 200 ppm as the maximum safe exposure level for carbon monoxide (CO). If an average breath has a mass of 1.286 g, what is the maximum mass of CO that can be inhaled at that maximum safe exposure level?

15. Which concentration is greater—15 ppm or 1,500 ppb?

16. Express the concentration 7,580 ppm in parts per billion.

17. What is the molarity of 0.500 L of a potassium chromate solution containing 0.0650 mol of K_2CrO_4?

18. What is the molarity of 4.50 L of a solution containing 0.206 mol of urea [$(NH_2)_2CO$]?

19. What is the molarity of a 2.66 L aqueous solution containing 56.9 g of NaBr?

20. If 3.08 g of $Ca(OH)_2$ is dissolved in enough water to make 0.875 L of solution, what is the molarity of the $Ca(OH)_2$?

21. What mass of HCl is present in 825 mL of a 1.25 M solution?

22. What mass of isopropyl alcohol (C_3H_8O) is dissolved in 2.050 L of a 4.45 M aqueous C_3H_8O solution?

23. What volume of 0.345 M NaCl solution is needed to obtain 10.0 g of NaCl?

24. How many milliliters of a 0.0015 M cocaine hydrochloride ($C_{17}H_{22}ClNO_4$) solution is needed to obtain 0.010 g of the solute?

25. Aqueous calcium chloride reacts with aqueous silver nitrate according to the following balanced chemical equation:

$$CaCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

How many moles of AgCl(s) are made if 0.557 L of 0.235 M $CaCl_2$ react with excess $AgNO_3$? How many grams of AgCl are made?

26. Sodium bicarbonate ($NaHCO_3$) is used to react with acid spills. The reaction with sulfuric acid (H_2SO_4) is as follows:

$$2NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(\ell) + 2CO_2(g)$$

If 27.6 mL of a 6.25 M H_2SO_4 solution were spilled, how many moles of $NaHCO_3$ would be needed to react with the acid? How many grams of $NaHCO_3$ is this?

27. The fermentation of glucose to make ethanol and carbon dioxide has the following overall chemical equation:

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

If 1.00 L of a 0.567 M solution of $C_6H_{12}O_6$ were completely fermented, what would be the resulting concentration of the C_2H_5OH solution? How many moles of CO_2 would be formed? How many grams is this? If each mole of CO_2 had a volume of 24.5 L, what volume of CO_2 is produced?

28. Aqueous sodium bisulfite gives off sulfur dioxide gas when heated:

$$2NaHSO_3(aq) \rightarrow Na_2SO_3(aq) + H_2O(\ell) + SO_2(g)$$

If 567 mL of a 1.005 M $NaHSO_3$ solution were heated until all the $NaHSO_3$ had reacted, what would be the resulting concentration of the Na_2SO_3 solution? How many moles of SO_2 would be formed? How many grams of SO_2 would be formed? If each mole of SO_2 had a volume of 25.78 L, what volume of SO_2 would be produced?

29. What is the concentration of a 1.0 M solution of $K^+(aq)$ ions in equivalents/liter?

30. What is the concentration of a 1.0 M solution of $SO_4^{2-}(aq)$ ions in equivalents/liter?

31. A solution having initial concentration of 0.445 M and initial volume of 45.0 mL is diluted to 100.0 mL. What is its final concentration?

32. A 50.0 mL sample of saltwater that is 3.0% m/v is diluted to 950 mL. What is its final mass/volume percent?

Answers

1. Solubility is the amount of a solute that can dissolve in a given amount of solute, typically 100 mL. The solubility of solutes varies widely.

2. The term *dilute* means relatively less solute and the term *concentrated* implies relatively more solute. Both are of limited usefulness because these are not accurate.
3. 76.5 g
4. 90.2 g

5. 12.6 g

6. unsaturated

7. 26.5%

8. 2.15%

9. 0.203%

10. 3.4%

11. 52.4 mL

12. 5.18 g

13. 0.00321 g

14. 2.57×10^{-4} g

15. 15 ppm

16. 7,580,000 ppb

17. 0.130 M
18. 0.0458 M

19. 0.208 M

20. 0.0475 M

21. 37.6 g

22. 548 g

23. 0.496 L

24. 20 mL

25. 0.262 mol; 37.5 g

26. 0.345 mol; 29.0 g

27. 1.13 M C_2H_5OH; 1.13 mol of CO_2; 49.7 g of CO_2; 27.7 L of CO_2
28. 0.503 M Na_2SO_3; 0.285 mol SO_2; 18.3 g SO_2; 471 L SO_2
29. 1.0 Eq/L
30. 2.0 Eq/L
31. 0.200 M

32. 0.16 % m/v

9.3: The Dissolution Process

Concept Review Exercise

1. Explain how the solvation process describes the dissolution of a solute in a solvent.

Answer

1. Each particle of the solute is surrounded by particles of the solvent, carrying the solute from its original phase.

Exercises

1. Describe what happens when an ionic solute like Na_2SO_4 dissolves in a polar solvent.

2. Describe what happens when a molecular solute like sucrose ($C_{12}H_{22}O_{11}$) dissolves in a polar solvent.

3. Classify each substance as an electrolyte or a nonelectrolyte. Each substance dissolves in H_2O to some extent.

    a. NH_4NO_3

b. CO_2
c. NH_2CONH_2
d. HCl

4. Classify each substance as an electrolyte or a nonelectrolyte. Each substance dissolves in H_2O to some extent.

a. $CH_3CH_2CH_2OH$
b. $Ca(CH_3CO_2)_2$
c. I_2
d. KOH

5. Will solutions of each solute conduct electricity when dissolved?

a. $AgNO_3$
b. $CHCl_3$
c. $BaCl_2$
d. Li_2O

6. Will solutions of each solute conduct electricity when dissolved?

a. CH_3COCH_3
b. $N(CH_3)_3$
c. $CH_3CO_2C_2H_5$
d. $FeCl_2$

Answers

1. Each ion of the ionic solute is surrounded by particles of solvent, carrying the ion from its associated crystal.

2. Each sucrose molecule is surrounded by solvent molecules (attracted to each other via intermolecular forces of attraction).

3. a. electrolyte
b. nonelectrolyte
c. nonelectrolyte
d. electrolyte

4.

a. nonelectrolyte
b. electrolyte
c. nonelectrolyte
d. electrolyte

5.

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

6.

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

9.4: Properties of Solutions

Concept Review Exercises

1. What are the colligative properties of solutions?

2. Explain how the following properties of solutions differ from those of the pure solvent: vapor pressure, boiling point, freezing point, and osmotic pressure.

Answers

1. Colligative properties are characteristics that a solution has that depend on the number, not the identity, of solute particles.

2. In solutions, the vapor pressure is lower, the boiling point is higher, the freezing point is lower, and the osmotic pressure is higher.

Exercises

1. In each pair of aqueous systems, which will have the lower vapor pressure?

    a. pure water or 1.0 M NaCl
    b. 1.0 M NaCl or 1.0 M $C_6H_{12}O_6$
    c. 1.0 M $CaCl_2$ or 1.0 M $(NH_4)_3PO_4$

2. In each pair of aqueous systems, which will have the lower vapor pressure?

    a. 0.50 M $Ca(NO_3)_2$ or 1.0 M KBr
    b. 1.5 M $C_{12}H_{22}O_{11}$ or 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$ or pure water

3. In each pair of aqueous systems, which will have the higher boiling point?

    a. pure water or a 1.0 M NaCl
    b. 1.0 M NaCl or 1.0 M $C_6H_{12}O_6$
    c. 1.0 M $CaCl_2$ or 1.0 M $(NH_4)_3PO_4$

4. In each pair of aqueous systems, which will have the higher boiling point?

    a. 0.50 M $Ca(NO_3)_2$ or 1.0 M KBr
    b. 1.5 M $C_{12}H_{22}O_{11}$ or 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$ or pure water

5. Estimate the boiling point of each aqueous solution. The boiling point of pure water is 100.0°C.

    a. 0.50 M NaCl
    b. 1.5 M Na_2SO_4
    c. 2.0 M $C_6H_{12}O_6$

6. Estimate the freezing point of each aqueous solution. The freezing point of pure water is 0.0°C.

    a. 0.50 M NaCl
    b. 1.5 M Na_2SO_4
    c. 2.0 M $C_6H_{12}O_6$

7. Explain why salt (NaCl) is spread on roads and sidewalks to inhibit ice formation in cold weather.

8. Salt (NaCl) and calcium chloride ($CaCl_2$) are used widely in some areas to minimize the formation of ice on sidewalks and roads. One of these ionic compounds is better, mole for mole, at inhibiting ice formation. Which is that likely to be? Why?

9. What is the osmolarity of each aqueous solution?

    a. 0.500 M NH_2CONH_2
    b. 0.500 M NaBr
    c. 0.500 M $Ca(NO_3)_2$

10. What is the osmolarity of each aqueous solution?

    a. 0.150 M KCl
    b. 0.450 M $(CH_3)_2CHOH$
    c. 0.500 M $Ca_3(PO_4)_2$

11. A 1.0 M solution of an unknown soluble salt has an osmolarity of 3.0 osmol. What can you conclude about the salt?

12. A 1.5 M NaCl solution and a 0.75 M $Al(NO_3)_3$ solution exist on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and the direction of solvent flow, if any, across the membrane.

Answers

1. a. 1.0 M NaCl
   b. 1.0 M NaCl
   c. 1.0 M $(NH_4)_3PO_4$

2.

    a. 1.0 M KBr
    b. 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$

3.

    a. 1.0 M NaCl
    b. 1.0 M NaCl
    c. 1.0 M $(NH_4)_3PO_4$

4.

    a. 1.0 M KBr
    b. 0.75 M $Ca(OH)_2$
    c. 0.10 M $Cu(NO_3)_2$

5.

    a. 100.5°C
    b. 102.3°C
    c. 101°C

6.

    a. -1.9°C
    b. -8.6°C
    c. -3.8°C

7. NaCl lowers the freezing point of water, so it needs to be colder for the water to freeze.

8. $CaCl_2$ splits up into 3 ions while NaCl splits up into 2 ions only. $CaCl_2$ will be more effective.

9.

a. 0.500 osmol
b. 1.000 osmol
c. 1.500 osmol

10.

a. 0.300 osmol
b. 0.450 osmol
c. 2.50 osmol

11. It must separate into three ions when it dissolves.

12. Both NaCl and $Al(NO_3)_3$ have 3.0 osmol. There will be no net difference in the solvent flow.

9.5: Chemical Equilibrium

Concept Review Exercises

1. What is chemical equilibrium?

2. What does the equilibrium constant tell us?

Answers

1. The rate of the forward reaction equals the rate of the reverse reaction.

2. The ratio of products and reactants when the system is at equilibrium.

EXERCISES

1. If the reaction $H_2 + I_2 \diamondsuit 2HI$ is at equilibrium, do the concentrations of HI, H_2, and I_2 have to be equal?

2. Do the concentrations at equilibrium depend upon how the equilibrium was reached?

3. What does it mean if the K_{eq} is > 1?

4. What does it mean if the K_{eq} is < 1?

5. Does the equilibrium state depend on the starting concentrations?

6. Write an expression for the equilibrium constant K equation.

a. $PCl_5\ (g) \rightleftarrows PCl_3\ (g) + Cl_2\ (g)$

b. $2\ O_3\ (g) \rightleftharpoons 3\ O_2\ (g)$

7. Benzene is one of the compounds used as octane enhancers in unleaded gasoline. It is manufactured by the catalytic conversion of acetylene to benzene: $3C_2H_2(g) \diamondsuit C_6H_6(g)$. Which value of *K* would make this reaction most useful commercially?

Explain your answer.

a. $K \approx 0.01$

b. $K \approx 1$

c. $K \approx 10$.

8. Tell whether the reactants or the products are favored at equilibrium:

a. $2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$ $K = 172$

b. $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ $K = 0.230$

c. $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$ $K = 4.6 \times 10^4$

d. $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $K = 0.050$

AnswerS

1. No, the concentrations are constant but the concentrations do not have to be equal.

2. No.

3. More products than reactants are present at equilibrium.

4. More reactants than products present at equilibrium.

5. No. The equilibrium ratio does not depend on the initial concentrations.

6.

a. $K = \frac{[PCl_3][Cl_2]}{[PCl_5]}$

b. $K = \frac{[O_2]^3}{[O_3]^2}$

7. The answer is c. $K \approx 10$.

Since $K = \frac{[C_6H_6]^2}{[C_2H_2]^3}$ ($K \approx 10$), this means that C_6H_6 predominates over C_2H_2. In such a case, the reaction would be commercially feasible if the rate to equilibrium is suitable.

8.

a. products

b. reactants

c. products

d. reactants

9.6: Le Chatelier's Principle

Concept Review Exercises

1. Define Le Chatelier's principle.

2. List the three factors types of changes that can disturb the equilibrium of a system.

Answers

1. Le Chatelier's principle states that a system at equilibrium is disturbed, it will respond in a way to minimize the disturbance.

2. temperature, change in amount of substance, change in pressure through change in volume

EXERCISES

1. How will each change affect the reaction?

$PCl_5(g)$ + heat ⇌ $PCl_3(g) + Cl_2(g)$

a. Addition of PCl_5
b. Addition of Cl_2
c. Removal of PCl_3
d. Increasing temperature
e. Decreasing temperature
f. Decreasing volume

2. How will each change affect the reaction?

$HNO_2(aq)$ ⇌ $H^+(aq) + NO_2^-(aq)$

a. Removal of HNO_2
b. Addition of HCl (i.e. adding more H^+)
c. Increasing volume
d. Decreasing volume
e. Removal of NO_2^-
f. Addition of OH^- (which will react with and remove H^+)

3. How will each change affect the reaction?

$CO_2(g) + C(s)$ ⇌ $2CO(g)$ $\Delta H=172.5kJ$

a. Addition of CO_2
b. Removal of CO_2
c. Increasing temperature
d. Decreasing temperature
e. Increasing volume
f. Addition of CO

4. How will each change affect the reaction?

$H_2(g) + I_2(g)$ ⇌ $2HI(g)$ $\Delta H=-9.48kJ$

a. Addition of H_2
b. Removal of H_2
c. Increasing temperature
d. Decreasing temperature
e. Increasing volume
f. Decreasing volume

AnswerS

1.

a. shift right
b. shift left
c. shift right
d. shift right
e. shift left
f. shift left

2.

a. shift left
b. shift left
c. no effect
d. no effect
e. shift right
f. shift right

3.

a. shift right
b. shift left
c. shift right
d. shift left
e. shift right
f. shift left

4.

a. shift right
b. shift left
c. shift left
d. shift right
e. no effect
f. no effect

9.7: Osmosis and Diffusion

Concept Review Exercises

1. What are some of the features of a semipermeable membrane?

2. What do the prefixes hyper, hypo, and iso mean?

Answers

1. A semipermeable membrane allows some substances to pass through but not others.

2. hyper – higher; hypo – lower; iso - same

EXERCISES

1. Two solutions are separated by a semipermeable membrane. Solution A contains 25.0 g of NaCl in 100.0 mL of water and solution B contains 35.0 g of NaCl in 100.0 mL of water.

a. Which one has a higher concentration?
b. Which way will water molecules flow?
c. Which volume will increase?
d. Which volume will decrease?
e. What will happen to the concentration of solution A?
f. What will happen to the concentration of solution B?

2. Two solutions with different concentrations and compositions are separated by a semipermeable membrane. The left-hand solution is a .50 M solution of $MgSO_4$, while the right-hand solution contains $CaCl_2$ at a concentration of .40 M. Determine the direction of the flow of solvent, left or right.

3. Given the following situations, wherein two tanks of different solutions are separated by a semipermeable membrane, determine the direction of the flow of solvent (water).

a. Solution A contains a 0.40 M concentration of $CaCl_2$, while Solution B contains a 0.45 M concentration of KI

b. Solution A contains a 1.00 M concentration of NH_4Cl, while Solution B contains a 1.00 M concentration of CH_2O

4. Cells are placed in a solution and the cells then undergo hemolysis. What can be said about the relative concentrations of solute in the cell and the solution?

5. Describe the relative concentrations inside and outside a red blood cell when crenation occurs.

6. A saltwater fish is placed in a freshwater tank. What will happen to the fish? Describe the flow of water molecules to explain the outcome.

7. What makes up the "head" region of a phospholipid? Is it hydrophobic or hydrophilic?

8. What makes up the "tail" region of a phospholipid? Is it hydrophobic or hydrophilic?

Answers

1. Two solutions are separated by a semipermeable membrane. Solution A contains 25.0 g of NaCl in 100.0 mL of water and solution B contains 35.0 g of NaCl in 100.0 mL of water.

a. Solution B
b. A → → B
c. B
d. A
e. increase
f. decrease

2. Water (solvent) flows from left to right.

3.

a. Water flows from Solution B to Solution A.

b. Water flows from Solution B to Solution A.

4. Cells contain fluid with higher concentration than solution outside the cell.

5. Cells contain fluid with a lower concentration than the solution outside the cell.

6. Water molecules will flow from the tank water into the fish because the fish has a higher concentration of salt. If the fish absorbs too much water, it will die.

7. The "head" region is a phosphate group and it is hydrophilic.

8. The "tail" is a hydrocarbon tail and it is hydrophobic.

Contributors

-

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

Additional Exercises

1. Calcium nitrate reacts with sodium carbonate to precipitate solid calcium carbonate:

$$Ca(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow CaCO_{3(s)} + NaNO_{3(aq)} \tag{14.E.1}$$

   a. Balance the chemical equation.
   b. How many grams of Na_2CO_3 are needed to react with 50.0 mL of 0.450 M $Ca(NO_3)_2$?
   c. Assuming that the Na_2CO_3 has a negligible effect on the volume of the solution, find the osmolarity of the $NaNO_3$ solution remaining after the $CaCO_3$ precipitates from solution.

2. The compound HCl reacts with sodium carbonate to generate carbon dioxide gas:

$$HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow H_2O_{(\ell)} + CO_{2(g)} + NaCl_{(aq)} \tag{14.E.2}$$

   a. Balance the chemical equation.
   b. How many grams of Na_2CO_3 are needed to react with 250.0 mL of 0.755 M HCl?
   c. Assuming that the Na_2CO_3 has a negligible effect on the volume of the solution, find the osmolarity of the NaCl solution remaining after the reaction is complete.

3. Estimate the freezing point of concentrated aqueous HCl, which is usually sold as a 12 M solution. Assume complete ionization into H^+ and Cl^- ions.

4. Estimate the boiling point of concentrated aqueous H_2SO_4, which is usually sold as an 18 M solution. Assume complete ionization into H^+ and HSO_4^- ions.

5. Seawater can be approximated by a 3.0% m/m solution of NaCl in water. Determine the molarity and osmolarity of seawater. Assume a density of 1.0 g/mL.

6. Human blood can be approximated by a 0.90% m/m solution of NaCl in water. Determine the molarity and osmolarity of blood. Assume a density of 1.0 g/mL.

7. How much water must be added to 25.0 mL of a 1.00 M NaCl solution to make a resulting solution that has a concentration of 0.250 M?

8. Sports drinks like Gatorade are advertised as capable of resupplying the body with electrolytes lost by vigorous exercise. Find a label from a sports drink container and identify the electrolytes it contains. You should be able to identify several simple ionic compounds in the ingredients list.

9. Occasionally we hear a sensational news story about people stranded in a lifeboat on the ocean who had to drink their own urine to survive. While distasteful, this act was probably necessary for survival. Why not simply drink the ocean water? (Hint: See Exercise 5 and Exercise 6 above. What would happen if the two solutions in these exercises were on opposite sides of a semipermeable membrane, as we would find in our cell walls?)

Answers

1. a. $Ca(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaNO_3(aq)$
   b. 2.39 g
   c. 1.80 osmol

2.

   a. $2HCl\ (aq) + Na_2CO_3(aq) \rightarrow H_2O(\ell) + CO_2(g) + 2NaCl\ (aq)$

   b. 10.0 g

   c. 1.51 M

3. –45.6°C

4. 118°C

5. 0.513 M; 1.026 osmol

6. molarity = 0.15 M; osmolarity = 0.31 M

7. 75.0 mL

8. magnesium chloride, calcium chloride (answers may vary)

9. The osmotic pressure of seawater is too high. Drinking seawater would cause water to go from inside our cells into the more concentrated seawater, ultimately killing the cells.

This page titled 14.E: Solutions (Exercises) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

- 9.1: Solutions by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 9.2: Concentration by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 9.3: The Dissolution Process by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 9.4: Properties of Solutions by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 9.5: Chemical Equilibrium by CK12 is licensed CK-12. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.
- 9.6: Le Chatelier's Principle by CK12 is licensed CK-12.
- 8.4: Osmosis and Diffusion by CK12 is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

14.S: Solutions (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

A **solution** is a homogeneous mixture. The major component is the **solvent**, while the minor component is the **solute**. Solutions can have any phase; for example, an **alloy** is a solid solution. Solutes are **soluble** or **insoluble**, meaning they dissolve or do not dissolve in a particular solvent. The terms **miscible** and **immiscible**, instead of soluble and insoluble, are used for liquid solutes and solvents. The statement *like dissolves like* is a useful guide to predicting whether a solute will dissolve in a given solvent.

The amount of solute in a solution is represented by the **concentration** of the solution. The maximum amount of solute that will dissolve in a given amount of solvent is called the **solubility** of the solute. Such solutions are **saturated**. Solutions that have less than the maximum amount are **unsaturated**. Most solutions are unsaturated, and there are various ways of stating their concentrations. **Mass/mass percent**, **volume/volume percent**, and **mass/volume percent** indicate the percentage of the overall solution that is solute. **Parts per million (ppm)** and **parts per billion (ppb)** are used to describe very small concentrations of a solute. **Molarity**, defined as the number of moles of solute per liter of solution, is a common concentration unit in the chemistry laboratory. **Equivalents** express concentrations in terms of moles of charge on ions. When a solution is diluted, we use the fact that the amount of solute remains constant to be able to determine the volume or concentration of the final diluted solution.

Dissolving occurs by **solvation**, the process in which particles of a solvent surround the individual particles of a solute, separating them to make a solution. For water solutions, the word **hydration** is used. If the solute is molecular, it dissolves into individual molecules. If the solute is ionic, the individual ions separate from each other, forming a solution that conducts electricity. Such solutions are called **electrolytes**. If the dissociation of ions is complete, the solution is a **strong electrolyte**. If the dissociation is only partial, the solution is a **weak electrolyte**. Solutions of molecules do not conduct electricity and are called **nonelectrolytes**.

Solutions have properties that differ from those of the pure solvent. Some of these are **colligative** properties, which are due to the number of solute particles dissolved, not the chemical identity of the solute. Colligative properties include **vapor pressure depression**, **boiling point elevation**, **freezing point depression**, and **osmotic pressure**. Osmotic pressure is particularly important in biological systems. It is caused by **osmosis**, the passage of solvents through certain membranes like cell walls. The **osmolarity** of a solution is the product of a solution's molarity and the number of particles a solute separates into when it dissolves. Osmosis can be reversed by the application of pressure; this reverse osmosis is used to make fresh water from saltwater in some parts of the world. Because of osmosis, red blood cells placed in hypotonic or hypertonic solutions lose function through either hemolysis or crenation. If they are placed in isotonic solutions, however, the cells are unaffected because osmotic pressure is equal on either side of the cell membrane.

This page titled 14.S: Solutions (Summary) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15: ACIDS AND BASES

Many of us are familiar with the group of chemicals called acids. But do you know what it takes for a compound to be an acid? Actually, there are several different definitions of acid that chemistry uses, and each definition is appropriate under different circumstances. Less familiar—but just as important to chemistry and ultimately to us—is the group of chemicals known as bases. Both acids and bases are important enough that we devote an entire chapter to them—their properties and their reactions.

15.1: Prelude to Acids and Bases

15.2: Arrhenius Definition of Acids and Bases

15.3: Brønsted-Lowry Definition of Acids and Bases

15.4: Water - Both an Acid and a Base

15.5: The Strengths of Acids and Bases

15.6: Buffers

15.E: Acids and Bases (Exercises)

15.S: Acids and Bases (Summary)

This page titled 15: Acids and Bases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.1: Prelude to Acids and Bases

One of the most concentrated acids in the body is stomach acid, which can be approximated as a 0.05 M hydrochloric acid solution. Special cells in the stomach wall secrete this acid, along with special enzymes, as part of the digestion process. In a laboratory, a 0.05 M solution of hydrochloric acid would dissolve some metals. How does the stomach survive the presence of such a reactive acid?

Actually, the stomach has several mechanisms for withstanding this chemical onslaught. First, the lining of the stomach is coated with a thin layer of mucus that contains some bicarbonate ions (HCO_3^-). These react with the hydrochloric acid to produce water, carbon dioxide, and harmless chloride ions. If any acid penetrates through the mucus, it can attack the surface layer of stomach cells, called the *gastric epithelium*. Cells in the gastric epithelium are being constantly shed, so damaged cells are quickly removed and replaced with healthy cells.

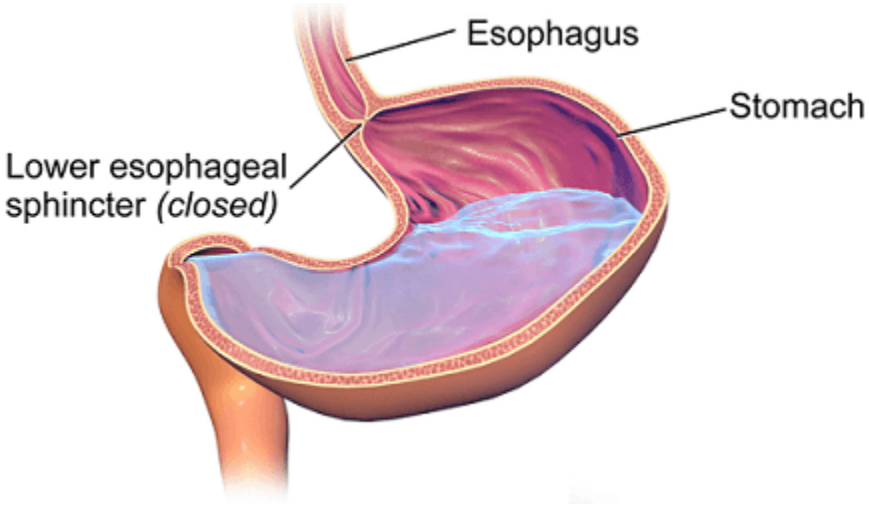


Figure 15.1.1: Stomach with acidic digestive juices. Image used with permissin (CC BY-SA 4.0; BruceBlaus).

However, if the gastric epithelium is destroyed faster than it can be replaced, the acid may reach the wall of the stomach, resulting in ulcers. If an ulcer grows large enough, it can expose blood vessels in the stomach wall, causing bleeding. In extreme situations, the loss of blood through a severe ulcer can threaten a person's health.

Ulcers can also result from the presence of a certain bacterium—*Helicobacter pylori*—in the stomach. The mechanism for this ulcer formation is not the same as that for ulcers caused by stomach acid and is not completely understood. However, there are two main treatments for ulcers: (1) *antacids* to react chemically with excess hydrochloric acid in the stomach and (2) *antibiotics* to destroy the *H. pylori* bacteria in the stomach.

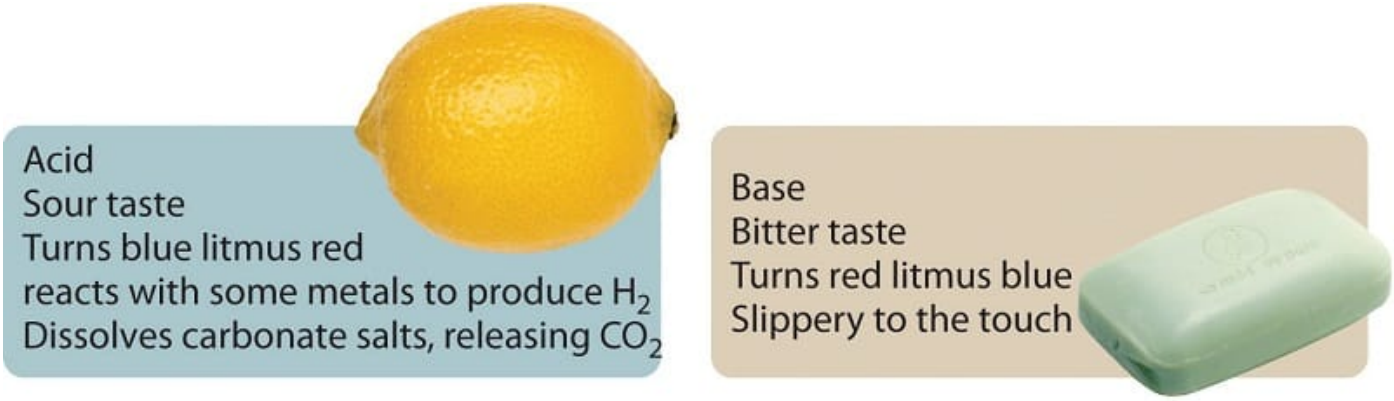
This page titled 15.1: Prelude to Acids and Bases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.2: Arrhenius Definition of Acids and Bases

Learning Objectives

- To recognize a compound as an Arrhenius acid or an Arrhenius base.

One way to define a class of compounds is by describing the various characteristics its members have in common. In the case of the compounds known as acids, the common characteristics include a sour taste, the ability to change the color of the vegetable dye *litmus* to red, and the ability to dissolve certain metals and simultaneously produce hydrogen gas. For the compounds called bases, the common characteristics are a slippery texture, a bitter taste, and the ability to change the color of litmus to blue. Acids and bases also react with each other to form compounds generally known as salts.



Although we include their tastes among the common characteristics of acids and bases, we never advocate tasting an unknown chemical!

Chemists prefer, however, to have definitions for acids and bases in chemical terms. The Swedish chemist Svante Arrhenius developed the first chemical definitions of acids and bases in the late 1800s. Arrhenius defined an acid as a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution. Many acids are simple compounds that release a hydrogen cation into solution when they dissolve. Similarly, Arrhenius defined a base as a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution. Many bases are ionic compounds that have the hydroxide ion as their anion, which is released when the base dissolves in water.

Table 15.2.1*: Formulas and Names for Some Acids and Bases*

Acids		Bases	
Formula	Name	Formula	Name
HCl(aq)	hydrochloric acid	NaOH(aq)	sodium hydroxide
HBr(aq)	hydrobromic acid	KOH(aq)	potassium hydroxide
HI(aq)	hydriodic acid	$Mg(OH)_2$(aq)	magnesium hydroxide
H_2S(aq)	hydrosulfuric acid	$Ca(OH)_2$(aq)	calcium hydroxide
$HC_2H_3O_2$(aq)	acetic acid	NH_3(aq)	ammonia
HNO_3(aq)	nitric acid		
HNO_2(aq)	nitrous acid		
H_2SO_4(aq)	sulfuric acid		

Acids		Bases	
$H_2SO_3(aq)$	sulfurous acid		
$HClO_3(aq)$	chloric acid		
$HClO_4(aq)$	perchloric acid		
$HClO_2(aq)$	chlorous acid		
$H_3PO_4(aq)$	phosphoric acid		
$H_3PO_3(aq)$	phosphorous acid		

Many bases and their aqueous solutions are named using the normal rules of ionic compounds that were presented previously; that is, they are named as hydroxide compounds. For example, the base sodium hydroxide (NaOH) is both an ionic compound and an aqueous solution. However, aqueous solutions of acids have their own naming rules. The names of *binary acids* (compounds with hydrogen and one other element in their formula) are based on the root of the name of the other element preceded by the prefix *hydro-* and followed by the suffix *-ic acid*. Thus, an aqueous solution of HCl [designated "HCl(aq)"] is called hydrochloric acid, $H_2S(aq)$ is called hydrosulfuric acid, and so forth. Acids composed of more than two elements (typically hydrogen and oxygen and some other element) have names based on the name of the other element, followed by the suffix *-ic acid* or *-ous acid*, depending on the number of oxygen atoms in the acid's formula. Other prefixes, like *per-* and *hypo-*, also appear in the names for some acids. Unfortunately, there is no strict rule for the number of oxygen atoms that are associated with the *-ic acid* suffix; the names of these acids are best memorized. Table 15.2.1 lists some acids and bases and their names. Note that acids have hydrogen written first, as if it were the cation, while most bases have the negative hydroxide ion, if it appears in the formula, written last.

> The name oxygen comes from the Latin meaning "acid producer" because its discoverer, Antoine Lavoisier, thought it was the essential element in acids. Lavoisier was wrong, but it is too late to change the name now.

Example 15.2.1

Name each substance.

a. HF(aq)
b. $Sr(OH)_2(aq)$

SOLUTION

a. This acid has only two elements in its formula, so its name includes the *hydro-* prefix. The stem of the other element's name, fluorine, is *fluor*, and we must also include the *-ic acid* ending. Its name is hydrofluoric acid.
b. This base is named as an ionic compound between the strontium ion and the hydroxide ion: strontium hydroxide.

Exercise 15.2.1

Name each substance.

a. $H_2Se(aq)$
b. $Ba(OH)_2(aq)$

Notice that one base listed in Table 15.2.1—ammonia—does not have hydroxide as part of its formula. How does this compound increase the amount of hydroxide ion in aqueous solution? Instead of dissociating into hydroxide ions, ammonia molecules react with water molecules by taking a hydrogen ion from the water molecule to produce an ammonium ion and a hydroxide ion:

$$NH_{3(aq)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{15.2.1}$$

Because this reaction of ammonia with water causes an increase in the concentration of hydroxide ions in solution, ammonia satisfies the Arrhenius definition of a base. Many other nitrogen-containing compounds are bases because they too react with water to produce hydroxide ions in aqueous solution.

Neutralization

As we noted previously, acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen ion containing compound and the base is a hydroxide ion containing compound, water is also a product. The general reaction is as follows:

$$\text{acid + base} \rightarrow \text{water + salt}$$

The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:

$$NaOH(aq) + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)} \tag{15.2.2}$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(\ell)} \tag{15.2.3}$$

Once a neutralization reaction is properly balanced, we can use it to perform stoichiometry calculations, such as the ones we practiced earlier.

Example 15.2.2

Nitric acid [HNO_3(aq)] can be neutralized by calcium hydroxide [$Ca(OH)_2$(aq)].

a. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.
b. For one reaction, 16.8 g of HNO_3 is present initially. How many grams of $Ca(OH)_2$ are needed to neutralize that much HNO_3?
c. In a second reaction, 805 mL of 0.672 M $Ca(OH)_2$ is present initially. What volume of 0.432 M HNO_3 solution is necessary to neutralize the $Ca(OH)_2$ solution?

SOLUTION

a. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of HNO_3 to provide H^+ ions. The balanced chemical equation is as follows:

$$Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(\ell)$$

The salt formed is calcium nitrate.

b. This calculation is much like the calculations we did in Chapter 6 "Quantities in Chemical Reactions". First we convert the mass of HNO_3 to moles using its molar mass of 1.01 + 14.00 + 3(16.00) = 63.01 g/mol; then we use the balanced chemical equation to determine the related number of moles of $Ca(OH)_2$ needed to neutralize it; and then we convert that number of moles of $Ca(OH)_2$ to the mass of $Ca(OH)_2$ using its molar mass of 40.08 + 2(1.01) + 2(16.00) = 74.10 g/mol.

$$16.8\ \text{g}\ \text{HNO}_3 \times \frac{1\ \text{mol}\ \text{HNO}_3}{63.01\ \text{g}\ \text{HNO}_3} \times \frac{1\ \text{mol}\ \text{Ca(OH)}_2}{2\ \text{mol}\ \text{HNO}_3} \times \frac{74.10\ \text{g}\ \text{Ca(OH)}_2}{1\ \text{mol}\ \text{Ca(OH)}_2} = 9.88\ \text{g}\ \text{Ca(OH)}_2\ \text{needed}$$

c. Having concentration information allows us to employ the skills we developed in Chapter 9. First, we use the concentration and volume data to determine the number of moles of $Ca(OH)_2$ present. Recognizing that 805 mL = 0.805 L,

$$0.672\ \text{M}\ \text{Ca(OH)}_2 = \frac{\text{mol}\ \text{Ca(OH)}_2}{0.805\ \text{L soln}}$$

0.672 M CaOH)$_2$ × (0.805 L soln) = mol Ca(OH)$_2$ = 0.541 mol Ca(OH)$_2$

We combine this information with the proper ratio from the balanced chemical equation to determine the number of moles of HNO_3 needed:

$$0.541 \text{ mol } Ca(OH)_2 \times \frac{2 \text{ mol } HNO_3}{1 \text{ mol } Ca(OH)_2} = 1.08 \text{ mol } HNO_3$$

Now, using the definition of molarity one more time, we determine the volume of acid solution needed:

$$0.432 \text{ M } HNO_3 = \frac{1.08 \text{ mol } HNO_3}{\text{volume of } HNO_3}$$

$$\text{volume of } HNO_3 = \frac{1.08 \text{ mol } HNO_3}{0.432 \text{ M } HNO_3} = 2.50 \text{ L} = 2.50 \times 10^3 \text{ mL } HNO_3$$

Exercise 15.2.2

Hydrocyanic acid [HCN(aq)] can be neutralized by potassium hydroxide [KOH(aq)].

a. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.
b. For one reaction, 37.5 g of HCN is present initially. How many grams of KOH are needed to neutralize that much HCN?
c. In a second reaction, 43.0 mL of 0.0663 M KOH is present initially. What volume of 0.107 M HCN solution is necessary to neutralize the KOH solution?

Hydrocyanic acid (HCN) is one exception to the acid-naming rules that specify using the prefix *hydro-* for binary acids (acids composed of hydrogen and only one other element).

Concept Review Exercises

1. Give the Arrhenius definitions of an acid and a base.
2. What is neutralization?

Answers

1. Arrhenius acid: a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution; Arrhenius base: a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.
2. the reaction of an acid and a base

Key Takeaway

- An Arrhenius acid increases the H^+ ion concentration in water, while an Arrhenius base increases the OH^- ion concentration in water.

This page titled 15.2: Arrhenius Definition of Acids and Bases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.3: Brønsted-Lowry Definition of Acids and Bases

Learning Objectives

1. Recognize a compound as a Brønsted-Lowry acid or a Brønsted-Lowry base.
2. Illustrate the proton transfer process that defines a Brønsted-Lowry acid-base reaction.

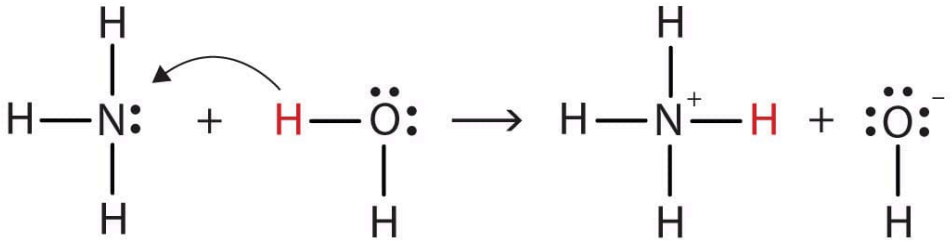
Ammonia (NH_3) increases the hydroxide ion concentration in aqueous solution by reacting with water rather than releasing hydroxide ions directly. In fact, the Arrhenius definitions of an acid and a base focus on hydrogen ions and hydroxide ions. Are there more fundamental definitions for acids and bases?

In 1923, the Danish scientist Johannes Brønsted and the English scientist Thomas Lowry independently proposed new definitions for acids and bases. Rather than considering both hydrogen and hydroxide ions, they focused on the hydrogen ion only. A Brønsted-Lowry acid is a compound that supplies a hydrogen ion in a reaction. A Brønsted-Lowry base, conversely, is a compound that accepts a hydrogen ion in a reaction. Thus, the Brønsted-Lowry definitions of an acid and a base focus on the movement of hydrogen ions in a reaction, rather than on the production of hydrogen ions and hydroxide ions in an aqueous solution.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$NH_{3(aq)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{15.3.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:



Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

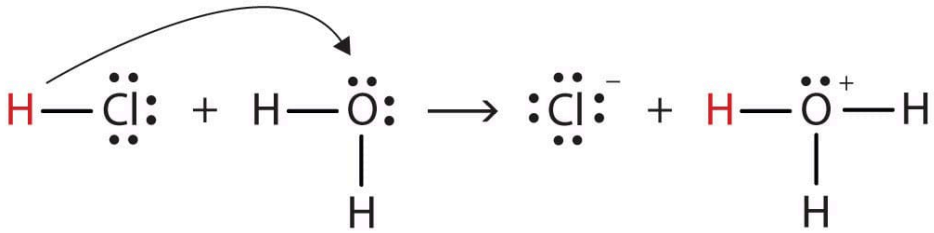
Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+.

With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$HCl_{(g)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)} \tag{15.3.2}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Example 15.3.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

SOLUTION

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$C_6H_5NH_2(aq) + H_2O(\ell) \rightarrow C_6H_5NH_3^+(aq) + OH^-(aq)$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

Exercise 15.3.1

Caffeine ($C_8H_{10}N_4O_2$) is a stimulant found in coffees and teas. When dissolved in water, it can accept a proton from a water molecule. Write the chemical equation for this process and identify the Brønsted-Lowry acid and base.

The Brønsted-Lowry definitions of an acid and a base can be applied to chemical reactions that occur in solvents other than water. The following example illustrates.

Example 15.3.2

Sodium amide ($NaNH_2$) dissolves in methanol (CH_3OH) and separates into sodium ions and amide ions (NH_2^-). The amide ions react with methanol to make ammonia and the methoxide ion (CH_3O^-). Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

SOLUTION

The equation for the reaction is between NH_2^- and CH_3OH to make NH_3 and CH_3O^- is as follows:

$$NH_2^-(solv) + CH_3OH(\ell) \rightarrow NH_3(solv) + CH_3O^-(solv)$$

The label *(solv)* indicates that the species are dissolved in some solvent, in contrast to *(aq)*, which specifies an aqueous (H_2O) solution. In this reaction, we see that the NH_2^- ion accepts a proton from a CH_3OH molecule to make an NH_3 molecule. Thus, as the proton acceptor, NH_2^- is the Brønsted-Lowry base. As the proton donor, CH_3OH is the Brønsted-Lowry acid.

Exercise 15.3.2

Pyridinium chloride (C_5H_5NHCl) dissolves in ethanol (C_2H_5OH) and separates into pyridinium ions ($C_5H_5NH^+$) and chloride ions. The pyridinium ion can transfer a hydrogen ion to a solvent molecule. Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

To Your Health: Brønsted-Lowry Acid-Base Reactions in Pharmaceuticals

There are many interesting applications of Brønsted-Lowry acid-base reactions in the pharmaceutical industry. For example, drugs often need to be water soluble for maximum effectiveness. However, many complex organic compounds are not soluble or are only slightly soluble in water. Fortunately, those drugs that contain proton-accepting nitrogen atoms (and there are a lot of them) can be reacted with dilute hydrochloric acid [HCl(aq)]. The nitrogen atoms—acting as Brønsted-Lowry bases—accept the hydrogen ions from the acid to make an ion, which is usually much more soluble in water. The modified drug molecules can then be isolated as chloride salts:

$$\mathrm{RN(sl\,aq) + H^+(aq) \rightarrow RNH^+(aq) \xrightarrow{Cl^-(aq)} RNHCl(s)} \tag{15.3.3}$$

where RN represents some organic compound containing nitrogen. The label *(sl aq)* means “slightly aqueous,” indicating that the compound RN is only slightly soluble. Drugs that are modified in this way are called *hydrochloride salts*. Examples include the powerful painkiller codeine, which is commonly administered as codeine hydrochloride. Acids other than hydrochloric acid are also used. Hydrobromic acid, for example, gives *hydrobromide salts*. Dextromethorphan, an ingredient in many cough medicines, is dispensed as dextromethorphan hydrobromide. The accompanying figure shows another medication as a hydrochloride salt.

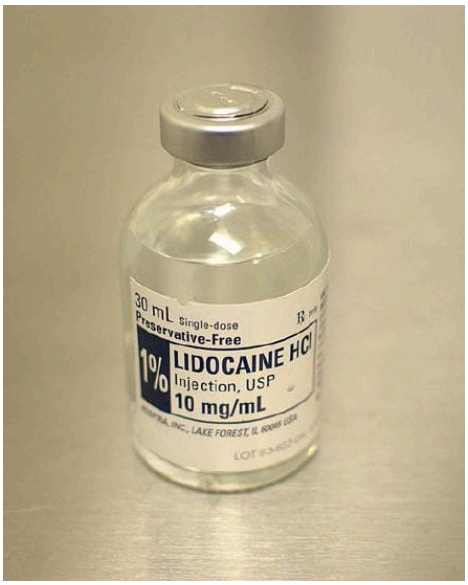


Figure used with permission from Wikipedia.

Concept Review Exercise

1. Give the definitions of a Brønsted-Lowry acid and a Brønsted-Lowry base.

Answer

1. A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

Key Takeaways

- A Brønsted-Lowry acid is a proton donor, and a Brønsted-Lowry base is a proton acceptor.
- Brønsted-Lowry acid-base reactions are essentially proton transfer reactions.

This page titled 15.3: Brønsted-Lowry Definition of Acids and Bases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.4: Water - Both an Acid and a Base

Learning Objectives

- To write chemical equations for water acting as an acid and as a base.

Water (H_2O) is an interesting compound in many respects. Here, we will consider its ability to behave as an acid or a base.

In some circumstances, a water molecule will accept a proton and thus act as a Brønsted-Lowry base. We saw an example in the dissolving of HCl in H_2O:

$$\mathrm{HCl} + \mathrm{H_2O}_{(\ell)} \rightarrow \mathrm{H_3O^+_{(aq)}} + \mathrm{Cl^-_{(aq)}} \tag{15.4.1}$$

In other circumstances, a water molecule can donate a proton and thus act as a Brønsted-Lowry acid. For example, in the presence of the amide ion (see Example 4 in Section 10.2), a water molecule donates a proton, making ammonia as a product:

$$H_2O_{(\ell)} + NH^-_{2(aq)} \rightarrow OH^-_{(aq)} + NH_{3(aq)} \tag{15.4.2}$$

In this case, NH_2^- is a Brønsted-Lowry base (the proton acceptor).

So, depending on the circumstances, H_2O can act as either a Brønsted-Lowry acid or a Brønsted-Lowry base. Water is not the only substance that can react as an acid in some cases or a base in others, but it is certainly the most common example—and the most important one. A substance that can either donate or accept a proton, depending on the circumstances, is called an amphiprotic compound.

A water molecule can act as an acid or a base even in a sample of pure water. About 6 in every 100 million (6 in 10^8) water molecules undergo the following reaction:

$$H_2O_{(\ell)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)} \tag{15.4.3}$$

This process is called the autoionization of water (Figure 15.4.1) and occurs in every sample of water, whether it is pure or part of a solution. Autoionization occurs to some extent in any amphiprotic liquid. (For comparison, liquid ammonia undergoes autoionization as well, but only about 1 molecule in a million billion (1 in 10^{15}) reacts with another ammonia molecule.)

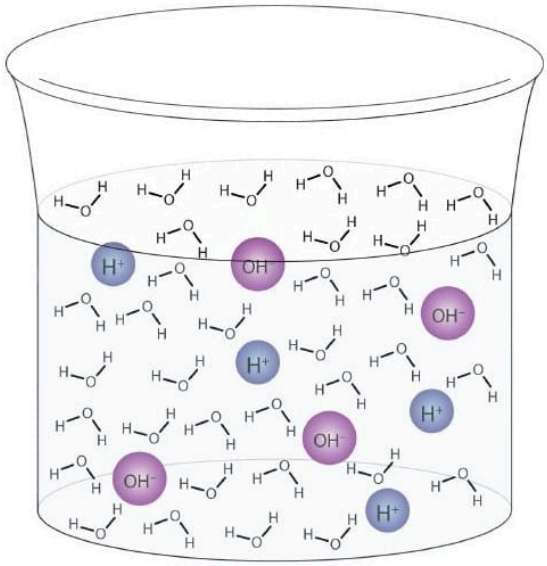


Figure 15.4.1 Autoionization. A small fraction of water molecules—approximately 6 in 100 million—ionize spontaneously into hydronium ions and hydroxide ions. This picture necessarily overrepresents the amount of autoionization that really occurs in pure water.

Example 15.4.1

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $H_2O(\ell) + NO_2^-(aq) \rightarrow HNO_2(aq) + OH^-(aq)$
2. $HC_2H_3O_2(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$

SOLUTION

1. In this reaction, the water molecule donates a proton to the NO_2^- ion, making $OH^-(aq)$. As the proton donor, H_2O acts as a Brønsted-Lowry acid.
2. In this reaction, the water molecule accepts a proton from $HC_2H_3O_2$, becoming $H_3O^+(aq)$. As the proton acceptor, H_2O is a Brønsted-Lowry base.

Exercise 15.4.2

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $HCOOH(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + HCOO^-(aq)$
2. $H_2O(\ell) + PO_4^{3-}(aq) \rightarrow OH^-(aq) + HPO_4^{2-}(aq)$

Concept Review Exercises

1. Explain how water can act as an acid.
2. Explain how water can act as a base.

Answers

1. Under the right conditions, H_2O can donate a proton, making it a Brønsted-Lowry acid.
2. Under the right conditions, H_2O can accept a proton, making it a Brønsted-Lowry base.

Key Takeaway

- Water molecules can act as both an acid and a base, depending on the conditions.

This page titled 15.4: Water - Both an Acid and a Base is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.5: The Strengths of Acids and Bases

Learning Objectives

- Describe the difference between strong and weak acids and bases.
- Describe how a chemical reaction reaches chemical equilibrium.
- Define the pH scale and use it to describe acids and bases.

Acids and bases do not all demonstrate the same degree of chemical activity in solution. Different acids and bases have different strengths.

Strong and Weak Acids

Let us consider the strengths of acids first. A small number of acids ionize completely in aqueous solution. For example, when HCl dissolves in water, every molecule of HCl separates into a hydronium ion and a chloride ion:

$$\mathrm{HCl\,(g) + H_2O\,(l) \xrightarrow{\sim 100\,\%} H_3O^+\,(aq) + Cl^-\,(aq)} \tag{15.5.1}$$

HCl(aq) is one example of a strong acid, which is a compound that is essentially 100% ionized in aqueous solution. There are very few strong acids. The important ones are listed in Table 15.5.1.

Table 15.5.1: Strong Acids and Bases (All in Aqueous Solution)

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	$Mg(OH)_2$
H_2SO_4	$Ca(OH)_2$
$HClO_4$	

By analogy, a strong base is a compound that is essentially 100% ionized in aqueous solution. As with acids, there are only a few strong bases, which are also listed in Table 15.5.1.

If an acid is not listed in Table 15.5.1, it is likely a weak acid, which is a compound that is not 100% ionized in aqueous solution. Similarly, a weak base is a compound that is not 100% ionized in aqueous solution. For example, acetic acid ($HC_2H_3O_2$) is a weak acid. The ionization reaction for acetic acid is as follows:

$$HC_2H_3O_{2(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + C_2H_3O^-_{2(aq)} \tag{15.5.2}$$

Depending on the concentration of $HC_2H_3O_2$, the ionization reaction may occur only for 1%–5% of the acetic acid molecules.

Looking Closer: Household Acids and Bases

Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for hydrochloric acid [HCl(aq)]. Vinegar has already been mentioned as a dilute solution of acetic acid [$HC_2H_3O_2$(aq)]. In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid ($HC_6H_7O_6$).

One of the more familiar household bases is ammonia (NH_3), which is found in numerous cleaning products. As we mentioned previously, ammonia is a base because it increases the hydroxide ion concentration by reacting with water:

$$NH_{3(aq)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{15.5.3}$$

Many soaps are also slightly basic because they contain compounds that act as Brønsted-Lowry bases, accepting protons from water and forming excess hydroxide ions. This is one reason that soap solutions are slippery.





Figure 15.5.1: (left) Bottles of alkaline drain cleaners containing sodium hydroxide can dissolve greases and hair. (right) Chemical burns caused by sodium hydroxide solution photographed 44 hours after exposure. Figures used with permission from Wikipedia.

Perhaps the most dangerous household chemical is the lye-based drain cleaner. Lye is a common name for sodium hydroxide,

although it is also used as a synonym for potassium hydroxide. Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and form a clog in a pipe. Unfortunately, lye can also attack tissues and other substances in our bodies. Thus, when we use lye-based drain cleaners, we must be very careful not to touch any of the solid drain cleaner or spill the water it was poured into. Safer, nonlye drain cleaners use peroxide compounds to react on the materials in the clog and clear the drain.

Chemical Equilibrium

The behavior of weak acids and bases illustrates a key concept in chemistry. Does the chemical reaction describing the ionization of a weak acid or base just stop when the acid or base is done ionizing? Actually, no. Rather, the reverse process—the reformation of the molecular form of the acid or base—occurs, ultimately at the same rate as the ionization process. For example, the ionization of the weak acid $HC_2H_3O_2$ (aq) is as follows:

$$HC_2H_3O_{2(aq)} + H_2O_{(\ell)} \rightarrow H_3O^+_{(aq)} + C_2H_3O^-_{2(aq)} \tag{15.5.4}$$

The reverse process also begins to occur:

$$H_3O^+_{(aq)} + C_2H_3O^-_{2(aq)} \rightarrow HC_2H_3O_{2(aq)} + H_2O_{(\ell)} \tag{15.5.5}$$

Eventually, there is a balance between the two opposing processes, and no additional change occurs. The chemical reaction is better represented at this point with a double arrow:

$$HC_2H_3O_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O^-_{2(aq)} \tag{15.5.6}$$

The $\rightleftharpoons$ implies that both the forward and reverse reactions are occurring, and their effects cancel each other out. A process at this point is considered to be at chemical equilibrium (or equilibrium). It is important to note that the processes do not stop. They balance out each other so that there is no further net change; that is, chemical equilibrium is a *dynamic equilibrium*.

Example $15.5.1$: Partial Ionization

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. HNO_2(aq)
b. C_5H_5N(aq)

SOLUTION

a. $HNO_2(aq) + H_2O(\ell) \diamond NO_2^-(aq) + H_3O^+(aq)$
b. $C_5H_5N(aq) + H_2O(\ell) \diamond C_5H_5NH^+(aq) + OH^-(aq)$

Exercise $15.5.1$

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. $HF_{(aq)}$
b. $AgOH_{(aq)}$

Hydrofluoric acid $HF_{(aq)}$ reacts directly with glass (very few chemicals react with glass). Hydrofluoric acid is used in glass etching.

Finally, you may realize that the autoionization of water is actually an equilibrium process, so it is more properly written with the double arrow:

$$H_2O_{(\ell)} + H_2O_{(\ell)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)} \tag{15.5.7}$$

The pH Scale

One qualitative measure of the strength of an acid or a base solution is the pH scale, which is based on the concentration of the hydronium (or hydrogen) ion in aqueous solution.

$$pH = -\log[H^+] \tag{15.5.8}$$

or

$$pH = -\log[H_3O^+] \tag{15.5.9}$$

A neutral (neither acidic nor basic) solution, one that has the same concentration of hydrogen and hydroxide ions, has a pH of 7. A pH below 7 means that a solution is acidic, with lower values of pH corresponding to increasingly acidic solutions. A pH greater than 7 indicates a basic solution, with higher values of pH corresponding to increasingly basic solutions. Thus, given the pH of several solutions, you can state which ones are acidic, which ones are basic, and which are more acidic or basic than others. Table 15.5.2 lists the pH of several common solutions. Notice that some biological fluids are nowhere near neutral.

Table 15.5.2: The pH Values of Some Common Solutions

Solution	pH
battery acid	0.3
stomach acid	1–2
lemon or lime juice	2.1
vinegar	2.8–3.0
Coca-Cola	3
wine	2.8–3.8
beer	4–5
coffee	5
milk	6
urine	6
pure H_2O	7

Solution	pH
(human) blood	7.3–7.5
sea water	8
antacid (milk of magnesia)	10.5
NH_3 (1 M)	11.6
bleach	12.6
NaOH (1 M)	14.0

Weak acids and bases are relatively common. You may notice from Table 15.5.2 that many food products are slightly acidic. They are acidic because they contain solutions of weak acids. If the acid components of these foods were strong acids, the food would likely be inedible.

Key Takeaways

- Acids and bases can be strong or weak depending on the extent of ionization in solution.
- Most chemical reactions reach equilibrium at which point there is no net change.
- The pH scale is used to succinctly communicate the acidity or basicity of a solution.

This page titled 15.5: The Strengths of Acids and Bases is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.6: Buffers

Learning Objectives

- To define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have a marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

The mechanism involves a buffer, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia (NH_3, a weak base) and ammonium chloride (NH_4Cl, a salt derived from that base).

Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base—a source of OH^-(aq) ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$HC_2H_3O_{2(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)} + C_2H_3O^-_{2(aq)} \tag{15.6.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$H^+_{(aq)} + C_2H_3O^-_{2(aq)} \rightarrow HC_2H_3O_{2(aq)} \tag{15.6.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure $15.6.1$ illustrates both actions of a buffer.

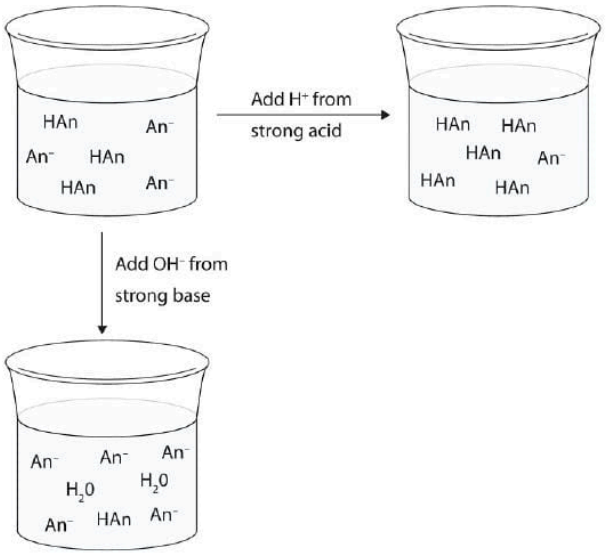


Figure 15.6.1: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)} \tag{15.6.3}$$

while the ammonium ion [NH_4^+(aq)] can react with any hydroxide ions introduced by strong bases:

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(\ell)} \tag{15.6.4}$$

Example 15.6.1

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

a. $HCHO_2$ and $NaCHO_2$
b. HCl and NaCl
c. CH_3NH_2 and CH_3NH_3Cl
d. NH_3 and NaOH

SOLUTION

a. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid—the formate ion (CHO_2^-). The combination of these two solutes would make a buffer solution.
b. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
c. Methylamine (CH_3NH_2) is like ammonia with one of its hydrogen atoms substituted with a CH_3 (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound CH_3NH_3Cl is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
d. Ammonia (NH_3) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

Exercise 15.6.1

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. H_3PO_4 and NaH_2PO_4
c. NH_3 and $(NH_4)_3PO_4$
d. NaOH and NaCl

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain capacity. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/– type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

This page titled 15.6: Buffers is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.E: Acids and Bases (Exercises)

1. Give two examples of Arrhenius acids.
2. Give two examples of Arrhenius bases.
3. List the general properties of acids.
4. List the general properties of bases.
5. Name each compound. (For acids, look up the name in Table 10.1.1. For bases, use the rules for naming ionic compounds from Chapter 3.)
    1. HBr(aq)
    2. $Ca(OH)_2$(aq)
    3. HNO_3(aq)
    4. $Fe(OH)_3$(aq)
6. Name each compound.
    1. HI(aq)
    2. $Cu(OH)_2$(aq)
    3. H_3PO_4(aq)
    4. CsOH(aq)
7. Write a balanced chemical equation for the neutralization of $Ba(OH)_2$(aq) with HNO_3(aq).
8. Write a balanced chemical equation for the neutralization of H_2SO_4(aq) with $Cr(OH)_3$(aq).
9. How many moles of sodium hydroxide (NaOH) are needed to neutralize 0.844 mol of acetic acid ($HC_2H_3O_2$)? (Hint: begin by writing a balanced chemical equation for the process.)
10. How many moles of perchloric acid ($HClO_4$) are needed to neutralize 0.052 mol of calcium hydroxide [$Ca(OH)_2$]? (Hint: begin by writing a balanced chemical equation for the process.)
11. Hydrazoic acid (HN_3) can be neutralized by a base.
    1. Write the balanced chemical equation for the reaction between hydrazoic acid and calcium hydroxide.
    2. How many milliliters of 0.0245 M $Ca(OH)_2$ are needed to neutralize 0.564 g of HN_3?
12. Citric acid ($H_3C_6H_5O_7$) has three hydrogen atoms that can form hydrogen ions in solution.
    1. Write the balanced chemical equation for the reaction between citric acid and sodium hydroxide.
    2. If an orange contains 0.0675 g of $H_3C_6H_5O_7$, how many milliliters of 0.00332 M NaOH solution are needed to neutralize the acid?
13. Magnesium hydroxide [$Mg(OH)_2$] is an ingredient in some antacids. How many grams of $Mg(OH)_2$ are needed to neutralize the acid in 158 mL of 0.106 M HCl(aq)? It might help to write the balanced chemical equation first.
14. Aluminum hydroxide [$Al(OH)_3$] is an ingredient in some antacids. How many grams of $Al(OH)_3$ are needed to neutralize the acid in 96.5 mL of 0.556 M H_2SO_4(aq)? It might help to write the balanced chemical equation first.

Answers

1. HCl and HNO_3 (answers will vary)

3. sour taste, react with metals, react with bases, and turn litmus red

5. 1. hydrobromic acid
   2. calcium hydroxide
   3. nitric acid
   4. iron(III) hydroxide

7. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O$

9. 0.844 mol

11. Part 1: $2HN_3(aq) + Ca(OH)_2 \rightarrow Ca(N_3)_2 + 2H_2O$

11. Part 2: 268 mL

13. 0.488 g

10.2 Brønsted-Lowry Definition of Acids and Bases

Exercises

1. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

   $HCl(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

2. Label each reactant as a Brønsted-Lowry acid or a Brønsted-Lowry base.

   $H_2O(\ell) + N_2H_4(aq) \rightarrow N_2H_5^+(aq) + OH^-(aq)$

3. Explain why a Brønsted-Lowry acid can be called a proton donor.

4. Explain why a Brønsted-Lowry base can be called a proton acceptor.

5. Write the chemical equation of the reaction of ammonia in water and label the Brønsted-Lowry acid and base.

6. Write the chemical equation of the reaction of methylamine (CH_3NH_2) in water and label the Brønsted-Lowry acid and base.

7. Demonstrate that the dissolution of HNO_3 in water is actually a Brønsted-Lowry acid-base reaction by describing it with a chemical equation and labeling the Brønsted-Lowry acid and base.

8. Identify the Brønsted-Lowry acid and base in the following chemical equation:

   $C_3H_7NH_2(aq) + H_3O^+(aq) \rightarrow C_3H_7NH_3^+(aq) + H_2O(\ell)$

9. Write the chemical equation for the reaction that occurs when cocaine hydrochloride ($C_{17}H_{22}ClNO_4$) dissolves in water and donates a proton to a water molecule. (When hydrochlorides dissolve in water, they separate into chloride ions and the appropriate cation.)

10. If codeine hydrobromide has the formula $C_{18}H_{22}BrNO_3$, what is the formula of the parent compound codeine?

Answers

1. HCl: Brønsted-Lowry acid; NH_3: Brønsted-Lowry base

3. A Brønsted-Lowry acid gives away an H^+ ion—nominally, a proton—in an acid-base reaction.

5. $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ (here NH_3 = Brønsted-Lowry base; H_2O = Brønsted-Lowry acid)

7. $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$ (here NO_3 = Brønsted-Lowry acid; H_2O = Brønsted-Lowry base)

9. $C_{17}H_{22}NO_{4+} + H_2O \rightarrow H_3O^+ + C_{17}H_{21}NO_4$

10.3 Water: Both an Acid and a Base

Exercises

1. Is $H_2O(\ell)$ acting as an acid or a base?

   $H_2O(\ell) + NH_4^+(aq) \rightarrow H_3O^+(aq) + NH_3(aq)$

2. Is $H_2O(\ell)$ acting as an acid or a base?

   $CH_3^-(aq) + H_2O(\ell) \rightarrow CH_4(aq) + OH^-(aq)$

3. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some $C_2H_3O_2^-$ solutions, the following reaction can occur:

   $C_2H_3O_2^-(aq) + H_2O(\ell) \rightarrow HC_2H_3O_2(aq) + OH^-(aq)$

   Is H_2O acting as an acid or a base in this reaction?

4. In the aqueous solutions of some salts, one of the ions from the salt can react with water molecules. In some NH_4^+ solutions, the following reaction can occur:

   $NH_4^+(aq) + H_2O \rightarrow NH_3(aq) + H_3O^+(aq)$

   Is H_2O acting as an acid or a base in this reaction?

5. Why is pure water considered neutral?

Answers

1. base

3. acid

5. When water ionizes, equal amounts of H^+ (acid) and OH^- (base) are formed, so the solution is neither acidic nor basic: $H_2O(\ell) \rightarrow H^+(aq) + OH^-(aq)$

[SIDE NOTE: It is rare to truly have pure water. Water exposed to air will usually be slightly acidic because dissolved carbon dioxide gas, or carbonic acid, decreases the pH slightly below 7. Alternatively, dissolved minerals, like calcium carbonate (limestone), can make water slightly basic.]

10.4 The Strengths of Acids and Bases

Concept Review Exercises

1. Explain the difference between a strong acid or base and a weak acid or base.
2. Explain what is occurring when a chemical reaction reaches equilibrium.
3. Define *pH*.

Answers

1. A strong acid or base is 100% ionized in aqueous solution; a weak acid or base is less than 100% ionized.
2. The overall reaction progress stops because the reverse process balances out the forward process.
3. pH is a measure of the hydrogen ion concentration.

Exercises

1. Name a strong acid and a weak acid. (Hint: use Table 10.4.1.)
2. Name a strong base and a weak base. (Hint: use Table 10.4.1.)
3. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. HF
    2. $HC_2H_3O_2$
    3. HCl
    4. $HClO_4$
4. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. H_2SO_4
    2. HSO_4^-
    3. HPO_4^{2-}
    4. HNO_3
5. Is each compound a strong base or a weak base? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. NH_3
    2. NaOH
    3. $Mg(OH)_2$
    4. $Cu(OH)_2$
6. Is each compound a strong base or a weak base? Assume all are in aqueous solution. (Hint: use Table 10.4.1.)
    1. KOH
    2. H_2O
    3. $Fe(OH)_2$
    4. $Fe(OH)_3$
7. Write the chemical equation for the equilibrium process for each weak acid in Exercise 3.
8. Write the chemical equation for the equilibrium process for each weak acid in Exercise 4.
9. Write the chemical equation for the equilibrium process for each weak base in Exercise 5.
10. Write the chemical equation for the equilibrium process for each weak base in Exercise 6.
11. Which is the stronger acid—HCl(aq) or HF(aq)?

12. Which is the stronger base—KOH(aq) or $Ni(OH)_2$(aq)?

13. Consider the two acids in Exercise 11. For solutions that have the same concentration, which one would you expect to have a lower pH?

14. Consider the two bases in Exercise 12. For solutions that have the same concentration, which one would you expect to have a higher pH?

15. Consider the list of substances in Table \PageIndex3\PageIndex3.2"The pH Values of Some Common Solutions". What is the most acidic substance on the list that you have encountered recently?

16. Consider the list of substances in Table \PageIndex3\PageIndex3.2"The pH Values of Some Common Solutions". What is the most basic substance on the list that you have encountered recently?

Answers

1. strong acid: HCl; weak acid: $HC_2H_3O_2$ (answers will vary)

3. 1. weak
   2. weak
   3. strong
   4. strong

5. 1. weak
   2. strong
   3. strong
   4. weak

7. 3a: $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$; 3b: $HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$

9. 5a: $NH_3(aq) + H_2O \rightleftharpoons NH_4^+(aq) + OH^-(aq)$; 5d: $Cu(OH)_2(aq) \rightleftharpoons Cu^{2+}(aq) + 2OH^-(aq)$

11. HCl(aq)

13. HCl(aq)

15. (answers will vary)

10.5: Buffers

Concept Review Exercise

1. Explain how a buffer prevents large changes in pH.

Answer

1. A buffer has components that react with both strong acids and strong bases to resist sudden changes in pH.

Exercises

1. Describe a buffer. What two related chemical components are required to make a buffer?

2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?

3. Which solute combinations can make a buffer? Assume all are aqueous solutions.

1. HCl and NaCl
2. HNO_2 and $NaNO_2$
3. NH_4NO_3 and HNO_3
4. NH_4NO_3 and NH_3

4. Which solute combinations can make a buffer? Assume all are aqueous solutions.

   1. H_3PO_4 and Na_3PO_4
   2. $NaHCO_3$ and Na_2CO_3
   3. $NaNO_3$ and $Ca(NO_3)_2$
   4. HN_3 and NH_3

5. For each combination in Exercise 3 that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.

6. For each combination in Exercise 4 that is a buffer, write the chemical equations for the reaction of the buffer components when a strong acid and a strong base is added.

7. The complete phosphate buffer system is based on four substances: H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}. What different buffer solutions can be made from these substances?

8. Explain why NaBr cannot be a component in either an acidic or a basic buffer.

9. Explain why $Mg(NO_3)_2$ cannot be a component in either an acidic or a basic buffer.

Answers

1. A buffer resists sudden changes in pH. It has a weak acid or base and a salt of that weak acid or base.

3.
   1. not a buffer
   2. buffer
   3. not a buffer
   4. buffer

5. 3b: strong acid: $H^+ + NO_2^- \rightarrow HNO_2$; strong base: $OH^- + HNO_2 \rightarrow H_2O + NO_2^-$; 3d: strong acid: $H^+ + NH_3 \rightarrow NH_4^+$; strong base: $OH^- + NH_4^+ \rightarrow H_2O + NH_3$

7. Buffers can be made by combining H_3PO_4 and $H_2PO_4^-$, $H_2PO_4^-$ and HPO_4^{2-}, and HPO_4^{2-} and PO_4^{3-}.

9. $Mg(NO_3)_2$ includes two types of ions, Mg^{2+} and NO_3^-. $Mg(OH)_2$ is strong base and completely dissociates (100% falls apart), so Mg^{2+} will not react with any added base (0% combines with OH^-). HNO_3 is strong acid and completely dissociates (100% falls apart), so NO_3^- will not react with any added acid (0% combines with H^+). Because $Mg(NO_3)_2$ does will not react with any added base or acid, it does not resist change in pH and is not a buffer.

10.6: Chapter Summary

Additional Exercises

1. The properties of a 1.0 M HCl solution and a 1.0 M $HC_2H_3O_2$ solution are compared. Measurements show that the hydrochloric acid solution has a higher osmotic pressure than the acetic acid solution. Explain why.

2. Of a 0.50 M HNO_3 solution and a 0.50 M $HC_2H_3O_2$ solution, which should have the higher boiling point? Explain why.

3. The reaction of sulfuric acid [$H_2SO_4(aq)$] with sodium hydroxide [NaOH(aq)] can be represented by two separate steps, with

only one hydrogen ion reacting in each step. Write the chemical equation for each step.

4. The reaction of aluminum hydroxide [$Al(OH)_3(aq)$] with hydrochloric acid [HCl(aq)] can be represented by three separate steps, with only one hydroxide ion reacting in each step. Write the chemical equation for each step.

5. A friend brings you a small sample of an unknown chemical. Assuming that the chemical is soluble in water, how would you determine if the chemical is an acid or a base?

6. A neutral solution has a hydrogen ion concentration of about 1×10^{-7} M. What is the concentration of the hydroxide ion in a neutral solution?

7. The Lewis definitions of an acid and a base are based on electron pairs, not protons. A Lewis acid is an electron pair acceptor, while a Lewis base is an electron pair donor. Use Lewis diagrams to show that

   $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

   is an acid-base reaction in the Lewis sense as well as in the Arrhenius and Brønsted-Lowry senses.

8. Given the chemical reaction

   $NH_3(g) + BF_3(g) \rightarrow NH_3—BF_3(s)$

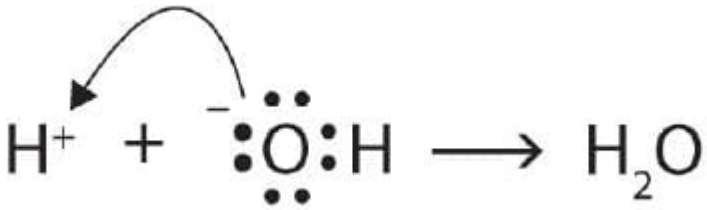
   show that the reaction illustrated by this equation is an acid-base reaction if we use the Lewis definitions of an acid and a base (see Exercise 7). The product contains a bond between the N and B atoms.

Answers

1. HCl is a strong acid and yields more ions in solution.

3. $H_2SO_4 + NaOH \rightarrow NaHSO_4 + H_2O$; $NaHSO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$

5. One way is to add it to $NaHCO_3$; if it bubbles, it is an acid. Alternatively, add the sample to litmus and look for a characteristic color change (red for acid, blue for base).

7. 

   The O atom is donating an electron pair to the H^+ ion, making the base an electron pair donor and the acid an electron pair acceptor.

This page titled 15.E: Acids and Bases (Exercises) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

15.S: Acids and Bases (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

The earliest chemical definition of an acid, the **Arrhenius definition**, says that an **acid** is a compound that increases the amount of hydrogen ion (H^+) in aqueous solution. An **Arrhenius base** is a compound that increases the amount of hydroxide ion (OH^-) in aqueous solution. While most bases are named as ionic hydroxide compounds, aqueous acids have a naming system unique to acids. Acids and bases react together in a characteristic chemical reaction called **neutralization**, in which the products are water and a salt. The principles of stoichiometry, along with the balanced chemical equation for a reaction between an acid and a base, can be used to determine how much of one compound will react with a given amount of the other.

A **Brønsted-Lowry acid** is any substance that donates a proton to another substance. A **Brønsted-Lowry base** is any substance that accepts a proton from another substance. The reaction of ammonia with water to make ammonium ions and hydroxide ions can be used to illustrate Brønsted-Lowry acid and base behavior.

Some compounds can either donate or accept protons, depending on the circumstances. Such compounds are called **amphiprotic**. Water is one example of an amphiprotic compound. One result of water being amphiprotic is that a water molecule can donate a proton to another water molecule to make a hydronium ion and a hydroxide ion. This process is called the **autoionization of water** and occurs in any sample of water.

Not all acids and bases are equal in chemical strength. A **strong acid** is an acid whose molecules are all dissociated into ions in aqueous solution. Hydrochloric acid is an example of a strong acid. Similarly, a **strong base** is a base whose molecules are dissociated into ions in aqueous solution. Sodium hydroxide is an example of a strong base. Any acid or base whose molecules are not all dissociated into ions in aqueous solution is a **weak acid** or a **weak base**. Solutions of weak acids and weak bases reach a **chemical equilibrium** between the un-ionized form of the compound and the dissociated ions. It is a dynamic equilibrium because acid and base molecules are constantly dissociating into ions and reassociating into neutral molecules.

The **pH** scale is a scale used to express the concentration of hydrogen ions in solution. A neutral solution, neither acidic nor basic, has a pH of 7. Acidic solutions have a pH lower than 7, while basic solutions have a pH higher than 7.

Buffers are solutions that resist dramatic changes in pH when an acid or a base is added to them. They contain a weak acid and a salt of that weak acid, or a weak base and a salt of that weak base. When a buffer is present, any strong acid reacts with the anion of the salt, forming a weak acid and minimizing the presence of hydrogen ions in solution. Any strong base reacts with the weak acid, minimizing the amount of additional hydroxide ions in solution. However, buffers only have limited **capacity**; there is a limit to the amount of strong acid or strong base any given amount of buffer will react with.

This page titled 15.S: Acids and Bases (Summary) is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16: EXAMPLES OF FINAL PROJECT

16.1: Amphetamines and the Brain
16.2: Great Pacific Garbage Patch
16.3: Sorting Recyclable Plastics by Density
16.4: Solubility and Pollution
16.5: Radiation Contamination from Fukushima
16.6: pH of Chesapeake Bay
16.7: 18-F fluorodeoxyglucose and Atherosclerotic Plaque
Student Project

This page titled 16: Examples of Final Project is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.1: Amphetamines and the Brain

This is a final project completed by Jordan Baur for Chemistry 150 in Spring 2020 at Modesto Junior College. When citing this project, he is the author.

For the final project I have decided to choose to focus on the effects of amphetamines on the brain. The reason for this is that I live in Modesto and there is a lot of drug abuse of methamphetamines and also in college I am increasingly surprised by the amount of people prescribed to and using variations of the same stimulant like Adderall, also a lot of people use it for off label purposes or even without a prescription. With more people being prescribed this drug, the risk of improper use and Adderall addiction has increased. Adderall is a combination drug made from amphetamine, which is the parent drug of methamphetamine. I would like to explore the effect of amphetamines on the brain.

Amphetamine is indicated for the treatment of attention-deficit/hyperactivity disorders (ADHD) as well as for the treatment of central nervous system disorders such as narcolepsy. (1) Amphetamine is also used off-label for the treatment of obesity, depression and chronic pain. (1)

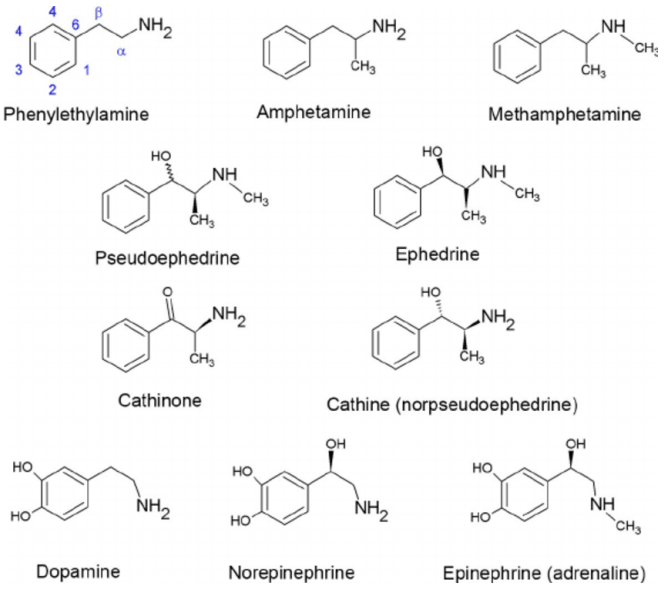


Figure \({16{.}1{.}\eatSpaces{1} \endSpaces}\): Comparison of Amphetamine and other neurotran

I have inserted a picture of the chemical structure of amphetamine, methamphetamine various other neurotransmitters to show the similarities in structures (Sanchez-Ramos, 2004, p. 133). In biology structure determines function. But also, simple changes in structure like the presence of a methyl group or a hydroxyl group, can have a huge impact on the effects on the body.

Because of the similarity in structure seen above, amphetamines bind to the receptor protein in the synapse for monoamines(dopamine, norepinephrine and adrenaline) and then are taken into the neuron via these transporter protein receptor. Amphetamines then disrupt the storage of these monoamines in the neuron. This leads to higher amounts of monoamines in the neuron. Then through mechanisms not fully understood, amphetamines then cause the monoamine transporter proteins to run in reverse. There is an increase release of monoamines from the same neuron that took up the amphetamine. This leads to very high levels of this neurotransmitters in the synaptic cleft. This is the basis of the action of amphetamines.

This drug affects reward centers and have a high rate of dependency and abuse. Amphetamine induced psychologic dependence which is manifested by elevated mood, increased wakefulness, concentration, physical performance and a feeling of well-being. With sustained use, the effects of rapid heartbeat and enhanced alertness diminish while adverse effects such as hallucinations and delusions may occur. (1)

It is shown that because of its effects as a stimulant (like cocaine) Adderall and other such amphetamines, do increase mood , attention , focus and maybe even motivation (reward centers). But these effects fail to lead to better performance on a variety of

neurocognitive tasks that measure short-term memory, reading comprehension and fluency. This shows that not only does it not benefit students academically but also may impair their performance. It will keep you awake and focused, but sleep is essential to overall health. Also attention and focus is something that can be improved upon over time. The drug however does improve mood but does so by flooding your synapses with neurotransmitters like dopamine. Users often fail to differentiate between how they feel on the drug vs their performance. If you are less miserable when trying to accomplish tasks, you then automatically think you are performing better.

Referring back to the picture with the structures. If you compare amphetamine to methamphetamine, meth has an extra group of carbon with 3 hydrogens (methyl group). This simple change in chemical structure makes the drug much more addictive and dangerous for use.

Learning Questions :

1. Why is methamphetamine more dangerous than Adderall?
2. How does a receptor on a synapse work? Based on shape, chemical composition etc ?
3. How does a substance that effects the nervous system, end up having an effect on the cardiovascular system ? (eg Adderall, increased heart rate)

References :

1- “Amphetamine.” *National Center for Biotechnology Information. PubChem Compound Database*, U.S. National Library of Medicine, pubchem.ncbi.nlm.nih.gov/compound/Amphetamine#section=Pharmacology-and-Biochemistry.

2- Berman, S M, et al. “Potential Adverse Effects of Amphetamine Treatment on Brain and Behavior: a Review.” *Molecular Psychiatry*, U.S. National Library of Medicine, Feb. 2009, www.ncbi.nlm.nih.gov/pmc/articles/PMC2670101/.

3- Pedersen, Traci. “Study: Adderall Fails to Boost Cognition in Healthy College Students.” *Psych Central*, 24 July 2018, psychcentral.com/news/2018/07/24/study-adderall-fails-to-boost-cognition-in-healthy-college-students/137238.html.

4- Reflections Recovery Center. “Adderall Vs Meth - Is Adderall the Legal Version of Meth?” *Reflections Recovery Center*, 22 July 2019, reflectionsrehab.com/blog/is-adderall-the-legal-version-of-meth/.

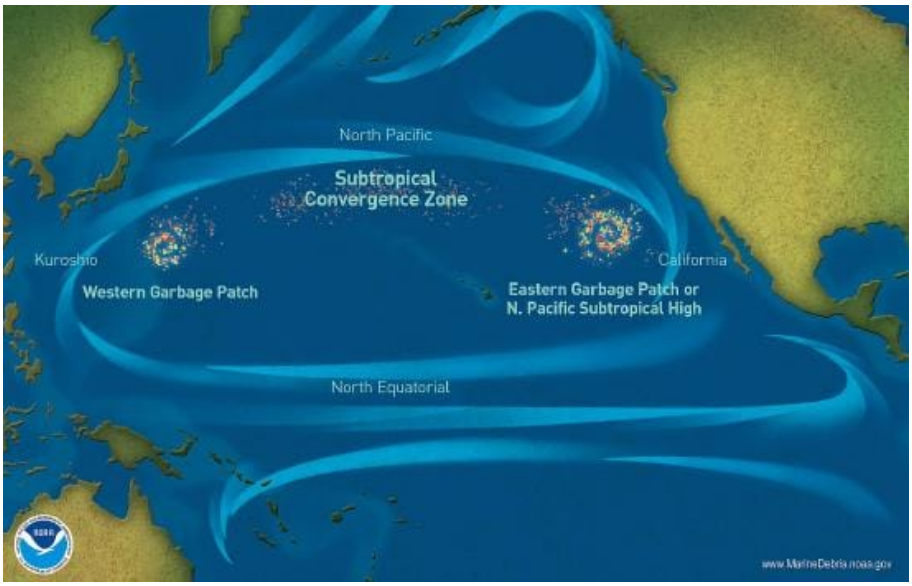
5- Sanchez-Ramos, Juan. (2004). Stimulant-induced movement disorders. Drug-induced movement disorders.. 295-308.

This page titled 16.1: Amphetamines and the Brain is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.2: Great Pacific Garbage Patch

The case study below was completed by Peron Scott Van Fleet as a final project for Chemistry 150 at Modesto Junior College in Spring 2020. He is the sole author of the content.

The Great Pacific garbage patch, also described as the Pacific trash vortex, is a collection of trash, collected by the currents of the ocean. “It spans waters from the West Coast of North America, to Japan.”(www.nationalgeographic.org) The one I will be talking about the most is the one closest to use, the Eastern Garbage Patch, located between the United States of America of Hawaii and California.



The main subject I am going to talk about is the plastic found in the Great Pacific Garbage Patch, the density for plastic as well, and how the plastic is harmful to the animals.

The most common plastic to find in the Great Pacific Garbage Patch is plastic 2, and plastic 1, and the plastic that is found is around the size of rice grans, they turn to plastic microfibers. The ones that are found together and still intact are Plastic bags, bottle caps, plastic water bottles, and styrofoam cups. They estimate an 80,000 metric tons in the patch, with 1.8 trillion plastic pieces of trash. If you don't know how many metric tons is, look at section Dimensional Analysis in the textbook.

There are 7 plastic types that are found in and out of your house. The most common types you see are bottle caps, and plastic water bottles, these are plastic 1. If you don't know the types here they are. Please look at the section on density from the textbook if you don't remember about density.

- Plastic 1: Polyethylene terephthalate (PET or PETE) “Is what’s used to make bottles for soda, water and other drinks, plastic peanut butter jars and containers for other food items.” Density is 1.38 g/cm3

- Plastic 2: High density polyethylene (HDPE) “They’re those used to make milk jugs, shampoo bottles, cleaning product containers and detergent bottles.” Density is 0.93 g/cm3

- Plastic 3: Polyvinyl chloride (PVC) It’s used for a huge array of household products. Plastic tubing, kids’ toys, plastic trays and furniture are often made out of PVC. Density is 1.38 g/cm3

- Plastic 4 : Low density polyethylene (LDPE) A lot of plastic wrappings are made of LDPE plastic. It’s also used to make grocery bags, and the bags that hold newspapers, among other things. Density is 0.910–0.940 g/cm3

- Plastic 5 : (Polypropylene) PP is used to make the food containers used for products like yogurt, and other dairy products. It’s made into straws, rope, carpet and bottle caps. Density is .90 g/cm3

- Plastic 6: (Polystyrene) PS Styrofoam products are made out of PS plastic, it used to make disposable coffee cups, packing peanuts, coolers and to-go food containers. Density is 1.06 g/cm3

- Plastic 7: Other Products stamped with a 7 are often made out of multiple plastic types or out of other types of plastic that can't easily be recycled. Density is ?

The website for this information on all 7 plastic types are from (https://millerrecycling.com/plastics-recycling-numbers) you can not get the density from there.

Fun fact: the first plastic based on a synthetic polymer was made from phenol and formaldehyde, with the first viable and cheap synthesis methods invented in 1907. Another fun fact: "plastic bags can take 20 years to decompose in the ocean, plastic bottles up to 450 years, and fishing line, 600 years." (blogs.ei.columbia.edu)

The ocean is one of the most abundant resources on this planet. "About 71 percent of the Earth's surface is water-covered, and the oceans hold about 96.5 percent of all Earth's water. So only 4.5 is fresh water, the other is salt water." (www.usgs.gov)

- About 3.5 percent of the weight of seawater comes from the dissolved **salts**

- The density of surface seawater ranges from about 1020 to 1029 kg/m3

- Salt in seawater salinity is about 35 parts per thousand, on average.

The Great Pacific Garbage Patch is a very dangerous thing for marine life as well as several different species of birds. Over 1 million marine animals (including mammals, fish, sharks, turtles, and birds) are killed each year due to plastic debris in the ocean. Major causes of death are caused by ingesting a material such as plastic. Sharp-edged plastic kills birds by punching holes in internal organs. Some seabirds eat so much plastic, there is little room left in their gut for food, which affects their body weight, jeopardizing their health. Or getting caught in material or trash. (Plastic bag, and more items)

Learning Question

For these questions look at the section on density and Dimensional Analysis from the textbook.

1. What is 80,000 metric tons in grams or g?
2. Compare plastic 1 density to plastic 2 density, to see which has the highest density. What is the difference in their densities?
3. Use the answer from the first question for the mass, and use plastic 1 density, to find the volume in mL.

References:

Website: seastewards.org

Article: Marine Debris and Plastics

Author: unknown

Date: unknown

http://seastewards.org/projects/healthy-oceans-initiative/marine-debris-and-plastics/

Website: nationalgeographic

Article: Ocean Trash: 5.25 Trillion Pieces and Counting, but Big Questions Remain

Author: Laura Parker

Last Updated: Feb. 22, 2019

https://www.nationalgeographic.org/article/ocean-trash-525-trillion-pieces-and-counting-big-questions-remain/

Website: nationalgeographic

Article:Great Pacific Garbage Patch

Author: Jeannie Evers

Last Updated: Feb. 22, 2019

https://www.nationalgeographic.org/encyclopedia/great-pacific-garbage-patch/

Website: conserveturtles.org

Article: Threats from Marine Debris

Author: unknown

Date: unknown

https://conserveturtles.org/information-sea-turtles-threats-marine-debris/

Website: oceanservice.noaa.gov

Article: Why is the ocean salty?

Author: NOAA

Date: 11/13/19

https://oceanservice.noaa.gov/facts/whysalty.html

Website: blogs.ei.columbia.edu/

Article: Our Oceans: A Plastic Soup

Author: RENEE CHO

Date: JANUARY 26, 2011

(https://blogs.ei.columbia.edu/2011/01/26/our-oceans-a-plastic-soup/)

Website: www.usgs.gov

Article: How Much Water is There on Earth?

Author: no author found

Date: no date found

(https://www.usgs.gov/special-topic/water-science-school/science/how-much-water-there-earth?qt-science_center_objects=0#qt-science_center_objects)

Website: www.forbes.com

Article: 300-Mile Swim Through The Great Pacific Garbage Patch Will Collect Data On Plastic Pollution

Author: Scott Snowden

Date: May 30, 2019

(https://www.forbes.com/sites/scottsnowden/2019/05/30/300-mile-swim-through-the-great-pacific-garbage-patch-will-collect-data-on-plastic-pollution/#66b968ea489f

This page titled 16.2: Great Pacific Garbage Patch is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

16.3: Sorting Recyclable Plastics by Density

In 2006 the amount of plastic bottles recycled reached a record high of 2,220,000,000 pounds. The amount of PET bottles recycled in 2006 increased more than 102 million pounds compared to 2005.



Figure 16.3.1 *Household items sorted by density for recycling*

HDPE bottle recycling increased in 2005 to 928 million pounds. Over 18,000 million pounds is produced. All plastic bottles were recycled at a rate of 24 percent in 2005 [1]. When different plastics are melted together, they tend to fractionate like oil and water, so the products that are made from recycled plastics tend to be structurally unsound unless the plastics are carefully sorted. The sorting is done by floatation in many commercial firms. Heavy plastics sink in a brine (salt) solution, while the lighter ones float and can be separated. Different brine solutions can be ued to separate all the major plastics[2]. Surprisingly, many recyclers depend on cheap labor to sort plastics by hand, and polypropylene, which is hard for laborers to recognize, is not recycled effectively (of over 19,000 million pounds produced, only about 132 million is recycled. Polypropylene is used in multicolored margarine and yogurt tubs, for example. The terms *heavy* and *light* are commonly used in two different ways. We refer to weight when we say that an adult is heavier than a child. On the other hand, something else is alluded to when we say that polyvinyl chloride (PVC) is heavier than polypropylene (PP). A small shaving of PVC would obviously weigh less than a roomful of PP, but PVC is heavier in the sense that a piece of given size weighs more than the same-size piece of PP.

What we are actually comparing is the *mass per unit volume*, that is, the **density**. In order to determine these densities, we might weigh a cubic centimeter of each type of plastic. If the PVC sample weighed 1.40 g and the PP 0.91 g, we could describe the density of PVC as 1.40 g cm^{-3} and that of PP as 0.91 g cm^{-3}. (Note that the negative exponent in the units cubic centimeters indicates a reciprocal. Thus 1 cm^{-3} = 1/cm^3 and the units for our densities could be written as $\frac{g}{cm^3}$, g/cm^3, or g cm^{-3}. In each case the units are read as grams per cubic centimeter, the *per* indicating division.) We often abbreviate "cm^3" as "cc", and 1 cm^3 = 1 mL exactly by definition.

Table 16.3.1 *Density of Plastics*

Plastic Symbol	name	Density/g/cm^3	Uses
5 PP	Polypropylene	0.90 - 0.92	yogurt cups, plasticware
4 LDPE	Low Density Polyethylene	0.91 - 0.93	squeezable bottles
2 HDPE	High Density Polyethylene	0.94 - 0.96	milk bottles, bags
6 PS	Polystyrene	1.03 - 1.06	egg cartons, packing peanuts

	Polyethylene Terephthalate	1.35 - 1.38	water bottles
	Polyvinyl Chloride	1.32 - 1.42	juice bottles, cling wrap
	Often Polycarbonate or Acrylonitrile butadiene styrene (ABS)	1.3 1.06	bottles auto body parts
Densities of many more materials are easily found.			

There are scores of plastics in common use, each with unique Properties.

In general it is not necessary to weigh exactly 1 cm^3 of a material in order to determine its density. We simply measure mass and volume and divide volume into mass:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \tag{16.3.1}$$

or

$$\rho = \frac{\text{m}}{\text{V}} \tag{16.3.2}$$

where ρ = density m = mass V = volume

Example 16.3.1: Density Calculation

Calculate the density of (a) a piece of PVC whose mass is 19.46 g and which, when submerged, increases the water level in a graduated cylinder by 13.9 ml; (b) a PVC cylinder of mass 12.990 g, radius 0.750 cm, and height 5.25 cm.

Solution

a) Since the submerged metal displaces its own volume,

$$\text{Density} = \rho = \frac{\text{m}}{\text{V}} = \frac{19.46g}{13.9mL} = 1.4 \text{ g/mL or } 1.4 \text{ g mL}^{-1}$$

b) The volume of the cylinder must be calculated first, using the formula

$$\text{V} = \pi r^2 h = 3.142 \times (0.750 \text{ cm})^2 * 5.25\text{cm} = 9.278\ 718\ 8 \text{ cm}^3$$

Then $\rho = \frac{\text{m}}{\text{V}} = \frac{12.990g}{9.2787188cm^3}$

= $1.430\frac{\text{g}}{\text{cm}^3}$

Note

Note that unlike mass or volume, the density of a substance is independent of the size of the sample. Thus density is a property by which one substance can be distinguished from another. A sample of PVC can be trimmed to any desired volume or adjusted to have any mass we choose, but its density will always be 1.40 g/cm^3 at 20°C. The densities of some common

pure substances are listed in the Table.

Tables and graphs are designed to provide a maximum of information in a minimum of space. When a physical quantity (number × units) is involved, it is wasteful to keep repeating the same units. Therefore it is conventional to use pure numbers in a table or along the axes of a graph. A pure number can be obtained from a quantity if we divide by appropriate units. For example, when divided by the units gram per cubic centimeter, the density of aluminum becomes a pure number 2.70:

$$\frac{\text{Density of aluminum}}{1\ \text{g cm}^{-3}} = \frac{2.70\ \text{g cm}^{-3}}{1\ \text{g cm}^{-3}} = 2.70 \tag{16.3.3}$$

Table 16.3.2 *Density of Several Substances at 20°C.*

Substance	Density/g cm^{-3}*
Helium gas	0.000 16
Dry air	0.001 185
Gasoline	0.66 → 0.69 (varies)
Kerosene	0.82
Benzene	0.880
Water	1.000
Carbon tetrachloride	1.595
Magnesium	1.74
Salt	2.16
Aluminum	2.70
Iron	7.87
Copper	8.96
Silver	10.5
Lead	11.34
Uranium	19.05
Gold	19.32

Therefore, a column in a table or the axis of a graph is conveniently labeled in the following form:

$$\frac{\text{Quantity}}{\text{units}} \tag{16.3.4}$$

This indicates the units that must be divided into the quantity to yield the pure number in the table or on the axis. This has been done in the second column of the Table.

Converting Densities

In our exploration of Density, notice that chemists may express densities differently depending on the subject. The density of pure substances may be expressed in kg/m^3 in some journals which insist on strict compliance with SI units; densities of soils may be expressed in lb/ft^3 in some agricultural or geological tables; the density of a cell may be expressed in mg/µL; and other units are in common use. It is easy to transform densities from one set of units to another, by multiplying the original quantity by one or more ***unity factors***:

Example 16.3.2: Density Conversion

Convert the density of water, 1 g/cm^3 to (a) lb/cm^3 and (b) lb/ft^3

a. The equality 454 g = 1 lb can be used to write two unity factors,

$\frac{454\ \text{g}}{1\ \text{lb}}$ or $\frac{1\ \text{lb}}{454}$

The given density can be multiplied by one of the unity factors to get the desired result. The correct conversion factor is chosen so that the units cancel:

$$\frac{1\ \text{g}}{\text{cm}^3} * \frac{1\ \text{lb}}{454\ \text{g}} = 0.002203 \frac{\text{lb}}{\text{cm}^3}$$

b. Similarly, the equalities 2.54 cm = 1 inch, and 12 inches = 1 ft can be use to write the unity factors:

$\frac{2.54\ \text{cm}}{1\ \text{in}}$, $\frac{1\ \text{in}}{2.54\ \text{cm}}$, $\frac{12\ \text{in}}{1\ \text{ft}}$ and $\frac{1\ \text{ft}}{12\ \text{in}}$

In order to convert the cm^3 in the denominator of 0.002203 $\frac{lb}{cm^3}$ to in^3, we need to multiply by the appropriate unity factor three times, or by the cube of the unity factor:

$$0.002203 \frac{\text{g}}{\text{cm}^3} \text{ x } \frac{2.54\ \text{cm}}{1\ \text{in}} \text{ x } \frac{2.54\ \text{cm}}{1\ \text{in}} \text{ x } \frac{2.54\ \text{cm}}{1\ \text{in}}$$

or

$$0.002203) \frac{\text{g}}{\text{cm}^3} \text{ x } (\frac{2.54\ \text{cm}}{1\ \text{in}})^3 = 0.0361 \frac{\text{lb}}{\text{in}^3}$$

This can then be converted to lb/ft^3:

$$0.0361 \frac{\text{lb}}{\text{in}^3} \text{ x } (\frac{12\ \text{in}}{1\ \text{ft}})^3 = 62.4 \frac{\text{lb}}{\text{ft}^3}$$

Note

It is important to notice that we have used conversion factors to convert from one unit to another unit of the ***same parameter***

From ChemPRIME: 1.8: Density

References

1. ↑ http://en.wikipedia.org/wiki/Plastic_recycling
2. ↑ Chemical & Engineering News, March 16, 2009, p. 30-33

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 16.3: Sorting Recyclable Plastics by Density is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, and Adam Hahn.

16.4: Solubility and Pollution

A solution may be gases, liquids, or solids and is composed of a solute and solvent. The solvent is usually the part of the solution that is present in the greatest amount. The solute is the substance that is dissolving in the solvent and is present in a smaller amount. Solubility is the amount of solute needed to form a saturated solution for a given quantity of solvent at a given temperature to produce a saturated solution.[1] Solubility versus temperature dependent graphs are often created and show that for most ionic substances their solubility increases with temperature. In addition, the solubility of ionic compounds at a particular temperature varies due to interactions between the solute and solvent. In addition, solubility is affected by the intermolecular bonding that occurs in the ionic compounds.

$$\text{Solubility} = \frac{\text{grams of solute}}{\text{100 grams of solvent}}$$

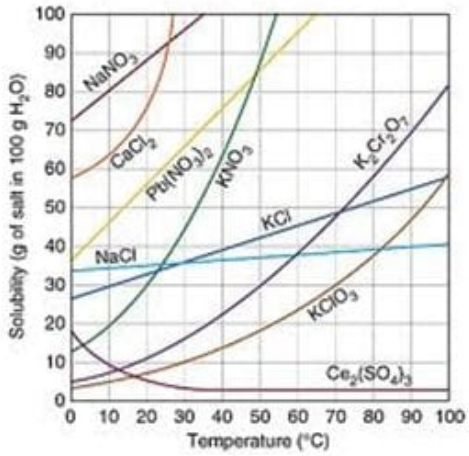


Figure 16.4.1. *Solubility curve of ionic compounds.*[2]

A saturated solution contains the maximum amount of solute for a given amount of solvent at a constant temperature. A saturated solution is in a state of dynamic equilibrium; no additional solute will dissolve if added to the solvent. An unsaturated solution contains less solute than a saturated solution; therefore it has the capacity to dissolve more solute. A supersaturated solution contains more solute than it can theoretically hold at a given temperature. A supersaturated solution is made by adding additional solute and heating the solution and slowly cooling undisturbed. Sometimes in order for crystallization to occur, a seed crystal must be introduced and the crystallization process is quite fast and dramatic.



Figure 16.4.2. *Crystallization of a supersaturated solution after the addition of a seed crystal.*[3]

The solubility of most solid and liquid solutions increases as the temperature increases. This is not true for gases; rather the solubility of gases decreases as the temperature increases. The reason for the decrease in solubility is related to the movement of the molecules. Temperature is a measure of the average kinetic energy of the molecules of a substance. An increase in temperature results in an increase in the kinetic energy of the molecules and an increase in thermal energy. The higher kinetic energy causes more motion in the molecules, which causes the intermolecular attraction to decrease and the gas molecules of solute can escape the solution.[1] This can easily be seen when an open can of soda in placed at room temperature or in a cool refrigerator. If an individual would taste the two cans, of soda the can at room temperature would taste very “flat” since more of the “fizz” which is carbon dioxide has escaped the can. Boiled water also tastes “flat” because more of the oxygen gas has been removed.

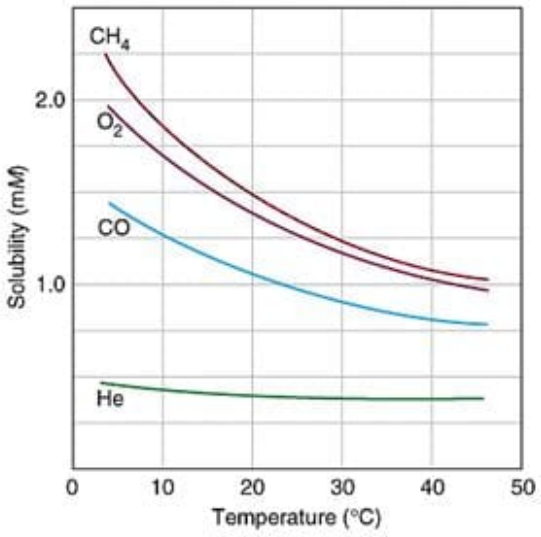


Figure 16.4.3. *A graph of the solubility of gases versus temperature, showing that as the temperature of the solution increases the solubility of a gas decreases.*[4]

Thermal pollution occurs when the temperature of a natural body of water, such as, river, lake or ocean is altered. This alters the natural environment and impacts the organisms that live or rely on an aquatic ecosystem. Thermal pollution is associated with electrical power plants and industrial factories that that use fossil fuels or nuclear power; this results in the release of large amounts of excess thermal energy.[5] Water absorbs the energy and is converted to steam, which is used to turn turbines and generate electricity. As some point, the water will need to be dumped back into nearby river, lake, or ocean and if not properly cooled, it can increase the temperature of the natural body of water. Thermal pollution can also involve putting colder the normal water into a nearby body of water.

Elevated water temperatures can also arise from deforestation and urbanization which causes changes in the landscape surrounding a body of water. When trees surrounding a body of water are removed, the now unshaded water can increase its temperature by as much as 10 ^{o}C. Thermal pollution can even result when the landscape away from the body of water is altered. These changes in landscape may cause erosion, changes in water volume, and pollutants that may change the composition of the water. For example, muddy water absorbs more energy from the sun than clear water.

Every response of an aquatic organism from incubation of the egg, to feeding activity, digestive and metabolic processes, reproduction, geographic distribution, and even survival occurs due to the thermal range by their immediate environment. The solubility of gases in water plays a major role in the survival of ectotherms, such as fish. Since an increase in temperature contributes to a decline in the solubility of gases in water; water that is of a higher temperature lacks oxygen. Therefore, many fish are only capable of existing in waters of a lower temperature where ample oxygen exists. Another problem that may occur due to thermal pollution is a limited metabolic range of many aquatic organisms. Many living organisms have metabolic enzymes that control the oxidation of organic food matter. Metabolic enzymes can only function in a narrow range of temperatures, and most enzymes show an optimum temperature in which they reach maximum rate of catalytic activity.[5] Small changes in temperature

can prove to be deadly for fish. Thermal pollution can also decrease the biodiversity of an ecosystem and allow for the creation of environment that can be invaded by alien aquatic species that may disrupt or destroy an entire ecosystem.[5]

Sea turtles in the San Diego Bay are example of how temperature may have affected the metabolic rate of an organism. In 1956, an electric power plant was built near the San Diego Bay, and is thought to be the reason for a colony of East Pacific green turtles living in the Bay. These turtles chose this location as their habitat; in spite of the fact, that it is one of the busiest harbors in the world. The power plant discharges the water that is 15 degrees above the normal temperature of the water in the Bay.[6] Researchers think that the warm water is the attraction for the turtles but due a decrease in oxygen in the water, the metabolic rate of the turtles is increased. The Eastern Pacific green turtles in this region are supersized, with some weighing as much as 550 pounds.[6] The sizes of these turtles are unprecedented in their enormity and are almost the double the normal size records of the same species in other habitats.[6] Carl Mayhugli, a Prescott College graduate, is working on his Masters on the movements and habitats of these turtles. He theorizes that the turtles are able to reach such massive size because the warmer water allows them to continue to forage for food when most sea turtles due to cooler water enter a state of hibernation called torpor.[6] The warm water allows eelgrass and algae to grow which provides a supplement to the turtle's normal diet of tube worms, sea hares, jellyfish and other slow-moving soft-bodied sea creatures.[6] The massive turtles are able to increase their food intake by increasing their herbivory diet. The power plant is set to be decommissioned in the next several years but the hope is they will continue to keep the bay as their nesting location.

From ChemPRIME: 10.15: Saturated and Supersaturated Solutions

References

1. Brown, T. Lemay, H., and Bursten, B. Chemistry the Central Science, 10th ed.; Pearson Education, Inc.: Upper Saddle River, NJ, 2006.

2. Solubility Curve image. http://www.chem.ufl.edu/~itl/2045/lectures/lec_i.html (accessed July 13, 2011).

3. Crystallization image. http://www.chem.ufl.edu/~itl/2045/lectures/lec_i.html (accessed July 13, 2011).

4. Solubility of Gases image. http://webserver.dmt.upm.es/~isidoro.../Exercise2.htm (accessed July 13, 2011).

5. EoEarth Organization. http://www.eoearth.org/article/Thermal_pollution?topic=49471(accessed July 13, 2011).

6. Bahnsen, C. "Saving the Green Giants" The Environmental Magazine 2009, August 31, 14-16.

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 16.4: Solubility and Pollution is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, and Adam Hahn.

16.5: Radiation Contamination from Fukushima

The 2011 Tōhoku earthquake and tsunami was the cause of a nuclear disaster at the Fukushima Daiichi nuclear power plant. The result was the release of radioactive isotopes. The total amount of iodine-131 and caesium-137 released into the atmosphere has been estimated to exceed 10% of the emissions from the Chernobyl disaster. [1][2] The radioactive isotopes were also released into the Pacific Ocean. The Japanese government has been unable to control the spread of radioactive material into the nation's food. Radioactive material has been detected in produce, tea leaves, milk, fish and beef. [3] As of June 2011, the International Atomic Energy Agency (IAEA) has not confirmed long term health effects to any person that had been reported as a result of radiation exposure from the nuclear accident. [4]

Emissions

The primary releases of radioactive nuclides have been iodine-131 and caesium-137. [5] However, strontium [6] and plutonium [7][8] have also been found. These elements have been found in the air, groundwater, ocean and agricultural products. [9] Iodine-131 has a half-life of 8 days while caesium-137 has a half-life of over 30 years. [10]

Iodine-131

The Austrian Meteorological Service report estimated the total amount of I-131 radiation released based on extrapolating data from several days of observation at worldwide radionuclide measuring facilities in Freiburg, Germany; Stockholm, Sweden; Takasaki, Japan and Sacramento, California. During the first 10 days of the accident. [11] The report's estimates of total I-131 ranged from 10 PBq to 700 PBq. [11]

Cesium-137

The Austrian Meteorological Service report estimated the total amount of caesium-137 radiation released based on extrapolating data from several days of observation at radionuclide measuring facilities in Freiburg, Germany; Stockholm, Sweden; Takasaki. Japan and Sacramento, California. The agency estimated an average being 5,000 TBq daily. [11] Over the course of the disaster, Chernobyl put out a total of 85,000 TBq of caesium-137. [1]

Air

The United Nations predicted that the initial radiation plume from the stricken Japanese reactors would reach the United States. Nuclear experts reported that radioactive isotopes in the plume would be in low concentrations and would have extremely minor health consequences in the United States. [12] Radioactive fallout with the isotopes of iodine-131, iodine-132, tellurium-132, iodine-133, caesium-134 and caesium-137 was detected in air filters at the University of Washington, Seattle, USA. [13] Westerly winds were dominant during most of the first week of the accident, depositing most of the radioactive material out to sea and away from the Japanese population centers, with some unfavorable wind directions depositing radioactive material over Tokyo. However, Tokyo did not receive levels of radiation dangerous to human health.

Water

TEPCO estimated that 520 tons radioactive water leaked into the sea before leaks in a pit in unit 2 were plugged, releasing 2,800 TBq of I-131, 940 TBq of Cs-134, 940 TBq of Cs-137. [14] The regulatory safe level for iodine-131 and caesium-137 in drinking water in Japan are 100 Bq/kg and 200 Bq/kg respectively. [15] The Japanese science ministry said that radioactive substances were detected in tap water in Tokyo, as well as Tochigi, Gunma, Chiba and Saitama prefectures. [16] IAEA reported that drinking water in Tokyo, Fukushima and Ibaraki had been above safe limits. [15] Tokyo drinking water exceeded the safe level for infants, prompting the government to distribute bottled water to families with infants. [17] Iodine-131 and was detected in 12 of 47 prefectures, of which the level in Tochigi was the highest at 110 Bq/kg. Caesium-137 was detected in 6 prefectures but always below 10 Bq/kg. [15] On 27 April, radiation levels in Tokyo's water supply fell to undetectable levels. [18]

Agricultural Products

The Japanese Ministry of Health, Labor and Welfare announced that levels of radioactivity exceeded legal limits and had been detected in milk produced in the Fukushima area. In Ibaraki, the IAEA confirmed that in some areas, iodine-131 was in milk and in freshly grown leafy vegetables, such as spinach and spring onions, therefore consumption was restricted. Iodine-131 was reported above safe levels and caesium-137 was reported at safe levels. Detection was reported in Ibaraki prefecture. [15] On 21 March,

levels of radioactivity in spinach grown in open air in Kitaibaraki city in Ibaraki, around Samples from Chiba, Ibaraki and Tochigi also had too high levels in celery, parsley, spinach and other leafy vegetables. In addition, certain samples of beef mainly showed concentrations of iodine-131 and/or caesium-134 and caesium-137 above the regulatory levels. [15]

From ChemPRIME: 19.0: Prelude to Nuclear Chemistry

References

1. Fukushima radioactive fallout nears Chernobyl levels". New Scientist. 24 March 2011. http://www.newscientist.com/article/dn20285-fukushima-radioactive-fallout-nears-chernobyl-levels.html. Retrieved 24 April 2011.

2. Peter Grier (April 26, 2011). "Was Chernobyl really worse than Fukushima?". The Christian Science Monitor.

3. Hiroko Tabuchi (July 18, 2011). "Radiation-Tainted Beef Spreads Through Japan's Markets". New York Times. http://www.nytimes.com/2011/07/19/world/asia/19beef.html?_r=1.

4. IAEA INTERNATIONAL FACT FINDING EXPERT MISSION OF THE FUKUSHIMA DAI-ICHI NPP ACCIDENT FOLLOWING THE GREAT EAST JAPAN EARTHQUAKE AND TSUNAMI". IAEA. June 2011. http://www-pub.iaea.org/MTCD/Meetings/PDFplus/2011/cn200/documentation/cn200_Final-Fukushima-Mission_Report.pdf. Retrieved 28.06.2011.

5. http://content.usatoday.com/communities/ondeadline/post/2011/03/report-radioactive-emissions-from-japan-plant-approach-chernobyl-levels/1

6. Govt to monitor radioactive strontium levels". NHK. 9 May 2011. http://www3.nhk.or.jp/daily/english/09_28.html. Retrieved 27 May 2011.

7. TEPCO says plutonium found on quake-damaged plant grounds

8. http://translate.google.com/translate?js=n&prev=_t&hl=en&ie=UTF-8&layout=2&eotf=1&sl=ja&tl=en&u=http%3A%2F%2Fjbpress.ismedia.jp%2Farticles%2I%2F7890%3Fpage%3D2 High concentrations of plutonium from the soil of rice field 50 km or more distant primary

9. http://www.abc.net.au/news/stories/2011/03/12/3162223.htm Radioactive steam released from nuclear plants

10. http://www-pub.iaea.org/MTCD/publica...S-2009_web.pdf INES - The International Nuclear and Radiological Event Scale pg. 16

11. Accident in the Japanese NPP Fukushima: Spread of Radioactivity" (PDF). http://www.zamg.ac.at/docs/aktuell/Japan2011-03-24_1600_E.pdf. Retrieved 24 April 2011.

12. Scientists Project Path of Radiation Plume By William J. Broad, 16 March 2011, The New York Times

13. Fission Products in Seattle Reveal Clues about Japan Nuclear Disaster". Technology Review, MIT. 25 March 2011. http://www.technologyreview.com/blog/arxiv/26571/.

14.. http://www.tepco.co.jp/en/press/corp-com/release/betu11_e/images/110421e2.pdf Outflow of radioactive water from the concrete crack near intake canal for Unit 2 of Fukushima Daiichi Nuclear Power Station

15. Japan Earthquake Update". IAEA. 19 March 2011. http://www.iaea.org/newscenter/news/2011/fukushimafull.html.

16. Tokyo tap water has trace level of radiation". The Japan Times. 20 March 2011. http://search.japantimes.co.jp/cgi-b...20110320a2.html. Retrieved 20 March 2011.

17. All clear on tap water, update". World Nuclear News. 24 March 2011. http://www.world-nuclear-news.org/RS_All_clear_on_tap_water_2403111.html

18. http://www.bloomberg.com/news/2011-04-27/tokyo-water-radiation-falls-to-zero-for-first-time-since-crisis.html http://www.bloomberg.com/news/2011-04-27/tokyo-water-radiation-falls-to-zero-for-first-time-since-crisis.html

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 16.5: Radiation Contamination from Fukushima is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, and Adam Hahn.

16.6: pH of Chesapeake Bay

pH is measured on a scale of 1 – 14 with 7 being neutral; (see figure 1.) The lower the number the more acidic a solution is, the higher the number the more basic it is. What does acidic or basic mean? Acidity means there are more hydronium ions than hydroxide ions and vice versa for basic solutions. Why care about pH and pOH? All organisms live in certain pH and pOH levels. Your body tends to be between a pH of 7.35 and 7.45. When the pH is higher or lower than the preferred range, your body does not function well. If your body is acidic (lower than pH of 7,) you can develop problems such as inflammation and wearing away of body tissue. This can lead to cancer, heart disease, or diabetes. Your body prefers a slightly basic (higher than 7 pH) environment. But if it gets too high, it can cause digestive and respiration problems. There is a great balancing act happening. The same is happening in the Chesapeake Bay.

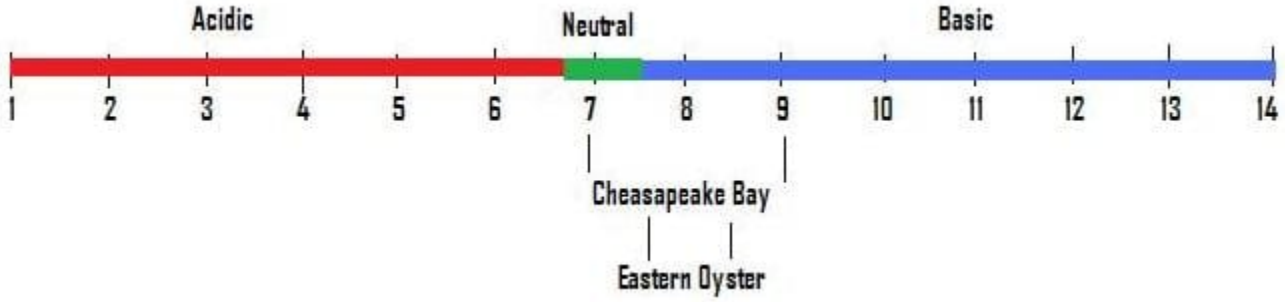


Figure 16.6.1: *pH Scale*

The Chesapeake tends to be between a pH of 7 and 9. When extra nutrients from farming (such as from fertilizers) leak into the water supply, algae can thrive and bloom (see Figure 2.) This causes an imbalance in the environment and creates a more alkaline (basic) pH water environment. A pH increase of 1 is more significant than it seems. pH is a log 10 scale, so a pH of 7 has a concentration of 10^{-7} hydronium ions. A pH of 8 has a concentration of 10^{-8}. This means that a pH of 8 has 10 times less [one tenth] hydronium ions as a pH of 7 and a pH of 9 has 100 times less [one hundredth] hydronium ions than a pH of 7. So, there are consequences with just a change of 1 pH unit. To figure out the pH and pOH levels from ion concentration, use the following formulas. pH = -log (H_3O^+) and pOH = -log (OH^-).



Figure 16.6.2: *Algae Bloom* http://drwaters.org/?itemCategory=33533&siteid=300&priorId=0

The Chesapeake Bay is home to the eastern oyster and other organisms. The eastern oyster can spawn in water with pH between

7.8 and 8.2 and had difficulties in pH that is lower than 6 or higher than 10. Larvae mortality rates increased rapidly when the pH is in the 9.00 to 9.50 range. The pH also affects respiration rate and pumping rate of water. When algae blooms occur, the oyster population suffers.



Figure 16.6.3: *Eastern Oyster from the Chesapeake Bay* http://www.chesapeakebay.net/fieldguide/critter/eastern_oyster

Practice Questions

Directions: Figure out the pH and pOH levels and answer the following questions.

1)What is the pH if the hydronium concentration is 10^{-5}?

pH = -log (H_3O^+)

pH = -log (10^{-5})

pH = 5

2)Would the Eastern Oyster survive in this pH? Explain.

The Eastern Oyster population would not survive in a pH of 5. The Eastern oyster needs a pH of 7.8 – 8.2 to spawn and anything under a pH of 6 makes the respiration rate decrease with spawning nearly impossible. This means that the oyster population would not increase and the oysters that were there would die off slowly due to not being able to function properly in that low of a pH.

3)What would the pOH be if the hydroxide concentration is 10<su>-8</sup>?

```
pOH = -log (OH)
```

pOH = -log (10^{-8})

pOH = 8

4)If pOH = 14 – pH, what is the pH level from your answer to number 3?

pOH = 14 – pH

8 = 14 – pH

pH + 8 = 14

pH = 14 – 8

pH = 6

5)Would the Eastern Oyster survive in this pH? Explain.

It would be possible for the Eastern Oyster to survive since they have more difficulties when the pH is lower than 6. They would be stressed, but not to the point of being unable to reproduce. This is the lowest the pH could be before the population would really start to decline. This could potentially be the “tipping point” of the population.

From ChemPRIME: 14.2: pH and pOH

References

http://www.acidalkalinediet.com/alkaline-diet/how-to-control-your-ideal-body-ph
http://www.dnr.state.md.us/bay/monit...ter/index.html
http://www.dnr.state.md.us/irc/docs/00000260_03.pdf

Contributors

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

This page titled 16.6: pH of Chesapeake Bay is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Ed Vitz, John W. Moore, Justin Shorb, Xavier Prat-Resina, Tim Wendorff, and Adam Hahn.

16.7: 18-F fluorodeoxyglucose and Atherosclerotic Plaque

This is a final project completed by Janet Jimenez for Chemistry 150 in Fall 2021 at Modesto Junior College. When citing this project, she is the author.

[^{18}F] fluorodeoxyglucose and Atherosclerotic Plaque

Radioactive elements are known to be dangerous and deadly, but what if I told you that they could also save lives. Ironic, right? By using the appropriate amount and proper handling, radioactive elements are of great help when trying to detect life threatening diseases in humans. An example is the positron emission tomography (PET) scan. The PET scan is one of few imaging techniques offered within the medical community that allow health care providers to see any abnormalities within the patient's body. All of this is possible because of a radioactive drug: Fluorodeoxyglucose. In this final project, we will see how chemistry is an active participant in nuclear medicine, and how the PET scan came to be. This project will primarily focus on how the PET scan is used to detect atherosclerotic plaque inflammation, something that was not thought of doing before.

The PET scan has been used by many health care providers to detect early signs of cancer, brain disorders, and heart disease. The way the PET scan works is by the use of a radioactive tracer called [^{18}F] fluorodeoxyglucose (Figure 1). If we take a closer look at its composition we see that it is made of $C_6H_{11}{}^{18}FO_5$. The carbon, hydrogen, and oxygen seem to look "normal" in the molecular formula, except for Fluorine. This is because the Fluorine represented in the molecular formula and in Figure 1, is an isotope of Fluorine (a radioactive one).

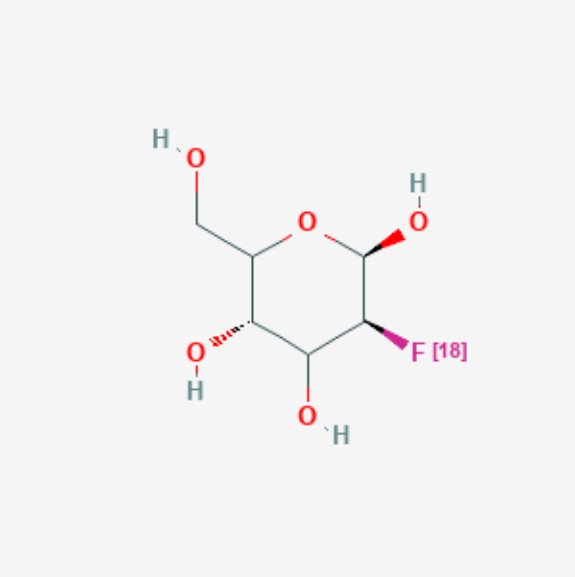


Figure 1 (National Center for Biotechnology Information)

However, to achieve the "stable" molecular compound listed in Figure 1, it must be synthesized. Figure 2 demonstrates how it is possible to achieve [^{18}F] fluorodeoxyglucose. Through a series of rearrangements, and the addition of acid and bases, the final product is formed. This product is now safe for humans use.

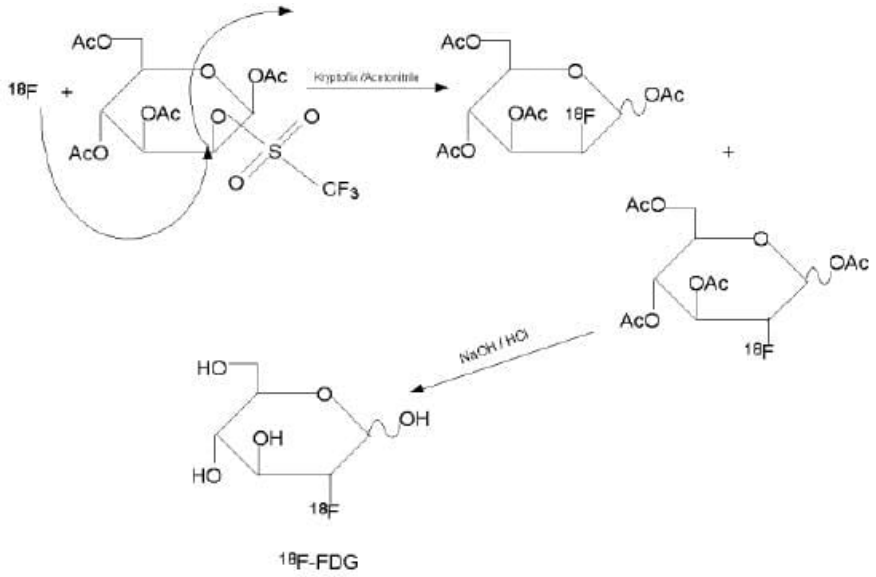


Figure 2 (Review of 18F-FDG Synthesis and Quality Control)

So now keeping the pictures above in mind, we now wonder, how does [^{18}F] fluorodeoxyglucose (^{18}FDG) interact with the human body? Well first you have to begin by knowing how it is introduced into the body. A very small amount of this radioactive substance is injected into the body. The body is then given some time to process this substance. From this point on, it is where the chemistry begins. "[^{18}F]-fluorodeoxyglucose (^{18}FDG) is a glucose analogue that is taken up by cells in proportion to their metabolic activity" (Rudd, 2002). Because of ^{18}FDG 's composition, cells are able to interact with it. "FDG is facilitated by the relatively long half-life of the ^{18}F isotope (109 min), which allows for tracer transport and typical whole body imaging... [it] is [then] transported from the plasma into cells... [where] FDG is phosphorylated, preventing its return to the circulation. At this point FDG is trapped within the cell and is not metabolized further" (Rudroff, 2014). Think of ^{18}FDG as a glowing donut; cells that have a lot of appetite (insatiable) will eat ^{18}FDG faster than those that are satiable. Now, the radioactive tracer has moved through your bloodstream, finding its way through the body, interacted with cells, and now has to display it has found its target (Figure 3).

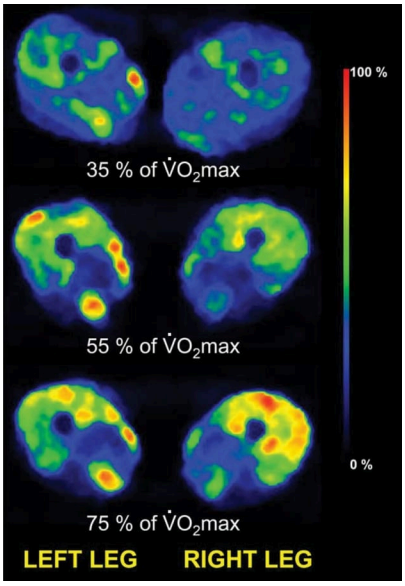


Figure 3

^{18}FDG has now made its way into an area (carotid artery) where there are a lot of insatiable cells. So, now there is a focused amount of ^{18}FDG at the carotid artery. It is here that the radioactive tracer begins detecting for atherosclerotic plaque

inflammation in the cells. As it detects the metabolic activity of the cells, an image is now displayed. This image then shows the area of interest.

The bright contrast is the metabolic activity of the cells around the plaque in the carotid artery. It was seen that there was a "high ^{18}FDG uptake in the brain, jaw muscles, and facial soft tissues" (Rudd, 2002). Thus, showing that a PET scan is able to detect the cells of interest.

When comparing individually the PET scan and the CT, the PET scan shows a more focalized view of the affected area. "The lower row (from left to right) demonstrates a low level of ^{18}FDG uptake in an asymptomatic carotid stenosis. The black arrow highlights the stenosis on the CT angiogram, and the white arrows demonstrate minimal ^{18}FDG accumulation at this site on the ^{18}FDG-PET and co-registered PET/CT images" (Rudd, 2002).

Throughout this project, radioactive tracers are explained and show how their involvement in nuclear medicine is crucial. Specifically the role of radioactive tracers in detecting atherosclerotic plaque inflammation. [^{18}F] fluorodeoxyglucose was shown to be an effective radioactive tracer and a more competent method of imaging in comparison to other imaging programs/ procedures. It is through the understanding and application of chemical properties that we are able to further advance within medicine. As we know, science is constantly evolving, chemistry is constantly evolving. We have advanced in ways we would have never imagined 50 years ago. So who knows, there just might be other effective ways of seeing/ prescreening for diseases that we have yet to discover.

Learning Questions:

1. What is an isotope? (Hint: you can find more information on isotopes in section 3.5 of the book)

2. Why is the radioactive tracer [^{18}F]-Fluorodeoxyglucose effective?

3. Suppose you need to do a PET scan on a patient, but the patient is diabetic. Would a PET scan be a good option? Why or why not?

References:

National Center for Biotechnology Information. "PubChem Compound Summary for CID 11469444, [18F]Fluorodeoxyglucose" *PubChem*,

https://pubchem.ncbi.nlm.nih.gov/com...rodeoxyglucose. Accessed 3 September, 2021.

Rudd, J.H.F., et al. "Imaging Atherosclerotic PLAQUE Inflammation with [18 f]-Fluorodeoxyglucose Positron Emission Tomography." *Circulation*, vol. 105, no. 23, 2002, pp. 2708–2711., doi:10.1161/01.cir.0000020548.60110.76.

Yu, S. "Review of 18F-FDG Synthesis and Quality Control." *Biomedical Imaging and Intervention Journal* 2 (2006): n. pag.

Rudroff, Thorsten, et al. "[18F]-Fdg Positron Emission Tomography—an Established Clinical Tool Opening a New Window into Exercise Physiology." *Journal of Applied Physiology*, vol. 118, no. 10, 2015, pp. 1181–1190., doi:10.1152/japplphysiol.01070.2014.

This page titled 16.7: 18-F fluorodeoxyglucose and Atherosclerotic Plaque is shared under a license and was authored, remixed, and/or curated by LibreTexts.

Student Project

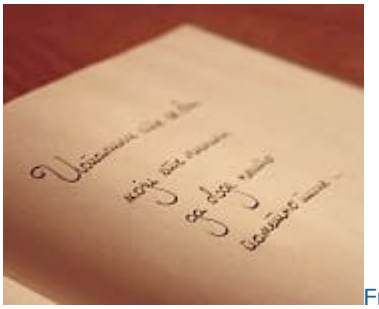
This page titled Student Project is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by LibreTexts.

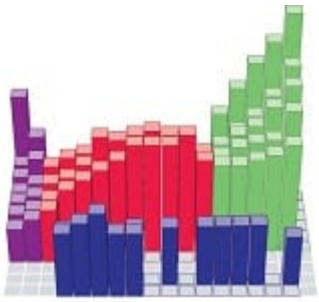
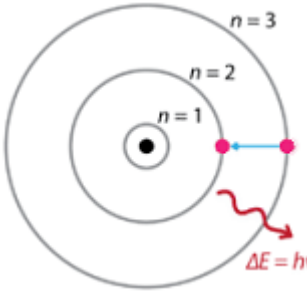
1: Unit 1

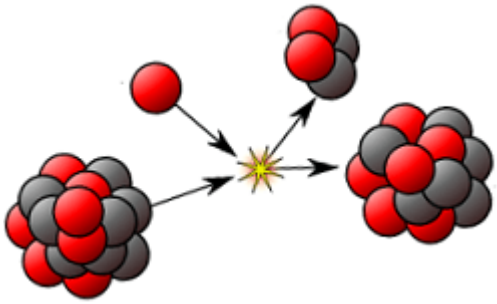
http://catdir.loc.gov/catdir/toc/eci...0060152l0.html

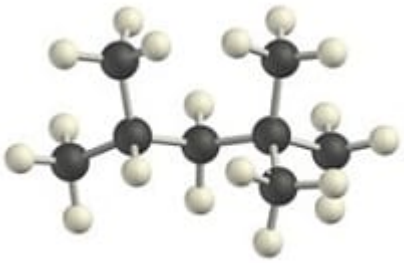
Chemistry is all around us "from the air we breathe to the food we eat" to the items at the supermarket that say “no chemicals added”. In fact, it is impossible to create something without using chemistry because chemistry consists of all matter. It allows us to answer questions as simple as why a candle goes out when a glass is placed over it to more complex questions such as does a candle actually burn in zero gravity? Chemistry is the study of matter and the changes it undergoes such as ice changing from the solid to liquid to gas phase. People have used it for things such as creating metal from an ore, dying fabric and making cheese. Chemistry deals with different substances and how they can interact with each other to create a product.

As you begin your study of college chemistry, those of you who do not intend to become professional chemists may well wonder why you need to study chemistry. You will soon discover that a basic understanding of chemistry is useful in a wide range of disciplines and career paths. You will also discover that an understanding of chemistry helps you make informed decisions about many issues that affect you, your community, and your world. A major goal of this text is to demonstrate the importance of chemistry in your daily life and in our collective understanding of both the physical world we occupy and the biological realm of which we are apart.

-  Front Matter
-  1: Introduction to Chemistry
-  2: The Mathematics of Chemistry

-  3: Atoms and the Periodic Table
- 

4: Valence Electrons and Bonding
-  5: Basics of Nuclear Science
-  6: Nuclear Weapons- Fission and Fusion
-  7: Nuclear Energy

-  8: Water Chemistry
-  9: Organic Chemistry
-  10: Pharmaceuticals
-  Back Matter

This page titled 1: Unit 1 is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Elizabeth Gordon.

- CHM101: Chemistry and Global Awareness (Gordon) by Elizabeth Gordon is licensed CC BY-NC-SA 4.0.

INDEX

A

Abegg's rule
4.5: Lewis Dot and Bonding
absolute asymmetric synthesis
3.5: Isotopes
acid rain
8.1: Water Cycle
Amphetamine
10.10: Amphetamines
Analgesics
10.12: Analgesics
anion
3.6: The Importance of Ions to a Chemist
anions
4.4: Predicting Ion Charges
Applied research
1.3: Pure and Applied Research
Arrhenius acid
4.13: Arrhenius Acids and Bases
Arrhenius base
4.13: Arrhenius Acids and Bases
atomic emission
4.2: Understanding Atomic Spectra
atomic mass
3.4: Atomic Mass and Atomic Number
3.5: Isotopes
atomic number
3.4: Atomic Mass and Atomic Number

B

becquerel (units)
5.6: Measuring Radiation
Bohr Diagram
4.1: Atomic Models of the Twentieth Century
Bohr's model of the hydrogen atom
4.1: Atomic Models of the Twentieth Century
Brachytherapy
5.9: Therapeutic Radiation
breeder reactor
7.11: New Generation Nuclear Reactors

C

cathode ray tube
3.3: Subatomic Particles - Electrons, Protons, and Neutrons
cation
3.6: The Importance of Ions to a Chemist
cations
4.4: Predicting Ion Charges
chain reaction
6.4: The Manhattan Project - Critical Mass and Bomb Construction
Chemotherapy
10.1: Chemotherapy
combination reaction
11.5: Some Types of Chemical Reactions
combustion reaction
11.5: Some Types of Chemical Reactions
concentration units
8.5: Water Concentration Units
conversion factor
2.2: Dimensional Analysis
conversion factors
2.2: Dimensional Analysis
critical mass
6.4: The Manhattan Project - Critical Mass and Bomb Construction
CT Scanning
5.8: Using Nuclear Science to Diagnose Disease
curie (units)
5.6: Measuring Radiation

D

Dalton's Atomic Theory
3.1: Basic Atomic Theory
decomposition reaction
11.5: Some Types of Chemical Reactions
democritus
3.1: Basic Atomic Theory
density
2.3: Density
drinking water
8.2: EPA's Primary and Secondary Drinking Water Standards

E

electromagnetic spectrum
5.2: The Electromagnetic Spectrum
Electrons, Protons and Neutrons
3.3: Subatomic Particles - Electrons, Protons, and Neutrons
enriched uranium
3.5: Isotopes
eutrophication
8.4: Chemical Contamination of Water
Exemplar
16.3: Sorting Recyclable Plastics by Density
16.4: Solubility and Pollution
16.5: Radiation Contamination from Fukushima
16.6: pH of Chesapeake Bay
External Beam Therapy
5.9: Therapeutic Radiation

F

fireworks
4.2: Understanding Atomic Spectra
fission
6.1: The History and Basics of Fission
fluoridation
8.7: Fluoridation of Water
free radical
4.10: Free Radicals and Health
4.9: Free Radicals and the Environment
fusion
6.1: The History and Basics of Fission
fusion bomb
6.7: Fusion

G

Graduated glassware
2.1: Measuring Matter
group
3.2: Development of the Modern Periodic Table

H

half life
5.7: Calculating Half-Life
hard water
8.4: Chemical Contamination of Water
hydrologic cycle
8.1: Water Cycle

I

IAEA
7.1: Sharing and Monitoring Nuclear Technology
ionizing power
5.4: Ionizing Radiation and Non-ionizing Radiation
ionizing radion
5.4: Ionizing Radiation and Non-ionizing Radiation
ions
3.6: The Importance of Ions to a Chemist
isotope
3.5: Isotopes
Isotopic Tests
5.8: Using Nuclear Science to Diagnose Disease

L

law of conservation of mass
11.2: The Law of Conservation of Matter
Lewis diagram
4.5: Lewis Dot and Bonding
4.8: Covalent Bonding and Formula Writing

M

Manhatten project
6.4: The Manhattan Project - Critical Mass and Bomb Construction
Material Safety Data Sheet (MSDS)
1.5.1: Applying Properties - MSDS/SDS
Maximum Contaminant Level (MCL)
8.2: EPA's Primary and Secondary Drinking Water Standards
Maximum Contaminant Level Goal (MCLG)
8.2: EPA's Primary and Secondary Drinking Water Standards
metal poisoning
4.3: Atomic Spectroscopy Applications
MSDS
1.5.1: Applying Properties - MSDS/SDS

N

neurotransmitters
10.8: Basics of Neurotransmitters

nonionizing radiation
5.4: Ionizing Radiation and Non-ionizing Radiation
nuclear fallout
6.8: Fallout and Weapons Testing
nuclear waste
7.10: Storing and Transporting Nuclear Waste Materials

O

Oconee Nuclear Power Plant
7.6: South Carolina Energy
octet rule
4.5: Lewis Dot and Bonding

P

penetration power
5.4: Ionizing Radiation and Non-ionizing Radiation
periodic table
1.6: All About the Elements
3.2: Development of the Modern Periodic Table
PET scan
5.8: Using Nuclear Science to Diagnose Disease
pharmaceuticals
10: Pharmaceuticals
plum pudding model
4.1: Atomic Models of the Twentieth Century
proton therapy
5.9: Therapeutic Radiation
Pure research
1.3: Pure and Applied Research

R

radioactive dating
5.7: Calculating Half-Life
radioactivity
5.1: The Discovery of Radiation
relative abundance
3.5: Isotopes
Roengten
5.1: The Discovery of Radiation
roentgen (units)
5.6: Measuring Radiation

S

scientific calculator
2.1: Measuring Matter
scientific method
1.2: Using the Scientific Method
scientific notation
2.1: Measuring Matter
scintillation counter
5.6: Measuring Radiation
SI units
2.1: Measuring Matter
sievert (units)
5.6: Measuring Radiation
solubility
4.11: Applications and Solubility of Covalent Compounds
4.7: Applications and Solubility of Ionic Compounds
Solubility of Covalent Compounds
4.11: Applications and Solubility of Covalent Compounds
stoichiometry
11.4: Quantitative Relationships Based on Chemical Equations

T

temperature
2.4: Temperature
Therapeutic Radiation
5.9: Therapeutic Radiation
Three Mile Island Nuclear Accident
7.7: Three Mile Island Nuclear Accident
transuranium elements
6.1: The History and Basics of Fission

V

volumetric glassware
2.1: Measuring Matter

W

water contamination
8.4: Chemical Contamination of Water
water cycle
8.1: Water Cycle
water treatment
8.6: Sanitation of Drinking Water

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

INDEX

A

Abegg's rule
4.5: Lewis Dot and Bonding

absolute asymmetric synthesis
3.5: Isotopes

acid rain
8.1: Water Cycle

Amphetamine
10.10: Amphetamines

Analgesics
10.12: Analgesics

anion
3.6: The Importance of Ions to a Chemist

anions
4.4: Predicting Ion Charges

Applied research
1.3: Pure and Applied Research

Arrhenius acid
4.13: Arrhenius Acids and Bases

Arrhenius base
4.13: Arrhenius Acids and Bases

atomic emission
4.2: Understanding Atomic Spectra

atomic mass
3.4: Atomic Mass and Atomic Number
3.5: Isotopes

atomic number
3.4: Atomic Mass and Atomic Number

B

becquerel (units)
5.6: Measuring Radiation

Bohr Diagram
4.1: Atomic Models of the Twentieth Century

Bohr's model of the hydrogen atom
4.1: Atomic Models of the Twentieth Century

Brachytherapy
5.9: Therapeutic Radiation

breeder reactor
7.11: New Generation Nuclear Reactors

C

cathode ray tube
3.3: Subatomic Particles - Electrons, Protons, and Neutrons

cation
3.6: The Importance of Ions to a Chemist

cations
4.4: Predicting Ion Charges

chain reaction
6.4: The Manhattan Project - Critical Mass and Bomb Construction

Chemotherapy
10.1: Chemotherapy

combination reaction
11.5: Some Types of Chemical Reactions

combustion reaction
11.5: Some Types of Chemical Reactions

concentration units
8.5: Water Concentration Units

conversion factor
2.2: Dimensional Analysis

conversion factors
2.2: Dimensional Analysis

critical mass
6.4: The Manhattan Project - Critical Mass and Bomb Construction

CT Scanning
5.8: Using Nuclear Science to Diagnose Disease

curie (units)
5.6: Measuring Radiation

D

Dalton's Atomic Theory
3.1: Basic Atomic Theory

decomposition reaction
11.5: Some Types of Chemical Reactions

democritus
3.1: Basic Atomic Theory

density
2.3: Density

drinking water
8.2: EPA's Primary and Secondary Drinking Water Standards

E

electromagnetic spectrum
5.2: The Electromagnetic Spectrum

Electrons, Protons and Neutrons
3.3: Subatomic Particles - Electrons, Protons, and Neutrons

enriched uranium
3.5: Isotopes

eutrophication
8.4: Chemical Contamination of Water

Exemplar
16.3: Sorting Recyclable Plastics by Density
16.4: Solubility and Pollution
16.5: Radiation Contamination from Fukushima
16.6: pH of Chesapeake Bay

External Beam Therapy
5.9: Therapeutic Radiation

F

fireworks
4.2: Understanding Atomic Spectra

fission
6.1: The History and Basics of Fission

fluoridation
8.7: Fluoridation of Water

free radical
4.10: Free Radicals and Health
4.9: Free Radicals and the Environment

fusion
6.1: The History and Basics of Fission

fusion bomb
6.7: Fusion

G

Graduated glassware
2.1: Measuring Matter

group
3.2: Development of the Modern Periodic Table

H

half life
5.7: Calculating Half-Life

hard water
8.4: Chemical Contamination of Water

hydrologic cycle
8.1: Water Cycle

I

IAEA
7.1: Sharing and Monitoring Nuclear Technology

ionizing power
5.4: Ionizing Radiation and Non-ionizing Radiation

ionizing radion
5.4: Ionizing Radiation and Non-ionizing Radiation

ions
3.6: The Importance of Ions to a Chemist

isotope
3.5: Isotopes

Isotopic Tests
5.8: Using Nuclear Science to Diagnose Disease

L

law of conservation of mass
11.2: The Law of Conservation of Matter

Lewis diagram
4.5: Lewis Dot and Bonding
4.8: Covalent Bonding and Formula Writing

M

Manhatten project
6.4: The Manhattan Project - Critical Mass and Bomb Construction

Material Safety Data Sheet (MSDS)
1.5.1: Applying Properties - MSDS/SDS

Maximum Contaminant Level (MCL)
8.2: EPA's Primary and Secondary Drinking Water Standards

Maximum Contaminant Level Goal (MCLG)
8.2: EPA's Primary and Secondary Drinking Water Standards

metal poisoning
4.3: Atomic Spectroscopy Applications

MSDS
1.5.1: Applying Properties - MSDS/SDS

N

neurotransmitters
10.8: Basics of Neurotransmitters

nonionizing radiation
5.4: Ionizing Radiation and Non-ionizing Radiation
nuclear fallout
6.8: Fallout and Weapons Testing
nuclear waste
7.10: Storing and Transporting Nuclear Waste Materials

O

Oconee Nuclear Power Plant
7.6: South Carolina Energy
octet rule
4.5: Lewis Dot and Bonding

P

penetration power
5.4: Ionizing Radiation and Non-ionizing Radiation
periodic table
1.6: All About the Elements
3.2: Development of the Modern Periodic Table
PET scan
5.8: Using Nuclear Science to Diagnose Disease
pharmaceuticals
10: Pharmaceuticals
plum pudding model
4.1: Atomic Models of the Twentieth Century
proton therapy
5.9: Therapeutic Radiation
Pure research
1.3: Pure and Applied Research

R

radioactive dating
5.7: Calculating Half-Life
radioactivity
5.1: The Discovery of Radiation
relative abundance
3.5: Isotopes
Roengten
5.1: The Discovery of Radiation
roentgen (units)
5.6: Measuring Radiation

S

scientific calculator
2.1: Measuring Matter
scientific method
1.2: Using the Scientific Method
scientific notation
2.1: Measuring Matter
scintillation counter
5.6: Measuring Radiation
SI units
2.1: Measuring Matter
sievert (units)
5.6: Measuring Radiation
solubility
4.11: Applications and Solubility of Covalent Compounds
4.7: Applications and Solubility of Ionic Compounds
Solubility of Covalent Compounds
4.11: Applications and Solubility of Covalent Compounds
stoichiometry
11.4: Quantitative Relationships Based on Chemical Equations

T

temperature
2.4: Temperature
Therapeutic Radiation
5.9: Therapeutic Radiation
Three Mile Island Nuclear Accident
7.7: Three Mile Island Nuclear Accident
transuranium elements
6.1: The History and Basics of Fission

V

volumetric glassware
2.1: Measuring Matter

W

water contamination
8.4: Chemical Contamination of Water
water cycle
8.1: Water Cycle
water treatment
8.6: Sanitation of Drinking Water

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.